ESSENTIAL CHEMISTRY FOR POISONS, POTIONS AND PHARMACEUTICALS



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Essential Chemistry for Poisons, Potions and Pharmaceuticals

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This text was compiled on 10/25/2026

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1: SCIENCE- ALCHEMY TO CHEMISTRY

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

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1.1: Prelude to Chemistry, Matter, and Measurement

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

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

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1.2: What is Chemistry?

Learning Objectives

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

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

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure 1.2.1). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

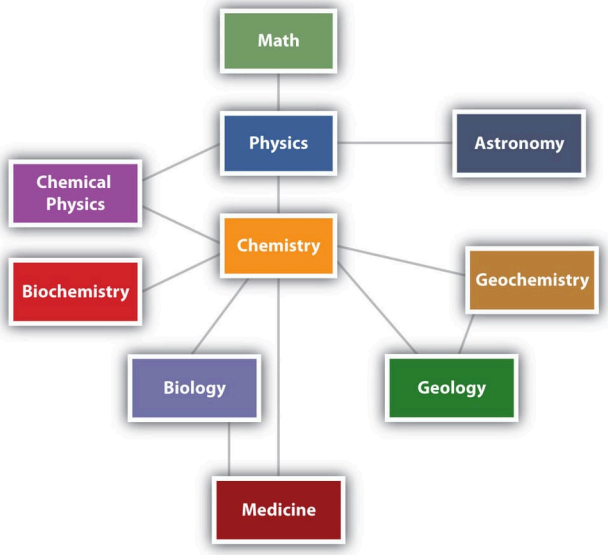


Figure 1.2.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

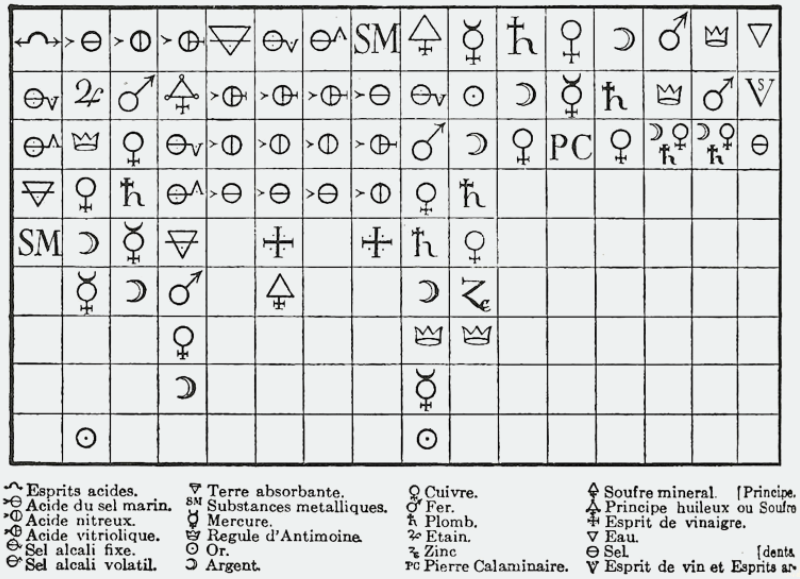
Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure 1.2.1 shows how many of the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

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

Alchemy

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

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



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

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

Exercise 1.2.1

Which fields of study are branches of science? Explain.

a. sculpture
b. astronomy

Answer a available at the end of this section.

Exercise 1.2.2

Which fields of study are branches of science?

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

Answer available at the end of this section.

How do scientists work? Generally, they follow a process called the scientific method. The scientific method is an organized

procedure for learning answers to questions and making explanations for observations. To find the answer to a question (for example, “Why do birds fly toward Earth’s equator during the cold months?”), a scientist goes through the following steps, which are also illustrated in Figure 1.2.2:

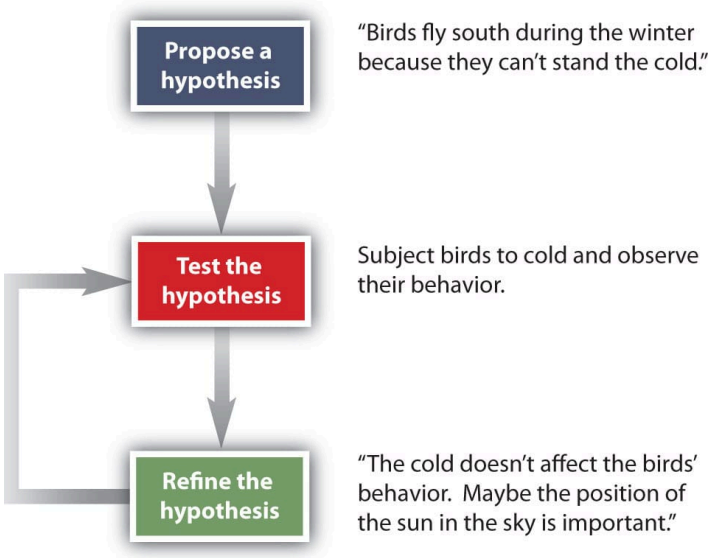


Figure 1.2.2: The General Steps of the Scientific Method. After an observation is made or a question is identified, a hypothesis is made and experiments are designed to test the hypothesis.

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

1. **Propose a hypothesis.** A scientist generates a testable idea, or hypothesis, to try to answer a question or explain an observation about how the natural universe works. Some people use the word *theory* in place of hypothesis, but the word hypothesis is the proper word in science. For scientific applications, the word theory is a general statement that describes a large set of observations and data. A theory represents the highest level of scientific understanding.
2. **Test the hypothesis.** A scientist evaluates the hypothesis by devising and carrying out experiments to test it. If the hypothesis passes the test, it may be a proper answer to the question. If the hypothesis does not pass the test, it may not be a good answer.
3. **Refine the hypothesis if necessary.** Depending on the results of experiments, a scientist may want to modify the hypothesis and then test it again. Sometimes the results show the original hypothesis to be completely wrong, in which case a scientist will have to devise a new hypothesis.

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

1.2.3Exerc

Define science and chemistry.

Answer available at the end of this section.

1.2.4Exerc

Name the steps of the scientific method.

Answer available at the end of this section.

Key Takeaways

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

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

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

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

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

3. Which of the following are examples of matter?

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

4. Which of the following are examples of matter?

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

5. Suggest a name for the science that studies the physics of rocks and the earth.

6. Suggest a name for the study of the physics of living organisms.

7. Engineering is the practical application of scientific principles and discoveries to develop things that make our lives easier. Is medicine science or engineering? Justify your answer.

8. Based on the definition of engineering in Exercise 7, would building a bridge over a river or road be considered science or engineering? Justify your answer.

9. When someone says, “I have a theory that excess salt causes high blood pressure,” does that person really have a theory? If it is not a theory, what is it?

10. When a person says, “My hypothesis is that excess calcium in the diet causes kidney stones,” what does the person need to do to determine if the hypothesis is correct?

11. Some people argue that many scientists accept many scientific principles on faith. Using what you know about the scientific method, how might you argue against that assertion?

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

Answers

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

2.

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

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

4.

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

5. geophysics

6. biophysics

7. Medicine is probably closer to a field of engineering than a field of science, but this may be arguable. Ask your doctor.

8. Engineering

9. In scientific terms, this person has a hypothesis.
10. Conduct experiments to determine if kidney stones contain calcium.
11. Science is based on reproducible facts, not blind belief.
12. No.

Answers

Exercise 1.2.1

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

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

Exercise 1.2.2

a, b and c only

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

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

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1.3: Science and Technology - The Roots of Knowledge

Learning Objectives

- Learn the definitions of science, chemistry, and technology.
- Know the principles of green chemistry.

Science is a practice of learning about the natural world. Natural sciences include physics, chemistry, biology, geology and astronomy. Science uses mathematics and logic, which are sometimes called "formal sciences". Natural science makes observations and experiments. Science produces accurate facts, scientific laws and theories. 'Science' also refers to the large amount of knowledge that has been found using this process.

People who study and research science are scientists. Scientists study things by looking at them very carefully, by measuring, and by doing experiments and tests. Scientists try to explain why things act the way that they do, and predict what will happen.

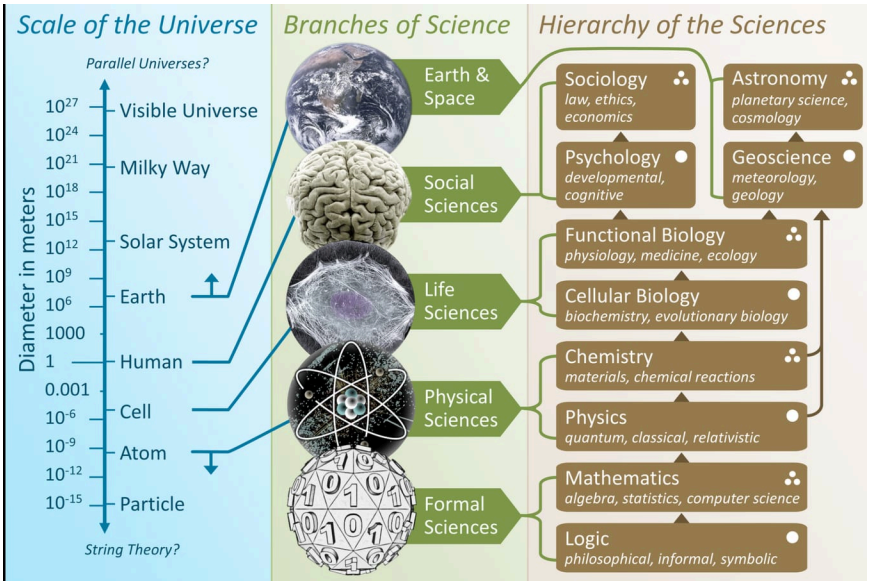


Figure 1.3.1: The scale of the universe mapped to the branches of science. (Source: Wikipedia)

Chemistry is the scientific discipline that involves the properties and behavior of matter, and the changes matter undergoes during a reaction with other substances.

In the scope of its subject, chemistry occupies an intermediate position between physics and biology. It is sometimes called *the central science*, because it provides a foundation for understanding both basic and applied scientific disciplines at a fundamental level (Figure 1.1.1). For example, chemistry explains aspects of plant chemistry (botany), the formation of igneous rocks (geology), how atmospheric ozone is formed and how environmental pollutants are degraded (ecology), the properties of the soil on the moon (astrophysics), how medications work (pharmacology), and how to collect DNA evidence at a crime scene (forensics).

Science is human understanding of how the universe and the things in it work. On the other hand, **technology** is the use of resources to solve a problem (such as knowledge, skills, processes, techniques, tools, and raw materials). Paleolithic technology is the oldest known, but people likely used technology long before then. The Paleolithic age began when hominids (early humans) started to use stones as tools for bashing, cutting, and scraping. The age ended when humans began to make small, fine tools (Mesolithic); it came to a close when they started to plant crops and develop other methods of agriculture (Neolithic).

Technology was once quite simple. It was passed on through word of mouth, until written word was invented. Writing allowed technology to develop much quicker. In the present day, people continue to understand more about the world and the universe. The use of the telescope by Galileo, Einstein's theory of relativity, lasers, and computing are all scientific discoveries. Technology is of great importance to science, medicine, and everyday life.

Green chemistry (also called **sustainable chemistry**) is a type of chemical research and engineering. It supports the design of products and processes that use the smallest amount of dangerous substances possible. In 1990, the Pollution Prevention Act was passed in the United States. This law sought new and original ways to handle pollution. The Pollution Prevention Act aims to avoid problems before they happen.

Green chemistry applies organic chemistry, inorganic chemistry, biochemistry, analytical chemistry, and even physical chemistry. While green chemistry seems to focus on just industrial applications, it does apply to other scientific disciplines. Green chemists aim to reduce the hazards and increase the efficiency of any chemical choice. Green chemistry is distinct from environmental chemistry, which focuses on chemical phenomena in the environment.

The 12 Principles of Green Chemistry

1. **Prevent waste.**
   - Create products without waste, or with minimal wastes, so that the wastes do not need to be taken care of afterwards.
2. **Design safer chemicals and products.**
   - Design chemicals to have little or no toxicity, without altering effectiveness.
3. **Design less hazardous chemical syntheses.**
   - Design a way of synthesizing products, without them being toxic to humans or the environment.
4. **Use renewable raw materials.**
   - Use renewable raw materials, like plant materials; rather than depleting materials, such as fossil fuels.
5. **Use catalysts, not stoichiometric reagents.**
   - Use catalysts, because of their ability to be able to be reused. Catalysts are less harmful than reagents.
6. **Avoid chemical derivatives.**
   - Chemical derivatives generate wastes that can be avoided.
7. **Maximize atom economy.**
   - Maximize the percentage of reactant atoms that convert to usable product atoms in a chemical reaction, so as to minimize or eliminate atomic waste.
8. **Use safer solvents and reaction conditions.**
   - Avoid using harsh solvents; if that cannot be avoided, then use benign chemicals.
9. **Increase energy efficiency.**
   - Use the normal ambient temperature and pressure wherever possible.
10. **Design for degradation.**
    - Design materials to break down into benign substances, by bacterial or other environmentally sound ways.
11. **Analyze in real time to prevent pollution.**
    - Monitor and control the formation of by-products during a reaction.
12. **Minimize the potential for accidents.**
    - Design chemicals to minimize the potential for accidents.

Summary

- Science is the human method to make discoveries about the natural world.
- Chemistry is the branch of science involved with the properties and behavior of matter, and the changes it undergoes during a reaction with other substances.
- Green chemistry applies organic chemistry, inorganic chemistry, biochemistry, analytical chemistry, and physical chemistry to reduce and prevent pollution.
- There are 12 guiding principles in the practice of green chemistry.

Contributors and Attributions

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1.4: Science- Reproducible, Testable, Tentative, Predictive, and Explanatory

Learning Objectives

- Describe the differences between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.
- Identify the components of the scientific method.

Although many have taken science classes throughout their course of studies, incorrect or misleading ideas about some of the most important and basic principles in science are still commonplace. Most students have heard of *hypotheses*, *theories*, and *laws*, but what do these terms really mean? Before you read this section, consider what you have learned about these terms previously, and what they mean to you. When reading, notice if any of the text contradicts what you previously thought. What do you read that supports what you thought?

What is a Fact?

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

What is a Hypothesis?

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

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

A hypothesis is tentative; it can be easily changed.

What is a Theory?

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

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

Note some key features of theories that are important to understand from this description:

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

Theories aren't likely to change.

What is a Law?

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

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

Laws vs Theories

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

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 1.4.1 *What is the difference between scientific law and theory?*

The Scientific Method

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

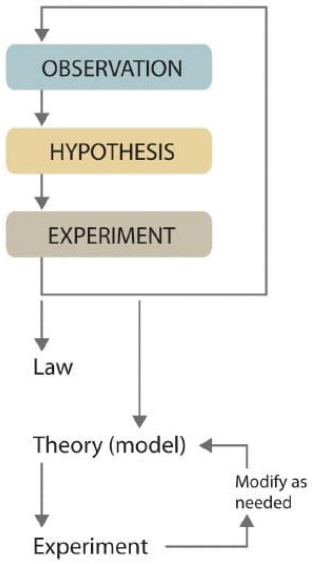


Figure 1.4.1 *The steps in the scientific method.*

Step 1: Make observations.

Observations can be qualitative or quantitative. **Qualitative observations** describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: "the outside air temperature is cooler during the winter season," "table salt is a crystalline solid," "sulfur crystals are yellow," and "dissolving a penny in dilute nitric acid forms a blue solution and a brown gas." **Quantitative observations** are measurements, which by definition consist of both a number and 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—the chemical name of which 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 which correspond to observed movements of the sun, moon, clouds, and shadows, is consistent with either of two hypotheses:

1. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun.
2. The sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. In the case of 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 supports or refutes 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 the case of validity, the scientist can proceed to step 5. In other cases, experiments may demonstrate that the hypothesis is incorrect or that it must be modified, thus requiring further experimentation.

Step 5: Development of 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 states what happens; it does not address the question of why.

One example of a law, the law of definite proportions (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. 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. A theory, in contrast, is incomplete and imperfect; it evolves with time to explain new facts as they are discovered.

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

Example 1.4.1

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

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

Solution

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

Exercise 1.4.1

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

1. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
2. Heat always flows from hot objects to cooler ones, not in the opposite direction.
3. The universe was formed by a massive explosion that propelled matter into a vacuum.
4. Michael Jordan is the greatest pure shooter ever to play professional basketball.
5. Limestone is relatively insoluble in water, but dissolves readily in dilute acid with the evolution of a gas.

Answer 1: available at the end of this section.

Summary

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

Contributors and Attributions

-

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Answers

Exercise 1.4.1

experiment
law
theory
hypothesis
observation

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1.5: Solving Society's Problems - Scientific Research

Learning Objective

- Know the meaning and examples of pure and applied research.

How did chemistry develop? What is happening in the field of chemistry today? What career can a chemistry degree apply to? These are all good questions that should be asked by students interested in chemistry. Research in chemistry (or any other field, for that matter) is interesting and challenging. But there are different directions a person can take as they explore research opportunities.

Types of Research

In science, two types of research are generally discussed: pure and applied. **Pure research** focuses on answering basic questions such as, "how do gases behave?" **Applied research** is involved in the process of developing specific preparations 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 (E*arth? Air? Fire? Water? All of the above?*). They weren't going to do anything with their knowledge—they just wanted to know.



Figure 1.5.1 *Ancient Greek philosophers.*

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

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 new products. Research on laundry detergents will probably not offer any new concepts of soap, but will help us develop materials that get our clothes cleaner, use less water, and create lower amounts of pollution.

A lot of research is done by petroleum companies. They want to find better ways to power vehicles, better lubricants to cut down on engine wear, and better ways to lower air pollution. These companies will use information that is readily available to come up with new products.



Figure 1.5.2 A *Gasoline pump*.

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 will 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 the development of treatments.

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 and Attributions

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1.6: Chemistry - A Study of Matter and Its Changes

Learning Objectives

- Separate physical from chemical properties.
- Label a change as chemical or physical.

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

Physical and Chemical Properties

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

Physical Property

A **physical property** is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets—a property called malleability. Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point. In Table 1.6.1, notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will “float.”

Table 1.6.1 Densities of Common Substances

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

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



Figure 1.6.1 A pencil (top) and a diamond ring (bottom). Both are a form of carbon, but exhibit very different physical properties.

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

Chemical Property

A chemical property of matter describes its "potential" to undergo some chemical change or reaction by virtue of its composition; as well as what elements, electrons, and bonding are present to give the potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, students of chemistry should be able to look at the formula of a compound and state a chemical property. For example, hydrogen has the potential to ignite and explode, given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with acids. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.



Figure 1.6.2 Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0;).

A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust (Figure 1.6.2). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used

in descriptions of chemical changes are *burn*, *rot*, *explode*, *decompose*, and *ferment*. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 1.6.2: Contrasting Physical and Chemical Properties

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

Example 1.6.1

Which of the following is a chemical property of iron?

1. Iron corrodes in moist air.
2. Density = 7.874 g/cm^3
3. Iron is soft when pure.
4. Iron melts at 1808 K.

Solution

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

Exercise 1.6.1A

Which of the following is a physical property of matter?

1. corrosiveness
2. pH (acidity)
3. density
4. flammability

Answer available at the end of this section.

Exercise 1.6.1B

Which of the following is a chemical property?

1. flammability
2. melting point
3. boiling point
4. density

Answer available at the end of this section.

Physical and Chemical Changes

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

Physical Change

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

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

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



Figure 1.6.3: Ice melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical, as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

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

Chemical Change

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

One good example of a chemical change is burning a candle. The act of burning paper actually results in the formation of new

chemicals (carbon dioxide and water, to be exact) from the burning of the wax. Another example of a chemical change is what occurs when natural gas is burned in a furnace. This time, on the left there is a molecule of methane, $\mathrm{CH_4}$, and two molecules of oxygen, $\mathrm{O_2}$; on the right are two molecules of water, $\mathrm{H_2O}$, and one molecule of carbon dioxide, $\mathrm{CO_2}$. In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.



Figure 1.6.4: The burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

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

Observations that help to indicate chemical change include:

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

Example 1.6.2

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

1. Boiling water.
2. A nail rusting.
3. A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
4. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

Solution

1. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written as: $\mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2O}\,(g)$
2. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $\mathrm{Fe} + \mathrm{O_2} \rightarrow \mathrm{Fe_2O_3}$

3. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been combined. No color *change,* or other evidence of chemical change, occurred.
4. Chemical: the formation of a precipitate and the color change from colorless to yellow indicates a chemical change.

Exercise 1.6.2

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

1. A mirror is broken.
2. An iron nail corroded in moist air.
3. Copper metal is melted.
4. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer 1: available at the end of this section.

Answer 2: available at the end of this section.

Answer 3: available at the end of this section.

Answer 4: available at the end of this section.

Summary

- A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points.
- A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change.
- Physical changes are changes that do not alter the identity of a substance.
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes. Observations that indicate a chemical change occurred include color change, temperature change, light given off, formation of bubbles, and formation of a precipitate.

Contributors and Attributions

-

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Answers

Exercise 1.6.1A

3.

Exercise 1.6.1B

1.

Exercise 1.6.2

physical change

Exercise 1.6.2

chemical change

Exercise 1.6.2

physical change

Exercise 1.6.2

chemical change

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

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

    a. Balsa is a very light wood.
    b. If held in a flame, magnesium metal burns in air.
    c. Mercury has a density of 13.6 g/mL.
    d. Human blood is red.

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

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

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

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

5. Give two examples of a heterogeneous mixture.

6. Give two examples of a homogeneous mixture.

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

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

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

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

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

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

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

    a. equal portions of salt and sand placed in a beaker and shaken up
    b. a combination of beeswax dissolved in liquid hexane
    c. hydrogen peroxide, a substance composed of hydrogen and oxygen

11. What word describes each phase change?

    a. solid to liquid
    b. liquid to gas
    c. solid to gas

12. What word describes each phase change?

a. liquid to solid
b. gas to liquid
c. gas to solid

Answers

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

2.

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

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

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

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

6. vinegar and rubbing alcohol (answers will vary)

7. a. element
   b. heterogeneous mixture
   c. compound

8.

   a. compound
   b. element
   c. heterogeneous mixture

9. a. solution
   b. solution
   c. compound

10.

   a. heterogeneous mixture
   b. solution
   c. compound

11.

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

12.

   a. freezing
   b. condensation
   c. deposition

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1.S: Chemistry, Matter, and Measurement (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.

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

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

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

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

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2: MEASUREMENTS

Learning Objectives

- Express quantities properly using a number and a unit.

A coffee maker's instructions tell you to fill the coffeepot with 4 cups of water and use 3 scoops of coffee. When you follow these instructions, you are measuring. When you visit a doctor's office, a nurse checks your temperature, height, weight, and perhaps blood pressure (Figure 2.1); the nurse is also measuring.

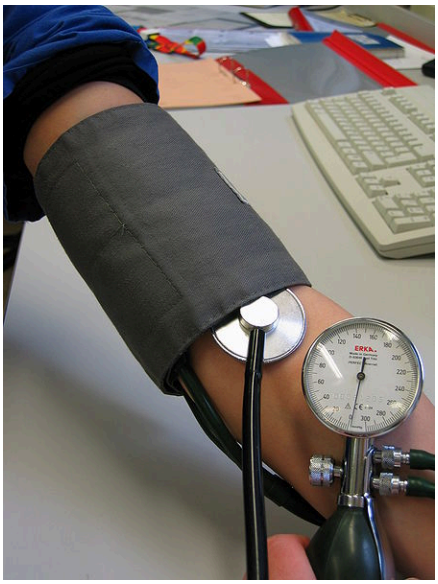


Figure 2.1: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GNU Free Documentation License; Pia von Lützau via Wikipedia).

Chemists measure the properties of matter using a variety of devices or measuring tools, many of which are similar to those used in everyday life. Rulers are used to measure length, balances (scales) are used to measure mass (weight), and graduated cylinders or pipettes are used to measure volume. Measurements made using these devices are expressed as quantities. A **quantity** is an amount of something and consists of a **number** and a **unit**. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5.2 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5.2.

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

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

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

> Both a number and a unit must be included to express a quantity properly.

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

Exercise 2.1

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

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

Answer a available at the end of this section.

Exercise 2.2

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

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

Answer a available at the end of this section.

Exercise 2.2

What are the two necessary parts of a quantity?

Answer available at the end of this section.

Key Takeaway

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

Answers

Exercise 2.1

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

Exercise 2.2

The number is 99, and the unit is bottles of soda.
The number is 60, and the unit is miles per hour.
The number 32, and the unit is fluid ounces
The number is 98.6, and the unit is degrees Fahrenheit

Exercise 2.2

The two necessary parts are the number and the unit.

2.1: The Classification of Matter
2.2: Expressing Numbers - Significant Figures
2.3: Expressing Numbers - Scientific Notation
2.4: The International System of Units

2.5: Converting Units
2.6: Dosage Calculations
2.7: Units, Measurement Uncertainty, and Significant Figures (Worksheet)
2.8: Chemistry, Matter, and Measurement (Exercises)

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2.1: The Classification of Matter

Learning Objectives

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

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

Physical and Chemical Properties

The properties that chemists use to describe matter fall into two general categories. **Physical properties** are characteristics that describe matter. They include characteristics such as size, shape, color, and mass. These characteristics can be observed or measured without changing the *identity* of the matter in question. **Chemical properties** are characteristics that describe how matter changes its chemical structure or composition. An example of a chemical property is flammability—a material's ability to burn—because burning (also known as combustion) changes the chemical composition of a material. The observation of chemical properties involves a *chemical change* of the matter in question, resulting in matter with a different *identity* and different physical and chemical properties.



Figure 2.1.1: (left) Ice Melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical as the the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (right) Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

Elements and Compounds

Any sample of matter that has the *same physical and chemical properties throughout* the sample is called a **substance**. There are two types of substances. A substance that cannot be broken down into chemically simpler components is called an **element**. Aluminum, which is used in soda cans and is represented by the symbol Al, is an element. A substance that can be broken down into chemically simpler components (because it consists of more than one element) is called a **compound**. Water is a compound composed of the elements hydrogen and oxygen and is described by the chemical formula, H_2O. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

Sometimes the word *pure* is used to describe a substance, but this is not absolutely necessary. By definition, any single substance, element or compound is *pure*.

The smallest part of an element that maintains the identity of that element is called an **atom**. Atoms are extremely tiny; to make a line of iron atoms that is 1 inch long, you would need approximately 217 million iron atoms. The smallest part of a compound that maintains the identity of that compound is called a **molecule**. Molecules are composed of two or more atoms that are attached together and behave as a unit. Scientists usually work with millions and millions of atoms and molecules at a time. When a scientist is working with large numbers of atoms or molecules at a time, the scientist is studying the *macroscopic viewpoint* of the universe. However, scientists can also describe chemical events on the level of individual atoms or molecules, which is referred to as the *microscopic viewpoint*. We will see examples of both macroscopic and microscopic viewpoints throughout this book (Figure 2.1.2).

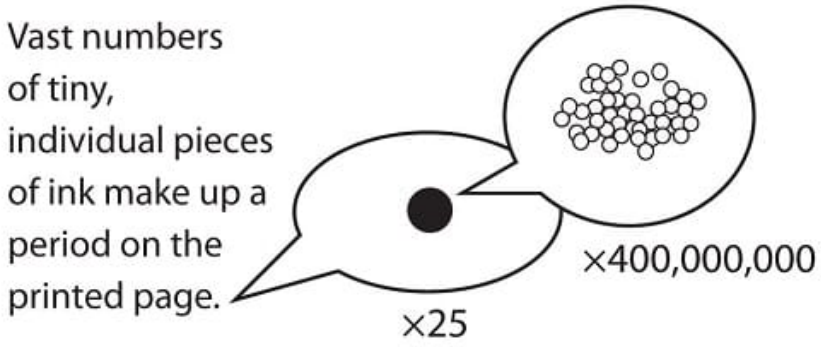


Figure 2.1.2: How Many Particles Are Needed for a Period in a Sentence? Although we do not notice it from a macroscopic perspective, matter is composed of microscopic particles so tiny that billions of them are needed to make a speck we can see with the naked eye. The ×25 and ×400,000,000 indicate the number of times the image is magnified.

Mixtures

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

Example 2.1.1

How would a chemist categorize each example of matter?

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

Answer a available at the end of this section.

Exercise 2.1.2

How would a chemist categorize each example of matter?

a. breakfast coffee
b. hydrogen
c. an egg

Answer a available at the end of this section.

Phases or Physical States of Matter

All matter can be further classified by one of three physical **states** or **phases**, solid, liquid or gas. These three descriptions each imply that the matter has certain physical properties when in these states. A solid has a definite shape and a definite volume. Liquids ordinarily have a definite volume but not a definite shape; they take the shape of their containers. Gases have neither a definite shape nor a definite volume, and they expand to fill their containers.

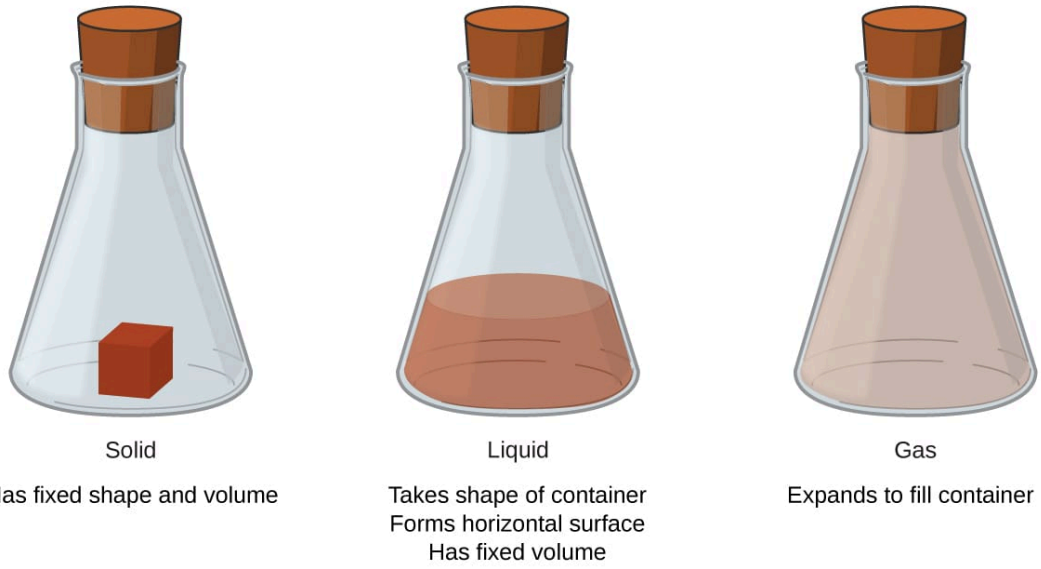


Figure 2.1.3: The three most common states or phases of matter are solid, liquid, and gas. (CC BY-4.0; OpenStax)

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

We encounter matter in each phase every day; in fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas) (Figure 2.1.2).



Figure 2.1.5: Boiling Water. When liquid water boils to make gaseous water, it undergoes a phase change. (CC BY-SA 3.0 Unported; Markus Schweiss via Wikipedia)

We know from our experience with water that substances can change from one phase to another if the conditions are right. Typically, varying the temperature of a substance (and, less commonly, the pressure exerted on it) can cause a **phase change**, a physical process in which a substance changes from one phase to another (Figure 2.1.5). Phase changes are identified by particular names depending on what phases are involved, as summarized in Table 2.1.1.

Table 2.1.1: 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 2.1.3 illustrates the relationships between the different ways matter can be classified.

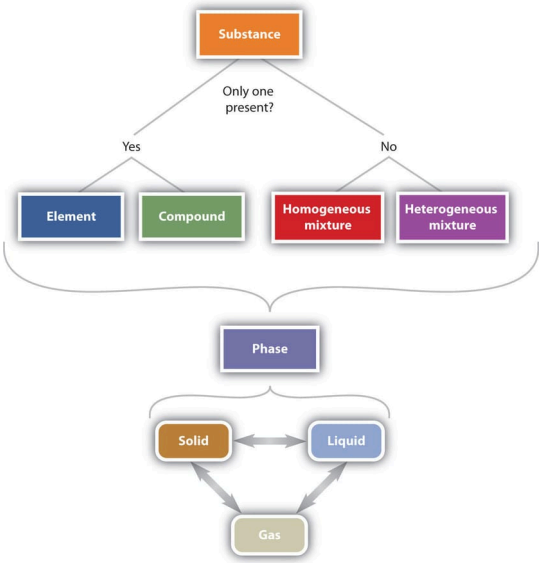


Figure 2.1.6: The Classification of Matter. Matter can be classified in a variety of ways, depending on its properties.

This table starts with a substance. If there is only one present, it can either be an element or a compound. If there is more than one present, it can either be a homogeneous mixture or a heterogeneous mixture. From these choices, the substance can be in a certain phase. It can either be solid, liquid, or gas, and each phase can go from one to another.

Concept Review Exercises

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

Answers

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

Key Takeaways

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

Exercise 2.1.3

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

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

Answer available at the end of this section.

Exercise 2.1.4

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

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

Answer available at the end of this section.

Exercise 2.1.5

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

Answer available at the end of this section.

Exercise 2.1.6

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

Answer available at the end of this section.

Exercise 2.1.7

Give two examples of a heterogeneous mixture.

Answer available at the end of this section.

Exercise 2.1.8

Give two examples of a homogeneous mixture.

Answer available at the end of this section.

Exercise 2.1.9

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

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

Answer available at the end of this section.

Exercise 2.1.10

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

1. sugar, a substance composed of carbon, hydrogen, and oxygen

2. hydrogen, the simplest chemical substance
3. dirt, a combination of rocks and decaying plant matter

Answer available at the end of this section.

Exercise 2.1.11

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

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

Answer available at the end of this section.

Exercise 2.1.12

What word describes each phase change?

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

Answer available at the end of this section.

Exercise 2.1.13

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

Answer available at the end of this section.

Answers

Example 2.1.1

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.

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

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

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

Exercise 2.1.2

homogeneous mixture or solution

element

heterogeneous mixture

Exercise 2.1.3

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

Exercise 2.1.4

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

4. physical property

Exercise 2.1.5

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

Exercise 2.1.6

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

Exercise 2.1.7

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

Exercise 2.1.8

vinegar and rubbing alcohol (answers will vary)

Exercise 2.1.9

1. element
2. heterogeneous mixture
3. compound

Exercise 2.1.10

1. compound
2. element
3. heterogeneous mixture

Exercise 2.1.11

1. heterogeneous mixture
2. solution
3. compound

Exercise 2.1.12

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

Exercise 2.1.13

1. freezing
2. condensation
3. deposition

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2.2: Expressing Numbers - Significant Figures

Learning Objectives

- Understand the importance of significant figures in measured numbers.
- Identify the number of significant figures in a reported value.
- Use significant figures correctly in arithmetical operations.

Scientists have established certain conventions for communicating the degree of **precision** of a measurement, which is dependent on the measuring device used. Imagine, for example, that you are using a meterstick to measure the width of a table. The centimeters (cm) marked on the meterstick, tell you how many *centimeters* wide the table is. Many metersticks also have markings for millimeters (mm), so we can measure the table to the nearest *millimeter*. Most metersticks do not have any smaller (or more precise) markings indicated, so you cannot report the measured width of the table any more precise than to the nearest millimeter. However, you can *estimate* one past the smallest marking, in this case the millimeter, to the next decimal place in the measurement (Figure 2.2.1).

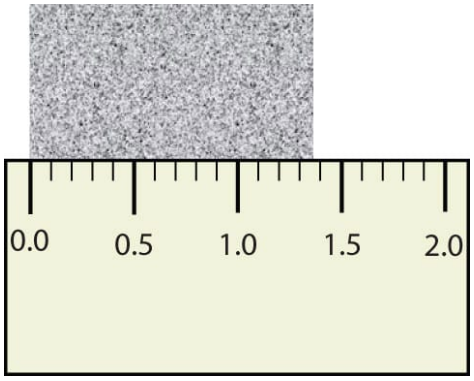


Figure 2.2.1: Measuring an Object to the Correct Number of Digits. How many digits should be reported for the length of this object?

The concept of **significant figures** takes this limitation into account. The significant figures of a measured quantity are defined as all the digits known with *certainty* (those indicated by the markings on the measuring device) ***and*** the first uncertain, or estimated, digit (one digit past the smallest marking on the measuring device). It makes no sense to report any digits after the first uncertain one, so it is the last digit reported in a measurement. Zeros are used when needed to place the significant figures in their correct positions. Thus, zeros are sometimes counted as significant figures but are sometimes only used as placeholders.

> "Sig figs" is a common abbreviation for significant figures.

Consider the earlier example of measuring the width of a table with a meterstick. If the table is measured and reported as being 1,357 mm wide, the number 1,357 has four significant figures. The 1 (thousands place), the 3 (hundreds place), and the 5 (tens place) are certain; the 7 (ones place) is assumed to have been estimated. It would make no sense to report such a measurement as 1,357.0 (five Sig Figs) or 1,357.00 (six Sig Figs) because that would suggest the measuring device was able to determine the width to the nearest tenth or hundredth of a millimeter, when in fact it shows only tens of millimeters and therefore the ones place was estimated.

On the other hand, if a measurement is reported as 150 mm, the 1 (hundreds) and the 5 (tens) are known to be significant, but how do we know whether the zero is or is not significant? The measuring device could have had marks indicating every 100 mm or marks indicating every 10 mm. How can you determine if the zero is significant (the estimated digit), or if the 5 is significant and the zero a value placeholder?

The **rules** for deciding which digits in a measurement are significant are as follows:

1. All nonzero digits are significant. In 1,357 mm, all the digits are significant.
2. *Sandwiched (or embedded) zeros*, those between significant digits, are significant. Thus, 405 g has *three* significant figures.
3. *Leading zeros*, which are zeros at the beginning of a decimal number less than 1, are not significant. In 0.000458 mL, the

first four digits are leading zeros and are not significant. The zeros serve only to put the digits 4, 5, and 8 in the correct decimal positions. This number has three significant figures.

4. *Trailing zeros*, which are zeros at the end of a number, are significant only if the number has a decimal point. Thus, in 1,500 m, the two trailing zeros are not significant because the number is written without a decimal point; the number has two significant figures. However, in 1,500.00 m, all six digits are significant because the number has a decimal point.

Example 2.2.1

How many significant figures does each number have?

a. 6,798,000
b. 6,000,798
c. 6,000,798.00
d. 0.0006798

Answer a available at the end of this section.

Exercise 2.2.1

How many significant figures does each number have?

a. 2.1828
b. 0.005505
c. 55,050
d. 5
e. 500

Answer a available at the end of this section.

Rounding off numbers

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

Consider the measurement 207.518m. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table 1.5.1.

Table 1.5.1: Rounding examples

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

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

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

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

Combining Numbers

For **addition or subtraction**, the rule is to stack all the numbers with their decimal points aligned and then limit (round to) the answer's significant figures to the rightmost column for which all the numbers have significant figures. Consider the following:

$$\begin{array}{r} 56.789 \\ +\quad 102.2 \\ +\,1{,}300.099 \\ \hline =\,1{,}459.088 \end{array}$$

↑ Limit to this column

The arrow points to the rightmost column in which all the numbers have significant figures—in this case, the tenths place. Therefore, we will limit our final answer to the tenths place. Is our final answer therefore 1,459.0? No, because when we drop digits from the end of a number, we also have to round the number. Notice that the first dropped digit, in the hundredths place, is 8. This suggests that the answer is actually closer to 1,459.1 than it is to 1,459.0, so we need to round up to **1,459.1**. The standard rules for rounding numbers are simple: If the first dropped digit is 5 or higher, round up. If the first dropped digit is lower than 5, do not round up.

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

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

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

Scientific notation provides a way of communicating significant figures without ambiguity. You simply include all the significant figures in the leading number. For example, the number 4,000 has one significant figure and should be written as the number 4×10^4. The number 450 has two significant figures and would be written in scientific notation as 4.5×10^2, whereas 450.0 has four significant figures and would be written as 4.500×10^2. In scientific notation, all reported digits are significant.

Example 2.2.2

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

a. 23.096 × 90.300
b. 125 × 9.000

c. 1,027 + 610.0 + 363.06

Answer a available at the end of this section.

Exercise 2.2.2

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

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

Answer a available at the end of this section.

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

Exercise 2.2.1

Explain why the concept of significant figures is important in scientific measurements.

Answer available at the end of this section.

Exercise 2.2.2

State the rules for determining the significant figures in a measurement.

Answer available at the end of this section.

Exercise 2.2.3

When do you round a number up, and when do you not round a number up?

Answer available at the end of this section.

Key Takeaways

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

Answers

Example 2.2.1

four (by rules 1 and 4)
seven (by rules 1 and 2)
nine (by rules 1, 2, and 4)
four (by rules 1 and 3)

Exercise 2.2.1

five

four

four

one

one

Example 2.2.2

The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the hundredths place) is greater than 5, we round up to 2,085.6, which in scientific notation is 2.0856×10^3.

The calculator gives 1,125 as the answer, but we limit it to three significant figures and convert into scientific notation: 1.13×10^3.

The calculator gives 2,000.06 as the answer, but because 1,027 has its farthest-right significant figure in the ones column, our answer must be limited to the ones position: 2,000 which in scientific notation is 2.000×10^3.

Exercise 2.2.2

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

$1437 = 1.437 \times 10^3$

$156 = 1.56 \times 10^2$

$2.14 = 2.14 \times 10^0$

Exercise 2.2.1

Significant figures represent all the known digits of a measurement plus the first estimated one. It gives information about how precise the measuring device and measurement is.

Exercise 2.2.2

All nonzero digits are significant; zeros between nonzero digits are significant; zeros at the end of a nondecimal number or the beginning of a decimal number are not significant; zeros at the end of a decimal number are significant.

Exercise 2.2.3

Round up only if the first digit dropped is 5 or higher.

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2.3: Expressing Numbers - Scientific Notation

Learning Objectives

- Express a large number or a small number in scientific notation.
- Convert a number in scientific notation to standard conventional form.

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

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

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

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

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

Recall that:

- $10^0 = 1$
- $10^1 = 10$
- $10^2 = 100$
- $10^3 = 1{,}000$
- $10^4 = 10{,}000$
- and so forth

The standard convention for expressing numbers in scientific notation is to write a single *nonzero* first digit, a decimal point, and the rest of the digits, excluding any trailing zeros (see rules for significant figures in the next section for more details on what to exclude). This number is followed by a multiplication sign and then by 10 raised to the power necessary to reproduce the original number. For example, although 1,500,000 can also be written as 15. × 10^5 (which would be 15. × 100,000), the convention is to have only one digit before the decimal point. How do we know to what power 10 is raised? The power is the number of places you have to move the decimal point to the *left* to place it after the first digit, so that the number being multiplied is *between 1 and 10*:

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

Example 2.3.1: Scientific Notation

Express each number in scientific notation.

a. 67,000,000,000
b. 1,689
c. 12.6

Answer a available at the end of this section.

Exercise 2.3.1

Express each number in scientific notation.

a. 1,492
b. 102,000,000
c. 101,325

Answer a available at the end of this section.

To change a number in **scientific notation** to **standard form**, we reverse the process, moving the decimal point to the right. Add zeros to the end of the number being converted, if necessary, to produce a number of the proper magnitude. Lastly, we drop the number 10 and its power.

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

Example 2.3.2

Express each number in standard, or conventional notation.

a. 5.27×10^4
b. 1.0008×10^6

Answer a available at the end of this section.

Exercise 2.3.2

Express each number in standard, or conventional notation.

a. 6.98×10^8
b. 1.005×10^2

Answer a available at the end of this section.

We can also use scientific notation to express numbers whose magnitudes are less than 1. For example, the quantity 0.006 centimeters can be expressed as follows:

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

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

Recall that:

- $10^{-1} = 1/10$
- $10^{-2} = 1/100$
- $10^{-3} = 1/1{,}000$
- $10^{-4} = 1/10{,}000$
- $10^{-5} = 1/100{,}000$
- and so forth

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

$$0.\underbrace{006} = 6 \times 10^{-3}$$

Note:

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

- Numbers that are greater than one have a positive power in scientific notation. If the decimal point is moved to the left n places, the power (n) of 10 is positive. $1{,}\underbrace{500{,}000} = 1.5 \times 10^{6}$
- Numbers that are less than one have a negative power in scientific notation. If the decimal point is moved to the right n places, the power (n) of 10 is negative. $0.\underbrace{006} = 6 \times 10^{-3}$

Example 2.3.3

Express each number in scientific notation.

a. 0.000006567
b. −0.0004004
c. 0.000000000000123

Answer a available at the end of this section.

Exercise 2.3.3

Express each number in scientific notation.

a. 0.000355
b. 0.314159
c. −0.051204

Answer a available at the end of this section.

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

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

Note

Changing a number in scientific notation to standard form:

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

Example 2.3.4

Change the number in scientific notation to standard form.

a. 6.22×10^{-2}
b. 9.9×10^{-9}

Answer a available at the end of this section.

Exercise 2.3.4

Change the number in scientific notation to standard form.

a. 9.98×10^{-5}
b. 5.109×10^{-8}

Answer a available at the end of this section.

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

Exercise 2.3.1

Why it is easier to use scientific notation to express very large or very small numbers?

Answer available at the end of this section.

Exercise 2.3.2

What is the relationship between how many places a decimal point moves and the power of 10 used in changing a conventional number into scientific notation?

Answer available at the end of this section.

Key Takeaway

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

Answers

Example 2.3.1: Scientific Notation

Moving the decimal point 10 places to the left gives 6.7×10^{10}.

The decimal point is assumed to be at the end of the number, so moving it three places to the left gives 1.689×10^{3}.

In this case, we need to move the decimal point only one place to the left, which yields 1.26×10^{1}.

Exercise 2.3.1

Moving the decimal point 3 places to the left gives 1.492×10^{3}.

The decimal point is assumed to be at the end of the number, so moving it 8 places to the left gives 1.02×10^{8}.

Moving the decimal point 5 places to the left yields 1.01325×10^{5}.

Example 2.3.2

Moving the decimal four places to the right and adding zeros give 52,700.

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

Exercise 2.3.2

Moving the decimal point eight places to the right and adding zeros give 698,000,000.

Moving the decimal point two places to the right gives 100.5

Example 2.3.3

Move the decimal point six places to the right to get 6.567×10^{-6}.

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

Move the decimal point 13 places to the right to get 1.23×10^{-13}.

Exercise 2.3.3

Moving the decimal point four places to the right gives 3.55×10^{-4}.

Moving the decimal point one place to the right gives 3.14159×10^{-1}.

Moving the decimal point one place to the right gives -5.1204×10^{-2}.

Example 2.3.4

0.0622

0.0000000099

Exercise 2.3.4

0.0000998

0.000000005109

Exercise 2.3.1

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

Exercise 2.3.2

The number of places the decimal point moves equals the power of 10. The power of 10 is positive if the decimal point moves to the left and negative if the decimal point moves to the right.

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2.4: The International System of Units

Learning Objectives

- Recognize the SI base units and explain the system of prefixes used with them.
- Define and calculate density.

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

Base SI Units

Base (or basic) units, are the fundamental units of SI. There are seven base units, which are listed in Table 2.4.1, Chemistry uses five of the base units: the mole for amount, the kilogram for mass, the meter for length, the second for time, and the kelvin for temperature. The degree Celsius (°C) is also commonly used for temperature. The numerical relationship between kelvins and degrees Celsius is as follows:

$$K = °C + 273 \tag{2.4.1}$$

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

Table 2.4.1: The Seven Base SI Units

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

The size of each base unit is defined by international convention. For example, the *kilogram* is defined as the quantity of mass of a special metal cylinder kept in a vault in France (Figure 2.4.1). The other base units have similar definitions and standards. The sizes of the base units are not always convenient for all measurements. For example, a meter is a rather large unit for describing the width of something as narrow as human hair. Instead of reporting the diameter of hair as 0.00012 m or as 1.2×10^{-4} m using scientific notation as discussed in section 1.4, SI also provides a series of **prefixes** that can be attached to the units, creating units that are larger or smaller by powers of 10.

Common prefixes and their multiplicative factors are listed in Table 2.4.2. (Perhaps you have already noticed that the base unit *kilogram* is a combination of a prefix, kilo- meaning 1,000 ×, and a unit of mass, the gram.) Some prefixes create a multiple of the original unit: 1 kilogram equals 1,000 grams, and 1 megameter equals 1,000,000 meters. Other prefixes create a fraction of the original unit. Thus, 1 centimeter equals 1/100 of a meter, 1 millimeter equals 1/1,000 of a meter, 1 microgram equals 1/1,000,000 of a gram, and so forth.

Table 2.4.2: Prefixes Used with SI Units

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

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



Figure 2.4.1: The Kilogram. The standard for the kilogram is a platinum-iridium cylinder kept in a special vault in France. Source: Photo reproduced by permission of the Bureau International des Poids et Mesures, who retain full internationally protected copyright.

The Difference Between Mass and Weight

The mass of a body is a measure of its inertial property or how much matter it contains. The weight of a body is a measure of the force exerted on it by gravity or the force needed to support it. Gravity on earth gives a body a downward acceleration of about 9.8 m/s^2. In common parlance, weight is often used as a synonym for mass in weights and measures. For instance, the verb "to weigh" means "to determine the mass of" or "to have a mass of." The incorrect use of weight in place of mass should be phased out, and the term mass used when mass is meant. The SI unit of mass is the kilogram (kg). In science and technology, the weight of a body in a particular reference frame is defined as the force that gives the body an acceleration equal to the local acceleration of free fall in that reference frame. Thus, the SI unit of the quantity weight defined in this way (force) is the newton (N).

Derived SI Units

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

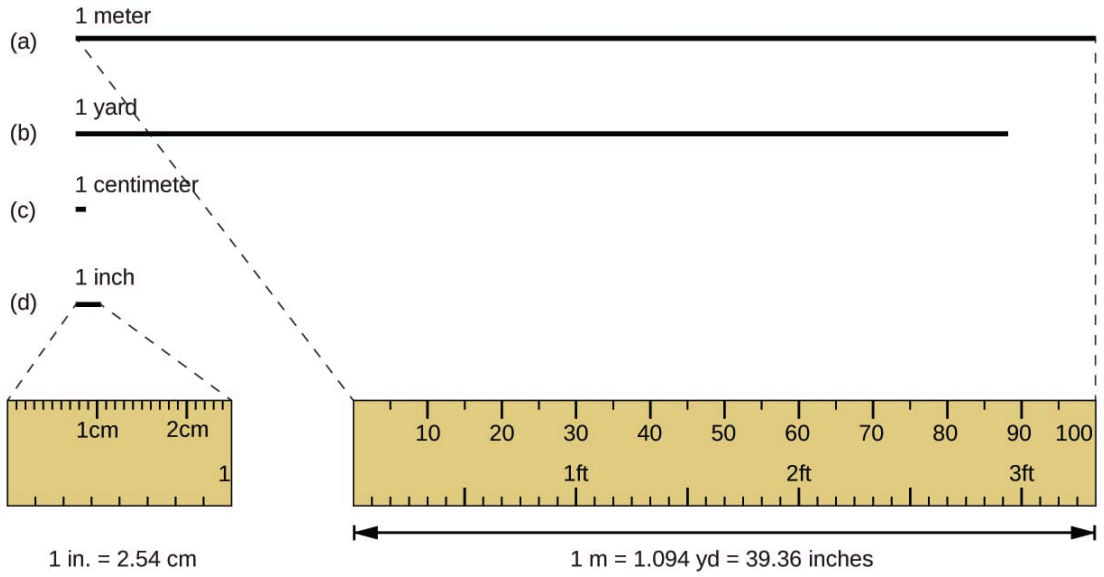


Figure 2.4.2: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd. (CC BY 4.0; OpenStax)

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

Approximate Equivalents to Some SI Units

- 1 m ≈ 39.36 in. ≈ 3.28 ft ≈ 1.09 yd
- 1 in. ≈ 2.54 cm
- 1 km ≈ 0.62 mi
- 1 kg ≈ 2.20 lb
- 1 lb ≈ 454 g
- 1 L ≈ 1.06 qt
- 1 qt ≈ 0.946 L

As shown in Figure 2.4.3, a liter is also 1,000 cm^3. By definition, there are 1,000 mL in 1 L, so 1 milliliter and 1 cubic centimeter represent the same volume.

$$1\ mL = 1\ cm^3 \tag{2.4.2}$$

Example 2.4.1

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

a. kiloliter

b. microsecond
c. decimeter
d. nanogram

Answer a available at the end of this section.

Exercise 2.4.1

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

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

Answer a available at the end of this section.

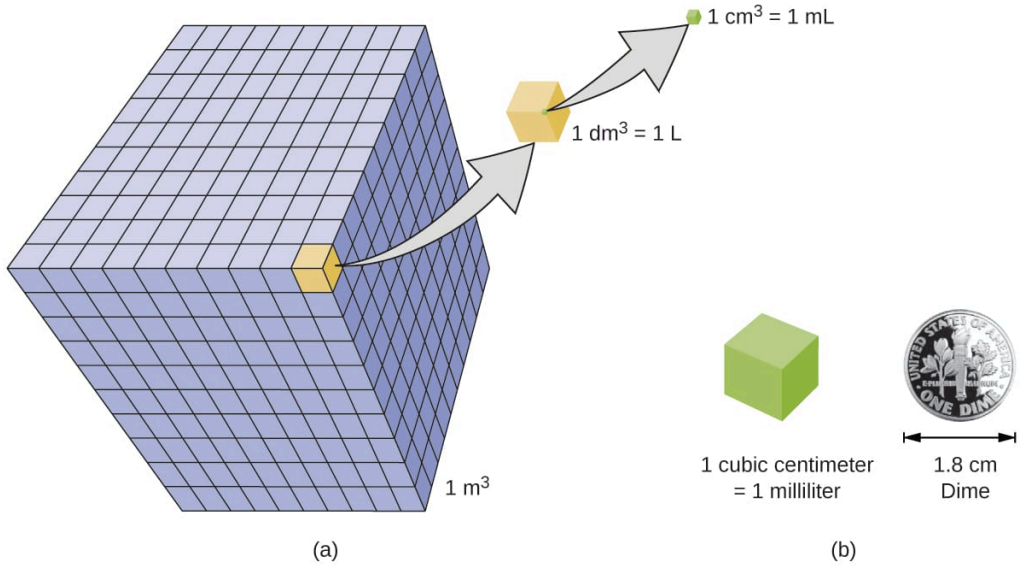


Figure 2.4.3: Units of Volume. (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). A liter (L) is defined as a cube 1 dm (1/10th of a meter) on a side. A milliliter (mL), 1/1,000th of a liter, is equal to 1 cubic centimeter. (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube. (CC BY 4.0; OpenStax)

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

To Your Health: Energy and Food

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

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

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

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

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{2.4.3}$$

Thus, the units of density are the units of mass divided by the units of volume: g/cm^3 or g/mL (for solids and liquids), g/L (for gases), kg/m^3, and so forth. For example, the density of water is about 1.00 g/cm^3, while the density of mercury is 13.6 g/mL. (Remember that 1 mL equals 1 cm^3.) Mercury is over 13 times as dense as water, meaning that it contains over 13 times the amount of matter in the same amount of space. The density of air at room temperature is about 1.3 g/L. Table 1.6.3 shows the densities of some common substances.

Table 2.4.2: Densities of Common Substances

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

Example 2.4.2: Density of Bone

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

Solution

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

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

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

Exercise 2.4.2: Density of Oxygen

What is the density of oxygen gas if a 15.0 L sample has a mass of 21.7 g?

Answer available at the end of this section.

Density can be used to convert between the mass and the volume of a substance. This will be discussed in the next section.

Concept Review Exercises

1. What is the difference between a base unit and a derived unit? Give two examples of each type of unit.

2. Do units follow the same mathematical rules as numbers do? Give an example to support your answer.

3. What is density?

Answers

1. Base units are the seven fundamental units of SI; derived units are constructed by making combinations of the base units; Two examples of base units: kilograms and meters (answers will vary); Two examples of derived units: grams per milliliter and joules (answers will vary).

2. yes; $\mathrm{mL} \times \frac{\mathrm{g}}{\mathrm{mL}} = \mathrm{g}$ (answers will vary)

3. Density is defined as the mass of an object divided by its volume

Key Takeaways

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

Answers

Example 2.4.1

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

Exercise 2.4.1

km (1,000 m)
mg (0.001 g)
ns (0.000000001 s)
cL (0.01L)

Exercise 2.4.2**: Density of Oxygen**

1.45 g/L

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2.5: Converting Units

Learning Objectives

- Convert a value reported in one unit to a corresponding value in a different unit.

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

Conversion Factors

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

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

or

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

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

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

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

$$\text{same quantity} \left\langle \frac{100 \text{ cm}}{1 \text{ m}} = 1 \right.$$

We know that 100 cm *is* 1 m, so we have the same quantity on the top and the bottom of our fraction, although it is expressed in different units. A fraction that has equivalent quantities in the numerator and the denominator but expressed in *different units* is called a **conversion factor**.

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

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

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

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

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

$$\frac{3.55}{1} \times \frac{100 \text{ cm}}{1} = 355 \text{ cm} \tag{2.5.1}$$

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

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

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

In the previous example (Equation 2.5.1), we used the fraction $\frac{100\text{ cm}}{1\text{ m}}$ as a conversion factor. Does the conversion factor $\frac{1\text{ m}}{100\text{ cm}}$ also equal 1? Yes, it does; it has the same quantity in the numerator as in the denominator (except that they are expressed in different units). Why did we not use *that* conversion factor? If we had used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what we would have gotten:

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

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure 2.5.1 shows a **concept map** for constructing a proper conversion.

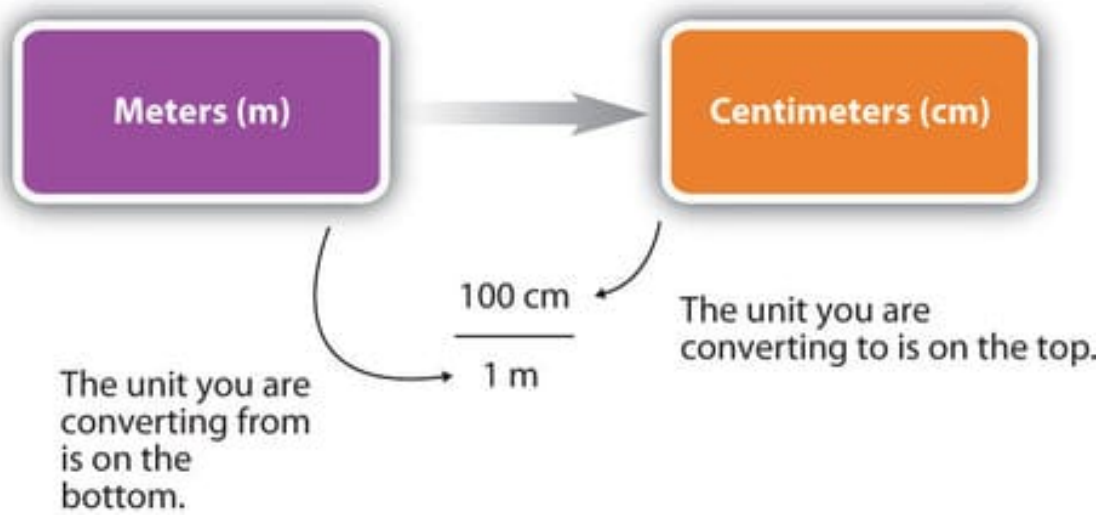


Figure 2.5.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another.

Meters converted to centimeters. A fraction showing 100 cm over 1 m. The unit you convert to is on top, and the unit you convert from is on bottom.

Significant Figures in Conversions

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

Example 2.5.1

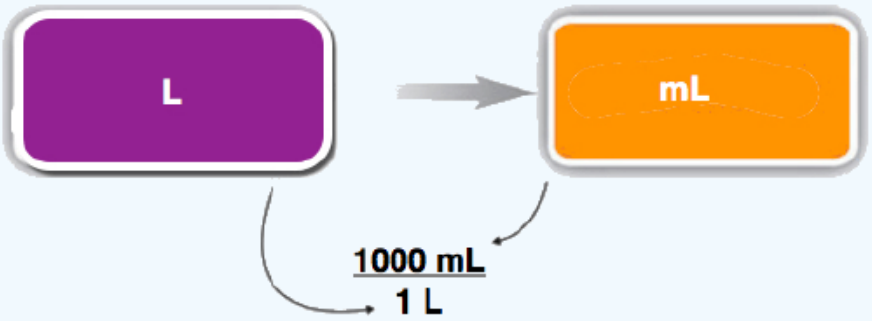
a. The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?
b. A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?

Solution

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

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

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



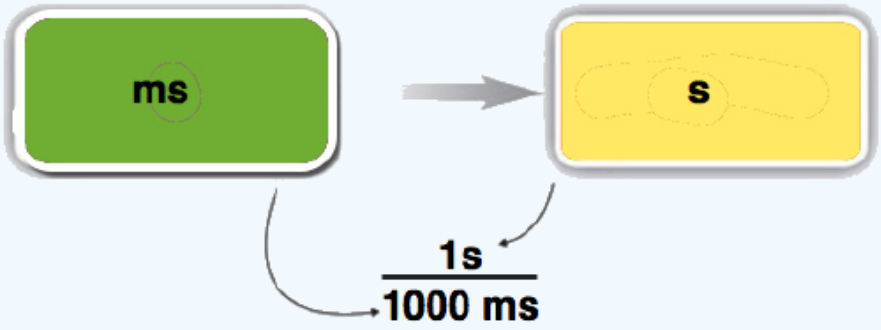
$$4.7\cancel{\text{L}} \times \frac{1,000\text{ mL}}{1\ \cancel{\text{L}}} = 4,700\text{ mL}$$

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

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

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

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



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

The conversion factor's numerical values do not affect our determination of the number of significant figures in the final answer.

Exercise 2.5.1

Perform each conversion.

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

Answer a available at the end of this section.

Conversion Factors From Different Units

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

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

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

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

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

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

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

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

Example 2.5.2: Mercury Thermometer

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

Solution

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

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

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

We have limited the final answer to three significant figures.

Exercise 2.5.2

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

Answer available at the end of this section.

Looking Closer: Density and the Body

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

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

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

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

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

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

Problem Solving With Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. You can either memorize the relationship between kilometers and millimeters, or you can do the conversion in two steps. Most people prefer to convert in steps.

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

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

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

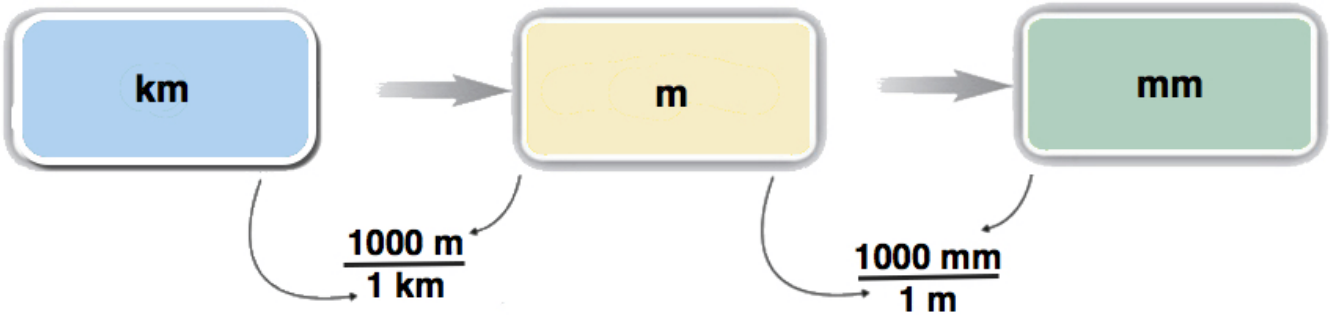
$$54{,}700 \cancel{\text{ m}} \times \frac{1{,}000 \text{ mm}}{1 \cancel{\text{ m}}} = 54{,}700{,}000 \text{ mm}$$

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

We have expressed the final answer in scientific notation.

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

Concept Map



Calculation

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

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

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

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

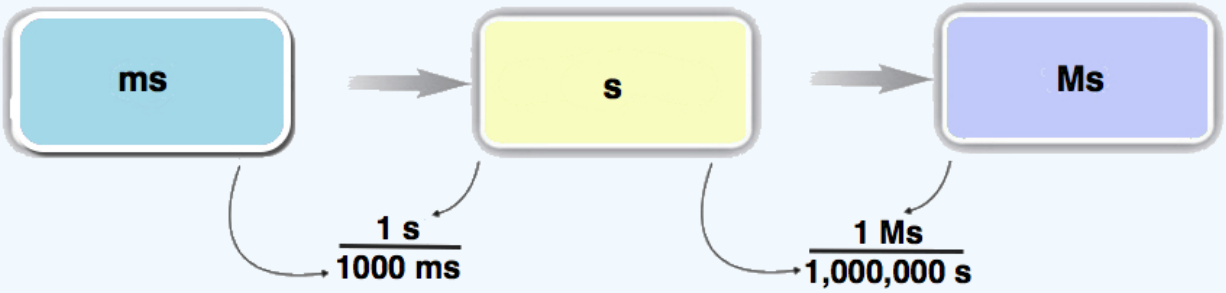
Example 2.5.3

Convert 58.2 ms to megaseconds in one multistep calculation.

Solution

First, convert the given unit (ms) to the base unit—in this case, seconds—and then convert seconds to the final unit, megaseconds:

Concept Map



Calculation

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

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

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

Exercise 2.5.3

Convert 43.007 mg to kilograms in one multistep calculation.

Answer available at the end of this section.

Career Focus: Pharmacist

A pharmacist dispenses drugs that have been prescribed by a doctor. Although that may sound straightforward, pharmacists in the United States must hold a doctorate in pharmacy and be licensed by the state in which they work. Most pharmacy programs require four years of education in a specialty pharmacy school.

Pharmacists must know a lot of chemistry and biology so they can understand the effects that drugs (which are chemicals, after all) have on the body. Pharmacists can advise physicians on the selection, dosage, interactions, and side effects of drugs. They can also advise patients on the proper use of their medications, including when and how to take specific drugs properly. Pharmacists can be found in drugstores, hospitals, and other medical facilities.

Curiously, an outdated name for pharmacist is *chemist*, which was used when pharmacists formerly did a lot of drug preparation, or *compounding*. In modern times, pharmacists rarely compound their own drugs, but their knowledge of the sciences, including chemistry, helps them provide valuable services in support of everyone's health.

Key Takeaway

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

Concept Review Exercises

1. How do you determine which quantity in a conversion factor goes in the denominator of the fraction?
2. State the guidelines for determining significant figures when using a conversion factor.
3. Write a concept map (a plan) for how you would convert 1.0×10^{12} *nano*liters (nL) to *kilo*liters (kL).

Answers

1. The unit you want to cancel from the numerator goes in the denominator of the conversion factor.
2. Exact numbers that appear in many conversion factors do not affect the number of significant figures; otherwise, the normal rules of multiplication and division for significant figures apply.
3. Concept Map: Convert the given (nanoliters, nL) to liters; then convert liters to kiloliters.



1.0 times 10 to the 12th nL converts to 1000 Liters converts to 1 kiloLiter. One Liter is 1.0 times 10 to the 9th nL. One kiloLiter is 1000 Liters.

Answers

Exercise 2.5.1

$$101,000 \cancel{\text{ns}} \times \frac{1 \text{ s}}{1,000,000,000 \cancel{\text{ns}}} = 0.000101 \text{ s}$$

$$32.08 \cancel{\text{kg}} \times \frac{1,000 \text{ g}}{1 \cancel{\text{kg}}} = 32,080 \text{ g}$$

Exercise 2.5.2

$$100.0 \cancel{\text{g}} \times \frac{1 \text{ L}}{1.3 \cancel{\text{g}}} = 76.92307692 \text{ L} \approx 77\text{L}$$

Because the density (1.3 g/L) has only 2 significant figures, we are rounding off the final answer to 2 significant figures.

Exercise 2.5.3

$$43.007 \cancel{\text{mg}} \times \frac{1\cancel{\text{g}}}{1,000 \cancel{\text{mg}}} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 0.000043007 \text{ kg}$$

$$= 4.3007 \times 10^{-5} \text{ kg}$$

.

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

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2.6: Dosage Calculations

Learning Objectives

- Calculate drug dosages using conversion factors.

To Your Health: Dosages

A medicine can be more harmful than helpful if it is not taken in the proper dosage. A dosage (or dose) is the specific amount of a medicine that is known to be therapeutic for an ailment in a patient of a certain size. Dosages of the active ingredient in medications are usually described by units of mass, typically grams or milligrams, and generally are equated with a number of capsules or teaspoonfuls to be swallowed or injected. The amount of the active ingredient in a medicine is carefully controlled so that the proper number of pills or spoonfuls contains the proper dose.

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

Effects are dose-dependent

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

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

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

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

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

Calculation of dose

Calculating drug dosages for humans based on the doses used in animal studies can be based on weight (e.g., mg/kg) or surface area (e.g., mg/m^2) based on weight. Drug dosage calculation is required if the physician's order is different from what is available.

Example 2.6.1

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

Solution

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

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

Hence, the nurse should administer 2 tablets.

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

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

Hence, 1.5 tablets should be taken.

Exercise 2.6.1

Calculate each of the following.

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

Answer a available at the end of this section.

Drug Dosage Calculation based on Body Weight. Many drugs (especially in children) are dosed according to body weight (mg/kg). These calculations are carried in 3-step conversions. The first step is to convert the body weight from *pounds (lbs)* to *kg*. The second step is to convert *kg* to *mg* (the total *mg dose* calculated based on body weight). And, finally, the *mg dose* is converted to the number of *tablets*. (as in Example 1.8.1).

Example 2.6.2

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

Solution

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

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

Hence, the nurse administers 3 tablets.

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

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

Hence, the nurse administers 2 tablets.

Exercise 2.6.2

Calculate each of the following.

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

Answer a available at the end of this section.

References

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

Contributors and Attributions

- Wikipedia
- Leticia Colmenares (UH-WCC)

Answers

Exercise 2.6.1

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

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

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

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

Exercise 2.6.2

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

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

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2.7: Units, Measurement Uncertainty, and Significant Figures (Worksheet)

Name: _____

Section: _____

Student ID#:_____

Work in groups on these problems. You should try to answer the questions without referring to your textbook. If you get stuck, try asking another group for help.

All scientists the world over use metric units. Since 1960, the metric system in use has been the *Système International d'Unités*, commonly called the SI units. These units facilitate international communication by discouraging use of units peculiar to one culture or another (e.g., pounds, inches, degrees Fahrenheit). But regardless of the units used, we want to have some confidence that our measured and calculated results bear a close relationship to the "true" values. Therefore, we need to understand the limits on our measured and calculated values. One way we convey this is by writing numerical answers with no more and no fewer than the number of digits that are justified by the limits of our ability to measure and know the quantity.

Learning Objective

- Know the units used to describe various physical quantities
- Become familiar with the prefixes used for larger and smaller quantities
- Master the use of unit conversion (dimensional analysis) in solving problems
- Appreciate the difference between precision and accuracy
- Understand the relationship between precision and the number of significant figures in a number

Success Criteria

- Associate units with physical quantities
- Replace prefixes by multiplying by appropriate numerical factors
- Be able to use dimensional analysis for unit conversions
- Report computed values to the correct number of significant figures.

Système International d'Unités (SI Units)

The SI units consist of seven base units and two supplementary units. For now, we will only use the four base units listed below. Later we will talk about two others. We will never use the seventh unit (candela), a unit for luminous intensity.

Table 2.7.1: Four of the Base units in the SI system

Quantity	Unit	Abbrev.
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K

Any other units can be constructed as a combination of fundamental units. For example, velocity could be measured in meters per second (written m/s or $m \times s^{-1}$), and area could be measured in units of meters squared (m^2). When a named unit is defined as a combination of base units, it is called a *derived unit*. For example, the SI unit of energy is the joule (J), which is defined

as a $kg \times m^2 \times s^{-2}$. Note that when a unit is named for some scientist (e.g., Joule, Herz, Kelvin) the written name of the unit *is not* capitalized, but the abbreviation *is* capitalized.

All metric units can be related to larger or smaller units for the same quantity by use of prefixes that imply multiplication of the stem unit by certain powers of 10. The following prefixes are important to know.

Table 2.7.2: SI Prefixes

Prefix	Abbrev.	$10^{\pm n}$	Example
Mega-	M	10^6	Megaherz (MHz)
Kilo-	k	10^3	kilogram (kg)
Deci-	d	10^{-1}	deciliter (dL)
Centi-	c	10^{-2}	centimeter (cm)
Milli-	m	10^{-3}	milliliter (mL)
Micro-	m	10^{-6}	microgram (μg)
Nano-	n	10^{-9}	nanometer (nm)
Pico-	p	10^{-12}	picosecond (ps)
Femto-	f	10^{-15}	femtosecond (fs)

Q1

Give the names and their abbreviations for the SI units of length, mass, time, and temperature.

Q2

The unit of volume is the liter (L). Why is this not a *base* SI unit? What kind of SI unit is it?

Q3

A student is asked to calculate the mass of calcium oxide produced by heating a certain amount of calcium carbonate. The student's answer of 90.32 is numerically correct, but the instructor marks it wrong. Why?

Q4

Write the number of seconds in a day (86,400 s) in exponential notation, using a coefficient that is greater than 1 and less than 10. (This form is called *scientific notation* and is generally the preferred form of exponential notation, as explained below).

Q5

The diameter of a helium atom is about 30 pm. Write this length in meters, using standard scientific notation.

Q7

A cubic container is 2.00 cm on each edge. What is its volume in liters? What is its volume in milliliters (mL)? Are your answers reasonable?

Dimensional Analysis

Units can actually help in setting up and solving many problems by using a method called *dimensional analysis* (also called the factor-label method). In dimensional analysis, a problem is typically viewed as a conversion of a given value in given units into a new value in certain desired units. Mathematically, such problems take on the general form

$$(\text{given quantity in } \cancel{\text{given units}}) \underbrace{\left(\frac{\text{wanted units}}{\cancel{\text{given units}}}\right)}_{\text{conversion factor} \equiv 1} = \text{wanted quantity in wanted units}$$

The factor "wanted units/given units" is a conversion factor, which is always a fractional expression of an equivalence relationship between two different units. In carrying out the multiplication and division, the given units cancel out, leaving the wanted units. To apply dimensional analysis, follow this general problem-solving strategy:

1. Identify and record what is know, with its given units;
2. identify what is to be calculated with its units;
3. identify the concepts and/or relationships that connect the given information with what needs to be calculated;
4. set up the solution using unit relationships as one or more conversion factors, such that all units except those desired for the answer cancel;
5. do the mathematics;
6. check or validate your answer by asking yourself if it is a reasonable result.

Example

How many inches is 2.00 cm, given that the inch is defined as exactly 2.54 cm?

1. We know the length in centimeters.
2. We want the length in inches.
3. 1 inch (in.) is exactly 2.54 cm (no uncertainty)
4. Possible conversion factors are 1 in/2.54 cm and 2.54 cm/1 in) We are starting with cm and want to end up with in, so the first conversion factor will do the job.

$$(2.00\,cm) \underbrace{\left(\frac{1\,in}{2.54\,cm}\right)}_{\text{conversion factor} \equiv 1} = ?$$

5. Do the mathematics.

$$(2.0\,\cancel{cm}) \underbrace{\left(\frac{1\,in.}{2.54\,\cancel{cm}}\right)}_{\text{conversion factor} \equiv 1} = 0.787\,in.$$

   Note that the centimeter units cancel, leaving the desired units of inches.
6. If 2.54 cm is an inch, then 2.00 cm should be a fraction of an inch. So, 0.787 in looks like a reasonable answer.

Q7

In general, how can you identify whether or not you have written the correct conversion factor for the problem?

Q8

One liter is 1.06 quarts (qt). Write two possible conversion factors from this relationship.

Q9.

The posted speed limit is 60 mi/hr. You are doing 120 km/hr in your Porsche convertible that you just bought in Germany. Are you speeding? Explain. [1.0 mi = 1.6 km]

Q10

In the gym, you slip on two 45-lb barbell plates to a bar that weighs 45 lb. What is the mass of the set-up in kilograms? [1.00 kg = 2.20 lbs]

Q11

A table top is 36 in long and 24 in wide. What is the area of the table top in square meters? [1 in = 2.54 cm, exactly]

Accuracy and Precision

Measured quantities always have some experimental error. Therefore, measured quantities are regarded as *inexact*. The **accuracy** of a measured quantity is its agreement with a standard or true value. In reality, we generally cannot know the true value of something we wish to measure. We gain confidence that our measured value is close to the truth by repeating the measurement many times. If our repeated measurements yield a set of data that differ very little from each other, we have some confidence that the average of these measured values is close to the true value. The repeatability of the measurements is called its **precision**. In general, we assume that greater precision in a set of numbers makes it more likely that the average value will be accurate. However, it is possible for a very precise set of values to be inaccurate. For example, a scientist could make the same error in each of a set of measurements, which could happen if a key measuring device were miscalibrated. Conversely, it is possible that a set of widely scattered values (poor precision) could have an average value that is very close to the true value, therefore resulting in high accuracy.

We express the precision of a number by writing all the repeatable digits and the first uncertain digit from a measurement or calculation. The retained digits are called the significant figures (sig. figs.) of the number. The following rules should be used to determine the number of significant figures of a number and to establish the correct number of significant figures in the answer to a calculation.

1. For decimal numbers with absolute value >1, all digits are significant.
   - 2.620 has 4 sig. figs.
   - 50.003 has 5 sig. figs.
2. If there is no decimal point, zeroes that set magnitude only are not significant.
   - $103,000$ has 3 sig. figs.
   - $103,000.$ has 6 sig. figs.
3. For decimal numbers with absolute value <1, start counting significant figures at the first non-zero digit to the right of the decimal point.
   - 0.0012 has 2 sig. figs.
   - 0.00070 has 2 sig. figs.
   - 2.0070 has 5 sig. figs.
4. Decimal numbers with absolute value <1 are frequently expressed with standard scientific exponential notation, particularly when the number is smaller than 0.1.
   - 1.2×10^{-3} has 2 sig. figs.
   - 7.0×10^{-4} has 2 sig. figs
5. In multiplication and division, the answer may have no more significant figures than the number in the chain with the fewest significant figures.

$$\frac{(9.97)(6.5)}{4.321} = 15 \text{ has 2 sig. figs.}$$

6. When adding or subtracting, the answer has the same number of decimal places as the number with the fewest decimal places. The number of significant figures for the result, then, is determined by the usual rules after establishing the appropriate number of decimal places.

$$3.0081 + 7.41 = 10.4181 \approx 10.42 \text{ has 2 decimal places and 2 sig. figs.}$$

   The rules for addition and subtraction may radically alter the number of significant figures for the answer in a chain of mathematical calculations, as the following shows.
7. Exact numbers, which are inherently integers or are set by definition, are not limited in their significant digits. Some exact numbers:

- All integer fractions: $\frac{1}{2}$, $\frac{1}{3}$, $\frac{7}{9}$
- Counted numbers: "15 people"
- Conversions *within* a unit system: 12 inches = 1 foot (exactly!)

Relationships between units in *different* unit systems are usually not exact:

- 2.2 lb. = 1.0 kg 2 sig. figs.
- 2.2046223 lb. = 1.0000000 kg 8 sig. figs.

But, the following inter-system conversion factors are now set by definition and are exact:

- 2.54 cm / 1 inch (exactly)
- 1 calorie / 4.184 Joules (exactly)

A way of getting around the ambiguity in significant figures for numbers like 103,000 is to use *standard scientific exponential notation*, consisting of a coefficient whose magnitude is greater than 1 and less than 10 multiplied by the appropriate power of ten. All digits in the coefficient are significant.

- 1.03×10^5 has 3 sig. figs.
- 1.030×10^5 has 4 sig. figs.
- 1.0300×10^5 has 5 sig. figs.
- 1.03000×10^5 has 6 sig. figs.

Q12

A one-gram standard reference weight is repeatedly weighed on an analytical balance. The readings from the balance are as follows:

- $1.003\ g$
- $0.9998\ g$
- $1.005\ g$
- $0.9995\ g$

Comment on the precision and accuracy of these data.

Q13

The same one-gram reference weight is weighed on another analytical balance. The readings from this balance are as follows:

- $0.9845\ g$
- $1.0114\ g$
- $0.9879\ g$
- $1.0208\ g$

Comment on the precision and accuracy of these data.

Q14

The same one-gram reference weight is weighed on a third analytical balance. The readings from this balance are as follows:

- $1.237\ g$
- $1.243\ g$
- $1.238\ g$
- $1.245\ g$

Comment on the precision and accuracy of these data.

Q15

How is precision represented in reporting a measured value?

Q16

How many significant figures are there in each of the following numbers?

- 0.0037
- 20.03
- 300
- $300.$
- $3.000x10^2$

Q17

Use your calculator to carry out the following calculations and report the answers to the correct number of significant figures:

- $x = (2)(39.0983) + (2)(51.996) + (7)(15.9994)$ (The first number in each multiplication is an integer)
- $x = \dfrac{1.44 \times 10^4}{2.40 \times 10^8}$
- $x = \dfrac{(3.5 \times 10^{-5})(6.2 \times 10^{12}}{3.3 \times 10^{-15}}$
- $x = \sqrt{(7.56 \times 10^{-5})(0.125)}$
- $x = \left[\dfrac{(0.5622)(3.20 + 8.111)}{621.25}\right]^{1/3}$

Q18

A supermarket in London is selling cod for 12.98 £/kg. If the rate of exchange is $1.6220 = 1.0000 £, what is the price in dollars per pound? 1.000 kg = 2.205 lb

Q19

A hollow metal sphere has an outer diameter (o.d.) of 4.366 cm and an inner diameter (i.d.) of 4.338 cm. What is the volume of the metal in the sphere? Express your answer to the proper number of significant figures. [$V_{sphere} = (4/3)\pi r^3$]

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2.8: Chemistry, Matter, and Measurement (Exercises)

1.1: What Is Chemistry?

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

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

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

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

3. Which of the following are examples of matter?

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

4. Which of the following are examples of matter?

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

5. Suggest a name for the science that studies the physics of rocks and the earth.

6. Suggest a name for the study of the physics of living organisms.

7. Engineering is the practical application of scientific principles and discoveries to develop things that make our lives easier. Is medicine science or engineering? Justify your answer.

8. Based on the definition of engineering in Exercise 7, would building a bridge over a river or road be considered science or engineering? Justify your answer.

9. When someone says, “I have a theory that excess salt causes high blood pressure,” does that person really have a theory? If it is not a theory, what is it?

10. When a person says, “My hypothesis is that excess calcium in the diet causes kidney stones,” what does the person need to do to determine if the hypothesis is correct?

11. Some people argue that many scientists accept many scientific principles on faith. Using what you know about the scientific method, how might you argue against that assertion?

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

Answers

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

2.

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

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

4.

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

5. geophysics

6. biophysics

7. Medicine is probably closer to a field of engineering than a field of science, but this may be arguable. Ask your doctor.

8. Engineering

9. In scientific terms, this person has a hypothesis.
10. Conduct experiments to determine if kidney stones contain calcium.
11. Science is based on reproducible facts, not blind belief.
12. No.

1.2: The Classification of Matter

Exercises

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

   a. Balsa is a very light wood.
   b. If held in a flame, magnesium metal burns in air.
   c. Mercury has a density of 13.6 g/mL.
   d. Human blood is red.

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

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

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

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

5. Give two examples of a heterogeneous mixture.

6. Give two examples of a homogeneous mixture.

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

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

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

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

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

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

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

    a. equal portions of salt and sand placed in a beaker and shaken up
    b. a combination of beeswax dissolved in liquid hexane
    c. hydrogen peroxide, a substance composed of hydrogen and oxygen

11. What word describes each phase change?

    a. solid to liquid
    b. liquid to gas
    c. solid to gas

12. What word describes each phase change?

    a. liquid to solid
    b. gas to liquid
    c. gas to solid

Answers

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

2.

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

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

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

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

6. vinegar and rubbing alcohol (answers will vary)

7. a. element
b. heterogeneous mixture
c. compound

8.

a. compound
b. element
c. heterogeneous mixture

9. a. solution
b. solution
c. compound

10.

a. heterogeneous mixture
b. solution
c. compound

11.

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

12.

a. freezing
b. condensation
c. deposition

1.3: Measurements

Exercises

1. Why are both parts of a quantity important when describing it?

2. Why are measurements an important part of any branch of science, such as chemistry?

3. You ask a classmate how much homework your chemistry professor assigned. Your classmate answers, “twenty.” Is that a proper answer? Why or why not?

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

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

Answers

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

2. Measurements are needed to carry out experiments.

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

4.

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

1.4: Expressing Numbers: Scientific Notation

Exercises

1. Why is scientific notation useful in expressing numbers?

2. What is the relationship between the power and the number of places a decimal point is moved when going from standard to scientific notation?

3. Express each number in scientific notation.

    a. 0.00064
    b. 5,230,000
    c. −56,200
    d. 0.000000000220
    e. 1.0

4. Express each number in scientific notation.

    a. 678
    b. −1,061
    c. 0.000560
    d. 0.0000003003
    e. 100,000,000

5. Express each number in standard form.

   a. 6.72×10^{4}
   b. 2.088×10^{-4}
   c. -3×10^{6}
   d. 9.98×10^{-7}

6. Express each number in standard form.

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

7. Complete the following table:

7. Complete the following table

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

8. Complete the following table:

8. Complete the following table

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

Answers

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

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

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

4.

   a. 6.78×10^{2}
   b. -1.061×10^{3}
   c. 5.60×10^{-4}

d. 3.003×10^{-7}
e. 1×10^{8}

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

6. Express each number in standard form.

   a. 905,000
   b. 0.0010
   c. 602,200,000,000,000,000,000,000
   d. 0.0000000000008834

7. *Answers to question 7: Complete the table*

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

8.

Answers to question 8: Complete the table

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

1.5: Expressing Numbers: Significant Figures

Exercises

1. Define *significant figures*. Why are they important?

2. Define the different types of zeros found in a number and explain whether or not they are significant.

3. How many significant figures are in each number?

   a. 140
   b. 0.009830
   c. 15,050
   d. 221,560,000
   e. 5.67×10^{3}
   f. 2.9600×10^{-5}

4. How many significant figures are in each number?

   a. 1.05
   b. 9,500

c. 0.0004505
d. 0.00045050
e. 7.210×10^{6}
f. 5.00×10^{-6}

5. Round each number to three significant figures.

a. 34,705
b. 34,750
c. 34,570

6. Round each number to three significant figures.

a. 34,705
b. 34,750
c. 34,570

7. Perform each operation and express the answer to the correct number of significant figures.

a. 467.88 + 23.0 + 1,306 = ?
b. 10,075 + 5,822.09 − 34.0 = ?
c. 0.00565 + 0.002333 + 0.0991 = ?

8. Perform each operation and express the answer to the correct number of significant figures.

a. 0.9812 + 1.660 + 8.6502 = ?
b. 189 + 3,201.8 − 1,100 = ?
c. 675.0 − 24 + 1,190 = ?

9. Perform each operation and express the answer to the correct number of significant figures.

a. 439 × 8,767 = ?
b. 23.09 ÷ 13.009 = ?
c. 1.009 × 876 = ?

10. Perform each operation and express the answer to the correct number of significant figures.

a. 3.00 ÷ 1.9979 = ?
b. 2,300 × 185 = ?
c. 16.00 × 4.0 = ?

11. Use your calculator to solve each equation. Express each answer in proper scientific notation and with the proper number of significant figures. If you do not get the correct answers, you may not be entering scientific notation into your calculator properly, so ask your instructor for assistance.

a. $(5.6 \times 10^{3}) \times (9.04 \times 10^{-7}) = ?$
b. $(8.331 \times 10^{-2}) \times (2.45 \times 10^{5}) = ?$
c. $983.09 \div (5.390 \times 10^{5}) = ?$
d. $0.00432 \div (3.9001 \times 10^{3}) = ?$

12. Use your calculator to solve each equation. Express each answer in proper scientific notation and with the proper number of significant figures. If you do not get the correct answers, you may not be entering scientific notation into your calculator properly, so ask your instructor for assistance.

a. $(5.2 \times 10^{6}) \times (3.33 \times 10^{-2}) = ?$
b. $(7.108 \times 10^{3}) \times (9.994 \times 10^{-5}) = ?$
c. $(6.022 \times 10^{7}) \div (1.381 \times 10^{-8}) = ?$

d. $(2.997 \times 10^{8}) \div (1.58 \times 10^{34}) = ?$

Answers

1. Significant figures represent all the known digits plus the first estimated digit of a measurement; they are the only values worth reporting in a measurement.

2. Leading zeros serve as decimal place holders only, hence are not significant. Sandwiched zeros are significant. Trailing zeros following non-zero numbers without a decimal point is not significant. Trailing zeros following non-zero digits with a decimal point are significant.

3. a. two
   b. four
   c. four
   d. five
   e. three
   f. five

4.

   a. three
   b. two
   c. four
   d. five
   e. four
   f. three

5.

   a. 34,700
   b. 34,800
   c. 34,600

6.

   a. 34,710
   b. 0.005411
   c. 8.904×10^{8}

7. a. 1,797
   b. 15,863
   c. 0.1071

8.

   a. 11.291
   b. 2,291
   c. 1,841

9.

   a. 3,850,000
   b. 1.775
   c. 884

10.

   a. 1.50
   b. 430,000
   c. 64

11.

a. 5.1×10^{-3}
b. 2.04×10^{4}
c. 1.824×10^{-3}
d. 1.11×10^{-6}

12.

a. 1.7×10^{5}
b. 7.104×10^{-1}
c. 4.361×10^{15}
d. 1.90×10^{-26}

1.6: The International System of Units

Exercises

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

5. Complete the following table

Unit	Abbreviation
centiliter	
	ms
	cm
	kL
micrometer	

6. Complete the following table:

6. Complete the following table

Unit	Abbreviation
microliter	
kilosecond	
	dL
	ns
millimeter	

7. What are some appropriate units for density?

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

Answers

1. second, meter, kilogram, and kelvin (answers will vary)

2. square meter (m^2), cubic meter (m^3), grams per milliliter (g/mL), milliliters per second (mL/s) answers will vary

3. 1,000; 100

4. 1,000,000; 1,000,000

5. *Answer to question 5: Complete the following table*

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

6.

Answers to question 6: Complete the following table

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

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

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

1.7: Converting Units

Exercises

1. Give the two conversion factors you can construct using each pair of units.

    a. meters and kilometers
    b. liters and microliters
    c. seconds and milliseconds

2. Give the two conversion factors you can construct using each pair of units.

    a. grams and centigrams
    b. millimeters and meters
    c. liters and megaliters

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

Answers

1. a. $\frac{1,000\text{ m}}{1\text{ km}}$; $\frac{1\text{ km}}{1,000\text{ m}}$
   b. $\frac{1,000,000\ \mu\text{L}}{1\text{ L}}$; $\frac{1\text{ L}}{1,000,000\ \mu\text{L}}$
   c. $\frac{1,000\text{ ms}}{1\text{ s}}$; $\frac{1\text{ s}}{1,000\text{ ms}}$

2. a. $\frac{1\text{ g}}{100\text{ cg}}$; $\frac{100\text{ cg}}{1\text{ g}}$
   b. $\frac{1,000\text{ mm}}{1\text{ m}}$; $\frac{1\text{ m}}{1,000\text{ mm}}$
   c. $\frac{1,000,000\text{ L}}{1\text{ ML}}$; $\frac{1\text{ ML}}{1,000,000\text{ L}}$

3. 5.62×10^4 m

4. 2.097×10^{-1} s

5. 4.41×10^7 µL

6. 9.0532×10^{-5} g

7. 1.096×10^{11} µg

8. 0.38 km; 3.8×10^{-1} km

9. 3.009×10^3 cL

10. 9.904×10^6 µm

11. 0.27 kg

12. 16,300 mL

13. 0.43 L

14. 11,200 mL

1.8: Dosage Calculations

Exercises

1. Vitamin C tablets can come in 500 mg tablets. How many of these tablets are needed to obtain 10 g of vitamin C?

2. Vitamin C tablets can come in 500 mg tablets. How many of these tablets are needed to obtain 10 g of vitamin C?

3. The recommended daily allowance (RDA) for magnesium for 19-30 yrs old+ men is 400 mg. Magnesium supplements come in 200 mg capsules. How many capsules need to be taken to meet RDA?

4. A 175 lb patient is to undergo surgery and will be given an intravenous anesthetic. The safe dosage of anesthetic is up 12 mg/kg of body weight. Determine the maximum dose of anesthetic that should be used. (Hint: 2.2 lbs = 1 kg)

5. The safe dosage of an IV antibiotic for children weighing more than 2.0kg is 60.mg per kilogram of body weight. How many mg should be administered to a child weighing 16 kg?

6. A drug dose of 1.5 mg/kg is ordered for a child weighing 70.4 lbs. How many mg of the drug should be administered? If the drug is available as 60 mg/2 mL. How many mL must the nurse administer?

Answers

1. 20 tablets

2. 11,000 mg; 11 g

3. 2 tablets

4. 955 mg

5. 960 mg

6. 48 mg; 1.6 mL

1.9: Chapter Summary

Exercises

1. A sample of urine has a density of 1.105 g/cm^3. What is the mass of 0.255 L of this urine?

2. The hardest bone in the body is tooth enamel, which has a density of 2.91 g/cm^3. What is the volume, in liters, of 75.9 g of tooth enamel?

3. Some brands of aspirin have 81 mg of aspirin in each tablet. If a person takes 8 tablets per day, how many grams of aspirin is that person consuming every day?

4. The US government has a recommended daily intake (RDI) of 5 µg of vitamin D per day. (The name *recommended daily allowance* was changed to RDI in 1997.) If milk contains 1.2 µg per 8 oz glass, how many ounces of milk are needed to supply the RDI of vitamin D?

5. The population of the United States, according to the 2000 census, was 281.4 million people.

    a. How many significant figures does this number have?
    b. What is the unit in this quantity?
    c. Express this quantity in proper scientific notation.

6. The United States produces 34,800,000,000 lb of sugar each year, and much of it is exported to other countries.

    a. How many significant figures does this number have?
    b. What is the unit in this quantity?
    c. Express this quantity in proper scientific notation.

7. Construct a conversion factor that can convert from one unit to the other in each pair of units.

    a. from millimeters to kilometers
    b. from kilograms to micrograms
    c. from centimeters to micrometers

8. Construct a conversion factor that can convert from one unit to the other in each pair of units.

    a. from kilometers to micrometers
    b. from decaliters to milliliters
    c. from megagrams to milligrams

9. What is the density of a dextrose solution if 355 mL of the solution has a mass of 406.9 g?

10. What is the density of a dental amalgam (an alloy used to fill cavities) if 1.005 kg of the material has a volume of 433 mL? Express your final answer in grams per milliliter.

For Exercises 11–16, see the accompanying table for the relationships between English and SI units.

For Exercises 11–16, see the accompanying table for the relationships between English and SI units.

1 m ≈ 39.36 in. ≈ 3.28 ft ≈ 1.09 yd
1 in. ≈ 2.54 cm
1 km ≈ 0.62 mi
1 kg ≈ 2.20 lb
1 lb ≈ 454 g
1 L ≈ 1.06 qt
1 qt ≈ 0.946 L

11. Approximately how many inches are in 4.76 m?

12. Approximately how many liters are in 1 gal, which is exactly 4 qt?

13. Approximately how many kilograms are in a person who weighs 170 lb?

14. The average distance between Earth and the sun is 9.3×10^7 mi. How many kilometers is that?

15. Show mathematically that 1 L equals 1 dm^3.

16. Show mathematically that 1 L equals 1,000 cm^3.

Answers

1. 282 g

2. 26.1 cm^3; 0.0261 L

3. 650 mg

4. 30 oz

5. a. four significant figures
   b. people
   c. 2.841×10^8 people

6.

   a. three significant figures
   b. pound (lb)
   c. 3.48×10^{10} lbs

7

   a. $\frac{1\text{ km}}{10^6\text{ mm}}$
   b. $\frac{10^9\ \mu\text{g}}{1\text{ kg}}$
   c. $\frac{10^4\ \mu\text{m}}{1\text{ cm}}$

8. Construct a conversion factor that can convert from one unit to the other in each pair of units.

   a. $\frac{10^9\ \mu\text{m}}{1\text{ km}}$
   b. $\frac{100\text{ mL}}{1\text{ dL}}$
   c. $\frac{10^9\text{ mg}}{1\text{ Mg}}$

9. 1.15 g/mL

10. 2.321 g/mL

11. 187 in.

12. 3.784 L

13. 77 kg

14. 1.5×10^8 km

15. $1\text{ L} = 0.001\text{ m}^3 \times \left(\frac{1\text{ dm}}{0.1\text{ m}}\right)^3 = 1\text{ dm}^3$

16. $1\text{ L} = 0.001\text{ m}^3 \times \left(\frac{1\text{ cm}}{0.01\text{ m}}\right)^3 = 1000\text{ cm}^3$

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3: ELEMENTS, ATOMS, AND THE PERIODIC TABLE

Just as a language has an alphabet from which words are built, chemistry has an alphabet from which matter is described. However, the chemical alphabet is larger than the one we use for spelling. You may have already figured out that the chemical alphabet consists of the chemical elements. Their role is central to chemistry, for they combine to form the millions and millions of known compounds.

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

3.2: The Elements

3.3: Atomic Theory

3.4: The Structure of Atoms

3.5: Nuclei of Atoms

3.6: Atomic Masses

3.7: Arrangements of Electrons

3.8: The Periodic Table

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

3.E.1: Elements, Atoms, and the Periodic Table (Additional Exercises)

3.S: Elements, Atoms, and the Periodic Table (Summary)

Template:HideTOC

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

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

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

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

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3.2: The Elements

Learning Objectives

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

An element is a substance that cannot be broken down into simpler chemical substances. There are about 90 naturally occurring elements known on Earth. Using technology, scientists have been able to create nearly 30 additional elements that do not occur in nature. Today, chemistry recognizes 118 elements—some of which were created an atom at a time. Figure 3.2.1 shows some of the chemical elements.



Figure 3.2.1: Samples of Elements. Gold is a yellowish solid, iron is a silvery solid, while mercury is a silvery liquid at room temperature. © Thinkstock

Abundance

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

Table 3.2.1: Elemental Composition of Earth

Earth's Crust		Earth (overall)	
Element	**Percentage**	**Element**	**Percentage**
oxygen	46.1	iron	34.6
silicon	28.2	oxygen	29.5
aluminum	8.23	silicon	15.2
iron	5.53	magnesium	12.7
calcium	4.15	nickel	2.4
sodium	2.36	sulfur	1.9
magnesium	2.33	all others	3.7
potassium	2.09		
titanium	0.565		

Earth's Crust		Earth (overall)	
Element	Percentage	Element	Percentage
hydrogen	0.14		
phosphorus	0.105		
all others	0.174		
Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 14–17.			

On the planet Earth, however, the situation is rather different. Oxygen makes up 46.1% of the mass of Earth's crust (the relatively thin layer of rock forming Earth's surface), mostly in combination with other elements, while silicon makes up 28.2%. Hydrogen, the most abundant element in the universe, makes up only 0.14% of Earth's crust. Table 3.2.1 lists the relative abundances of elements on Earth as a whole and in Earth's crust. Table 3.2.2 lists the relative abundances of elements in the human body. If you compare Table 3.2.1 and Table 3.2.2, you will find disparities between the percentage of each element in the human body and on Earth. Oxygen has the highest percentage in both cases, but carbon, the element with the second highest percentage in the body, is relatively rare on Earth and does not even appear as a separate entry in Table 3.2.1; carbon is part of the 0.174% representing "other" elements. How does the human body concentrate so many apparently rare elements?

Table 3.2.2: Elemental Composition of a Human Body

Element	Percentage by Mass
oxygen	61
carbon	23
hydrogen	10
nitrogen	2.6
calcium	1.4
phosphorus	1.1
sulfur	0.20
potassium	0.20
sodium	0.14
chlorine	0.12
magnesium	0.027
silicon	0.026
iron	0.006
fluorine	0.0037
zinc	0.0033
all others	0.174
Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 7–24.	

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

LOOKING CLOSER: PHOSPHOROUS, THE CHEMICAL BOTTLENECK

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

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

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

Names and Symbols

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

Today, new elements are usually named after famous scientists.

The names of the elements can be cumbersome to write in full, especially when combined to form the names of compounds. Therefore, each element name is abbreviated as a one- or two-letter chemical symbol. By convention, the first letter of a chemical symbol is a capital letter, while the second letter (if there is one) is a lowercase letter. The first letter of the symbol is usually the first letter of the element's name, while the second letter is some other letter from the name. Some elements have symbols that derive from earlier, mostly Latin names, so the symbols may not contain any letters from the English name. Table 3.2.3 lists the names and symbols of some of the most familiar elements.

Table 3.2.3: Element Names and Symbols

aluminum	Al	magnesium	Mg
argon	Ar	manganese	Mn
arsenic	As	mercury	Hg*
barium	Ba	neon	Ne
bismuth	Bi	nickel	Ni
boron	B	nitrogen	N
bromine	Br	oxygen	O
calcium	Ca	phosphorus	P
carbon	C	platinum	Pt
chlorine	Cl	potassium	K*
chromium	Cr	silicon	Si
***The symbol comes from the Latin name of element.**			
†The symbol for tungsten comes from its German name—*wolfram*.			

copper	Cu*	silver	Ag*
fluorine	F	sodium	Na*
gold	Au*	strontium	Sr
helium	He	sulfur	S
hydrogen	H	tin	Sn*
iron	Fe	tungsten	W†
iodine	I	uranium	U
lead	Pb*	zinc	Zn
lithium	Li	zirconium	Zr
***The symbol comes from the Latin name of element.**			
†The symbol for tungsten comes from its German name—*wolfram*.			

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

Example 3.2.1

Write the chemical symbol for each element without consulting Table 3.2.3 "Element Names and Symbols".

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

Answer a available at the end of this section.

Exercise 3.2.1

Write the chemical symbol for each element without consulting Table 3.2.3.

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

Answer a available at the end of this section.

Example 3.2.2

What element is represented by each chemical symbol?

a. Na

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

Answer a available at the end of this section.

Exercise 3.2.2

What element is represented by each chemical symbol?

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

Answer a available at the end of this section.

Concept Review Exercises

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

Answers

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

Key Takeaways

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

Answers

Example 3.2.1

Br
B
C
Ca
Au

Exercise 3.2.1

Mn
Mg
Ne
N
Ag

Example 3.2.2

sodium
mercury

phosphorus
potassium
iodine

Exercise 3.2.2

lead
tin
uranium
oxygen
fluorine

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3.3: Atomic Theory

Learning Objectives

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

Take some aluminum foil. Cut it in half. Now you have two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil.

It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can you take this exercise, at least in theory? Can you continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? (Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century BC.)

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 3.3.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

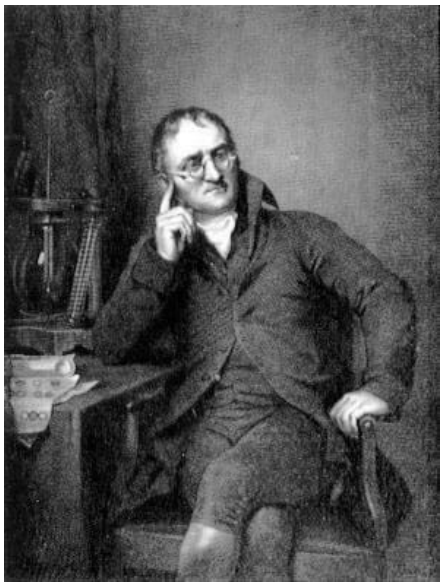


Figure 3.3.1 John Dalton was an English scientist who enunciated the modern atomic theory.

Most elements in their pure form exist as **individual atoms**. For example, a macroscopic chunk of iron (Fe) metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called **diatomic molecules**. In representing a diatomic molecule, we use the symbol of the element and include the ***subscript*** "2" to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), nitrogen (N_2), fluorine (F_2), oxygen (O_2), iodine (I_2), chlorine (Cl_2) and bromine (Br_2). A few other elements can exist as 3-atom molecules like ozone (O_3) and 4-atom molecules like phosphorus (P_4). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is S_8 (Figure 3.3.2).

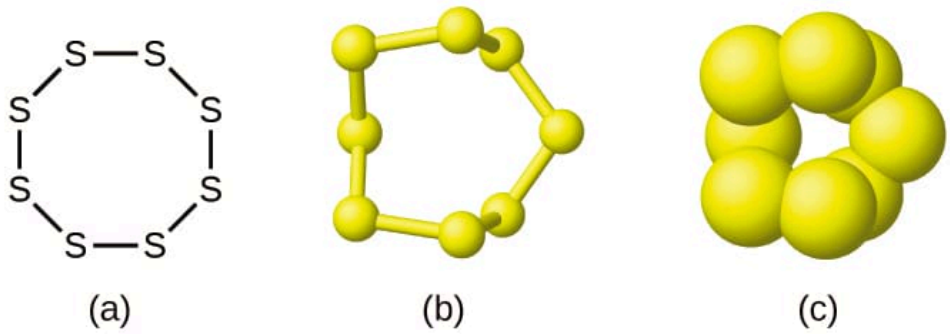


Figure 3.3.2: A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

Figure A shows eight sulfur atoms, symbolized with the letter S, that are bonded to each other to form an octagon. Figure B shows a 3-D, ball-and-stick model of the arrangement of the sulfur atoms. The shape is clearly not octagonal as it is represented in the structural formula. Figure C is a space-filling model that shows each sulfur atom is partially embedded into the sulfur atom it bonds with.

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example, H_2 and 2H represent distinctly different species. H_2 is a molecular formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together. The expression 2H, on the other hand, indicates two separate hydrogen atoms that are **not** combined as a unit. The expression $2H_2$ represents two molecules of diatomic hydrogen (Figure 3.3.3).

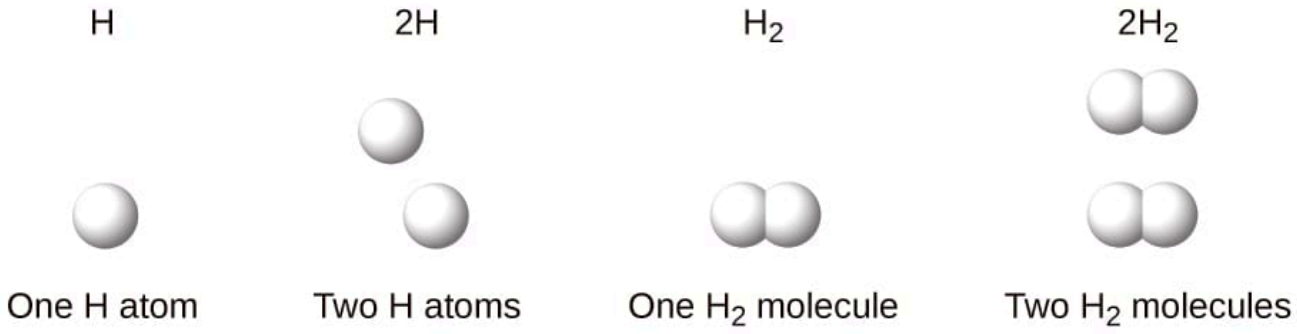


Figure 3.3.3: The symbols H, 2H, H_2, and $2H_2$ represent very different entities.

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

Example 3.3.1

Write the chemical formula of the following elements:

a. oxygen
b. carbon
c. potassium

Solution

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

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

3.3.1 Exerc

Write the chemical formula of the following elements:

a. hydrogen
b. nitrogen
c. neon

Answer a available at the end of this section.

Looking Closer: Atomic Theory

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

Concept Review Exercises

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

Answers

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

Key Takeaways

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

Answers

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

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

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

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3.4: The Structure of Atoms

Learning Objectives

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

There have been several minor but important modifications to Dalton's atomic theory. For one thing, Dalton considered atoms to be indivisible. We know now that atoms not only can be divided but also are composed of three different kinds of particle, *subatomic particles*, with their own properties that are different from the chemical properties of atoms.

Subatomic Particles

The first subatomic particle was identified in 1897 and called the electron. It is an extremely tiny particle, with a mass of about 9.109×10^{-31} kg. Experiments with magnetic fields showed that the electron has a negative electrical charge. By 1920, experimental evidence indicated the existence of a second particle. A proton has the same amount of charge as an electron, but its charge is positive, not negative. Another major difference between a proton and an electron is mass. Although still incredibly small, the mass of a proton is 1.673×10^{-27} kg, which is almost 2,000 times greater than the mass of an electron. Because opposite charges attract each other (while like charges repel each other), protons attract electrons (and vice versa).

Finally, additional experiments pointed to the existence of a third particle. Evidence produced in 1932 established the existence of the neutron, a particle with about the same mass as a proton but with no electrical charge, it is *neutral*. We understand now that all atoms can be broken down into subatomic particles: protons, neutrons, and electrons. Table 3.4.1 lists some of their important characteristics and the symbols used to represent each particle.

Table 3.4.1: Properties of the Subatomic Particles

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

The Nucleus

How are these subatomic particles arranged? Between 1909 and 1911, Ernest Rutherford, a Cambridge physicist, and his associates Hans Geiger and Ernest Marsden performed experiments that provided strong evidence concerning the internal structure of an atom. They took a very thin metal foil, such as gold or platinum, and aimed a beam of positively charged particles (called alpha particles, which are combinations of two protons and two neutrons) from a radioactive source toward the foil. Surrounding the foil was a detector—either a scintillator (a material that glows when hit by such particles) or some unexposed film (which is exposed where the particles hit it). The detector allowed the scientists to determine the distribution of the alpha particles after they interacted with the foil. Figure 3.4.1 shows a diagram of the experimental setup.

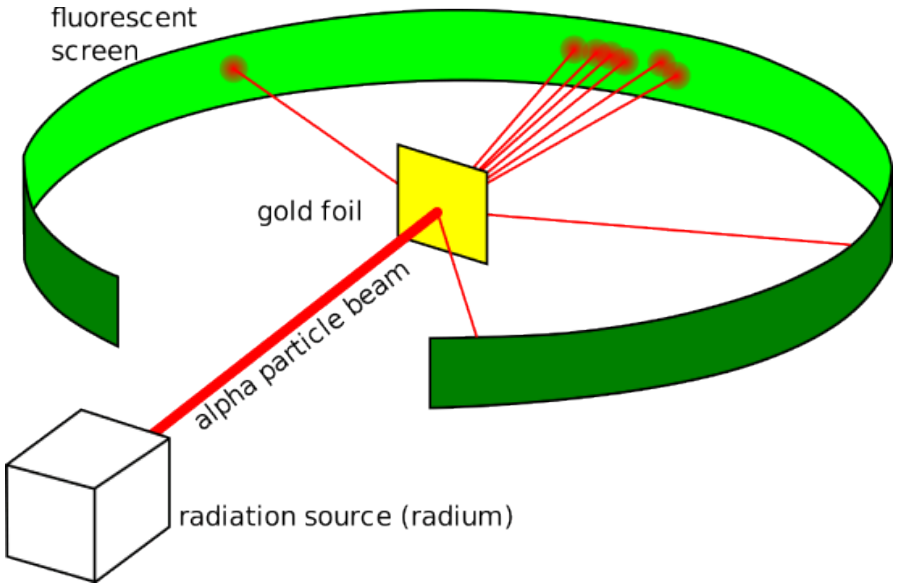


Figure 3.4.1: The Geiger-Marsden Experimental Setup. Experiments using this setup were used to investigate the structure of atoms. In this experiment, most of the particles traveled straight through the foil, but some alpha particles were deflected off to one side. Some were even deflected back toward the source. This was unexpected. Rutherford once said, "It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." (CC SA-BY 3.0; Kurzon).

Rutherford proposed the following model to explain these experimental results. Protons and neutrons are concentrated in a central region he called the nucleus (plural, *nuclei*) of the atom. Electrons are outside the nucleus and orbit about it because they are attracted to the positive charge in the nucleus. Most of the mass of an atom is in the nucleus, while the orbiting electrons account for an atom's size. As a result, an atom consists largely of empty space. Rutherford called his description the "planetary model" of the atom. Figure 3.4.2 shows how this model explains the experimental results.

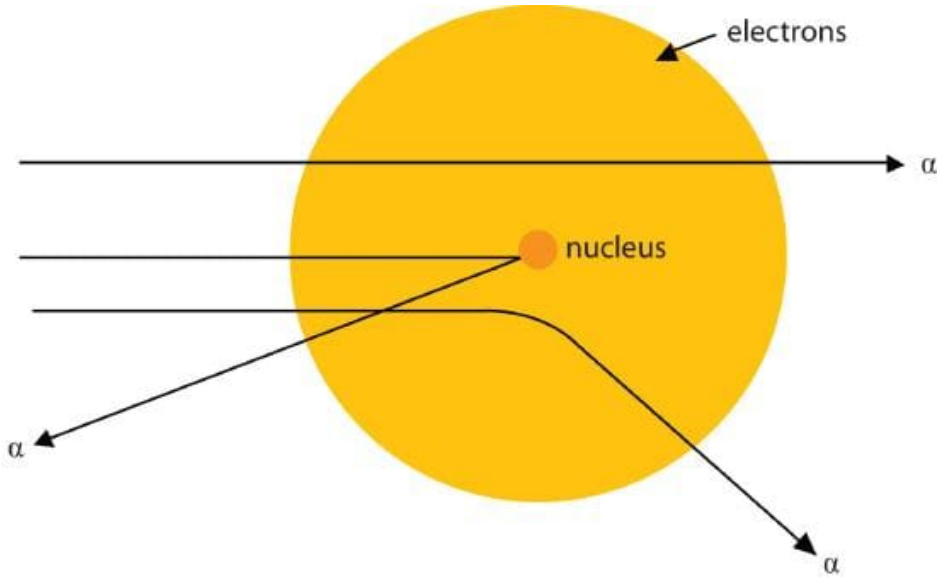


Figure 3.4.2: Rutherford's Metal Foil Experiments. Rutherford explained the results of the metal foil experiments by proposing that most of the mass and the positive charge of an atom are located in its nucleus, while the relatively low-mass electrons orbit about the nucleus. Most alpha particles go straight through the empty space, a few particles are deflected, and fewer still ricochet back toward the source. The nucleus is much smaller proportionately than depicted here.

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

Rutherford's model is essentially the same model that we use today to describe atoms but with one important modification. The planetary model suggests that electrons occupy certain specific, circular orbits about the nucleus. We know now that this model is

overly simplistic. A better description is that electrons form fuzzy clouds around nuclei. Figure 3.4.3 shows a more modern version of our understanding of atomic structure.

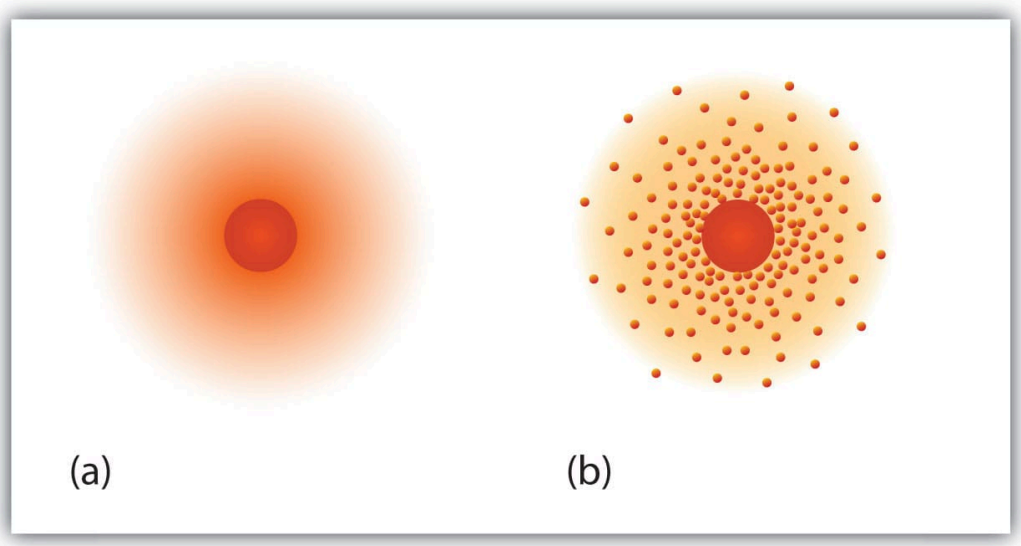


Figure 3.4.3: A Modern Depiction of Atomic Structure. A more modern understanding of atoms, reflected in these representations of the electron in a hydrogen atom, is that electrons occupy regions of space about the nucleus; they are not in discrete orbits like planets around the sun. (a) The darker the color, the higher the probability that an electron will be at that point. (b) In a two-dimensional cross section of the electron in a hydrogen atom, the more crowded the dots, the higher the probability that an electron will be at that point. In both (a) and (b), the nucleus is in the center of the diagram.

Concept Review Exercises

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

Answers

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

Key Takeaways

- Atoms are composed of three main subatomic particles: protons, neutrons, and electrons.
- Protons and neutrons are grouped together in the nucleus of an atom, while electrons orbit about the nucleus.

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3.5: Nuclei of Atoms

Learning Objectives

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

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

Atomic Number

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

For example, hydrogen has the atomic number 1; all hydrogen atoms have 1 proton in their nuclei. Helium has the atomic number 2; all helium atoms have 2 protons in their nuclei. There is no such thing as a hydrogen atom with 2 protons in its nucleus; a nucleus with 2 protons would be a helium atom. The atomic number *defines* an element. Table $3.5.1$ lists some common elements and their atomic numbers. Based on its **atomic number**, you can determine the **number of protons in the nucleus** of an atom. The largest atoms have over 100 protons in their nuclei.

Example $3.5.1$

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

a. aluminum
b. iron
c. carbon

Answer a available at the end of this section.

Exercise $3.5.1$

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

a. sodium
b. oxygen
c. chlorine

Answer a available at the end of this section.

How many electrons are in an atom? Previously we said that for an electrically neutral atom, the **number of electrons equals the number of protons**, so the total opposite charges cancel. Thus, the **atomic number** of an element also gives the **number of electrons** in an atom of that element. (Later we will find that some elements may gain or lose electrons from their atoms, so those atoms will no longer be electrically neutral. Thus we will need a way to differentiate the number of electrons for those elements.)

Example 3.5.2

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

a. sulfur
b. tungsten
c. argon

Answer a available at the end of this section.

Exercise 3.5.2

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

a. magnesium
b. potassium
c. iodine

Answer a available at the end of this section.

Table 3.5.1: Some Common Elements and Their Atomic Numbers

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

Isotopes

How many neutrons are in atoms of a particular element? At first it was thought that the number of neutrons in a nucleus was also characteristic of an element. However, it was found that atoms of the same element can have *different* numbers of neutrons. Atoms of the same element (i.e., same atomic number, Z) that have different numbers of neutrons are called isotopes. For example, 99% of the carbon atoms on Earth have 6 neutrons and 6 protons in their nuclei; about 1% of the carbon atoms have 7 neutrons in their nuclei. Naturally occurring carbon on Earth, therefore, is actually a mixture of isotopes, albeit a mixture that is 99% carbon with 6 neutrons in each nucleus.

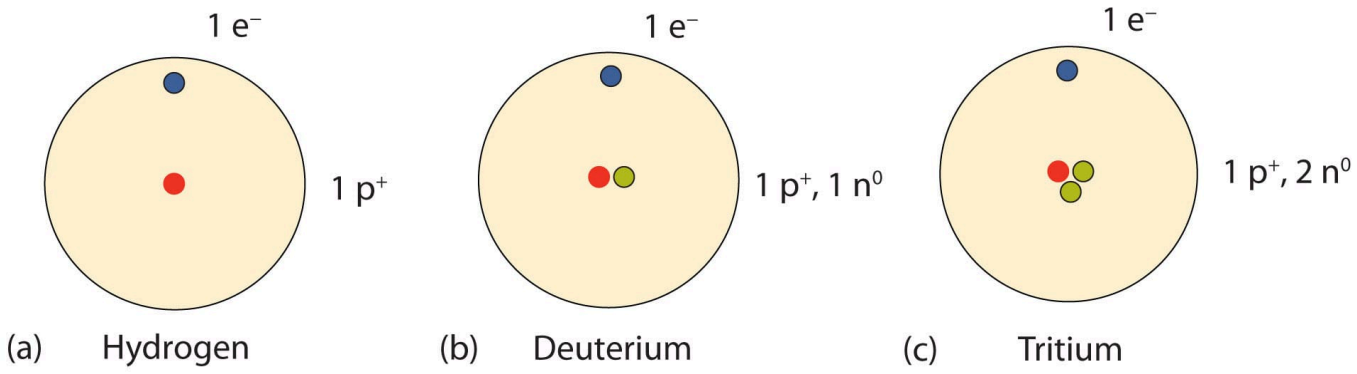


Figure 3.5.1: Isotopes of Hydrogen. Most hydrogen atoms have only a proton in the nucleus (a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. Figure 3.5.1 compares the three isotopes of hydrogen.

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

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

A simple way of indicating the mass number of a particular isotope is to list it as a superscript on the left side of an element's symbol. Atomic numbers are often listed as a subscript on the left side of an element's symbol. Thus, we might see

$$\begin{matrix}\text{mass number} \longrightarrow 56 \\ \text{atomic number} \longrightarrow 26\end{matrix}\text{Fe} \tag{3.5.1}$$

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

Example 3.5.3

How many protons and neutrons are in each atom?

a. $^{35}_{17}\text{Cl}$

b. $^{127}_{53}\mathrm{I}$

Answer a available at the end of this section.

Exercise 3.5.3

How many protons and neutrons are in each atom?

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

Answer a available at the end of this section.

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

Summary

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

Almost all of the mass of an atom is from the total protons and neutrons contained within a tiny (and therefore very dense) nucleus. The majority of the volume of an atom is the surrounding space in which the electrons reside. A representation of a carbon-12 atom is shown below in Figure 3.5.2.

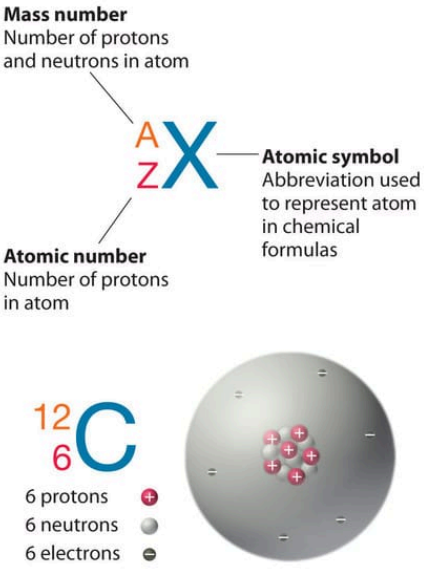


Figure 3.5.2: Formalism used for identifying specific nuclide (any particular kind of nucleus)

The atomic symbol, X, is the abbreviation used to represent an atom in chemical formulas. The mass number, A, is the number of protons and neutrons in the atom, which is to the left of the X. The atomic number, Z, is the number of protons in the atom, which is to the left of the X and below the A.

Concept Review Exercises

1. Why is the atomic number so important to the identity of an atom?

2. What is the relationship between the number of protons and the number of electrons in an atom?
3. How do isotopes of an element differ from each other?
4. What is the mass number of an element?

Answers

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

Key Takeaways

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

Answers

Example 3.5.1

According to Table 2.4.1, aluminum has an atomic number of 13. Therefore, every aluminum atom has 13 protons in its nucleus.
Iron has an atomic number of 26. Therefore, every iron atom has 26 protons in its nucleus.
Carbon has an atomic number of 6. Therefore, every carbon atom has 6 protons in its nucleus.

Exercise 3.5.1

Sodium has 11 protons in its nucleus.
Oxygen has 8 protons in its nucleus.
Chlorine has 17 protons in its nucleus

Example 3.5.2

The atomic number of sulfur is 16. Therefore, in a neutral atom of sulfur, there are 16 electrons.
The atomic number of tungsten is 74. Therefore, in a neutral atom of tungsten, there are 74 electrons.
The atomic number of argon is 18. Therefore, in a neutral atom of argon, there are 18 electrons.

Exercise 3.5.2

Mg has 12 electrons.
K has 19 electrons.
I has 53 electrons.

Example 3.5.3

In $^{35}_{17}\mathrm{Cl}$ there are 17 protons, and 35 – 17 = 18 neutrons in each nucleus.
In $^{127}_{53}\mathrm{I}$ there are 53 protons, and 127 – 53 = 74 neutrons in each nucleus.

Exercise 3.5.3

In $^{197}_{79}\mathrm{Au}$ there are 79 protons, and 197 – 79 = 118 neutrons in each nucleus.
In $^{23}_{11}\mathrm{Na}$ there are 11 protons, and 23 – 11 = 12 neutrons in each nucleus.

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3.6: Atomic Masses

Learning Objectives

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

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

Atomic Mass Unit

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

$$1\ \text{u} = \frac{1}{12}\ \text{the mass of}\ {}^{12}\text{C}\ \text{atom} \tag{3.6.1}$$

It is equal to 1.661×10^{-24} g.

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

Atomic Mass is the Weighted Average Mass of Isotopes

As stated above, most elements occur naturally as a mixture of two or more isotopes. Listed below (Table 3.6.1) are the naturally occurring isotopes of selected elements along with the percent natural abundance of each.

Table 3.6.1: Atomic Masses and Percent Abundances of Some Natural Isotopes

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

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

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

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

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

$$0.7577\,(34.969\,u) + 0.2423\,(36.966\,u) = 35.453\,u$$

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

$$\text{Atomic mass} = (\%1)(\text{mass 1}) + (\%2)(\text{mass 2}) + \cdots$$

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

$$\text{Atomic mass} = (\%1)(\text{mass 1}) + (\%2)(\text{mass 2}) + (\%3)(\text{mass 3})$$

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

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

Example 3.6.1

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

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

Solution

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

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

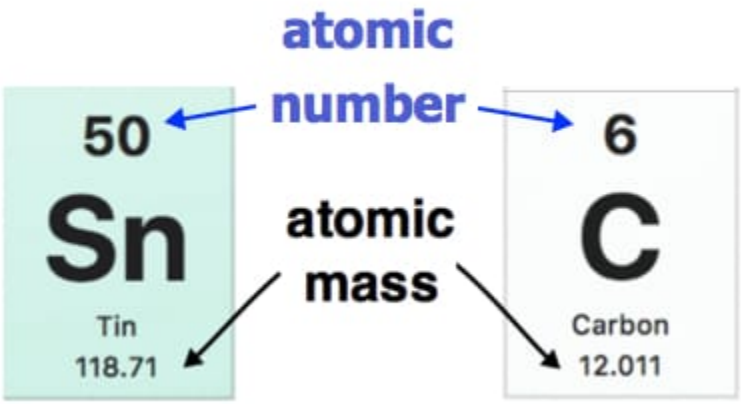
3.6.1 Exerc

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

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

Answer available at the end of this section.

The **atomic mass** of each element is found under the element symbol in the **periodic table.** Examples are shown below. The atomic mass of tin (Sn) is 118.71 u while the atomic mass of carbon (C) is 12.011 u. On the other hand, the **atomic number (Z)** of each element is found **above** the atomic symbol.



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

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

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

Example 3.6.2: Mass of Carbon

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

Solution

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

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

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

Exercise 3.6.2: Mass of Tin

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

Answer available at the end of this section.

Concept Review Exercises

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

Answers

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

Key Takeaway

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

Answers

63.546 u

Exercise 3.6.2: Mass of Tin

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

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3.7: Arrangements of Electrons

Learning Objectives

- Describe how electrons are grouped within atoms.

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

The modern theory of electron behavior is called quantum mechanics. It makes the following statements about electrons in atoms:

- Electrons in atoms can have only certain specific energies. We say that the energies of the electrons are quantized.
- Electrons are organized according to their energies into sets called **shells** (labeled by the principle quantum number, ***n***). Generally the higher the energy of a shell, the farther it is (on average) from the nucleus. Shells do not have specific, fixed distances from the nucleus, but an electron in a higher-energy shell will spend more time farther from the nucleus than does an electron in a lower-energy shell.
- Shells are further divided into subsets of electrons called **subshells**. The first shell has only one subshell, the second shell has two subshells, the third shell has three subshells, and so on. The subshells of each shell are labeled, in order, with the letters s, *p*, *d*, and *f*. Thus, the first shell has only a single *s* subshell (called **1s**), the second shell has **2s** and **2*p*** subshells, the third shell has **3s**, **3*p***, and **3*d*** and so forth.

Table 3.7.1: Shells and Subshells

Shell	Number of Subshells	Names of Subshells
1	1	*1s*
2	2	*2s* and *2p*
3	3	*3s*, *3p* and *3d*
4	4	*4s*, *4p*, *4d* and *4f*

- Different subshells hold a different maximum number of electrons. **Any s subshell can hold up to 2 electrons; *p*, 6; *d*, 10; and *f*, 14.**

Table 3.7.2: Number of Electrons

Subshell	Maximum Number of Electrons
s	2
p	6
d	10
f	14

It is the arrangement of electrons into shells and subshells that most concerns us here, so we will focus on that.

We use numbers to indicate which shell an electron is in. As shown in Table 3.7.1, the first shell, closest to the nucleus and with the lowest-energy electrons, is shell 1. This first shell has only one subshell, which is labeled 1s and can hold a maximum of 2 electrons. We combine the shell and subshell labels when referring to the organization of electrons about a nucleus and use a superscript to indicate how many electrons are in a subshell. Thus, because a hydrogen atom has its single electron in the s subshell of the first shell, we use $\mathbf{1s^1}$ to describe the electronic structure of hydrogen. This structure is called an electron configuration. Electron configurations are shorthand descriptions of the arrangements of electrons in atoms. The electron configuration of a hydrogen atom is spoken out loud as “one-ess-one.”

Helium atoms have 2 electrons. Both electrons fit into the 1s subshell because s subshells can hold up to 2 electrons; therefore,

the electron configuration for helium atoms is **$1s^2$** (spoken as “one-ess-two”).

The 1s subshell cannot hold 3 electrons (because an *s* subshell can hold a maximum of 2 electrons), so the electron configuration for a lithium atom cannot be $1s^3$. Two of the lithium electrons can fit into the 1s subshell, but the third electron must go into the second shell. The second shell has two subshells, *s* and *p*, which fill with electrons in that order. The 2*s* subshell holds a maximum of 2 electrons, and the 2*p* subshell holds a maximum of 6 electrons. Because lithium’s final electron goes into the 2*s* subshell, we write the electron configuration of a lithium atom as **$1s^2 2s^1$.** The shell diagram for a lithium atom is shown below. The shell closest to the nucleus (first shell) has 2 dots representing the 2 electrons in **1s**, while the outermost shell (**2s**) has 1 electron.

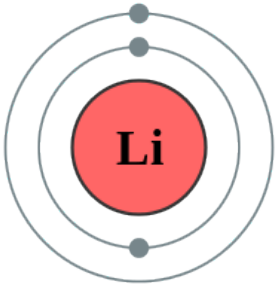


Figure 3.7.1: Shell diagram of lithium (Li) atom.

The next largest atom, beryllium, has 4 electrons, so its electron configuration is $1s^2 2s^2$. Now that the 2*s* subshell is filled, electrons in larger atoms start filling the 2*p* subshell. Thus, the electron configurations for the next six atoms are as follows:

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

With neon, the 2*p* subshell is completely filled. Because the second shell has only two subshells, atoms with more electrons now must begin the third shell. The third shell has three subshells, labeled *s*, *p*, and *d*. The *d* subshell can hold a maximum of 10 electrons. The first two subshells of the third shell are filled in order—for example, the electron configuration of aluminum, with 13 electrons, is $1s^2 2s^2 2p^6 3s^2 3p^1$. However, a curious thing happens after the 3*p* subshell is filled: the 4*s* subshell begins to fill before the 3*d* subshell does. In fact, the exact ordering of subshells becomes more complicated at this point (after argon, with its 18 electrons), so we will not consider the electron configurations of larger atoms. A fourth subshell, the *f* subshell, is needed to complete the electron configurations for all elements. An *f* subshell can hold up to 14 electrons.

Electron filling always starts with **1s**, the subshell closest to the nucleus. Next is **2s, 2*p*, 3s, 3*p*, 4s, 3*d*, 4*p*, 5s, 4d, 5p, 6s,** etc., shown in the electron shell filling order diagram in Figure 3.7.2. Follow each arrow in order **from top to bottom**. The subshells you reach along each arrow give the ordering of filling of subshells in larger atoms.

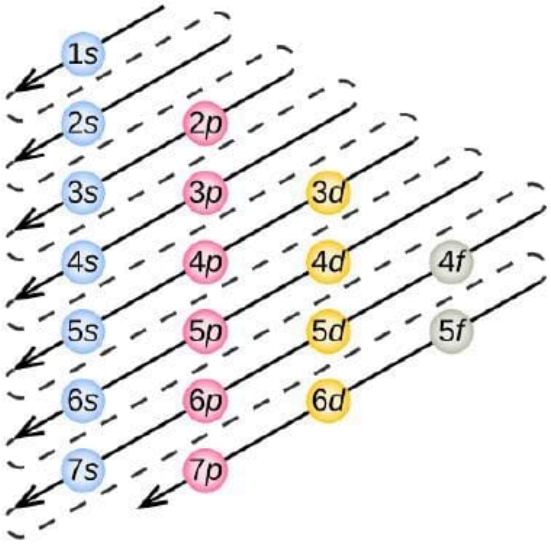


Figure 3.7.2: The order of electron filling in an atom.

Example 3.7.1: Electronic Configuration of Phosphorus Atoms

Using Figure 3.7.2 as your guide, write the electron configuration of a neutral phosphorus atom. The atomic number of P is 15.

Solution

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

Exercise 3.7.1: Electronic Configuration of Chlorine Atoms

Using Figure 3.7.2 as your guide, write the electron configuration of a neutral chlorine atom. The atomic number of Cl is 17.

Answer available at the end of this section.

Since the arrangement of the periodic table is based on the electron configurations, Figure 3.7.3 provides an alternative method for determining the electron configuration. The filling order simply **begins at the top left,** with hydrogen (Z=1) and includes each subshell as you proceed in increasing atomic number (*Z)* order.

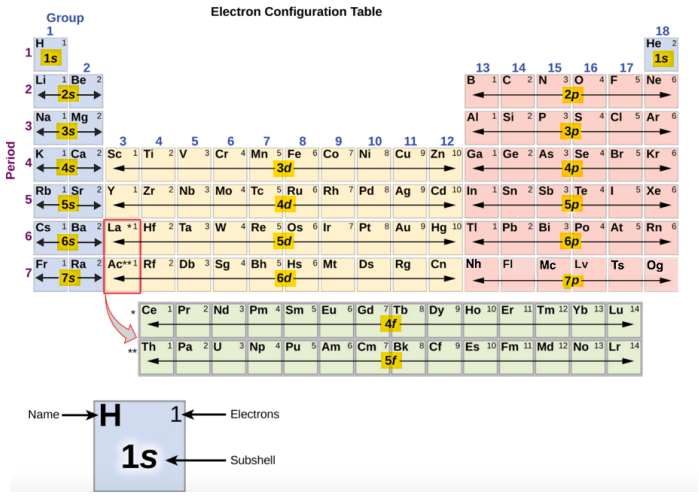


Figure 3.7.3: This periodic table shows the electron configuration for each subshell. By "building up" from hydrogen, this table can be used to determine the electron configuration for any atom on the periodic table.

For example, the first row (Period 1) contains H and He only, because only two electrons are required to fill the 1s subshell. The second row s-block, contains only two elements, Li and Be, to fill the 2s subshell. This is followed by the second row p-block, containing 6 elements (B through Ne) since six electrons are required to fill the 2p subshell. The third row is similar to the second row elements. Two electrons are needed (Na and Mg) to fill the 3s subshell and six electrons are required (Al through Ar) to complete the 3p subshell. After filling the 3*p* block up to Ar, we see the next subshell will be 4s (K, Ca), followed by the 3*d* subshell, which are filled by ten electrons (Sc through Zn). The 4p subshell is filled next by six electrons (Ga through Kr). As you can see, the periodic table shown in Figure 3.7.3 provides a simple way to remember the order of filling the subshells in determining the electron configuration. The order of filling subshells is the same: **1s, 2s, 2*p*, 3s, 3*p*, 4s, 3*d*, 4*p*, 5s, 4d, 5p, 6s,** etc.

Example 3.7.2: Aluminum

Using Figure 3.7.3 as your guide, write the electron configuration of neutral aluminum atom. The atomic number of Al is 13.

Solution

Aluminum has 13 electrons.

Start at Period 1 of the periodic table, Figure 3.7.3. Place two electrons in the 1s subshell (**$1s^2$**).

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

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

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

3.7.2Exerc

Using Figure 3.7.3 as your guide, write the electron configuration of the atom that has 20 electrons

Answer available at the end of this section.

Valence Electrons

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

In the second period elements, the two electrons in the $1s$ sublevel are called **inner-shell electrons** and are not involved directly in the element's reactivity or in the formation of compounds. Lithium has a single electron in the second principal energy level and so we say that lithium has one valence electron. Beryllium has two valence electrons. How many valence electrons does boron have? You must recognize that the second principal energy level consists of both the $2s$ and the $2p$ sublevels and so the answer is three. In fact, the number of valence electrons goes up by one for each step across a period until the last element is reached. Neon, with its configuration ending in $2s^2 2p^6$, has eight valence electrons.

The alkali metal sodium (atomic number 11) has one more electron than the neon atom. This electron must go into the lowest-energy subshell available, the 3s orbital, giving a $1s^2 2s^2 2p^6 3s^1$ configuration. The electrons occupying the outermost shell orbital(s) (highest value of *n*) are called valence electrons, and those occupying the inner shell orbitals are called **core electrons** (Figure \PageIndex4). Since the core electron shells correspond to noble gas electron configurations, we can abbreviate electron configurations by writing the noble gas that matches the core electron configuration, along with the valence electrons in a condensed format. For our sodium example, the symbol [Ne] represents core electrons, ($1s^2 2s^2 2p^6$) and our abbreviated or condensed configuration is $[Ne]3s^1$.



Figure 3.7.4: A core-abbreviated electron configuration (right) replaces the core electrons with the noble gas symbol whose configuration matches the core electron configuration of the other element.

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

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

A chemical reaction results from electron removal, electron addition, or electron sharing of the valence electrons of the different atoms. The path a specific element will take depends on where the electrons are in the atom and how many there are. Thus, it is convenient to separate electrons into two groups. Valence shell electrons (or, more simply, the *valence electrons*) are the electrons in the highest-numbered shell, or valence shell, while core electrons are the electrons in lower-numbered shells. We can see from the electron configuration of a carbon atom—$1s^2 2s^2 2p^2$—that it has 4 valence electrons ($2s^2 2p^2$) and 2 core electrons ($1s^2$). You will see in the next chapters that the chemical properties of elements are determined by the number of valence electrons.

Example 3.7.3

Examine the electron configuration of neutral phosphorus atoms in Example 3.7.1, $1s^2 2s^2 2p^6 3s^2 3p^3$ and write the abbreviated notation.

Solution

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

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

The 10 inner shell (core) electrons, $1s^2 2s^2 2p^6$ can be replaced by [Ne] (see Figure 3.7.3). Abbreviated notation is : $[Ne]3s^2 3p^3$

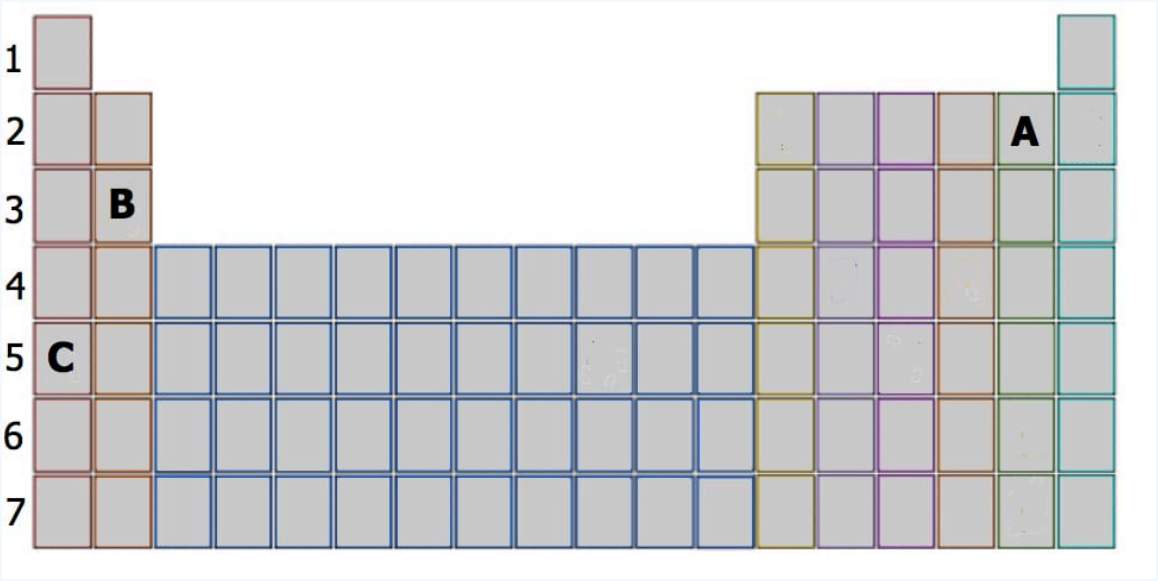
Exercise 3.7.3

Examine the electron configuration of neutral calcium atom (Exercise 3.7.2), $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$, and write the abbreviated notation.

Answer available at the end of this section.

Example 3.7.4

Based on their respective locations in the periodic table (use Figure 3.7.3), determine the number of valence electrons and the valence shell configuration of elements A, B and C.



Solution

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

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

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

3.7.4Exerc

Using the location of Na is the periodic table (Figure 3.7.3), draw the shell diagram of sodium atom.

Answer available at the end of this section.

Concept Review Exercises

1. How are electrons organized in atoms?
2. What information does an electron configuration convey?
3. What is the difference between core electrons and valence electrons?

Answers

1. Electrons are organized into shells and subshells around nuclei.
2. The electron configuration states the arrangement of electrons in shells and subshells.
3. Valence electrons are in the highest-numbered shell; all other electrons are core electrons.

Key Takeaway

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

Exercises

Answers

Exercise 3.7.1: Electronic Configuration of Chlorine Atoms

A neutral chlorine atom has 17 electrons. Two electrons can go into the 1s subshell, 2 can go into the 2*s* subshell, and 6 can go into the 2*p* subshell. That leaves 7 electrons. Of those 7 electrons, 2 can go into the 3s subshell, and the remaining 5 electrons can go into the 3*p* subshell. Thus, the electron configuration of neutral chlorine atoms is $1s^2 2s^2 2p^6 3s^2 3p^5$.

Start at Period 1 of Figure 3.7.3. Place two electrons in the 1s subshell (**$1s^2$**).

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

Proceed to Period 3 (left to right direction). Place the next two electrons in the 3s subshell (**$3s^2$**) and the next six electron in the 3p subshell (**$3p^6$**).

Proceed to Period 4. Place the remaining two electrons in the 4s subshell (**$4s^2$).**

The electron configuration is **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$**

Exercise 3.7.3

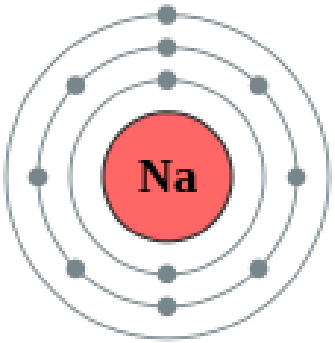
The highest-numbered shell is the fourth shell $4s^2$, which has **2 electrons** in the *4s* subshell. Hence, Calcium has **2 valence electrons**.

The 18 inner-shell (core) electrons, $1s^2 2s^2 2p^6 3s^2 3p^6$, can be replaced by [Ar], see Figure 3.7.3. The abbreviated notation is: [Ar]$4s^2$

Sodium (Na) is the first element in the 3rd row (Period 3) in the periodic table. This means that the first shell and second shells of Na atom are filled to the maximum number of electrons.

The first shell **(1s)** *is filled with* ***2 electrons***. The second shell (2s and 2p) has a **total of *8 electrons***. And, the third (last) shell has ***1 electron***.

The shell diagram of the Na atom is shown below. The shell nearest the nucleus (first shell) has 2 electrons (2 dots), the second shell has 8 electrons and the last (outermost) shell has 1 electron. (**2.8.1**)



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3.8: The Periodic Table

Learning Objectives

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

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

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

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

The periodic table is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern version is shown in Figure 3.8.1. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.

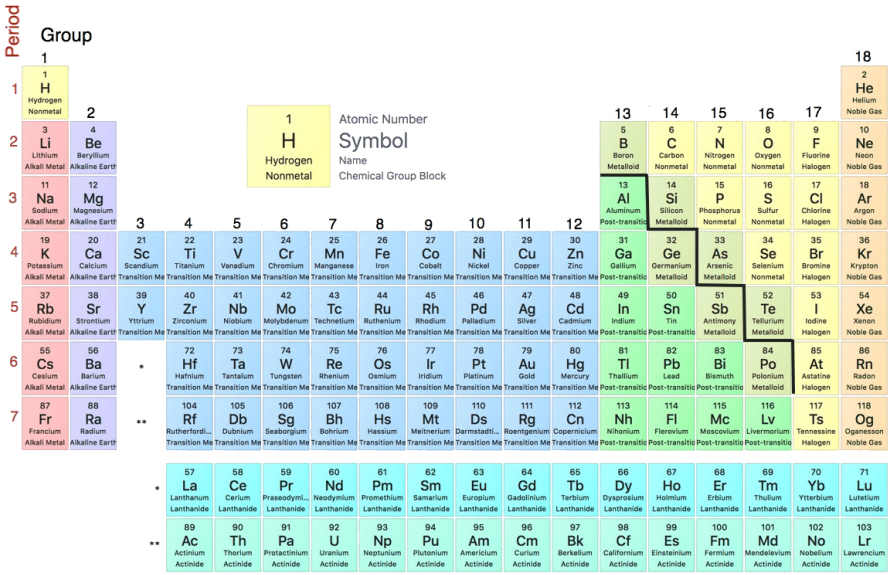


Figure 3.8.1: Modern Periodic Table. (Public Domain; PubChem modified Leticia Colmenares).

Features of the Periodic Table

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

Each row of elements on the periodic table is called a **period**. Periods have different lengths; the first period has only 2 elements

(hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Metals, Nonmetals and Metalloids

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a semimetal, as shown in Figure 3.8.2. A metal is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A nonmetal is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 3.8.2, **metals** occupy the left three-fourths of the periodic table, while **nonmetals** (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **semimetals** (or **metalloids**). Elements adjacent to the **bold zigzag** line in the right-hand portion of the periodic table have semimetal properties.

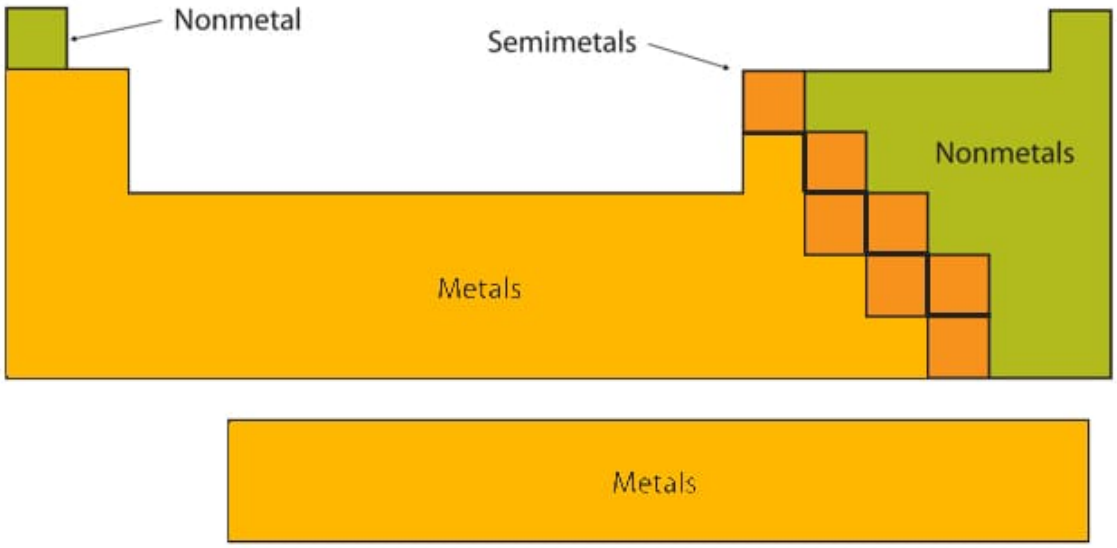


Figure 3.8.2: Types of Elements. Elements are either metals, nonmetals, or semimetals. Each group is located in a different part of the periodic table.

Example 3.8.1

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

Solution

The atomic number of selenium is 34, which places it in period 4 and group 16. In Figure 2.7.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal. Note, however, that because selenium is close to the metal-nonmetal dividing line, it would not be surprising if selenium were similar to a semimetal in some of its properties.

3.8.1 Exerc

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

Answer available at the end of this section.

Representative, Transition and Inner-transition

Another way to categorize the elements of the periodic table is shown in Figure 3.8.3. The first two columns on the left and the last six columns on the right are called the **main group** or **representative** elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the inner transition metals. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

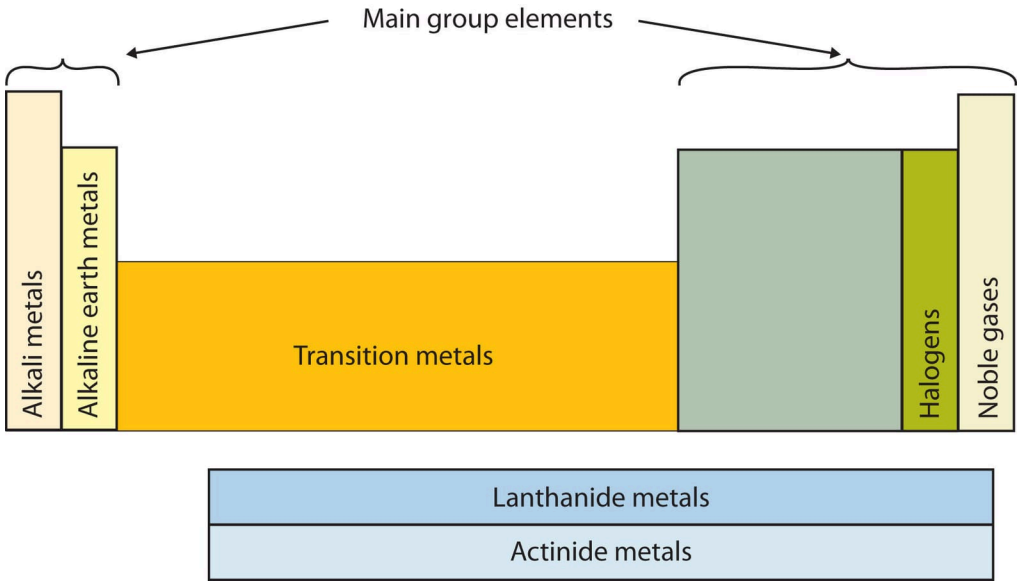


Figure 3.8.3: Special Names for Sections of the Periodic Table. Some sections of the periodic table have special names. The elements lithium, sodium, potassium, rubidium, cesium, and francium are collectively known as alkali metals.

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

To Your Health: Transition Metals in the Body

Most of the elemental composition of the human body consists of main group elements. The first element appearing on the list that is not a main group element is iron, at 0.006 percentage by mass. Because iron has relatively massive atoms, it would appear even lower on a list organized in terms of percent by *atoms* rather than percent by mass.

Iron is a transition metal. Transition metals have interesting chemical properties, partially because some of their electrons are in *d* subshells. The chemistry of iron makes it a key component in the proper functioning of red blood cells.

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

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

Other transition metals have important functions in the body, despite being present in low amounts. Zinc is needed for the body's immune system to function properly, as well as for protein synthesis and tissue and cell growth. Copper is also needed for several proteins to function properly in the body. Manganese is needed for the body to metabolize oxygen properly. Cobalt is a necessary component of vitamin B-12, a vital nutrient. These last three metals are not listed explicitly in Table 2.1.2, so they are present in the body in very small quantities. However, even these small quantities are required for the body to function properly.

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group.

Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins. For example, the elements of Group 1 are known as the alkali metals, Group 2 are the alkaline earth metals, Group 17 are the halogens, and Group 18 are the noble gases.

Group 1: The Alkali Metals

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

Group 2: The Alkaline Earth Metals

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

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for "salt forming," which reflects that all the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth's ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

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

To Your Health: Radon

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

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

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

Why do elements in a given group have similar properties?

The periodic table is organized on the basis of similarities in elemental properties, but what explains these similarities? It turns out that the shape of the periodic table reflects the filling of subshells with electrons, as shown in Figure 3.8.4. Starting with the first period and going from left to right, the table reproduces the order of filling of the electron subshells in atoms. Furthermore, elements in the same *group* share the same valence shell electron configuration. For example, all elements in the first column have a single s electron in their valence shells, so their electron configurations can be described as ns^1 (where n represents the shell number). This last observation is crucial. Chemistry is largely the result of interactions between the valence electrons of different atoms. Thus, atoms that have **the same valence shell electron configuration** will have **similar chemistry.**

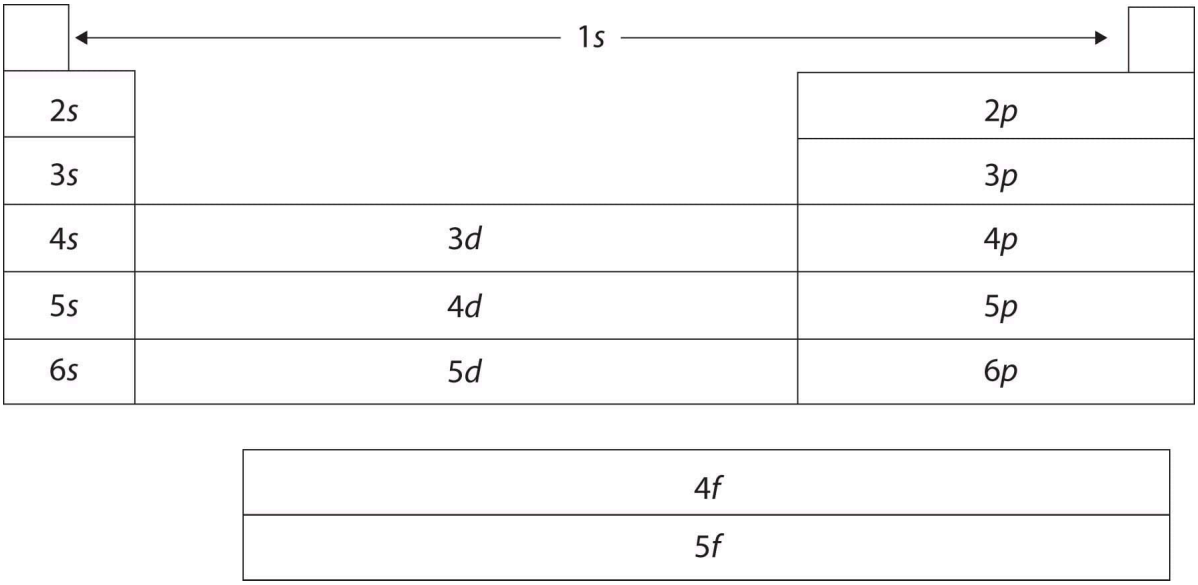


Figure 3.8.4 The Shape of the Periodic Table. The shape of the periodic table reflects the order in which electron shells and subshells fill with electrons.

Example 3.8.1

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

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

Answer a available at the end of this section.

Exercise 3.8.1

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

a. the halogens

b. the column of elements headed by oxygen

Answer a available at the end of this section.

Valence Electrons and Group Number

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

Table 3.8.1. The Group number and the number of valence electrons.

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

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

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

Atomic Radius

The periodic table is useful for understanding atomic properties that show periodic trends. One such property is the atomic radius (Figure 3.8.5). The atomic radius is defined as one-half the distance between the nuclei of identical atoms that are bonded together. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be $74\ \text{pm}$. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2} = 37\ \text{pm}$.

As mentioned earlier, the higher the shell number, the farther from the nucleus the electrons in that shell are likely to be. In other words, the size of an atom is generally determined by the number of the valence electron shell. Therefore, as we go down a column on the periodic table, the atomic radius increases. As we go *across* a period on the periodic table, however, electrons are being added to the *same* valence shell; meanwhile, more protons are being added to the nucleus, so the positive charge of the nucleus is increasing. The increasing positive charge attracts the electrons more strongly, pulling them closer to the nucleus. Consequently, as we go across a period, from left to right, the atomic radius decreases. These trends are seen clearly in Figure 3.8.5

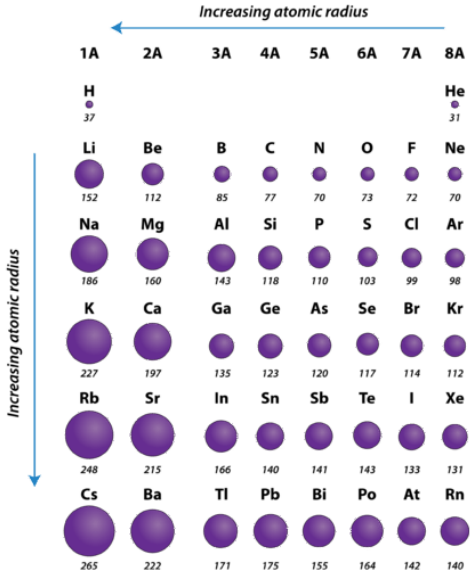


Figure 3.8.5 Trends on the Periodic Table. Atomic radii of the representative elements measured in picometers. The relative sizes of the atoms show several trends with regard to the structure of the periodic table. Atoms become larger going down a group and going from right to left across a period.

Example 3.8.2

Using the periodic table (rather than Figure 3.8.5), which atom is larger?

a. N or Bi
b. Mg or Cl

Answer a available at the end of this section.

Exercise 3.8.2

Using the periodic table (rather than Figure 3.8.5), which atom is larger?

a. Li or F
b. Na or K

Answer a available at the end of this section.

Career Focus: Clinical Chemist

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

Concept Review Exercises

1. How are the elements organized into the periodic table?
2. Looking at the periodic table, where do the following elements appear?

a. the metals
b. the nonmetals
c. the halogens
d. the transition metals
e. the noble gases

3. Describe the trends in atomic radii as related to an element's position on the periodic table.

Answers

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

Key Takeaways

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

Answers

metal

Example 3.8.1

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

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

Exercise 3.8.1

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

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

Example 3.8.2

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

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

Exercise 3.8.2

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

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

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3.E: ELEMENTS, ATOMS, AND THE PERIODIC TABLE (EXERCISES)

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

2.1: The Elements

Exercises

1. Which of the following substances are elements?

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

2. Which of the following substances are elements?

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

3. Write the chemical symbol for each element.

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

4. Write the chemical symbol for each element.

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

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

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

7. Complete the following table.

Element Symbol	Element Name
F	
Fe	
I	
Cr	
C	

Element Symbol	Element Name
P	

8. Complete the following table.

Element Symbol	Element Name
Mg	
Mn	
Ca	
Cl	
K	
Pt	

Answers

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

2.

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

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

4.

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

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

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

7.

Element Symbol	Element Name
F	fluorine
Fe	iron

Element Symbol	Element Name
I	iodine
Cr	chromium
C	carbon
P	phosphorus

8.

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

2.2: Atomic Theory

Exercises

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

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

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

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

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

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

Answers

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

2.

    a. yes

b. no
c. no
d. yes

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

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

2.3: The Structure of Atoms

Exercises

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

Answers

1. An electron is smaller.

2. An atom of lead is larger.

3. proton; electron

4. neutron; inside the nucleus

5. most; inside

6. least; outside

7. Electrons are in orbit about the nucleus.

8. Electrons do not have specific circular orbits about the nucleus.

9. Most of the alpha particles went through the metal sheet because atoms are mostly empty space.

10. minority; mass

2.4: Nuclei of Atoms

Exercises

1. How many protons are in the nucleus of each element?

   a. radon
   b. tungsten
   c. chromium
   d. beryllium

2. How many protons are in the nucleus of each element?

   a. sulfur
   b. uranium
   c. calcium
   d. lithium

3. What are the atomic numbers of the elements in Exercise 1?

4. What are the atomic numbers of the elements in Exercise 2?

5. How many electrons are in neutral atoms of the elements in Exercise 1?

6. How many electrons are in neutral atoms of the elements in Exercise 2?

7. Complete the following table.

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

8. Complete the following table.

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

9. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.

   a. $^{131}\mathrm{I}$
   b. $^{40}\mathrm{K}$
   c. $^{201}\mathrm{Hg}$
   d. $^{19}\mathrm{F}$

10. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.

    a. $^{3}\mathrm{H}$
    b. $^{133}\mathrm{Cs}$
    c. $^{56}\mathrm{Fe}$

d. ^{207}Pb

11. What is the mass number of a gallium atom that has 38 neutrons in it?

12. What is the mass number of a uranium atom that has 143 neutrons in it?

13. Complete each sentence.

    a. ^{48}Ti has _____ neutrons.
    b. ^{40}Ar has _____ neutrons.
    c. ^{3}H has _____ neutrons.

14. Complete each sentence.

    a. ^{18}O has _____ neutrons.
    b. ^{60}Ni has _____ neutrons.
    c. ^{127}I has _____ neutrons.

Answers

1. a. 86
   b. 74
   c. 24
   d. 4

2.

   a. 16
   b. 92
   c. 20
   d. 3

3. 86, 74, 24, and 4

4. 16, 92, 20, 3

5. 86, 74, 24, and 4

6. 16, 92, 20, 3

7.

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

8.

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

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

10.

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

11. 69

12. 235

13.

   a. 26
   b. 22
   c. 2

14.

   a. 10
   b. 32
   c. 74

2.5: Atomic Masses

Exercises

1. What is the atomic mass of zinc in atomic mass units?
2. What is the atomic mass of barium in atomic mass units?
3. What is the average mass of a single magnesium atom in grams?
4. What is the average mass of a single calcium atom in grams?
5. What is the mass of 1.000×10^{24} aluminum atoms in grams?
6. What is the mass of 5.000×10^{23} carbon atoms in grams?

7. Which has more mass—1 tungsten atom or 11 oxygen atoms?

8.Which has more mass—1 tungsten atom or 11 oxygen atoms?

9. Determine the atomic mass of lithium, given the isotopic composition: 92.4% lithium-7 (mass 7.016 u) and 7.60% lithium-6 (mass 6.015 u).

10. Determine the atomic mass of neon, given the isotopic composition: 90.48% neon-20 (mass 19.992 u), 0.27% neon-21 (mass 20.994 u), and 9.25% neon-22 (mass 21.991 u).

11. Determine the atomic mass of magnesium, given the isotopic composition: 78.70% magnesium-24 (mass 23.98 u), 10.13% magnesium-25 (mass 24.99 u) and 11.17% magnesium-26 (mass 25.98).

12. Determine the atomic mass of bromine, given the isotopic composition: 50.69% bromine-79 (mass 78.9183 u), 49.31% and

bromine-81 (mass 80.9163 u).

Answers

1. 65.4 u

2. 137.33 u

3. 4.0038×10^{-23} g

4. 6.657×10^{-23} g

5. 44.81 g

6. 9.974 g

7. 1 tungsten atom

8. 1 magnesium atom

9. 6.94 u

10. 20.18 u

11. 24.30 u

12. 79.90 u

2.6: Arrangements of Electrons

Exercises

1. What is the maximum number of electrons that can fit in an s subshell? Does it matter what shell the s subshell is in?
2. What is the maximum number of electrons that can fit in a *p* subshell? Does it matter what shell the *p* subshell is in?
3. What is the maximum number of electrons that can fit in a *d* subshell? Does it matter what shell the *d* subshell is in?
4. What is the maximum number of electrons that can fit in an *f* subshell? Does it matter what shell the *f* subshell is in?
5. What is the electron configuration of a carbon atom?
6. What is the electron configuration of a sulfur atom?
7. What is the valence shell electron configuration of a calcium atom?
8. What is the valence shell electron configuration of a selenium atom?
9. What atom has the electron configuration $1s^2 2s^2 2p^5$?
10. What atom has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^3$?
11. Draw a representation of the electronic structure of an oxygen atom.
12. Draw a representation of the electronic structure of a phosphorus atom.
13. A potassium atom has ___ core electrons and ___ valence electrons.

14. A silicon atom has ____ core electrons and ____ valence electrons.

Answers

1. 2; no

2. 6; no

3. 10; no

4. 14; no

5. $1s^2 2s^2 2p^2$

6. $1s^2 2s^2 2p^6 3s^2 3p^4$

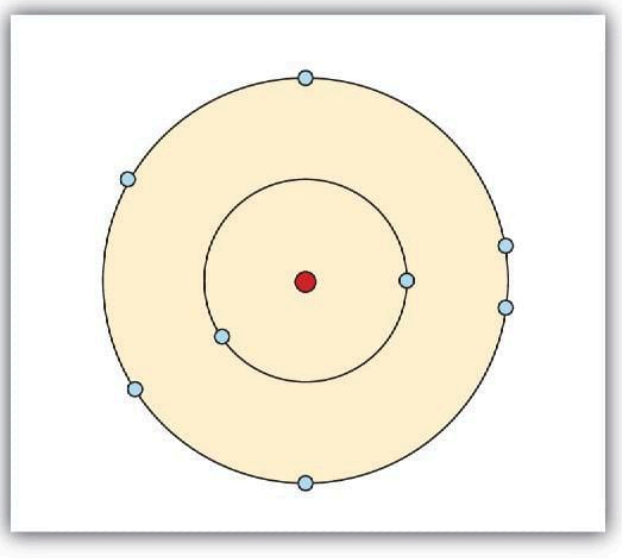
7. $4s^2$

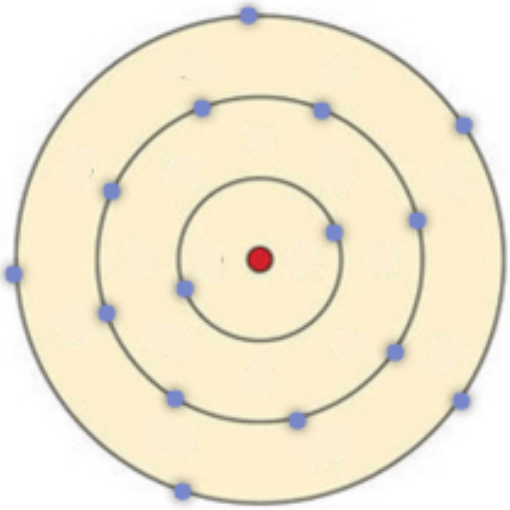
8. $4s^2 4p^4$

9. fluorine

10. phosphorus

11.





12.

13. 18; 1

14. 10; 4

2.7: The Periodic Table

Exercises

1. Which elements have chemical properties similar to those of magnesium?

    a. sodium
    b. fluorine
    c. calcium
    d. barium
    e. selenium

2. Which elements have chemical properties similar to those of lithium?

    a. sodium
    b. calcium
    c. beryllium
    d. barium
    e. potassium

3. Which elements have chemical properties similar to those of chlorine?

    a. sodium
    b. fluorine
    c. calcium
    d. iodine
    e. sulfur

4. Which elements have chemical properties similar to those of carbon?

    a. silicon
    b. oxygen
    c. germanium

d. barium
e. argon

5. Which elements are alkali metals?

a. sodium
b. magnesium
c. aluminum
d. potassium
e. calcium

6. Which elements are alkaline earth metals?

a. sodium
b. magnesium
c. aluminum
d. potassium
e. calcium

7. Which elements are halogens?

a. oxygen
b. fluorine
c. chlorine
d. sulfur
e. carbon

8. Which elements are noble gases?

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

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

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

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

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

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

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

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

a. H and He
b. N and F

c. Cl and Br
d. Al and B

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

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

Answers

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

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

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

4.
   a. yes
   b. no
   c. yes
   d. no
   e. no

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

6.
   a. no
   b. yes
   c. no
   d. no
   e. yes

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

8.

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

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

10.

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

11.

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

12.

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

13. metal; transition metals

14. semimetal; main group elements

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

Q2.1.1

Which of the following substances are elements?

1. sodium
2. milk
3. gold
4. water
5. air
6. liquefied nitrogen

Hint: what is the definition of an element?

Answer:

1. element
2. not an element
3. element
4. not an element
5. not an element

6. element

Solution:

At first we need to know the definition of an element: a chemical element is a **PURE** chemical substance consisting of a **SINGLE** type of atom. Also, we need to know the compositions of these substances. The substances that can be represented with a SINGLE chemical symbol (maybe with subscript or superscript).

1. Sodium is Na → an element
2. Milk is an organic solution containing lipids, proteins, salts minerals, vitamins, carbohydrates and other miscellaneous compounds → not an element
3. Gold is Au → an element
4. Water is H_2O → not an element
5. Air is composed of many gases such O_2, N_2 → not an element
6. Liquefied nitrogen is still N_2 → an element.

Q2.1.2

Which of the following substances are elements?

1. paper
2. electricity
3. neon
4. carbon
5. wood
6. concrete

Hint: what is the definition of an element?

Answer:

1. not an element
2. not an element
3. element
4. element
5. not an element
6. not an element element

Solution:

At first we need to know the definition of an element: a chemical element is a **PURE** chemical substance consisting of a **SINGLE** type of atom. Also, we need to know the compositions of these substances. The substances that can be represented with a SINGLE chemical symbol (can have subscript or superscript).

1. Paper is mostly cellulose composed of carbon, hydrogen and oxygen $[(C_6H_{10}O_5)_n]$ → not an element
2. Electricity is a form of energy not atoms → not an element
3. Neon is a noble gas with chemical symbol Ne → an element
4. Carbon is C → an element
5. Wood is majorly composed with cellulose and lignin → not an element
6. Concrete is composed of water and coarse granular material → not an element.

Q2.1.3

Write the chemical symbol for each element.

1. silver
2. sulfur
3. nitrogen

4. neon

Hint: what is the first or first two letters of the name of the element? Sometimes it's the Latin name of the element. However, sometimes it's not the first 2 letters.

Answer:

1. Ag
2. S
3. N
4. Ne

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element with the **first letter ALWAYS in uppercase and the second ALWAYS letter in lowercase**. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.1.4

Write the chemical symbol for each element.

1. bromine
2. oxygen
3. lithium
4. boron

Hint: what is the first or first two letters of the name of the element? Sometimes it's the Latin name of the element. However, sometimes it's not the first 2 letters.

Answer:

1. Br
2. O
3. Li
4. B

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element with the **first letter ALWAYS in uppercase and the second letter ALWAYS in lowercase**. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.1.5

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

Hint: How would this notation confuse the readers?

Answer: If "CO" was written as the chemical symbol for cobalt, it would be the same as the chemical formula for carbon monoxide. This ambiguity would create a lot problems such as in labeling and might lead to life-threatening danger.

Solution:

The chemical symbol is **ALWAYS** written with the **first letter in uppercase and the second letter in lowercase.**

Q2.1.6

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

Hint: How would this notation confuse the readers?

Answer: If "NO" was written as the chemical symbol for cobalt, it would be the same as the chemical formula for nitrogen monoxide. This ambiguity would create a lot problems such as in labeling and might lead to life-threatening danger.

Solution:

The chemical symbol is **ALWAYS** written with the **first letter in uppercase and the second letter in lowercase.**

Q2.1.7

Complete the following table.

Element Symbol	Element Name
F	
Fe	
I	
Cr	
C	
P	

Hint: The element symbols are usually the first or first two letters of the element names.

Answer

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

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element. Also, metals tend to end in -ium; halogens gases (group 17 gases) end in -ine; noble gases except helium end in -on; gases other than halogens and noble gases tend to end in -gen. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.1.8

Complete the following table.

Element Symbol	Element Name
Mg	
Mn	
Ca	
Cl	
K	
Pt	

Hint: The element symbols are usually the first or first two letters of the element names.

Answer

Element Symbol	Element Name
Mg	Magnesium
Mn	Manganese
Ca	Calcium
Cl	Chlorine
K	Potassium
Pt	Platinum

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element. Also, metals tend to end in -ium; halogens gases (group 17 gases) end in -ine; noble gases except helium end in -on; gases other than halogens and noble gases tend to end in -gen. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.2.1

Which of the following elements exist as diatomic molecules?

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

Hint: Which atoms don't have an octet?

Answer:

a. No
b. Yes
c. Yes
d. No

Solution:

Diatomic molecules are molecules composed of only two atoms. If an atom has full valence shell, the atom is very stable and hardly react with other atoms. When an atom doesn't have a complete valence shell, it wants to pair up the unpaired electrons in valence shell to reach a stabler status. So, if an atom has unpaired valence electrons, it will form one or more covalent bonds with

another atom which also has unpaired valence electrons to form an octet. Metal atoms form metallic bonds and exist as large continuous lattice structure.

a. Helium is in the first period and has 2 valence electrons, so its valence shell is full→ monatomic
b. Hydrogen is also in the first period but has only 1 valence electrons, so it has an unpaired electron in the valence shell. 2 Hydrogen atoms form 1 covalent bond → diatomic
c. Iodine is in group VII so it has 7 valence electrons leaving one valence electron unpaired. 2 Iodine atoms form 1 covalent bond → diatomic
d. First of all, gold is metal. Second, gold is a transition metal so it has more than one unpaired valence electron → not diatomic

Q2.2.2

Which of the following elements exist as diatomic molecules?

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

Hint: Which atoms don't have an octet?

Answer:

a. Yes
b. No
c. No
d. Yes

Solution

Diatomic molecules are molecules composed of only two atoms. If an atom has full valence shell, the atom is very stable and hardly react with other atoms. When an atom doesn't have a complete valence shell, it wants to pair up the unpaired electrons in valence shell to reach a stabler status. So, if an atom has unpaired valence electrons, it will form one or more covalent bonds with another atom which also has unpaired valence electrons to form an octet. Metal atoms form metallic bonds and exist as large continuous lattice structure.

a. Chlorine is in group VII so it has 7 valence electrons leaving one valence electron unpaired. 2 Chlorine atoms form 1 covalent bond → diatomic
b. Potassium is in group I so it has 1 unpaired valence electron. However, even 2 potassium atoms pair up, they cannot form an octet → not diatomic
c. Silver is a transition metal so it has more than one unpaired valence electron → not diatomic
d. Oxygen is in group VI so it has 6 valence electrons so it has 2 unpaired valence electron. 2 Chlorine atoms form 2 covalent bonds → diatomic

Q2.1.3

Why is it proper to represent the elemental form of helium as He but improper to represent the elemental form of hydrogen as H?

Hint: How many atoms are there in these 2 gas molecules?

Answer: Because Hydrogen molecules are diatomic while Helium molecules are monatomic.

Solution: If an atom has unpaired valence electrons, it will form one or more covalent bonds with another atom which also has unpaired valence electrons to form an octet. Hydrogen is a group I element, so it has 1 valence electron. Therefore, it wants to form one covalent bond with, in this case, another hydrogen atom to complete its duet (since hydrogen only has the first shell). On the other hand, Helium already has a complete valence shell so it is stable and exist in a monatomic form.

Q2.1.4

Why is it proper to represent the elemental form of chlorine as Cl_2 but improper to represent the elemental form of calcium as Ca_2?

Hint: How many atoms are there in these 2 gas molecules?

Answer: Because Chlorine molecules are diatomic while Calcium molecules exist as large continuous lattice structure.

Solution: If an atom has unpaired valence electrons, it will form one or more covalent bonds with another atom which also has unpaired valence electrons to form an octet. Chlorine is a group VII element, so it has 7 valence electrons. Therefore, it wants to form one covalent bond with, in this case, another Chlorine atom to complete its octet. Calcium is a group II element, but it's metal. So, calcium atoms form metallic bonds and exist in a vast network structure called lattice system.

Additional Exercises

1. If the atomic radius of sodium atoms is 1.86×10^{-10} m, how many sodium atoms are needed to make a line that is 1.00 cm in length?
2. If the atomic radius of osmium atoms is 1.34×10^{-10} m, how many osmium atoms are needed to make a line that is 5.85 m in length?
3. What might be the electron configuration of K^+, an atom that has lost an electron?
4. What might be the electron configuration of Cl^-, an atom that has gained an additional electron?
5. The electron configuration of the Ti atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$. What is the valence shell electron configuration of Ti?
6. The electron configuration of the Ge atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$. What is the valence shell electron configuration of Ge?
7. What is the mass of an electron in atomic mass units?
8. In a footnote in this chapter, an alpha particle was defined as a particle with 2 protons and 2 neutrons. What is the mass, in grams, of an alpha particle? (Hint: what element does an alpha particle resemble?)
9. A sample of the nonexistent element mythium consists of 25.59% of an isotope with mass number 580, 32.74% of an isotope with mass number 582, and 41.67% of an isotope with mass number 581. What is the atomic mass of mythium?
10. Because the distribution of isotopes is different on different planets in the solar system, the average atomic mass of any element differs from planet to planet. Assume that on Mercury, a rather hot planet, there is more deuterium left in the atmosphere than on Earth, so that 92.55% of the hydrogen on Mercury is ^1H, while the remainder is ^2H. What is the atomic mass of hydrogen on Mercury?
11. The compound that sodium makes with chlorine has sodium and chlorine atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with sodium.
12. The compound that magnesium makes with oxygen has magnesium to oxygen atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with magnesium.

3.E.1: Elements, Atoms, and the Periodic Table (Additional Exercises)

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3.E.1: Elements, Atoms, and the Periodic Table (Additional Exercises)

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

2.1: The Elements

Exercises

1. Which of the following substances are elements?

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

2. Which of the following substances are elements?

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

3. Write the chemical symbol for each element.

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

4. Write the chemical symbol for each element.

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

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

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

7. Complete the following table.

Element Symbol	Element Name
F	
Fe	
I	

Element Symbol	Element Name
Cr	
C	
P	

8. Complete the following table.

Element Symbol	Element Name
Mg	
Mn	
Ca	
Cl	
K	
Pt	

Answers

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

2.

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

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

4.

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

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

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

7.

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

8.

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

2.2: Atomic Theory

Exercises

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

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

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

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

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

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

Answers

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

2.

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

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

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

2.3: The Structure of Atoms

Exercises

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

Answers

1. An electron is smaller.

2. An atom of lead is larger.

3. proton; electron

4. neutron; inside the nucleus

5. most; inside

6. least; outside

7. Electrons are in orbit about the nucleus.

8. Electrons do not have specific circular orbits about the nucleus.

9. Most of the alpha particles went through the metal sheet because atoms are mostly empty space.

10. minority; mass

2.4: Nuclei of Atoms

Exercises

1. How many protons are in the nucleus of each element?

   a. radon
   b. tungsten
   c. chromium
   d. beryllium

2. How many protons are in the nucleus of each element?

   a. sulfur
   b. uranium
   c. calcium
   d. lithium

3. What are the atomic numbers of the elements in Exercise 1?

4. What are the atomic numbers of the elements in Exercise 2?

5. How many electrons are in neutral atoms of the elements in Exercise 1?

6. How many electrons are in neutral atoms of the elements in Exercise 2?

7. Complete the following table.

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

8. Complete the following table.

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

9. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.

   a. $^{131}\mathrm{I}$
   b. $^{40}\mathrm{K}$
   c. $^{201}\mathrm{Hg}$
   d. $^{19}\mathrm{F}$

10. State the number of protons, neutrons, and electrons in neutral atoms of each isotope.

    a. $^{3}\mathrm{H}$
    b. $^{133}\mathrm{Cs}$
    c. $^{56}\mathrm{Fe}$

d. ^{207}Pb

11. What is the mass number of a gallium atom that has 38 neutrons in it?

12. What is the mass number of a uranium atom that has 143 neutrons in it?

13. Complete each sentence.

    a. ^{48}Ti has _____ neutrons.
    b. ^{40}Ar has _____ neutrons.
    c. ^{3}H has _____ neutrons.

14. Complete each sentence.

    a. ^{18}O has _____ neutrons.
    b. ^{60}Ni has _____ neutrons.
    c. ^{127}I has _____ neutrons.

Answers

1. a. 86
   b. 74
   c. 24
   d. 4

2.

   a. 16
   b. 92
   c. 20
   d. 3

3. 86, 74, 24, and 4

4. 16, 92, 20, 3

5. 86, 74, 24, and 4

6. 16, 92, 20, 3

7.

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

8.

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

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

10.

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

11. 69

12. 235

13.

   a. 26
   b. 22
   c. 2

14.

   a. 10
   b. 32
   c. 74

2.5: Atomic Masses

Exercises

1. What is the atomic mass of zinc in atomic mass units?
2. What is the atomic mass of barium in atomic mass units?
3. What is the average mass of a single magnesium atom in grams?
4. What is the average mass of a single calcium atom in grams?
5. What is the mass of 1.000×10^{24} aluminum atoms in grams?
6. What is the mass of 5.000×10^{23} carbon atoms in grams?

7. Which has more mass—1 tungsten atom or 11 oxygen atoms?

8.Which has more mass—1 tungsten atom or 11 oxygen atoms?

9. Determine the atomic mass of lithium, given the isotopic composition: 92.4% lithium-7 (mass 7.016 u) and 7.60% lithium-6 (mass 6.015 u).

10. Determine the atomic mass of neon, given the isotopic composition: 90.48% neon-20 (mass 19.992 u), 0.27% neon-21 (mass 20.994 u), and 9.25% neon-22 (mass 21.991 u).

11. Determine the atomic mass of magnesium, given the isotopic composition: 78.70% magnesium-24 (mass 23.98 u), 10.13% magnesium-25 (mass 24.99 u) and 11.17% magnesium-26 (mass 25.98).

12. Determine the atomic mass of bromine, given the isotopic composition: 50.69% bromine-79 (mass 78.9183 u), 49.31% and

bromine-81 (mass 80.9163 u).

Answers

1. 65.4 u

2. 137.33 u

3. 4.0038×10^{-23} g

4. 6.657×10^{-23} g

5. 44.81 g

6. 9.974 g

7. 1 tungsten atom

8. 1 magnesium atom

9. 6.94 u

10. 20.18 u

11. 24.30 u

12. 79.90 u

2.6: Arrangements of Electrons

Exercises

1. What is the maximum number of electrons that can fit in an s subshell? Does it matter what shell the s subshell is in?
2. What is the maximum number of electrons that can fit in a *p* subshell? Does it matter what shell the *p* subshell is in?
3. What is the maximum number of electrons that can fit in a *d* subshell? Does it matter what shell the *d* subshell is in?
4. What is the maximum number of electrons that can fit in an *f* subshell? Does it matter what shell the *f* subshell is in?
5. What is the electron configuration of a carbon atom?
6. What is the electron configuration of a sulfur atom?
7. What is the valence shell electron configuration of a calcium atom?
8. What is the valence shell electron configuration of a selenium atom?
9. What atom has the electron configuration $1s^2 2s^2 2p^5$?
10. What atom has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^3$?
11. Draw a representation of the electronic structure of an oxygen atom.
12. Draw a representation of the electronic structure of a phosphorus atom.
13. A potassium atom has ____ core electrons and ____ valence electrons.

14. A silicon atom has ____ core electrons and ____ valence electrons.

Answers

1. 2; no

2. 6; no

3. 10; no

4. 14; no

5. $1s^2 2s^2 2p^2$

6. $1s^2 2s^2 2p^6 3s^2 3p^4$

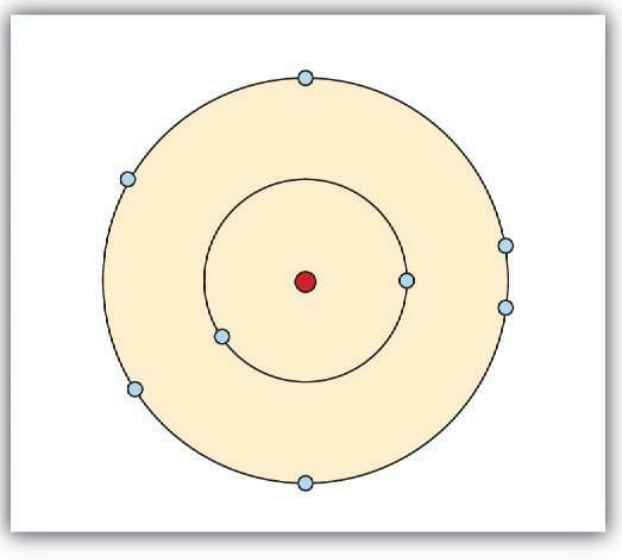
7. $4s^2$

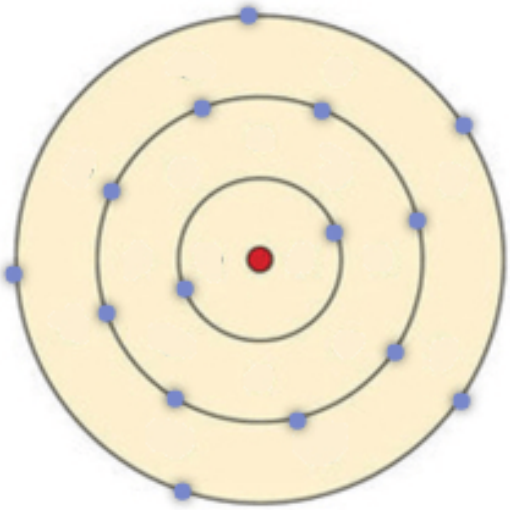
8. $4s^2 4p^4$

9. fluorine

10. phosphorus

11.





12.

13. 18; 1

14. 10; 4

2.7: The Periodic Table

Exercises

1. Which elements have chemical properties similar to those of magnesium?

    a. sodium
    b. fluorine
    c. calcium
    d. barium
    e. selenium

2. Which elements have chemical properties similar to those of lithium?

    a. sodium
    b. calcium
    c. beryllium
    d. barium
    e. potassium

3. Which elements have chemical properties similar to those of chlorine?

    a. sodium
    b. fluorine
    c. calcium
    d. iodine
    e. sulfur

4. Which elements have chemical properties similar to those of carbon?

    a. silicon
    b. oxygen
    c. germanium

d. barium
e. argon

5. Which elements are alkali metals?

a. sodium
b. magnesium
c. aluminum
d. potassium
e. calcium

6. Which elements are alkaline earth metals?

a. sodium
b. magnesium
c. aluminum
d. potassium
e. calcium

7. Which elements are halogens?

a. oxygen
b. fluorine
c. chlorine
d. sulfur
e. carbon

8. Which elements are noble gases?

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

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

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

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

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

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

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

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

a. H and He
b. N and F

c. Cl and Br
d. Al and B

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

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

Answers

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

2.

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

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

4.

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

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

6.

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

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

8.

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

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

10.

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

11.

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

12.

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

13. metal; transition metals

14. semimetal; main group elements

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

Q2.1.1

Which of the following substances are elements?

1. sodium
2. milk
3. gold
4. water
5. air
6. liquefied nitrogen

Hint: what is the definition of an element?

Answer:

1. element
2. not an element
3. element
4. not an element
5. not an element

6. element

Solution:

At first we need to know the definition of an element: a chemical element is a **PURE** chemical substance consisting of a **SINGLE** type of atom. Also, we need to know the compositions of these substances. The substances that can be represented with a SINGLE chemical symbol (maybe with subscript or superscript).

1. Sodium is Na → an element
2. Milk is an organic solution containing lipids, proteins, salts minerals, vitamins, carbohydrates and other miscellaneous compounds → not an element
3. Gold is Au → an element
4. Water is H_2O → not an element
5. Air is composed of many gases such O_2, N_2 → not an element
6. Liquefied nitrogen is still N_2 → an element.

Q2.1.2

Which of the following substances are elements?

1. paper
2. electricity
3. neon
4. carbon
5. wood
6. concrete

Hint: what is the definition of an element?

Answer:

1. not an element
2. not an element
3. element
4. element
5. not an element
6. not an element element

Solution:

At first we need to know the definition of an element: a chemical element is a **PURE** chemical substance consisting of a **SINGLE** type of atom. Also, we need to know the compositions of these substances. The substances that can be represented with a SINGLE chemical symbol (can have subscript or superscript).

1. Paper is mostly cellulose composed of carbon, hydrogen and oxygen $[(C_6H_{10}O_5)_n]$ → not an element
2. Electricity is a form of energy not atoms → not an element
3. Neon is a noble gas with chemical symbol Ne → an element
4. Carbon is C → an element
5. Wood is majorly composed with cellulose and lignin → not an element
6. Concrete is composed of water and coarse granular material → not an element.

Q2.1.3

Write the chemical symbol for each element.

1. silver
2. sulfur
3. nitrogen

4. neon

Hint: what is the first or first two letters of the name of the element? Sometimes it's the Latin name of the element. However, sometimes it's not the first 2 letters.

Answer:

1. Ag
2. S
3. N
4. Ne

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element with the **first letter ALWAYS in uppercase and the second ALWAYS letter in lowercase**. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.1.4

Write the chemical symbol for each element.

1. bromine
2. oxygen
3. lithium
4. boron

Hint: what is the first or first two letters of the name of the element? Sometimes it's the Latin name of the element. However, sometimes it's not the first 2 letters.

Answer:

1. Br
2. O
3. Li
4. B

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element with the **first letter ALWAYS in uppercase and the second letter ALWAYS in lowercase**. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.1.5

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

Hint: How would this notation confuse the readers?

Answer: If "CO" was written as the chemical symbol for cobalt, it would be the same as the chemical formula for carbon monoxide. This ambiguity would create a lot problems such as in labeling and might lead to life-threatening danger.

Solution:

The chemical symbol is **ALWAYS** written with the **first letter in uppercase and the second letter in lowercase.**

Q2.1.6

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

Hint: How would this notation confuse the readers?

Answer: If "NO" was written as the chemical symbol for cobalt, it would be the same as the chemical formula for nitrogen monoxide. This ambiguity would create a lot problems such as in labeling and might lead to life-threatening danger.

Solution:

The chemical symbol is **ALWAYS** written with the **first letter in uppercase and the second letter in lowercase.**

Q2.1.7

Complete the following table.

Element Symbol	Element Name
F	
Fe	
I	
Cr	
C	
P	

Hint: The element symbols are usually the first or first two letters of the element names.

Answer

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

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element. Also, metals tend to end in -ium; halogens gases (group 17 gases) end in -ine; noble gases except helium end in -on; gases other than halogens and noble gases tend to end in -gen. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.1.8

Complete the following table.

Element Symbol	Element Name
Mg	
Mn	
Ca	
Cl	
K	
Pt	

Hint: The element symbols are usually the first or first two letters of the element names.

Answer

Element Symbol	Element Name
Mg	Magnesium
Mn	Manganese
Ca	Calcium
Cl	Chlorine
K	Potassium
Pt	Platinum

Solution:

Usually, the chemical symbol of an element is the first or first 2 letters of the English name or Latin name of the element. Also, metals tend to end in -ium; halogens gases (group 17 gases) end in -ine; noble gases except helium end in -on; gases other than halogens and noble gases tend to end in -gen. However, that's not always the case, such as Tin (Stannum in Latin) is Sn. This is the time a Periodic Table comes in handy.

Q2.2.1

Which of the following elements exist as diatomic molecules?

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

Hint: Which atoms don't have an octet?

Answer:

a. No
b. Yes
c. Yes
d. No

Solution:

Diatomic molecules are molecules composed of only two atoms. If an atom has full valence shell, the atom is very stable and hardly react with other atoms. When an atom doesn't have a complete valence shell, it wants to pair up the unpaired electrons in valence shell to reach a stabler status. So, if an atom has unpaired valence electrons, it will form one or more covalent bonds with

another atom which also has unpaired valence electrons to form an octet. Metal atoms form metallic bonds and exist as large continuous lattice structure.

a. Helium is in the first period and has 2 valence electrons, so its valence shell is full→ monatomic
b. Hydrogen is also in the first period but has only 1 valence electrons, so it has an unpaired electron in the valence shell. 2 Hydrogen atoms form 1 covalent bond → diatomic
c. Iodine is in group VII so it has 7 valence electrons leaving one valence electron unpaired. 2 Iodine atoms form 1 covalent bond → diatomic
d. First of all, gold is metal. Second, gold is a transition metal so it has more than one unpaired valence electron → not diatomic

Q2.2.2

Which of the following elements exist as diatomic molecules?

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

Hint: Which atoms don't have an octet?

Answer:

a. Yes
b. No
c. No
d. Yes

Solution

Diatomic molecules are molecules composed of only two atoms. If an atom has full valence shell, the atom is very stable and hardly react with other atoms. When an atom doesn't have a complete valence shell, it wants to pair up the unpaired electrons in valence shell to reach a stabler status. So, if an atom has unpaired valence electrons, it will form one or more covalent bonds with another atom which also has unpaired valence electrons to form an octet. Metal atoms form metallic bonds and exist as large continuous lattice structure.

a. Chlorine is in group VII so it has 7 valence electrons leaving one valence electron unpaired. 2 Chlorine atoms form 1 covalent bond → diatomic
b. Potassium is in group I so it has 1 unpaired valence electron. However, even 2 potassium atoms pair up, they cannot form an octet → not diatomic
c. Silver is a transition metal so it has more than one unpaired valence electron → not diatomic
d. Oxygen is in group VI so it has 6 valence electrons so it has 2 unpaired valence electron. 2 Chlorine atoms form 2 covalent bonds → diatomic

Q2.1.3

Why is it proper to represent the elemental form of helium as He but improper to represent the elemental form of hydrogen as H?

Hint: How many atoms are there in these 2 gas molecules?

Answer: Because Hydrogen molecules are diatomic while Helium molecules are monatomic.

Solution: If an atom has unpaired valence electrons, it will form one or more covalent bonds with another atom which also has unpaired valence electrons to form an octet. Hydrogen is a group I element, so it has 1 valence electron. Therefore, it wants to form one covalent bond with, in this case, another hydrogen atom to complete its duet (since hydrogen only has the first shell). On the other hand, Helium already has a complete valence shell so it is stable and exist in a monatomic form.

Q2.1.4

Why is it proper to represent the elemental form of chlorine as Cl_2 but improper to represent the elemental form of calcium as Ca_2?

Hint: How many atoms are there in these 2 gas molecules?

Answer: Because Chlorine molecules are diatomic while Calcium molecules exist as large continuous lattice structure.

Solution: If an atom has unpaired valence electrons, it will form one or more covalent bonds with another atom which also has unpaired valence electrons to form an octet. Chlorine is a group VII element, so it has 7 valence electrons. Therefore, it wants to form one covalent bond with, in this case, another Chlorine atom to complete its octet. Calcium is a group II element, but it's metal. So, calcium atoms form metallic bonds and exist in a vast network structure called lattice system.

Additional Exercises

1. If the atomic radius of sodium atoms is 1.86×10^{-10} m, how many sodium atoms are needed to make a line that is 1.00 cm in length?
2. If the atomic radius of osmium atoms is 1.34×10^{-10} m, how many osmium atoms are needed to make a line that is 5.85 m in length?
3. What might be the electron configuration of K^+, an atom that has lost an electron?
4. What might be the electron configuration of Cl^-, an atom that has gained an additional electron?
5. The electron configuration of the Ti atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$. What is the valence shell electron configuration of Ti?
6. The electron configuration of the Ge atom is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$. What is the valence shell electron configuration of Ge?
7. What is the mass of an electron in atomic mass units?
8. In a footnote in this chapter, an alpha particle was defined as a particle with 2 protons and 2 neutrons. What is the mass, in grams, of an alpha particle? (Hint: what element does an alpha particle resemble?)
9. A sample of the nonexistent element mythium consists of 25.59% of an isotope with mass number 580, 32.74% of an isotope with mass number 582, and 41.67% of an isotope with mass number 581. What is the atomic mass of mythium?
10. Because the distribution of isotopes is different on different planets in the solar system, the average atomic mass of any element differs from planet to planet. Assume that on Mercury, a rather hot planet, there is more deuterium left in the atmosphere than on Earth, so that 92.55% of the hydrogen on Mercury is ^1H, while the remainder is ^2H. What is the atomic mass of hydrogen on Mercury?
11. The compound that sodium makes with chlorine has sodium and chlorine atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with sodium.
12. The compound that magnesium makes with oxygen has magnesium to oxygen atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with magnesium.

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3.S: Elements, Atoms, and the Periodic Table (Summary)

To ensure that you understand the material in this chapter, you should review the meanings of the following bold terms and ask yourself how they relate to the topics in the chapter.

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

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

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

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

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

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

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

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

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4: CHEMICAL REACTIONS AND EQUATIONS

Chemistry is largely about chemical changes. Indeed, if there were no chemical changes, chemistry as such would not exist! Chemical changes are a fundamental part of chemistry. Because chemical changes are so central, it may be no surprise that chemistry has developed some special ways of presenting them.

4.1: Prelude to Chemical Reactions

4.2: The Chemical Equation

4.3: Types of Chemical Reactions - Single and Double Replacement Reactions

4.4: Ionic Equations - A Closer Look

4.5: Composition, Decomposition, and Combustion Reactions

4.6: Neutralization Reactions

4.E: Chemical Reactions and Equations (Exercises)

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4.1: Prelude to Chemical Reactions

The space shuttle—and any other rocket-based system—uses chemical reactions to propel itself into space and maneuver itself when it gets into orbit. The rockets that lift the orbiter are of two different types. The three main engines are powered by reacting liquid hydrogen with liquid oxygen to generate water. Then there are the two solid rocket boosters, which use a solid fuel mixture that contains mainly ammonium perchlorate and powdered aluminum. The chemical reaction between these substances produces aluminum oxide, water, nitrogen gas, and hydrogen chloride. Although the solid rocket boosters each have a significantly lower mass than the liquid oxygen and liquid hydrogen tanks, they provide over 80% of the lift needed to put the shuttle into orbit—all because of chemical reactions.



Figure 4.1.1 The Space Shuttle. (Public Domain; NASA).

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4.2: The Chemical Equation

Learning Objectives

- Define *chemical equation*.
- Identify the parts of a chemical equation.

A chemical reaction expresses a chemical change. For example, one chemical property of hydrogen is that it will react with oxygen to make water. We can write that as follows:

$$\text{hydrogen reacts with oxygen to make water}$$

We can represent this chemical change more succinctly as:

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water}$$

where the + sign means that the two substances interact chemically with each other and the → symbol implies that a chemical reaction takes place. But substances can also be represented by chemical formulas. Remembering that hydrogen and oxygen both exist as diatomic molecules, we can rewrite our chemical change as:

$$\mathrm{H_2 + O_2 \longrightarrow H_2O}$$

This is an example of a chemical equation, which is a concise way of representing a **chemical reaction**. The initial substances are called **reactants** and the final substances are called **products**.

Unfortunately, it is also an *incomplete* chemical equation. The law of conservation of matter says that matter cannot be created or destroyed. In chemical equations, the number of atoms of each element in the reactants must be the same as the number of atoms of each element in the products. If we count the number of hydrogen atoms in the reactants and products, we find two hydrogen atoms. But if we count the number of oxygen atoms in the reactants and products, we find that there are two oxygen atoms in the reactants, but only one oxygen atom in the products.

What can we do? Can we change the subscripts in the formula for water so that it has two oxygen atoms in it? No—you *cannot* change the formulas of individual substances, because the chemical formula for a given substance is characteristic of that substance. What you *can* do, however, is to change the number of molecules that react or are produced. We do this one element at a time, going from one side of the reaction to the other, changing the number of molecules of a substance until all elements have the same number of atoms on each side.

To accommodate the two oxygen atoms as reactants, let us assume that we have two water molecules as products:

$$\mathrm{H_2 + O_2 \longrightarrow 2\,H_2O}$$

The 2 in front of the formula for water is called a **coefficient**. Now there are the same number of oxygen atoms in the reactants as there are in the product. But in satisfying the need for the same number of oxygen atoms on both sides of the reaction, we have also changed the number of hydrogen atoms on the product side, so the number of hydrogen atoms is no longer equal. No problem, simply go back to the reactant side of the equation, and add a coefficient in front of the H_2. The coefficient that works is 2:

$$\mathrm{2\,H_2 + O_2 \longrightarrow 2\,H_2O}$$

There are now four hydrogen atoms in the reactants and also four atoms of hydrogen in the product. There are two oxygen atoms in the reactants and two atoms of oxygen in the product. The law of conservation of matter has been satisfied. When the reactants and products of a chemical equation have the same number of atoms of all elements present, we say that an equation is **balanced**. All proper chemical equations are balanced. If a substance does not have a coefficient written in front of it, it is assumed to be 1. Also, the convention is to use all whole numbers when balancing chemical equations. This sometimes makes us do a bit more "back and forth" work when balancing a chemical equation.

Example 4.2.1

Write and balance the chemical equation for each given chemical reaction.

1. Hydrogen and chlorine react to make HCl
2. Ethane, $\mathrm{C_2H_6}$, reacts with oxygen to make carbon dioxide and water.

Solution

1. Let us start by simply writing a chemical equation in terms of the formulas of the substances, remembering that both elemental hydrogen and chlorine are diatomic:

$$\mathrm{H_2 + Cl_2 \longrightarrow HCl}$$

There are two hydrogen atoms and two chlorine atoms in the reactants and one of each atom in the product. We can fix this by including the coefficient 2 on the product side:

$$\mathrm{H_2 + Cl_2 \longrightarrow 2\,HCl}$$

Now there are two hydrogen atoms and two chlorine atoms on both sides of the chemical equation, so it is balanced.

2. Start by writing the chemical equation in terms of the substances involved:

$$\mathrm{C_2H_6 + O_2 \longrightarrow CO_2 + H_2O}$$

We have two carbon atoms on the left, so we need two carbon dioxide molecules on the product side, so that each side has two carbon atoms; that element is balanced. We have six hydrogen atoms in the reactants, so we need six hydrogen atoms in the products. We can get this by having three water molecules:

$$\mathrm{C_2H_6 + O_2 \longrightarrow 2\,CO_2 + 3\,H_2O}$$

Now we have seven oxygen atoms in the products (four from the CO2 and three from the H2O). This means we need seven oxygen atoms in the reactants. However, because oxygen is a diatomic molecule, we can only get an even number of oxygen atoms at a time. We can achieve this by multiplying the other coefficients by 2:

$$\mathrm{2\,C_2H_6 + O_2 \longrightarrow 4\,CO_2 + 6\,H_2O}$$

By multiplying everything else by 2, we do not unbalance the other elements, and we now get an even number of oxygen atoms in the product—14. We can get 14 oxygen atoms on the reactant side by having 7 oxygen molecules:

$$\mathrm{2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O}$$

As a check, recount everything to determine that each side has the same number of atoms of each element. This chemical equation is now balanced.

Exercise 4.2.1

Write and balance the chemical equation that represents nitrogen and hydrogen reacting to produce ammonia, NH3.

Answer available at the end of this section.

Many chemical equations also include phase labels for the substances: (s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for aqueous (i.e., dissolved in water). Special conditions, such as temperature, may also be listed above the arrow. For example,

$$\mathrm{2\,NaHCO_3\,(s) \overset{200^\circ C}{\rightarrow} Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(l)}$$

Key Takeaways

- A chemical equation is a concise description of a chemical reaction.
- Proper chemical equations are balanced.

Answers

Exercise 4.2.1

\[\ce{N2 + 3H2 → 2NH3 \nonumber \nonumber \]

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4.3: Types of Chemical Reactions - Single and Double Replacement Reactions

Learning Objectives

- Recognize chemical reactions as single-replacement reactions and double-replacement reactions.
- Use the periodic table, an activity series, or solubility rules to predict whether single-replacement reactions or double-replacement reactions will occur.

Up until now, we have presented chemical reactions as a topic, but we have not discussed how the products of a chemical reaction can be predicted. Here we will begin our study of certain types of chemical reactions that allow us to predict what the products of the reaction will be.

A **single-replacement reaction** is a chemical reaction in which one element is substituted for another element in a compound, generating a new element and a new compound as products. Presented below:

$$2\,\mathrm{HCl}\,(\mathrm{aq}) + \mathrm{Zn}\,(\mathrm{s}) \longrightarrow \mathrm{ZnCl_2}\,(\mathrm{aq}) + \mathrm{H_2}\,(\mathrm{g})$$

is an example of a single-replacement reaction. The hydrogen atoms in HCl are replaced by Zn atoms, and in the process a new element—hydrogen—is formed. Another example of a single-replacement reaction is

$$2\,\mathrm{NaCl}\,(\mathrm{aq}) + \mathrm{F_2}\,(\mathrm{g}) \longrightarrow 2\,\mathrm{NaF}\,(\mathrm{s}) + \mathrm{Cl_2}\,(\mathrm{g})$$

Here the negatively charged ion changes from chloride to fluoride. A typical characteristic of a single-replacement reaction is that there is one element as a reactant and another element as a product.

Not all proposed single-replacement reactions will occur between two given reactants. This is most easily demonstrated with fluorine, chlorine, bromine, and iodine. Collectively, these elements are called the *halogens* and are in the next-to-last column on the periodic table (Figure $4.3.1$). The elements on top of the column will replace the elements below them on the periodic table, but not the other way around. Thus, the reaction represented by

$$\mathrm{CaI_2}\,(\mathrm{s}) + \mathrm{Cl_2}\,(\mathrm{g}) \longrightarrow \mathrm{CaCl_2}\,(\mathrm{s}) + \mathrm{I_2}\,(\mathrm{s})$$

will occur; but the reaction

$$\mathrm{CaF_2}\,(\mathrm{s}) + \mathrm{Br_2}(\ell) \longrightarrow \mathrm{CaBr_2}\,(\mathrm{s}) + \mathrm{F_2}\,(\mathrm{g})$$

will not, because bromine is below fluorine on the periodic table. This is just one of many ways the periodic table helps us to understand chemistry.

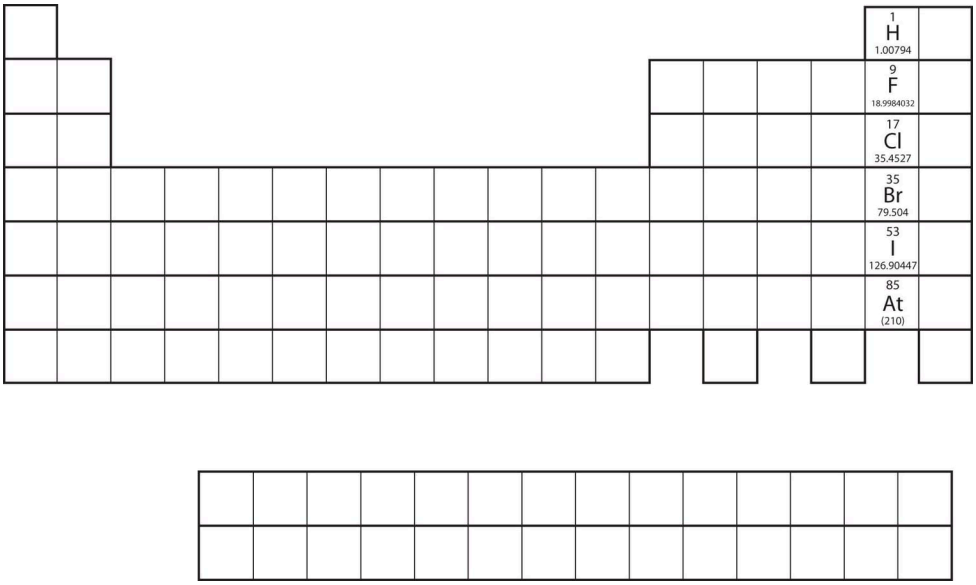


Figure 4.3.1: Halogens on the Periodic Table. The halogens are the elements in the next-to-last column on the periodic table.

Example 4.3.1

Will a single-replacement reaction occur? If so, identify the products.

a. $MgCl_2 + I_2 \rightarrow ?$
b. $CaBr_2 + F_2 \rightarrow ?$

Solution

a. Because iodine is below chlorine on the periodic table, a single-replacement reaction will not occur.
b. Because fluorine is above bromine on the periodic table, a single-replacement reaction will occur, and the products of the reaction will be CaF_2 and Br_2.

Exercise 4.3.1

Will a single-replacement reaction occur? If so, identify the products.

$$FeI_2 + Cl_2 \longrightarrow$$

Answer available at the end of this section.

Chemical reactivity trends are easy to predict when replacing anions in simple ionic compounds—simply use their relative positions on the periodic table. However, when replacing the cations, the trends are not as straightforward. This is partly because there are so many elements that can form cations; an element in one column on the periodic table may replace another element nearby, or it may not. A list called the **activity series** does the same thing the periodic table does for halogens: it lists the elements that will replace elements below them in single-replacement reactions. A simple activity series is shown below.

Activity Series for Cation Replacement in Single-Replacement Reactions

- Li
- K
- Ba
- Sr

- Ca
- Na
- Mg
- Al
- Mn
- Zn
- Cr
- Fe
- Ni
- Sn
- Pb
- H_2
- Cu
- Hg
- Ag
- Pd
- Pt
- Au

Using the activity series is similar to using the positions of the halogens on the periodic table. An element on top will replace an element below it in compounds undergoing a single-replacement reaction. Elements will not replace elements above them in compounds.

Example 4.3.2

Use the activity series to predict the products, if any, of each equation.

1. $FeCl_2 + Zn \rightarrow ?$
2. $HNO_3 + Au \rightarrow ?$

Solution

1. Because zinc is above iron in the activity series, it will replace iron in the compound. The products of this single-replacement reaction are $ZnCl_2$ and Fe.
2. Gold is below hydrogen in the activity series. As such, it will not replace hydrogen in a compound with the nitrate ion. No reaction is predicted.

Exercise 4.3.2

Use the activity series to predict the products, if any, of this equation.

$$\mathrm{AlPO_4 + Mg \longrightarrow}$$

Answer available at the end of this section.

A **double-replacement reaction** occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is

$$\mathrm{CuCl_2\,(aq) + 2\,AgNO_3\,(aq) \longrightarrow Cu(NO_3)_2\,(aq) + 2\,AgCl\,(s)}$$

There are two equivalent ways of considering a double-replacement equation: either the cations are swapped, or the anions are swapped. (You cannot swap both; you would end up with the same substances you started with.) Either perspective should allow you to predict the proper products, as long as you pair a cation with an anion, and not a cation with a cation or an anion with an anion.

Example 4.3.3

Predict the products of this double-replacement equation:

$$\mathrm{BaCl_2 + Na_2SO_4 \longrightarrow}$$

Solution

Thinking about the reaction as either switching the cations or switching the anions, we would expect the products to be BaSO4 and NaCl.

Exercise 4.3.3

Predict the products of this double-replacement equation:

$$\mathrm{KBr + AgNO_3 \longrightarrow}$$

Answer available at the end of this section.

Predicting whether a double-replacement reaction occurs is somewhat more difficult than predicting a single-replacement reaction. However, there is one type of double-replacement reaction that we can predict: the precipitation reaction. A **precipitation reaction** occurs when two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve; this new compound falls out of solution as a solid **precipitate.** The formation of a solid precipitate is the driving force that makes the reaction proceed.

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble, or insoluble). Table 4.3.1 lists some general solubility rules. We need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules. If a compound is soluble, we use the (aq) label with it, indicating that it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of solution. If everything is soluble, then no reaction will be expected.

Table 4.3.1: Some Useful Solubility Rules (soluble)

These compounds generally dissolve in water (are soluble):	Exceptions:
All compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and $NH4^+$	None
All compounds of $NO3^-$ and $C2H3O2^-$	None
Compounds of Cl^-, Br^-, I^-	Ag^+, $Hg2^{2+}$, Pb^{2+}
Compounds of $SO4^2$	$Hg2^{2+}$, Pb^{2+}, Sr^{2+}, Ba^{2+}

Table 4.3.2: Some Useful Solubility Rules (insoluble)

These compounds generally do not dissolve in water (are insoluble):	Exceptions:
Compounds of $CO3^{2-}$ and $PO4^{3-}$	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and $NH4^+$
Compounds of OH^-	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, $NH4^+$, Sr^{2+}, and Ba^{2+}

For example, consider the possible double-replacement reaction between Na2SO4 and SrCl2. The solubility rules say that all

ionic sodium compounds are soluble and all ionic chloride compounds are soluble, except for Ag^{+}, Hg_2^{2+}, and Pb^{2+}, which are not being considered here. Therefore, Na_2SO_4 and $SrCl_2$ are both soluble. The possible double-replacement reaction products are NaCl and $SrSO_4$. Are these soluble? NaCl is (by the same rule we just quoted), but what about $SrSO_4$? Compounds of the sulfate ion are generally soluble, but Sr^{2+} is an exception: we expect it to be insoluble—a precipitate. Therefore, we expect a reaction to occur, and the balanced chemical equation would be:

$$\mathrm{Na_2SO_4\,(aq) + SrCl_2\,(aq) \longrightarrow 2\,NaCl\,(aq) + SrSO_4\,(s)}$$

You would expect to see a visual change corresponding to $SrSO_4$ precipitating out of solution (Figure 4.3.2).



Figure 4.3.2: Double-Replacement Reactions. Some double-replacement reactions are obvious because you can see a solid precipitate coming out of solution. Source: Photo courtesy of Choij, http://commons.wikimedia.org/wiki/File:Copper_solution.jpg(opens in new window).

Example 4.3.4

Will a double-replacement reaction occur? If so, identify the products.

1. $Ca(NO_3)_2$ + KBr → ?
2. NaOH + $FeCl_2$ → ?

Solution

1. According to the solubility rules, both $Ca(NO_3)_2$ and KBr are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)—namely, $CaBr_2$ and KNO_3. However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and $FeCl_2$ are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $Fe(OH)_2$. NaCl is soluble, but, according to the solubility rules, $Fe(OH)_2$ is not. Therefore, a reaction would occur, and $Fe(OH)_2$(s) would precipitate out of solution. The balanced chemical equation is

$$\mathrm{2\,NaOH\,(aq) + FeCl_2\,(aq) \longrightarrow 2\,NaCl\,(aq) + Fe(OH)_2\,(s)}$$

Exercise 4.3.4

$$\mathrm{Sr(NO_3)_2 + KCl \longrightarrow}$$

Answer available at the end of this section.

Key Takeaways

- A single-replacement reaction replaces one element for another in a compound.
- The periodic table or an activity series can help predict whether single-replacement reactions occur.
- A double-replacement reaction exchanges the cations (or the anions) of two ionic compounds.
- A precipitation reaction is a double-replacement reaction in which one product is a solid precipitate.
- Solubility rules are used to predict whether some double-replacement reactions will occur.

Answers

Exercise 4.3.1

Yes; $FeCl_2$ and I_2

Exercise 4.3.2

$Mg_3(PO_4)_2$ and Al

Exercise 4.3.3

KNO_3 and AgBr

Exercise 4.3.4

No reaction; all possible products are soluble.

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4.4: Ionic Equations - A Closer Look

Learning Objectives

- Write ionic equations for chemical reactions between ionic compounds.
- Write net ionic equations for chemical reactions between ionic compounds.

For single-replacement and double-replacement reactions, many of the reactions included ionic compounds—compounds between metals and nonmetals, or compounds that contained recognizable polyatomic ions. Now, we take a closer look at reactions that include ionic compounds.

One important aspect about ionic compounds that differs from molecular compounds has to do with dissolution in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process—for example, with NaCl:

$$\mathrm{NaCl\,(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+\,(aq) + Cl^-\,(aq)}$$

When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 4.4.1). This process is called **dissociation;** we say that the ions *dissociate*.

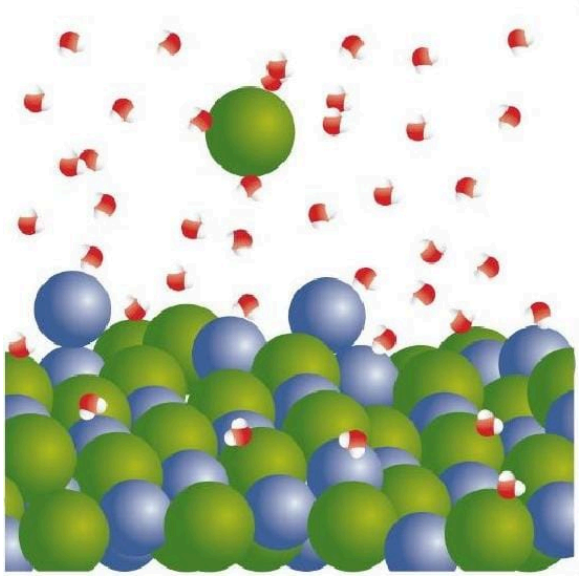


Figure 4.4.1 Ionic Solutions. When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. This behavior was first suggested by the Swedish chemist Svante August Arrhenius [1859–1927] as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry. Keep in mind that when the ions separate, *all* the ions separate. Thus, when $CaCl_2$ dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from one another:

$$CaCl_2(s) \overset{H_2O}{\rightarrow} Ca^{2+}(aq) + Cl^-(aq) + Cl^-(aq)$$

$$CaCl_2(s) \overset{H_2O}{\rightarrow} Ca^{2+}(aq) + 2Cl^-(aq)$$

That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions), and they do not stick together to make Cl_2^- or Cl_2^{2-}. They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

Example 4.4.1

Write the chemical equation that represents the dissociation of each ionic compound.

1. KBr
2. Na_2SO_4

Solution

1. $KBr(s) \rightarrow K^+(aq) + Br^-(aq)$
2. Not only do the two sodium ions go their own way, but the sulfate ion stays together as the sulfate ion. The dissolving equation is $Na_2SO_4(s) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq)$

Exercise 4.4.1

Write the chemical equation that represents the dissociation of $(NH_4)_2S$.

Answer available at the end of this section.

When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when NaCl(aq) reacts with $AgNO_3(aq)$ in a double-replacement reaction to precipitate AgCl(s) and form $NaNO_3(aq)$, the complete ionic equation includes NaCl, $AgNO_3$, and $NaNO_3$ written as separate ions:

$$\mathrm{Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)}$$

This is more representative of what is occurring in the solution.

Example 4.4.1

Write the complete ionic equation for each chemical reaction.

1. $KBr(aq) + AgC_2H_3O_2(aq) \rightarrow KC_2H_3O_2(aq) + AgBr(s)$
2. $MgSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + BaSO_4(s)$

Solution

For any ionic compound that is aqueous, we will write the compound as separated ions.

1. The complete ionic equation is $K^+(aq) + Br^-(aq) + Ag^+(aq) + C_2H_3O_2^-(aq) \rightarrow K^+(aq) + C_2H_3O_2^-(aq) + AgBr(s)$
2. The complete ionic equation is $Mg^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + BaSO_4(s)$

Exercise 4.4.1

Write the complete ionic equation for

$$CaCl_2\,(aq) + Pb(NO_3)_2\,(aq) \longrightarrow Ca(NO_3)_2\,(aq) + PbCl_2\,(s)$$

Answer available at the end of this section.

You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in

$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO3^-(aq)$

the $Ag^+(aq)$ and $Cl^-(aq)$ ions become AgCl(s), but the $Na^+(aq)$ ions and the $NO3^-(aq)$ ions stay as $Na^+(aq)$ ions and $NO3^-(aq)$ ions. These two ions are examples of spectator ions—ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:

$$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow AgCl(s) + \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)}$$

What remains when the spectator ions are removed is called the **net ionic equation**, which represents the actual chemical change occurring between the ionic compounds:

$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

It is important to reiterate that the spectator ions are still present in solution, but they do not experience any net chemical change, so they are not written in a net ionic equation.

Example 4.4.1

Write the net ionic equation for each chemical reaction.

1. $K^+(aq) + Br^-(aq) + Ag^+(aq) + C2H3O2^-(aq) \rightarrow K^+(aq) + C2H3O2^-(aq) + AgBr(s)$
2. $Mg^{2+}(aq) + SO4^{2-}(aq) + Ba^{2+}(aq) + 2NO3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO3^-(aq) + BaSO4(s)$

Solution

1. In the first equation, the $K^+(aq)$ and $C2H3O2^-(aq)$ ions are spectator ions, so they are canceled:

$$\cancel{K^+(aq)} + Br^-(aq) + Ag^+(aq) + \cancel{C_2H_3O_2^-(aq)} \rightarrow K^+(aq) + \cancel{C_2H_3O_2^-(aq)} + AgBr(s)$$

The net ionic equation is

$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s)$

2. In the second equation, the $Mg^{2+}(aq)$ and $NO3^-(aq)$ ions are spectator ions, so they are canceled:

$$\cancel{Mg^{2+}(aq)} + SO_4^{2-}(aq) + Ba^{2+}(aq) + \cancel{2NO_3^-(aq)} \rightarrow Mg^{2+}(aq) + \cancel{2NO_3^-(aq)} + BaSo_4(s)$$

The net ionic equation is

$SO4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO4(s)$

Exercise 4.4.1

Write the net ionic equation for

$CaCl_2(aq) + Pb(NO_3)_2(aq) \rightarrow Ca(NO_3)_2(aq) + PbCl_2(s)$

Answer available at the end of this section.

Chemistry is Everywhere: Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of compound that can be dissolved in a sample of water. For example, you can dissolve a maximum of 36.0 g of NaCl in 100 g of water at room temperature, but you can dissolve only 0.00019 g of AgCl in 100 g of water. We consider NaCl soluble but AgCl insoluble.

One place where solubility is important is in the tank-type water heater found in many homes in the United States. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate ($CaCO_3$). However, $CaCO_3$ has the relatively unusual property of being less soluble in hot water than in cold water. So as the water heater operates by heating water, $CaCO_3$ can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.



Figure 4.4.2 Tank water heater © Thinkstock. Most homes in the United States have a tank-type water heater like this one.

Another place where solubility versus insolubility is an issue is the Grand Canyon. We usually think of rock as insoluble. But it is actually ever so slightly soluble. This means that over a period of about two billion years, the Colorado River carved rock from the surface by slowly dissolving it, eventually generating a spectacular series of gorges and canyons. And all because of solubility!



Figure 4.4.3: The Grand Canyon was formed by water running through rock for billions of years, very slowly dissolving it. Note the Colorado River is still present in the lower part of the photo. (Sonaal Bangera via unsplash)

Key Takeaways

- Ionic compounds that dissolve separate into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

Answers

Exercise 4.4.1

$(NH_4)_2S(s) \rightarrow 2NH_4^+(aq) + S^{2-}(aq)$

Exercise 4.4.1

$Ca^{2+}(aq) + 2Cl^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq) + PbCl_2(s)$

Exercise 4.4.1

$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

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4.5: Composition, Decomposition, and Combustion Reactions

Learning Objectives

- Recognize composition, decomposition, and combustion reactions.
- Predict the products of a combustion reaction.

Three classifications of chemical reactions will be reviewed in this section. Predicting the products in some of them may be difficult, but the reactions are still easy to recognize.

A **composition reaction** (sometimes also called a *combination reaction* or a *synthesis reaction*) produces a single substance from multiple reactants. A single substance as a product is the key characteristic of the composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction. In the reaction

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

water is produced from hydrogen and oxygen. Although there are two molecules of water being produced, there is only one substance—water—as a product. So this is a composition reaction.

A **decomposition reaction** starts from a single substance and produces more than one substance; that is, it decomposes. The key characteristics of a decomposition reaction are: one substance as a reactant and more than one substance as the products. For example, in the decomposition of sodium hydrogen carbonate (also known as sodium bicarbonate):

$$2\,NaHCO_3\,(s) \longrightarrow Na_2CO_3\,(s) + CO_2\,(g) + H_2O(\ell)$$

sodium carbonate, carbon dioxide, and water are produced from the single substance sodium hydrogen carbonate.

Composition and decomposition reactions are difficult to predict; however, they should be easy to recognize.

Example 4.5.1: Identifying Reactions

Identify each equation as a composition reaction, a decomposition reaction, or neither.

a. $Fe_2O_3 + 3\,SO_3 \longrightarrow Fe_2(SO_4)_3$
b. $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$
c. $(NH_4)_2Cr_2O_7 \longrightarrow Cr_2O_3 + 4\,H_2O + N_2$

Solution

a. In this equation, two substances combine to make a single substance. This is a composition reaction.
b. Two different substances react to make two new substances. This does not fit the definition of either a composition reaction or a decomposition reaction, so it is neither. In fact, you may recognize this as a double-replacement reaction.
c. A single substance reacts to make multiple substances. This is a decomposition reaction.

Exercise 4.5.1

Identify the equation as a composition reaction, a decomposition reaction, or neither.

$$C_3H_8 \longrightarrow C_3H_4 + 2\,H_2$$

Answer available at the end of this section.

A **combustion reaction** occurs when a reactant combines with oxygen, many times from the atmosphere, to produce oxides of all other elements as products; any nitrogen in the reactant is converted to elemental nitrogen, N2. Many reactants, called *fuels*, contain mostly carbon and hydrogen atoms, reacting with oxygen to produce CO2 and H2O. For example, the balanced chemical equation for the combustion of methane, CH4, is as follows:

$$\mathrm{CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O}$$

Kerosene can be approximated with the formula $\mathrm{C_{12}H_{26}}$, and its combustion equation is:

$$\mathrm{2\,C_{12}H_{26} + 37\,O_2 \longrightarrow 24\,CO\text{-}2 + 26\,H_2O}$$

Sometimes fuels contain oxygen atoms, which must be counted when balancing the chemical equation. One common fuel is ethanol, $\mathrm{C_2H_5OH}$, whose combustion equation is:

$$\mathrm{C_2H_5OH + 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O}$$

If nitrogen is present in the original fuel, it is converted to $\mathrm{N_2}$, not to a nitrogen-oxygen compound. Thus, for the combustion of the fuel dinitroethylene, whose formula is $\mathrm{C_2H_2N_2O_4}$, we have:

$$\mathrm{2\,C_2H_2N_2O_4 + O_2 \longrightarrow 4\,CO_2 + 2\,H_2O + 2\,N_2}$$

Example 4.5.2: Combustion Reactions

Complete and balance each combustion equation.

1. the combustion of propane ($\mathrm{C_3H_8}$)
2. the combustion of ammonia ($\mathrm{NH_3}$)

Solution

1. The products of the reaction are CO2 and H2O, so our unbalanced equation is

   $$\mathrm{C_3H_8 + O_2 \longrightarrow CO_2 + H_2O}$$

   Balancing (and you may have to go back and forth a few times to balance this), we get

   $$\mathrm{C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O}$$

2. The nitrogen atoms in ammonia will react to make N2, while the hydrogen atoms will react with O2 to make H2O:

   $$\mathrm{NH_3 + O_2 \longrightarrow N_2 + H_2O}$$

   To balance this equation without fractions (which is the convention), we get

   $$\mathrm{4\,NH_3 + 3\,O_2 \longrightarrow 2\,N_2 + 6\,H_2O}$$

Exercise 4.5.2

Complete and balance the combustion equation for cyclopropanol (\ce{C3H6O}\)).

Answer available at the end of this section.

Key Takeaways

- A composition reaction produces a single substance from multiple reactants.
- A decomposition reaction produces multiple products from a single reactant.
- Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products (although nitrogen atoms react to make N2).

Answers

Exercise 4.5.1

decomposition

Exercise 4.5.2

$$C_3H_6O + 4\,O_2 \longrightarrow 3\,CO_2 + 3\,H_2O$$



Figure 4.5.1 A 20 lb (9.1 kg) steel propane cylinder.

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4.6: Neutralization Reactions

Learning Objectives

- Identify an acid and a base.
- Identify a neutralization reaction and predict its products.

In Chapter 3, we defined an acid as an ionic compound that contains H^+ as the cation. This is slightly incorrect, but until additional concepts were developed, a better definition needed to wait. Now we can redefine an acid: an **acid** is any compound that increases the amount of hydrogen ion (H^+) in an aqueous solution. The chemical opposite of an acid is a base. The equivalent definition of a base is that a **base** is a compound that increases the amount of hydroxide ion (OH^-) in an aqueous solution. These original definitions were proposed by Arrhenius (the same person who proposed ion dissociation) in 1884, so they are referred to as the **Arrhenius definition** of an acid and a base, respectively.

You may recognize that, based on the description of a hydrogen atom, an H^+ ion is a hydrogen atom that has lost its lone electron; that is, H^+ is simply a proton. Do we really have bare protons moving about in aqueous solution? No. What is more likely is that the H^+ ion has attached itself to one (or more) water molecule(s). To represent this chemically, we define the **hydronium ion** as H_3O^+, which represents an additional proton attached to a water molecule. We use the hydronium ion as the more logical way that a hydrogen ion appears in an aqueous solution, although in many chemical reactions H^+ and H_3O^+ are treated equivalently.

The reaction of an acid and a base is called a **neutralization reaction**. Although acids and bases have their own unique chemistries, the acid and base cancel each other's chemistry to produce a rather innocuous substance—water. In fact, the general reaction between an acid and a base is:

$$\text{acid} + \text{base} \longrightarrow \text{water} + \text{salt}$$

where the term **salt** is generally used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base. (In chemistry, the word *salt* refers to more than just table salt.) For example, the balanced chemical equation for the reaction between HCl(aq) and KOH(aq) is

$$\mathrm{HCl\,(aq)} + \mathrm{KOH\,(aq)} \longrightarrow \mathrm{H_2O(\ell)} + \mathrm{KCl\,(aq)}$$

where the salt is KCl. By counting the number of atoms of each element, we find that only one water molecule is formed as a product. However, in the reaction between HCl(aq) and $Mg(OH)_2$(aq), additional molecules of HCl and H_2O are required to balance the chemical equation:

$$\mathrm{2\,HCl\,(aq)} + \mathrm{Mg(OH)_2\,(aq)} \longrightarrow \mathrm{2\,H_2O(\ell)} + \mathrm{MgCl_2\,(aq)}$$

Here, the salt is $MgCl_2$. (This is one of several reactions that take place when a type of antacid—a base—is used to treat stomach acid.)

Example 4.6.1

Write the neutralization reactions between each acid and base.

a. HNO_3(aq) and $Ba(OH)_2$(aq)
b. H_3PO_4(aq) and $Ca(OH)_2$(aq)

Solution

First, we will write the chemical equation with the formulas of the reactants and the expected products; then we will balance the equation.

a. The expected products are water and barium nitrate, so the initial chemical reaction is

$$\mathrm{HNO_3\,(aq) + Ba(OH)_2\,(aq) \longrightarrow H_2O(\ell) + Ba(NO_3)_2\,(aq)}$$

To balance the equation, we need to realize that there will be two H2O molecules, so two HNO3 molecules are required:

$$\mathrm{2\,HNO_3\,(aq) + Ba(OH)_2\,(aq) \longrightarrow 2\,H_2O(\ell) + Ba(NO_3)_2\,(aq)}$$

This chemical equation is now balanced.

b. The expected products are water and calcium phosphate, so the initial chemical equation is

$$\mathrm{H_3PO_4\,(aq) + Ca(OH)_2\,(aq) \longrightarrow H_2O(\ell) + Ca_3(PO_4)_2\,(s)}$$

According to the solubility rules, Ca3(PO4)2 is insoluble, so it has an (s) phase label. To balance this equation, we need two phosphate ions and three calcium ions. We end up with six water molecules to balance the equation:

$$\mathrm{2\,H_3PO_4\,(aq) + 3\,Ca(OH)_2\,(aq) \longrightarrow 6\,H_2O(\ell) + Ca_3(PO_4)_2\,(s)}$$

This chemical equation is now balanced.

Exercise 4.6.1

Write the neutralization reaction between H2SO4(aq) and Sr(OH)2(aq).

Answer

$$\mathrm{H_2SO_4\,(aq) + Sr(OH)_2\,(aq) \longrightarrow 2\,H_2O(\ell) + SrSO_4\,(aq)}$$

Neutralization reactions are one type of chemical reaction that proceeds even if one reactant is not in the aqueous phase. For example, the chemical reaction between HCl(aq) and Fe(OH)3(s) still proceeds according to the equation:

$$\mathrm{3\,HCl\,(aq) + Fe(OH)_3\,(s) \longrightarrow 3\,H_2O(\ell) + FeCl_3\,(aq)}$$

even though Fe(OH)3 is not soluble. When one realizes that Fe(OH)3(s) is a component of rust, this explains why some cleaning solutions for rust stains contain acids—the neutralization reaction produces products that are soluble and wash away. (Washing with acids like HCl is one way to remove rust and rust stains, but HCl must be used with caution!)

Complete and net ionic reactions for neutralization reactions will depend on whether the reactants and products are soluble, even if the acid and base react. For example, in the reaction of HCl(aq) and NaOH(aq),

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \longrightarrow H_2O(\ell) + NaCl\,(aq)}$$

the complete ionic reaction is

$$\mathrm{H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O(\ell) + Na^+\,(aq) + Cl^-\,(aq)}$$

The Na^+(aq) and Cl^-(aq) ions are spectator ions, so we can remove them to have

$$\mathrm{H^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O(\ell)}$$

as the net ionic equation. If we wanted to write this in terms of the hydronium ion, $H3O^+$(aq), we would write it as

$$\mathrm{H_3O^+\,(aq) + OH^-\,(aq) \longrightarrow 2\,H_2O(\ell)}$$

With the exception of the introduction of an extra water molecule, these two net ionic equations are equivalent.

However, for the reaction between HCl(aq) and Cr(OH)2(s), because chromium(II) hydroxide is insoluble, we cannot separate it into ions for the complete ionic equation:

$$2\,\mathrm{H^+(aq)} + 2\,\mathrm{Cl^-(aq)} + \mathrm{Cr(OH)_2(s)} \longrightarrow 2\,\mathrm{H_2O}(\ell) + \mathrm{Cr^{2+}(aq)} + 2\,\mathrm{Cl^-(aq)}$$

The chloride ions are the only spectator ions here, so the net ionic equation is

$$2\,\mathrm{H^+(aq)} + \mathrm{Cr(OH)_2(s)} \longrightarrow 2\,\mathrm{H_2O}(\ell) + \mathrm{Cr^{2+}(aq)}$$

Example 4.6.2

Oxalic acid, $H_2C_2O_4(s)$, and $Ca(OH)_2(s)$ react very slowly. What is the net ionic equation between these two substances if the salt formed is insoluble? (The anion in oxalic acid is the oxalate ion, $C_2O_4^{2-}$.)

Solution

The products of the neutralization reaction will be water and calcium oxalate:

$H_2C_2O_4(s) + Ca(OH)_2(s) \rightarrow 2H_2O(\ell) + CaC_2O_4(s)$

Because nothing is dissolved, there are no substances to separate into ions, so the net ionic equation is the equation of the three solids and one liquid.

Exercise 4.6.2

What is the net ionic equation between $HNO_3(aq)$ and $Ti(OH)_4(s)$?

Answer

$4H^+(aq) + Ti(OH)_4(s) \rightarrow 4H_2O(\ell) + Ti^{4+}(aq)$

Key Takeaways

- The Arrhenius definition of an acid is a substance that increases the amount of H^+ in an aqueous solution.
- The Arrhenius definition of a base is a substance that increases the amount of OH^- in an aqueous solution.
- Neutralization is the reaction of an acid and a base, which forms water and a salt.
- Net ionic equations for neutralization reactions may include solid acids, solid bases, solid salts, and water.

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4.E: Chemical Reactions and Equations (Exercises)

Exercises (The Chemical Equation)

1. From the statement “nitrogen and hydrogen react to produce ammonia,” identify the reactants and the products.
2. From the statement “sodium metal reacts with water to produce sodium hydroxide and hydrogen,” identify the reactants and the products.
3. From the statement “magnesium hydroxide reacts with nitric acid to produce magnesium nitrate and water,” identify the reactants and the products.
4. From the statement “propane reacts with oxygen to produce carbon dioxide and water,” identify the reactants and the products.
5. Write and balance the chemical equation described by Exercise 1.
6. Write and balance the chemical equation described by Exercise 2.
7. Write and balance the chemical equation described by Exercise 3.
8. Write and balance the chemical equation described by Exercise 4. The formula for propane is C_3H_8.
9. Balance: $\_\_NaClO_3 \rightarrow \_\_NaCl + \_\_O_2$
10. Balance: $\_\_N_2 + \_\_H_2 \rightarrow \_\_N_2H_4$
11. Balance: $\_\_Al + \_\_O_2 \rightarrow \_\_Al_2O_3$
12. Balance: $\_\_C_2H_4 + \_\_O_2 \rightarrow \_\_CO_2 + \_\_H_2O$
13. How would you write the balanced chemical equation in Exercise 10 if all substances were gases?
14. How would you write the balanced chemical equation in Exercise 12 if all the substances except water were gases and water itself were a liquid?

Answers

1. reactants: nitrogen and hydrogen; product: ammonia

3. reactants: magnesium hydroxide and nitric acid; products: magnesium nitrate and water

5. $N_2 + 3H_2 \rightarrow 2NH_3$

7. $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$

9. $2NaClO_3 \rightarrow 2NaCl + 3O_2$

11. $4Al + 3O_2 \rightarrow 2Al_2O_3$

13. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Exercises (Types of Chemical Reactions - Single and Double Displacement Reactions)

1. What are the general characteristics that help you recognize single-replacement reactions?

2. What are the general characteristics that help you recognize double-replacement reactions?

3. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
   a. $Zn + Fe(NO_3)_2 \rightarrow ?$
   b. $F_2 + FeI_3 \rightarrow ?$

4. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
   a. $Li + MgSO_4 \rightarrow ?$
   b. $NaBr + Cl_2 \rightarrow ?$

5. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
   a. $Sn + H_2SO_4 \rightarrow ?$
   b. $Al + NiBr_2 \rightarrow ?$

6. Assuming that each single-replacement reaction occurs, predict the products and write each balanced chemical equation.
   a. $Mg + HCl \rightarrow ?$
   b. $HI + Br_2 \rightarrow ?$

7. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
   a. $FeCl_2 + Br_2 \rightarrow ?$
   b. $Fe(NO_3)_3 + Al \rightarrow ?$

8. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
   a. $Zn + Fe_3(PO_4)_2 \rightarrow ?$
   b. $Ag + HNO_3 \rightarrow ?$

9. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
   a. $NaI + Cl_2 \rightarrow ?$
   b. $AgCl + Au \rightarrow ?$

10. Use the periodic table or the activity series to predict if each single-replacement reaction will occur and, if so, write a balanced chemical equation.
    a. $Pt + H_3PO_4 \rightarrow ?$
    b. $Li + H_2O \rightarrow ?$ (Hint: treat H_2O as if it were composed of H^+ and OH^- ions.)

11. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
    a. $Zn(NO_3)_2 + NaOH \rightarrow ?$
    b. $HCl + Na_2S \rightarrow ?$

12. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
    a. $Ca(C_2H_3O_2)_2 + HNO_3 \rightarrow ?$
    b. $Na_2CO_3 + Sr(NO_2)_2 \rightarrow ?$

13. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
    a. $Pb(NO_3)_2 + KBr \rightarrow ?$
    b. $K_2O + MgCO_3 \rightarrow ?$

14. Assuming that each double-replacement reaction occurs, predict the products and write each balanced chemical equation.
    a. $Sn(OH)_2 + FeBr_3 \rightarrow ?$
    b. $CsNO_3 + KCl \rightarrow ?$

15. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
    a. $Pb(NO_3)_2 + KBr \rightarrow ?$
    b. $K_2O + Na_2CO_3 \rightarrow ?$

16. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
    a. $Na_2CO_3 + Sr(NO_2)_2 \rightarrow ?$
    b. $(NH_4)_2SO_4 + Ba(NO_3)_2 \rightarrow ?$

17. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
    a. $K_3PO_4 + SrCl_2 \rightarrow ?$
    b. $NaOH + MgCl_2 \rightarrow ?$

18. Use the solubility rules to predict if each double-replacement reaction will occur and, if so, write a balanced chemical equation.
    a. $KC_2H_3O_2 + Li_2CO_3 \rightarrow ?$
    b. $KOH + AgNO_3 \rightarrow ?$

Answers

1. One element replaces another element in a compound.

3.

    a. $Zn + Fe(NO_3)_2 \rightarrow Zn(NO_3)_2 + Fe$
    b. $3F_2 + 2FeI_3 \rightarrow 3I_2 + 2FeF_3$

5.

    a. $Sn + H_2SO_4 \rightarrow SnSO_4 + H_2$
    b. $2Al + 3NiBr_2 \rightarrow 2AlBr_3 + 3Ni$

7.

    a. No reaction occurs.
    b. $Fe(NO_3)_3 + Al \rightarrow Al(NO_3)_3 + Fe$

9.

    a. $2NaI + Cl_2 \rightarrow 2NaCl + I_2$
    b. No reaction occurs.

11.

    1. $Zn(NO_3)_2 + 2NaOH \rightarrow Zn(OH)_2 + 2NaNO_3$
    2. $2HCl + Na_2S \rightarrow 2NaCl + H_2S$

13.

a. $Pb(NO_3)_2 + 2KBr \rightarrow PbBr_2 + 2KNO_3$
b. $K_2O + MgCO_3 \rightarrow K_2CO_3 + MgO$

15.

a. $Pb(NO_3)_2 + 2KBr \rightarrow PbBr_2(s) + 2KNO_3$
b. No reaction occurs.

17.

a. $2K_3PO_4 + 3SrCl_2 \rightarrow Sr_3(PO_4)_2(s) + 6KCl$
b. $2NaOH + MgCl_2 \rightarrow 2NaCl + Mg(OH)_2(s)$

Exercises (Ionic Equations - A Closer Look)

1. Write a chemical equation that represents NaBr(s) dissociating in water.
2. Write a chemical equation that represents $SrCl_2$(s) dissociating in water.
3. Write a chemical equation that represents $(NH_4)_3PO_4$(s) dissociating in water.
4. Write a chemical equation that represents $Fe(C_2H_3O_2)_3$(s) dissociating in water.
5. Write the complete ionic equation for the reaction of $FeCl_2$(aq) and $AgNO_3$(aq). You may have to consult the solubility rules.
6. Write the complete ionic equation for the reaction of $BaCl_2$(aq) and Na_2SO_4(aq). You may have to consult the solubility rules.
7. Write the complete ionic equation for the reaction of KCl(aq) and $NaC_2H_3O_2$(aq). You may have to consult the solubility rules.
8. Write the complete ionic equation for the reaction of $Fe_2(SO_4)_3$(aq) and $Sr(NO_3)_2$(aq). You may have to consult the solubility rules.
9. Write the net ionic equation for the reaction of $FeCl_2$(aq) and $AgNO_3$(aq). You may have to consult the solubility rules.
10. Write the net ionic equation for the reaction of $BaCl_2$(aq) and Na_2SO_4(aq). You may have to consult the solubility rules.
11. Write the net ionic equation for the reaction of KCl(aq) and $NaC_2H_3O_2$(aq). You may have to consult the solubility rules.
12. Write the net ionic equation for the reaction of $Fe_2(SO_4)_3$(aq) and $Sr(NO_3)_2$(aq). You may have to consult the solubility rules.
13. Identify the spectator ions in Exercises 9 and 10.
14. Identify the spectator ions in Exercises 11 and 12.

Answers

1. $\mathrm{NaBr(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+(aq) + Br^-(aq)}$

3. $\mathrm{(NH_4)_3\,PO_4(\ s)} \xrightarrow{\mathrm{H_2O}} \mathrm{3NH_4^+(aq) + PO_4^{3-}(aq)}$

5. $\mathrm{Fe^{2+}(aq) + 2Cl^-(aq) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow Fe^{2+}(aq) + 2NO_3^-(aq) + 2AgCl(s)}$

7. $\mathrm{K^+(aq) + Cl^-(aq) + Na^+(aq) + C_2H_3O_2^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + K^+(aq) + C_2H_3O_2^-(aq)}$

9. $\mathrm{2Cl^-(aq) + 2Ag^+(aq) \rightarrow 2AgCl(s)}$

11. There is no overall reaction.

13. In Exercise 9, $\mathrm{Fe^{2+}(aq)}$ and $\mathrm{NO_3^-(aq)}$ are spectator ions; in Exercise 10, $\mathrm{Na^+(aq)}$ and $\mathrm{Cl^-(aq)}$ are spectator ions.

Exercises (Composition, Decomposition, and Combustion Reactions)

1. Which is a composition reaction and which is not?
   a. $\mathrm{NaCl + AgNO3 \rightarrow AgCl + NaNO3}$
   b. $\mathrm{CaO + CO2 \rightarrow CaCO3}$

2. Which is a composition reaction and which is not?
   a. $\mathrm{H2 + Cl2 \rightarrow 2HCl}$
   b. $\mathrm{2HBr + Cl2 \rightarrow 2HCl + Br2}$

3. Which is a composition reaction and which is not?
   a. $\mathrm{2SO2 + O2 \rightarrow 2SO3}$
   b. $\mathrm{6C + 3H2 \rightarrow C6H6}$

4. Which is a composition reaction and which is not?
   a. $\mathrm{4Na + 2C + 3O2 \rightarrow 2Na2CO3}$
   b. $\mathrm{Na2CO3 \rightarrow Na2O + CO2}$

5. Which is a decomposition reaction and which is not?
   a. $\mathrm{HCl + NaOH \rightarrow NaCl + H2O}$
   b. $\mathrm{CaCO3 \rightarrow CaO + CO2}$

6. Which is a decomposition reaction and which is not?
   a. $\mathrm{3O2 \rightarrow 2O3}$
   b. $\mathrm{2KClO3 \rightarrow 2KCl + 3O2}$

7. Which is a decomposition reaction and which is not?
   a. $\mathrm{Na2O + CO2 \rightarrow Na2CO3}$
   b. $\mathrm{H2SO3 \rightarrow H2O + SO2}$

8. Which is a decomposition reaction and which is not?
   a. $\mathrm{2C7H5N3O6 \rightarrow 3N2 + 5H2O + 7CO + 7C}$
   b. $\mathrm{C6H12O6 + 6O2 \rightarrow 6CO2 + 6H2O}$

9. Which is a combustion reaction and which is not?
   a. $\mathrm{C6H12O6 + 6O2 \rightarrow 6CO2 + 6H2O}$
   b. $\mathrm{2Fe2S3 + 9O2 \rightarrow 2Fe2O3 + 6SO2}$

10. Which is a combustion reaction and which is not?
    a. $\mathrm{CH4 + 2F2 \rightarrow CF4 + 2H2}$
    b. $\mathrm{2H2 + O2 \rightarrow 2H2O}$

11. Which is a combustion reaction and which is not?
    a. $\mathrm{P4 + 5O2 \rightarrow 2P2O5}$
    b. $\mathrm{2Al2S3 + 9O2 \rightarrow 2Al2O3 + 6SO2}$

12. Which is a combustion reaction and which is not?
    a. $\mathrm{C2H4 + O2 \rightarrow C2H4O2}$
    b. $\mathrm{C2H4 + Cl2 \rightarrow C2H4Cl2}$

13. Is it possible for a composition reaction to also be a combustion reaction? Give an example to support your case.

14. Is it possible for a decomposition reaction to also be a combustion reaction? Give an example to support your case.

15. Complete and balance each combustion equation.
    a. $C_4H_9OH + O_2 \rightarrow ?$
    b. $CH_3NO_2 + O_2 \rightarrow ?$

16. Complete and balance each combustion equation.
    a. $B_2H_6 + O_2 \rightarrow ?$ (The oxide of boron formed is B_2O_3.)
    b. $Al_2S_3 + O_2 \rightarrow ?$ (The oxide of sulfur formed is SO_2.)
    c. $Al_2S_3 + O_2 \rightarrow ?$ (The oxide of sulfur formed is SO_3.)

Answers

1.

    a. not composition
    b. composition

3.

    a. composition
    b. composition

5.

    a. not decomposition
    b. decomposition

7.

    a. not decomposition
    b. decomposition

9.

    a. combustion
    b. combustion

11.

    a. combustion
    b. combustion

13. Yes; $2H_2 + O_2 \rightarrow 2H_2O$ (answers will vary)

15.

    a. $C_4H_9OH + 6O_2 \rightarrow 4CO_2 + 5H_2O$
    b. $4CH_3NO_2 + 3O_2 \rightarrow 4CO_2 + 6H_2O + 2N_2$

Exercises (Neutralization Reactions)

1. What is the Arrhenius definition of an acid?

2. What is the Arrhenius definition of a base?

3. Predict the products of each acid-base combination listed. Assume that a neutralization reaction occurs.
   a. HCl and KOH
   b. H_2SO_4 and KOH
   c. H_3PO_4 and $Ni(OH)_2$

4. Predict the products of each acid-base combination listed. Assume that a neutralization reaction occurs.
   a. HBr and $Fe(OH)_3$
   b. HNO_2 and $Al(OH)_3$
   c. $HClO_3$ and $Mg(OH)_2$

5. Write a balanced chemical equation for each neutralization reaction in Exercise 3.

6. Write a balanced chemical equation for each neutralization reaction in Exercise 4.

7. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.
   a. $HI(aq) + KOH(aq) \rightarrow ?$
   b. $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow ?$

8. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.
   a. $HNO_3(aq) + Fe(OH)_3(s) \rightarrow ?$
   b. $H_3PO_4(aq) + CsOH(aq) \rightarrow ?$

9. Write the net ionic equation for each neutralization reaction in Exercise 7.

10. Write the net ionic equation for each neutralization reaction in Exercise 8.

11. Write the complete and net ionic equations for the neutralization reaction between $HClO_3(aq)$ and $Zn(OH)_2(s)$. Assume the salt is soluble.

12. Write the complete and net ionic equations for the neutralization reaction between $H_2C_2O_4(s)$ and $Sr(OH)_2(aq)$. Assume the salt is insoluble.

13. Explain why the net ionic equation for the neutralization reaction between HCl(aq) and KOH(aq) is the same as the net ionic equation for the neutralization reaction between $HNO_3(aq)$ and RbOH.

14. Explain why the net ionic equation for the neutralization reaction between HCl(aq) and KOH(aq) is different from the net ionic equation for the neutralization reaction between HCl(aq) and AgOH.

15. Write the complete and net ionic equations for the neutralization reaction between HCl(aq) and KOH(aq) using the hydronium ion in place of H^+. What difference does it make when using the hydronium ion?

16. Write the complete and net ionic equations for the neutralization reaction between $HClO_3(aq)$ and $Zn(OH)_2(s)$ using the hydronium ion in place of H^+. Assume the salt is soluble. What difference does it make when using the hydronium ion?

Answers

1. An Arrhenius acid increases the amount of H^+ ions in an aqueous solution.

3.

   a. KCl and H_2O
   b. K_2SO_4 and H_2O

c. $Ni_3(PO_4)_2$ and H_2O

5.

a. $HCl + KOH \rightarrow KCl + H_2O$
b. $H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$
c. $2H_3PO_4 + 3Ni(OH)_2 \rightarrow Ni_3(PO_4)_2 + 6H_2O$

7.

a. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(\ell)$
b. $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$

9.

a. $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$
b. $2H^+(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2OH^-(aq) \rightarrow BaSO_4(s) + 2H_2O(\ell)$

11. Complete ionic equation:

$$2H^+(aq) + 2ClO_3^-(aq) + Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn^{2+}(aq) + 2ClO_3^-(aq) + 2H_2O(\ell)$$

Net ionic equation:

$$2H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(\ell)$$

13. Because the salts are soluble in both cases, the net ionic reaction is just $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$.

15. Complete ionic equation:

$$H_3O^+(aq) + Cl^-(aq) + K^+(aq) + OH^-(aq) \rightarrow 2H_2O(\ell) + K^+(aq) + Cl^-(aq)$$

Net ionic equation:

$$H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(\ell)$$

17. The difference is simply the presence of an extra water molecule as a product.

Exercises (Oxidation-Reduction Reactions)

1. Is the reaction

$$2K(s) + Br_2(\ell) \rightarrow 2KBr(s)$$

an oxidation-reduction reaction? Explain your answer.

2. Is the reaction

$$NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

an oxidation-reduction reaction? Explain your answer.

3. In the reaction

$2Ca(s) + O_2(g) \rightarrow 2CaO$

indicate what has lost electrons and what has gained electrons.

4. In the reaction

$16Fe(s) + 3S_8(s) \rightarrow 8Fe_2S_3(s)$

indicate what has lost electrons and what has gained electrons.

5. In the reaction

$2Li(s) + O_2(g) \rightarrow Li_2O_2(s)$

indicate what has been oxidized and what has been reduced.

6. In the reaction

$2Ni(s) + 3I_2(s) \rightarrow 2NiI_3(s)$

indicate what has been oxidized and what has been reduced.

7. What are two different definitions of oxidation?

8. What are two different definitions of reduction?

9. Assign oxidation numbers to each atom in each substance.
   a. P_4
   b. SO_2
   c. SO_2^{2-}
   d. $Ca(NO_3)_2$

10. Assign oxidation numbers to each atom in each substance.
   a. PF_5
   b. $(NH_4)_2S$
   c. Hg
   d. Li_2O_2 (lithium peroxide)

11. Assign oxidation numbers to each atom in each substance.
   a. CO
   b. CO_2
   c. $NiCl_2$
   d. $NiCl_3$

12. Assign oxidation numbers to each atom in each substance.
   a. NaH (sodium hydride)
   b. NO_2
   c. NO_2^-
   d. $AgNO_3$

13. Assign oxidation numbers to each atom in each substance.
   a. CH_2O
   b. NH_3
   c. Rb_2SO_4
   d. $Zn(C_2H_3O_2)_2$

14. Assign oxidation numbers to each atom in each substance.
    a. C_6H_6
    b. $B(OH)_3$
    c. Li_2S
    d. Au

15. Identify what is being oxidized and reduced in this redox equation by assigning oxidation numbers to the atoms.

    $2NO + Cl_2 \rightarrow 2NOCl$

16. Identify what is being oxidized and reduced in this redox equation by assigning oxidation numbers to the atoms.

    $Fe + SO_3 \rightarrow FeSO_3$

17. Identify what is being oxidized and reduced in this redox equation by assigning oxidation numbers to the atoms.

    $2KrF_2 + 2H_2O \rightarrow 2Kr + 4HF + O_2$

18. Identify what is being oxidized and reduced in this redox equation by assigning oxidation numbers to the atoms.

    $SO_3 + SCl_2 \rightarrow SOCl_2 + SO_2$

19. Identify what is being oxidized and reduced in this redox equation by assigning oxidation numbers to the atoms.

    $2K + MgCl_2 \rightarrow 2KCl + Mg$

20. Identify what is being oxidized and reduced in this redox equation by assigning oxidation numbers to the atoms.

    $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

Answers

1. Yes; both K and Br are changing oxidation numbers.

3. Ca has lost electrons, and O has gained electrons.

5. Li has been oxidized, and O has been reduced.

7. loss of electrons; increase in oxidation number

9.

    a. P: 0
    b. S: +4; O: −2
    c. S: +2; O: −2
    d. Ca: 2+; N: +5; O: −2

11.

    a. C: +2; O: −2
    b. C: +4; O: −2
    c. Ni: +2; Cl: −1
    d. Ni: +3; Cl: −1

13.

    a. C: 0; H: +1; O: −2
    b. N: −3; H: +1

c. Rb: +1; S: +6; O: −2
d. Zn: +2; C: 0; H: +1; O: −2

15. N is being oxidized, and Cl is being reduced.

17. O is being oxidized, and Kr is being reduced.

19.K is being oxidized, and Mg is being reduced.

Additional Exercises

1. Chemical equations can also be used to represent physical processes. Write a chemical reaction for the boiling of water, including the proper phase labels.

2. Chemical equations can also be used to represent physical processes. Write a chemical reaction for the freezing of water, including the proper phase labels.

3. Explain why

$$4Na(s) + 2Cl_2(g) \rightarrow 4NaCl(s)$$

should not be considered a proper chemical equation.

4. Explain why

$$H_2(g) + 1/2O_2(g) \rightarrow H_2O(\ell)$$

should not be considered a proper chemical equation.

5. Does the chemical reaction represented by

$$3Zn(s) + 2Al(NO_3)_3(aq) \rightarrow 3Zn(NO_3)_2(aq) + 2Al(s)$$

proceed as written? Why or why not?

6. Does the chemical reaction represented by

$$2Au(s) + 2HNO_3(aq) \rightarrow 2AuNO_3(aq) + H_2(g)$$

proceed as written? Gold is a relatively useful metal for certain applications, such as jewelry and electronics. Does your answer suggest why this is so?

7. Explain what is wrong with this double-replacement reaction.

$$NaCl(aq) + KBr(aq) \rightarrow NaK(aq) + ClBr(aq)$$

8. Predict the products of and balance this double-replacement reaction.

$$Ag_2SO_4(aq) + SrCl_2(aq) \rightarrow ?$$

9. Write the complete and net ionic equations for this double-replacement reaction.

$$BaCl_2(aq) + Ag_2SO_4(aq) \rightarrow ?$$

10. Write the complete and net ionic equations for this double-replacement reaction.

$$Ag_2SO_4(aq) + SrCl_2(aq) \rightarrow ?$$

11. Identify the spectator ions in this reaction. What is the net ionic equation?

$NaCl(aq) + KBr(aq) \rightarrow NaBr(aq) + KCl(aq)$

12. Complete this reaction and identify the spectator ions. What is the net ionic equation?

$3H_2SO_4(aq) + 2Al(OH)_3(s) \rightarrow ?$

13. Can a reaction be a composition reaction and a redox reaction at the same time? Give an example to support your answer.

14. Can a reaction be a combustion reaction and a redox reaction at the same time? Give an example to support your answer.

15. Can a reaction be a decomposition reaction and a redox reaction at the same time? Give an example to support your answer.

16. Can a reaction be a combustion reaction and a double-replacement reaction at the same time? Give an example to support your answer.

17. Why is CH_4 not normally considered an acid?

18. Methyl alcohol has the formula CH_3OH. Why would methyl alcohol not normally be considered a base?

19. What are the oxidation numbers of the nitrogen atoms in these substances?

a. N_2
b. NH_3
c. NO
d. N_2O
e. NO_2
f. N_2O_4
g. N_2O_5
h. $NaNO_3$

20. What are the oxidation numbers of the sulfur atoms in these substances?

a. SF_6
b. Na_2SO_4
c. K_2SO_3
d. SO_3
e. SO_2
f. S_8
g. Na_2S

21. Disproportion is a type of redox reaction in which the same substance is both oxidized and reduced. Identify the element that is disproportionating and indicate the initial and final oxidation numbers of that element.

$2CuCl(aq) \rightarrow CuCl_2(aq) + Cu(s)$

22. Disproportion is a type of redox reaction in which the same substance is both oxidized and reduced. Identify the element that is disproportionating and indicate the initial and final oxidation numbers of that element.

$3Cl_2(g) + 6OH^-(aq) \rightarrow 5Cl^-(aq) + ClO_3^-(aq) + 3H_2O(\ell)$

Answers

1. $H_2O(\ell) \rightarrow H_2O(g)$

3. The coefficients are not in their lowest whole-number ratio.

5. No; zinc is lower in the activity series than aluminum.

7. In the products, the cation is pairing with the cation, and the anion is pairing with the anion.

9. Complete ionic equation: $Ba^{2+}(aq) + 2Cl^{-}(aq) + 2Ag^{+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2AgCl(s)$

Net ionic equation: The net ionic equation is the same as the complete ionic equation.

11. Each ion is a spectator ion; there is no overall net ionic equation.

13. Yes; $H_2 + Cl_2 \rightarrow 2HCl$ (answers will vary)

15, Yes; $2HCl \rightarrow H_2 + Cl_2$ (answers will vary)

17. It does not increase the H^+ ion concentration; it is not a compound of H^+.

19.

a. 0
b. −3
c. +2
d. +1
e. +4
f. +4
g. +5
h. +5

21. Copper is disproportionating. Initially, its oxidation number is +1; in the products, its oxidation numbers are +2 and 0, respectively.

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5: STOICHIOMETRY AND THE MOLE

We have already established that quantities are important in science, especially in chemistry. It is important to make accurate measurements of a variety of quantities when performing experiments. However, it is also important to be able to relate one measured quantity to another, unmeasured quantity. In this chapter, we will consider how we manipulate quantities to relate them to each other.

5.1: Introduction

5.2: Stoichiometry

5.3: The Mole

5.4: The Mole in Chemical Reactions

5.5: Mole-Mass and Mass-Mass Calculations

5.6: Yields

5.7: Limiting Reagents

5.E: Stoichiometry and the Mole (Exercises)

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5.1: Introduction

At Contrived State University in Anytown, Ohio, a new building was dedicated in March 2010 to house the College of Education. The 100,000-square-foot building has enough office space to accommodate 86 full-time faculty members and 167 full-time staff. In a fit of monetary excess, the university administration offered to buy new furniture (desks and chairs) and computer workstations for all faculty and staff members moving into the new building. However, to save on long-term energy and materials costs, the university offered to buy only 1 laser printer per 10 employees, with the plan to network the printers together.



Figure 5.1.1 Outfitting a New Building. Source: Photo courtesy of Benjamin Benschneider, Cleveland State University.

How many laser printers did the administration have to buy? It is rather simple to show that 26 laser printers are needed for all the employees. However, what if a chemist was calculating quantities for a chemical reaction? Interestingly enough, similar calculations can be performed for chemicals as well as laser printers. In filling a new office building with furniture and equipment, managers do calculations similar to those performed by scientists doing chemical reactions.

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5.2: Stoichiometry

Learning Objectives

- Define *stoichiometry*.
- Relate quantities in a balanced chemical reaction on a molecular basis.

Consider a classic recipe for pound cake: 1 pound of eggs, 1 pound of butter, 1 pound of flour, and 1 pound of sugar. (That's why it's called "pound cake.") If you have 4 pounds of butter, how many pounds of sugar, flour, and eggs do you need? You would need 4 pounds each of sugar, flour, and eggs.

Now suppose you have 1.00 g H_2. If the chemical reaction follows the balanced chemical equation

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

then what mass of oxygen do you need to make water?

Curiously, this chemical reaction question is very similar to the pound cake question. Both of them involve relating a quantity of one substance to a quantity of another substance or substances. The relating of one chemical substance to another using a balanced chemical reaction is called **stoichiometry**. Using stoichiometry is a fundamental skill in chemistry; it greatly broadens your ability to predict what will occur and, more importantly, how much is produced.

Let us consider a more complicated example. A recipe for pancakes calls for 2 cups (c) of pancake mix, 1 egg, and 1/2 c of milk. We can write this in the form of a chemical equation:

2 c mix + 1 egg + 1/2 c milk → 1 batch of pancakes

If you have 9 c of pancake mix, how many eggs and how much milk do you need? It might take a little bit of work, but eventually you will find you need 4½ eggs and 2¼ c milk.

How can we formalize this? We can make a conversion factor using our original recipe and use that conversion factor to convert from a quantity of one substance to a quantity of another substance. This is similar to the way we constructed a conversion factor between feet and yards in Chapter 2. Because one recipe's worth of pancakes requires 2 c of pancake mix, 1 egg, and 1/2 c of milk, we actually have the following mathematical relationships that relate these quantities:

2 c pancake mix ⇔ 1 egg ⇔ 1/2 c milk

where ⇔ is the mathematical symbol for "is equivalent to." This does not mean that 2 c of pancake mix equal 1 egg. However, *as far as this recipe is concerned*, these are the equivalent quantities needed for a single recipe of pancakes. So, any possible quantities of two or more ingredients must have the same numerical ratio as the ratios in the equivalence.

We can deal with these equivalences in the same way we deal with equalities in unit conversions: we can make conversion factors that essentially equal 1. For example, to determine how many eggs we need for 9 c of pancake mix, we construct the conversion factor:

$$\frac{1\,egg}{2\,c\,pancake\,mix}$$

This conversion factor is, in a strange way, equivalent to 1 because the recipe relates the two quantities. Starting with our initial quantity and multiplying by our conversion factor,

$$\cancel{9\,c\,pancake\,mix} \times \frac{1\,egg}{\cancel{2\,c\,pancake\,mix}} = 4.5\,eggs$$

Note how the units *cups pancake mix* canceled, leaving us with units of *eggs*. This is the formal, mathematical way of getting our amounts to mix with 9 c of pancake mix. We can use a similar conversion factor for the amount of milk:

$$9\,c\,\cancel{pancake\ mix} \times \frac{\frac{1}{2}c\,milk}{2\,c\,\cancel{pancake\ mix}} = 2.25\,c\,milk$$

Again, units cancel, and new units are introduced.

A balanced chemical equation is nothing more than *a recipe for a chemical reaction*. The difference is that a balanced chemical equation is written in terms of atoms and molecules, not cups, pounds, and eggs.

For example, consider the following chemical equation:

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

We can interpret this as, literally, "two hydrogen molecules react with one oxygen molecule to make two water molecules." That interpretation leads us directly to some equivalencies, just as our pancake recipe did:

2H2 molecules ⇔ 1O2 molecule ⇔ 2H2O molecules

These equivalences allow us to construct conversion factors:

$$\frac{2\,molecules\,H_2}{1\,molecule\,O_2}\ \frac{2\,molecules\,H_2}{2\,molecules\,H_2O}\ \frac{1\,molecule\,H_2}{2\,molecules\,H_2O}$$

and so forth. These conversions can be used to relate quantities of one substance to quantities of another. For example, suppose we need to know how many molecules of oxygen are needed to react with 16 molecules of H2. As we did with converting units, we start with our given quantity and use the appropriate conversion factor:

$$\cancel{16\,molecules\,H_2} \times \frac{1\,molecules\,O_2}{\cancel{2\,molecules\,H_2}} = 8\,molecules\,O_2$$

Note how the unit *molecules H2* cancels algebraically, just as any unit does in a conversion like this. The conversion factor came directly from the coefficients in the balanced chemical equation. This is another reason why a properly balanced chemical equation is important.

Example 5.2.1

How many molecules of SO3 are needed to react with 144 molecules of Fe2O3 given this balanced chemical equation?

$$Fe_2O_3 + 3\,SO_3 \longrightarrow Fe_2(SO_4)_3$$

Solution

We use the balanced chemical equation to construct a conversion factor between Fe2O3 and SO3. The number of molecules of Fe2O3 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO3 go on the top. Thus, the appropriate conversion factor is

$$\frac{3\,molecules\,SO_3}{1\,molecule\,Fe_2O_3}$$

Starting with our given amount and applying the conversion factor, the result is

$$144\ \cancel{molecules}\ Fe_2O_3 \times \frac{3\,molecules\,SO_3}{1\ \cancel{molecule}\ Fe_2O_3} = 432\,molecules\,SO_3$$

We need 432 molecules of SO3 to react with 144 molecules of Fe2O3.

Exercise 5.2.1

How many molecules of H_2 are needed to react with 29 molecules of N_2 to make ammonia if the balanced chemical equation is:

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3?$$

Answer available at the end of this section.

Chemical equations also allow us to make conversions regarding the number of atoms in a chemical reaction, because a chemical formula lists the number of atoms of each element in a compound. The formula H_2O indicates that there are two hydrogen atoms and one oxygen atom in each molecule, and these relationships can be used to make conversion factors:

$$\frac{2\,atoms\,H}{1\,molecule\,H_2O} \quad \frac{1\,molecule\,H_2O}{1\,atom\,O}$$

Conversion factors like this can also be used in stoichiometry calculations.

Example 5.2.2

How many molecules of NH_3 can you make if you have 228 atoms of H_2?

Solution

From the formula, we know that one molecule of NH_3 has three H atoms. Use that fact as a conversion factor:

$$228\ \cancel{atoms\,H} \times \frac{1\,molecule\,NH_3}{3\ \cancel{atoms\,H}} = 76\,NH_3$$

Exercise 5.2.2

How many molecules of $Fe_2(SO_4)_3$ can you make from 777 atoms of S?

Answer available at the end of this section.

Summary

Quantities of substances can be related to each other using balanced chemical equations.

Answers

Exercise 5.2.1

87 molecules

Exercise 5.2.2

259 molecules

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5.3: The Mole

Learning Objectives

- Describe the unit *mole*.
- Relate the mole quantity of substance to its mass.

So far, we have been talking about chemical substances in terms of individual atoms and molecules. Yet we do not typically deal with substances an atom or a molecule at a time; we work with millions, billions, and trillions of atoms and molecules at a time. We need a way to deal with macroscopic, rather than microscopic, amounts of matter. We need a unit of amount that relates quantities of substances on a scale that we can interact with.

Chemistry uses a unit called mole. A **mole** (mol) is a number of things equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \times 10^{23} \text{ things}$$

Understand that a mole means a specific number of things, just like a dozen means a certain number of things—twelve, in the case of a dozen. But a mole is a much larger number of things. These things can be atoms, or molecules, or eggs; however, in chemistry, we usually use the mole to refer to the amounts of atoms or molecules. Although the number of things in a mole is known to eight decimal places, it is usually fine to use only two or three decimal places in calculations. The numerical value of things in a mole is often called *Avogadro's number* (N_A). *Avogadro's number* is also known as the *Avogadro constant*, after Amadeo Avogadro, an Italian chemist who first proposed its importance.

Example 5.3.1

How many molecules are present in 2.76 mol of H_2O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O, we can also determine the number of atoms in the sample.

$$2.76 \ \cancel{mol\ H_2O} \times \frac{6.022 \times 10^{23} molecules\ H_2O}{\cancel{mol\ H_2O}} = 1.66 \times 10^{24} molecules\ H_2O$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24} \ \cancel{molecules\ H_2O} \times \frac{3\ atoms}{1\ molecule} = 4.99 \times 10^{24}\ atoms$$

Exercise 5.3.1

How many molecules are present in 4.61×10^{-2} mol of O_2?

Answer available at the end of this section.

How big is a mole? It is very large. Suppose you had a mole of dollar bills that need to be counted. If everyone on earth (about 6 billion people) counted one bill per second, it would take about 3.2 million years to count all the bills. A mole of sand would fill a cube about 32 km on a side. A mole of pennies stacked on top of each other would have about the same diameter as our galaxy, the Milky Way. A mole is a lot of things—but atoms and molecules are very tiny. One mole of carbon atoms would make a cube

that is 1.74 cm on a side, small enough to carry in your pocket.

Why is the mole unit so important? It represents the link between the microscopic and the macroscopic, especially in terms of mass. *A mole of a substance has the same mass in grams as one unit (atom or molecules) has in atomic mass units*. The mole unit allows us to express amounts of atoms and molecules in visible amounts that we can understand.

For example, we already know that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022×10^{23} atoms:

$$12 \text{ g C} = 6.022 \times 10^{23} \text{ atoms C}$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50×10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25} \cancel{atoms\,C} \times \frac{12.0000\,g\,C}{6.022 \times 10^{23} \cancel{atoms\,C}} = 299\,g\,C$$

But it also goes beyond carbon. Previously we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the number of grams in 1 mol of atoms of that element, while the molecular mass of a compound is the number of grams in 1 mol of molecules of that compound. Sometimes these masses are called **molar masses** to emphasize the fact that they are the mass for 1 mol of things. (The term *molar* is the adjective form of mole and has nothing to do with teeth.)

Here are some examples. The mass of a hydrogen atom is 1.0079 u; the mass of 1 mol of hydrogen atoms is 1.0079 g. Elemental hydrogen exists as a diatomic molecule, H_2. One molecule has a mass of 1.0079 + 1.0079 = 2.0158 u, while 1 mol H_2 has a mass of 2.0158 g. A molecule of H_2O has a mass of about 18.01 u; 1 mol H_2O has a mass of 18.01 g. A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g. In each of these moles of substances, there are 6.022×10^{23} units: 6.022×10^{23} atoms of H, 6.022×10^{23} molecules of H_2 and H_2O, 6.022×10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

Example 5.3.2: Sugar

What is the molar mass of sugar ($C_6H_{12}O_6$)?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula; but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table.

Solutions to Example 5.3.2

6 C = 6 × 12.011	= 72.066
12 H = 12 × 1.0079	= 12.0948
6 O = 6 × 15.999	= 95.994
TOTAL	= 180.155 g/mol

Per convention, the unit *grams per mole* is written as a fraction.

Exercise 5.3.2

What is the molar mass of $AgNO_3$?

Answer available at the end of this section.

Knowing the molar mass of a substance, we can calculate the number of moles in a certain mass of a substance and vice versa, as these examples illustrate. The molar mass is used as the conversion factor.

Example 5.3.3

What is the mass of 3.56 mol of $HgCl_2$? The molar mass of $HgCl_2$ is 271.49 g/mol.

Solution

Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$3.56 \ \cancel{mol\ HgCl_2} \times \frac{271.49\ g\ HgCl_2}{\cancel{mol\ HgCl_2}} = 967\ g\ HgCl_2$$

Exercise 5.3.3

What is the mass of 33.7 mol of H_2O?

Answer available at the end of this section.

Example 5.3.4

How many moles of H_2O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H_2O as a conversion factor from mass to moles. The molar mass of water is

$$(1.0079 + 1.0079 + 15.999) = 18.015 \text{ g/mol}.$$

However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0 \ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.015\ \cancel{g\ H_2O}} = 13.32\ mol\ H_2O$$

Exercise 5.3.4

How many moles are present in 35.6 g of H_2SO_4 (molar mass = 98.08 g/mol)?

Answer available at the end of this section.

Other conversion factors can be combined with the definition of mole—density, for example.

Example 5.3.5

The density of ethanol is 0.789 g/mL. How many moles are in 100.0 mL of ethanol? The molar mass of ethanol is 46.08 g/mol.

Solution

Here, we use density to convert from volume to mass and then use the molar mass to determine the number of moles.

$$100 \cancel{ml}\ ethanol \times \frac{0.789\,g}{\cancel{ml}} \times \frac{1\,mol}{46.08\,\cancel{g}} = 1.71\,mol\,ethanol$$

Exercise 5.3.5

If the density of benzene, C_6H_6, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer available at the end of this section.

Summary

The mole is a key unit in chemistry. The molar mass of a substance, in grams, is numerically equal to one atom's or molecule's mass in atomic mass units.

Answers

Exercise 5.3.1

2.78×10^{22} molecules

Exercise 5.3.2

169.87 g/mol

Exercise 5.3.3

607 g

Exercise 5.3.4

0.363 mol

Exercise 5.3.5

0.201 mol

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5.4: The Mole in Chemical Reactions

Learning Objectives

- Balance a chemical equation in terms of moles.
- Use the balanced equation to construct conversion factors in terms of moles.
- Calculate moles of one substance from moles of another substance using a balanced chemical equation.

Consider this balanced chemical equation:

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

We interpret this as "two molecules of hydrogen react with one molecule of oxygen to make two molecules of water." The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:

$$100\,H_2 + 50\,O_2 \longrightarrow 100\,H_2O$$

This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:

$$5{,}000\,H_2 + 2{,}500\,O_2 \longrightarrow 5{,}000\,H_2O$$

Again, this is not conventional, but it is still balanced. Suppose we use a much larger number:

$$12.044 \times 10^{23} H_2 + 6.022 \times 10^{23} O_2 \rightarrow 12.044 \times 10^{23} H_2O$$

These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro's number, while the second number is Avogadro's number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. We can just as easily read this chemical equation as "two moles of hydrogen react with one mole of oxygen to make two moles of water." All balanced chemical reactions are balanced in terms of moles.

Example 5.4.1

Interpret this balanced chemical equation in terms of moles.

$$P_4 + 5\,O_2 \longrightarrow P_4O_{10}$$

Solution

The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as "one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide."

Exercise 5.4.1

Interpret this balanced chemical equation in terms of moles.

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$$

Answer available at the end of this section.

In Section 4.1, we stated that a chemical equation is simply a recipe for a chemical reaction. As such, chemical equations also give us equivalents—equivalents between the reactants and the products. However, now we understand that *these equivalents are expressed in terms of moles*. Consider the chemical equation

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

This chemical reaction gives us the following equivalents:

$$2 \text{ mol } H_2 \Leftrightarrow 1 \text{ mol } O_2 \Leftrightarrow 2 \text{ mol } H_2O$$

Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance. If, for example, we want to know how many moles of oxygen will react with 17.6 mol of hydrogen, we construct a conversion factor between 2 mol of H_2 and 1 mol of O_2 and use it to convert from moles of one substance to moles of another:

$$17.6\ \cancel{mol\ H_2} \times \frac{1\ mol\ O_2}{2\ \cancel{mol\ H_2}} = 8.80\ mol\ O_2$$

Note how the mol H_2 unit cancels, and mol O_2 is the new unit introduced. This is an example of a **mole-mole calculation,** when you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the numbers are small, but numbers won't always be so simple!

Example 5.4.2

For the balanced chemical equation

$$2\,C_4H_{10}\,(g) + 13\,O_2 \longrightarrow 8\,CO_2\,(g) + 10\,H_2O(\ell)$$

if 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution

We are relating an amount of oxygen to an amount of carbon dioxide, so we need the equivalence between these two substances. According to the balanced chemical equation, the equivalence is

$$13 \text{ mol } O_2 \Leftrightarrow 8 \text{ mol } CO_2$$

We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

$$154\ \cancel{mol\ O_2} \times \frac{8\ mol\ CO_2}{13\ \cancel{mol\ O_2}} = 94.8\ mol\ CO_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they do not contribute to the number of significant figures in the final answer.

Exercise 5.4.2

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer available at the end of this section.

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or liters—but

moles. Any stoichiometry problem will likely need to work through the mole unit at some point, especially if you are working with a balanced chemical reaction.

Summary

Balanced chemical reactions are balanced in terms of moles. A balanced chemical reaction gives equivalents in moles that allow stoichiometry calculations to be performed.

Answers

Exercise 5.4.1

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

Exercise 5.4.2

118 mol

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5.5: Mole-Mass and Mass-Mass Calculations

Learning Objectives

- From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Suppose we know we have 123.2 g of Cl_2. How can we determine how many moles of $AlCl_3$ we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of $AlCl_3$, we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2 \ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} = 1.738\ mol\,Cl_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of $AlCl_3$. The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2\,mol\,AlCl_3}{3\,mol\,Cl_2}$$

Using this conversion factor with the molar quantity we calculated above, we get

$$1.738\ \cancel{mol\,Cl_2} \times \frac{2\,mol\,AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\,AlCl_3$$

So, we will get 1.159 mol of $AlCl_3$ if we react 123.2 g of Cl_2.

In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

$$123.2\ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} \times \frac{2\,mol\,AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\,AlCl_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

Example 5.5.1

How many moles of HCl will be produced when 249 g of AlCl3 are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

We will do this in two steps: convert the mass of AlCl3 to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of AlCl3 is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$249\ \cancel{g\ AlCl_3} \times \frac{1\,mol\ AlCl_3}{133.33\ \cancel{g\ AlCl_3}} = 1.87\,mol\ AlCl_3$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of AlCl3 and the number of moles of HCl:

$$frac6\,mol\ HCl2\ \cancel{mol\ AlCl_3}$$

Applying this conversion factor to the quantity of AlCl3, we get

$$1.87\ \cancel{mol\ AlCl_3} \times \frac{6\,mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.61\,mol\ HCl$$

Alternatively, we could have done this in one line:

$$249\ \cancel{g\ AlCl_3} \times \frac{1\ \cancel{mol\ AlCl_3}}{133.33\ \cancel{g\ AlCl_3}} \times \frac{6\,mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.60\,mol\ HCl$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

Exercise 5.5.1

How many moles of Al2O3 will be produced when 23.9 g of H2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer available at the end of this section.

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

Example 5.5.2

How many grams of NH3 will be produced when 33.9 mol of H2 are reacted according to this chemical equation?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Solution

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have

$$33.9\ \cancel{mol\ H_2} \times \frac{2\ mol\ NH_3}{3\ \cancel{mol\ H_2}} = 22.6\ mol\ NH_3$$

Now, using the molar mass of NH_3, which is 17.03 g/mol, we get

$$22.6\ \cancel{mol\ NH_3} \times \frac{17.03\ g\ NH_3}{1\ \cancel{mol\ NH_3}} = 385\ g\ NH_3$$

Exercise 5.5.2

How many grams of N_2 are needed to produce 2.17 mol of NH_3 when reacted according to this chemical equation?

$$\mathrm{N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g)}$$

Answer available at the end of this section.

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO_3 that can be produced by the reaction of 45.3 g of SO_2 and O_2:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

First, we convert the given amount, 45.3 g of SO_2, to moles of SO_2 using its molar mass (64.06 g/mol):

$$45.3\ \cancel{g\ SO_2} \times \frac{1\ mol\ SO_2}{64.06\ \cancel{g\ SO_2}} = 0.707\ mol\ SO_2$$

Second, we use the balanced chemical reaction to convert from moles of SO_2 to moles of SO_3:

$$0.707\ \cancel{mol\ SO_2} \times \frac{2\ mol\ SO_3}{2\ \cancel{mol\ SO_2}} = 0.707\ mol\ SO_3$$

Finally, we use the molar mass of SO_3 (80.06 g/mol) to convert to the mass of SO_3:

$$0.707\ \cancel{mol\ SO_3} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

We can also perform all three steps sequentially, writing them on one line as

$$45.3\ \cancel{g\ SO_2} \times \frac{1\ mol\ SO_2}{64.06\ \cancel{g\ SO_2}} \times \frac{2\ mol\ SO_3}{2\ \cancel{mol\ SO_2}} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO_3, which is what we are looking for, as our final answer.

Example 5.5.3

What mass of Mg will be produced when 86.4 g of K are reacted?

$$\mathrm{MgCl_2\,(s) + 2\,K\,(s) \longrightarrow Mg\,(s) + 2\,KCl\,(s)}$$

Solution

We will simply follow the steps

$$\text{mass K} \rightarrow \text{mol K} \rightarrow \text{mol Mg} \rightarrow \text{mass Mg}$$

In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4\ \cancel{g\,K} \times \frac{1\ mol\ K}{39.09\ \cancel{g\,K}} \times \frac{1\ \cancel{mol\ Mg}}{2\ \cancel{mol\ K}} \times \frac{24.31\ g\ Mg}{1\ \cancel{mol\ Mg}} = 26.87\ g\ Mg$$

Exercise 5.5.3

What mass of H_2 will be produced when 122 g of Zn are reacted?

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)}$$

Answer available at the end of this section.

Summary

- Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.
- Mass quantities of one substance can be related to mass quantities using a balanced chemical equation.
- In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

Answers

Exercise 5.5.1

0.442 mol

Exercise 5.5.2

30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)

Exercise 5.5.3

3.77 g

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5.6: Yields

Learning Objective

- Define and determine theoretical yields, actual yields, and percent yields.

In all the previous calculations we have performed involving balanced chemical equations, we made two assumptions:

1. The reaction goes exactly as written.
2. The reaction proceeds completely.

In reality, such things as side reactions occur that make some chemical reactions rather messy. For example, in the actual combustion of some carbon-containing compounds, such as methane, some CO is produced as well as CO_2. However, we will continue to ignore side reactions, unless otherwise noted. The second assumption, that the reaction proceeds completely, is more troublesome. Many chemical reactions do not proceed to completion as written, for a variety of reasons (some of which we will consider in Chapter 13). When we calculate an amount of product assuming that all the reactant reacts, we calculate the **theoretical yield**, an amount that is theoretically produced as calculated using the balanced chemical reaction.

In many cases, however, this is not what really happens. In many cases, less—sometimes, much less—of a product is made during the course of a chemical reaction. The amount that is actually produced in a reaction is called the **actual yield**. By definition, the actual yield is less than or equal to the theoretical yield. If it is not, then an error has been made.

Both theoretical yields and actual yields are expressed in units of moles or grams. It is also common to see something called a percent yield. The **percent yield** is a comparison between the actual yield and the theoretical yield and is defined as

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{5.6.1}$$

It does not matter whether the actual and theoretical yields are expressed in moles or grams, as long as they are expressed in the same units. However, the percent yield always has units of percent. Proper percent yields are between 0% and 100%. Again, if percent yield is greater than 100%, an error has been made.

Example 5.6.1

A worker reacts 30.5 g of Zn with nitric acid and evaporates the remaining water to obtain 65.2 g of $Zn(NO_3)_2$. What are the theoretical yield, the actual yield, and the percent yield?

$$\mathrm{Zn\,(s) + 2\,HNO_3\,(aq) \longrightarrow Zn(NO_3)_2\,(aq) + H_2\,(g)}$$

Solution

A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of Zn (65.39 g/mol) and $Zn(NO_3)_2$ (189.41 g/mol). In three steps, the mass-mass calculation is:

$$30.5\ \cancel{g\ Zn} \times \frac{1\ \cancel{mol\ Zn}}{65.39\ \cancel{g\ Zn}} \times \frac{1\ \cancel{mol\ Zn(NO_3)_2}}{1\ \cancel{mol\ Zn}} \times \frac{189.41\ g\ Zn(NO_3)_2}{1\ \cancel{mol\ Zn(NO_3)_2}} = 88.3\ g\ Zn(NO_3)_2$$

Thus, the theoretical yield is 88.3 g of $Zn(NO_3)_2$. The actual yield is the amount that was actually made, which was 65.2 g of $Zn(NO_3)_2$. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100 (Equation 5.6.1):

$$\frac{65.2\ g\ Zn(NO_3)_2}{88.3\ g\ Zn(NO_3)_2} \times 100\% = 73.8\%$$

The worker achieved almost three-fourths of the possible yield.

Exercise 5.6.1

A synthesis produced 2.05 g of NH_3 from 16.5 g of N_2. What is the theoretical yield and the percent yield?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

*Technically, this is a reversible reaction (with double arrows), but for this exercise consider it irreversible (single arrow).

Answer

theoretical yield = 20.1 g; percent yield = 10.2%

Chemistry is Everywhere: Actual Yields in Drug Synthesis and Purification

Many drugs are the product of several steps of chemical synthesis. Each step typically occurs with less than 100% yield, so the overall percent yield might be very small. The general rule is that the overall percent yield is the product of the percent yields of the individual synthesis steps. For a drug synthesis that has many steps, the overall percent yield can be very tiny, which is one factor in the huge cost of some drugs. For example, if a 10-step synthesis has a percent yield of 90% for each step, the overall yield for the entire synthesis is only 35%. Many scientists work every day trying to improve percent yields of the steps in the synthesis to decrease costs, improve profits, and minimize waste.

Even purifications of complex molecules into drug-quality purity are subject to percent yields. Consider the purification of impure albuterol. Albuterol ($C_{13}H_{21}NO_2$; accompanying figure) is an inhaled drug used to treat asthma, bronchitis, and other obstructive pulmonary diseases. It is synthesized from norepinephrine, a naturally occurring hormone and neurotransmitter. Its initial synthesis makes very impure albuterol that is purified in five chemical steps. The details of the steps do not concern us; only the percent yields do:

A child using an albuterol inhaler, a container of albuterol medication, and a molecular model of albuterol are shown in three combined images.

impure albuterol → intermediate A	percent yield = 70%
intermediate A → intermediate B	percent yield = 100%
intermediate B → intermediate C	percent yield = 40%
intermediate C → intermediate D	percent yield = 72%
intermediate D → purified albuterol	percent yield = 35%
overall percent yield = 70% × 100% × 40% × 72% × 35% = 7.5%	

That is, only about *one-fourteenth* of the original material was turned into the purified drug. This demonstrates one reason why some drugs are so expensive—a lot of material is lost in making a high-purity pharmaceutical.



Figure 5.6.1 A child using an albuterol inhaler, the container of albuterol medication, and a molecular model of the albuterol molecule. Source: Photo on far left © Thinkstock. Photo in center courtesy of Intropin, http://commons.wikimedia.org/wiki/Fi...te_%281%29.JPG.

Summary

Theoretical yield is the calculated yield using the balanced chemical reaction. Actual yield is what is actually obtained in a chemical reaction. Percent yield is a comparison of the actual yield with the theoretical yield.

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5.7: Limiting Reagents

Learning Objectives

- Identify a limiting reagent from a set of reactants.
- Calculate how much product will be produced from the limiting reagent.
- Calculate how much reactant(s) remains when the reaction is complete.

In addition to the assumption that reactions proceed all the way to completion, one additional assumption we have made about chemical reactions is that all the reactants are present in the proper quantities to react to products; this is not always the case. In Figure 5.7.2 we are taking hydrogen atoms and oxygen atoms (left) to make water molecules (right). However, there are not enough oxygen atoms to use up all the hydrogen atoms. We run out of oxygen atoms and cannot make any more water molecules, so the process stops when we run out of oxygen atoms.

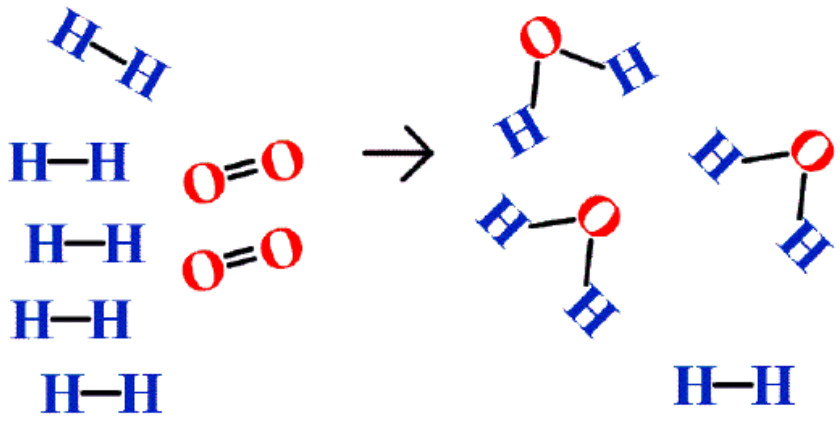


Figure 5.7.2: Making Water. In this scenario for making water molecules, we run out of O atoms before we use up all the H atoms. Similar situations exist for many chemical reactions when one reactant runs out before the other.

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reagent; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reagent and which is in excess.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reagent. What we need to do is determine an amount of one product (either moles or mass), assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reagent. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

For example, consider this reaction:

$$4As(s) + 3O_2(g) \rightarrow 2As_2O_3(s)$$

Suppose we start a reaction with 50.0 g of As and 50.0 g of O_2. Which one is the limiting reagent? We need to perform two mole-mass calculations, each assuming that each reactant reacts completely. Then we compare the amount of the product produced by each and determine which is less.

The calculations are as follows:

$$50.0\ \cancel{g\,As} \times \frac{1\ \cancel{mol\,As}}{74.92\ \cancel{g\,As}} \times \frac{2\,mol\,As_2O_3}{4\ \cancel{mol\,As}} = 0.334\,mol\,As_2O_3$$

$$50.0\ \cancel{g\,O_2} \times \frac{1\ \cancel{mol\,O_2}}{32.00\ \cancel{g\,O_2}} \times \frac{2\,mol\,As_2O_3}{3\ \cancel{mol\,O_2}} = 1.04\,mol\,As_2O_3$$

Comparing these two answers, it is clear that 0.334 mol of As2O3 is less than 1.04 mol of As2O3, so arsenic is the limiting reagent. If this reaction is performed under these initial conditions, the arsenic will run out before the oxygen runs out. We say that the oxygen is "in excess."

Identifying the limiting reagent, then, is straightforward. However, there are usually two associated questions: (1) what mass of product (or products) is then actually formed? and (2) what mass of what reactant is left over? The first question is straightforward to answer: simply perform a conversion from the number of moles of product formed to its mass, using its molar mass. For As2O3, the molar mass is 197.84 g/mol; knowing that we will form 0.334 mol of As2O3 under the given conditions, we will get

$$0.334\ \cancel{mol\,As_2O_3} \times \frac{197.84\,g\,As_2}{1\ \cancel{mol\,As_2O_3}} = 66.1\,g\,As_2O_3$$

The second question is somewhat more convoluted to answer. First, we must do a mass-mass calculation relating the limiting reagent (here, As) to the other reagent (O2). Once we determine the mass of O2 that reacted, we subtract that from the original amount to determine the amount left over. According to the mass-mass calculation,

$$50.0\ \cancel{g\,As} \times \frac{1\ \cancel{mol\,As}}{74.92\ \cancel{g\,As}} \times \frac{3\ \cancel{mol\,O_2}}{4\ \cancel{mol\,As}} \times \frac{32.00\,g\,O_2}{1\ \cancel{mol\,O_2}} = 16.0\,g\,O_2\ reacted$$

Because we reacted 16.0 g of our original O2, we subtract that from the original amount, 50.0 g, to get the mass of O2 remaining:

50.0 g O2 – 16.0 g O2 reacted = 34.0 g O2 left over

You must remember to perform this final subtraction to determine the amount remaining; a common error is to report the 16.0 g as the amount remaining.

Example 5.7.1

A 5.00 g quantity of Rb is combined with 3.44 g of MgCl2 according to this chemical reaction:

$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of Mg is formed, and what mass of what reactant is left over?

Solution

Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less.

$$5.00\ \cancel{g\,Rb} \times \frac{1\ \cancel{mol\,Rb}}{85.47\ \cancel{g\,Rb}} \times \frac{1\ \cancel{mol\,Mg}}{2\ \cancel{mol\,Rb}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 0.711\,g\,Mg$$

$$3.44\ \cancel{g\,MgCl_2} \times \frac{1\ \cancel{mol\,MgCl_2}}{95.21\ \cancel{g\,MgCl_2}} \times \frac{1\ \cancel{mol\,Mg}}{1\ \cancel{mol\,MgCl_2}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 0.878\,g\,Mg$$

The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of MgCl2 reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount.

$$5.00\ \cancel{g\ Rb} \times \frac{1\ \cancel{mol\ Rb}}{85.47\ \cancel{g\ Rb}} \times \frac{1\ \cancel{mol\ MgCl_2}}{2\ \cancel{mol\ Rb}} \times \frac{95.21\ g\ Mg}{1\ \cancel{mol\ MgCl_2}} = 2.78\ g\ MgCl_2\ \ reacted$$

Because we started with 3.44 g of MgCl2, we have

3.44 g MgCl2 – 2.78 g MgCl2 reacted = 0.66 g MgCl2 left

Exercise 5.7.1

Given the initial amounts listed, what is the limiting reagent, and what is the mass of the leftover reagent?

$$\underbrace{22.7\ g}_{MgO(s)} + \underbrace{17.9\ g}_{H_2S} \rightarrow MgS(s) + H_2O(l)$$

Answer available at the end of this section.

Summary

The limiting reagent is the reactant that produces the least amount of product. Mass-mass calculations can determine how much product is produced and how much of the other reactants remain.

Answers

Exercise 5.7.1

H2S is the limiting reagent; 1.5 g of MgO are left over.

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5.E: Stoichiometry and the Mole (Exercises)

Exercises (Stoichiometry)

1. Think back to the pound cake recipe. What possible conversion factors can you construct relating the components of the recipe?

2. Think back to the pancake recipe. What possible conversion factors can you construct relating the components of the recipe?

3. What are all the conversion factors that can be constructed from the balanced chemical reaction: $2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$?

4. What are all the conversion factors that can be constructed from the balanced chemical reaction $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$?

5. Given the chemical equation : $Na(s) + H_2O(\ell) \rightarrow NaOH(aq) + H_2(g)$
   a. Balance the equation.
   b. How many molecules of H_2 are produced when 332 atoms of Na react?

6. Given the chemical equation: $S(s) + O_2(g) \rightarrow SO_3(g)$
   a. Balance the equation.
   b. How many molecules of O_2 are needed when 38 atoms of S react?

7. For the balanced chemical equation:

   $6H^+(aq) + 2MnO_4^-(aq) + 5H_2O_2(\ell) \rightarrow 2Mn^{2+}(aq) + 5O_2(g) + 8H_2O(\ell)$

   how many molecules of H_2O are produced when 75 molecules of H_2O_2 react?

8. For the balanced chemical reaction

   $2C_6H_6(\ell) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(\ell)$

   how many molecules of CO_2 are produced when 56 molecules of C_6H_6 react?

9. Given the balanced chemical equation

   $Fe_2O_3(s) + 3SO_3(g) \rightarrow Fe_2(SO_4)_3$

   how many molecules of $Fe_2(SO_4)_3$ are produced if 321 atoms of S are reacted?

10. For the balanced chemical equation

    $CuO(s) + H_2S(g) \rightarrow CuS + H_2O(\ell)$

    how many molecules of CuS are formed if 9,044 atoms of H react?

11. For the balanced chemical equation

    $Fe_2O_3(s) + 3SO_3(g) \rightarrow Fe_2(SO_4)_3$

    suppose we need to make 145,000 molecules of $Fe_2(SO_4)_3$. How many molecules of SO_3 do we need?

12. One way to make sulfur hexafluoride is to react thioformaldehyde, CH_2S, with elemental fluorine:

$CH_2S + 6F_2 \rightarrow CF_4 + 2HF + SF_6$

If 45,750 molecules of SF_6 are needed, how many molecules of F_2 are required?

13. Construct the three independent conversion factors possible for these two reactions:
    a. $2H_2 + O_2 \rightarrow 2H_2O$
    b. $H_2 + O_2 \rightarrow H_2O_2$

Why are the ratios between H_2 and O_2 different?

The conversion factors are different because the stoichiometries of the balanced chemical reactions are different.

14. Construct the three independent conversion factors possible for these two reactions:
    a. $2Na + Cl_2 \rightarrow 2NaCl$
    b. $4Na + 2Cl_2 \rightarrow 4NaCl$

What similarities, if any, exist in the conversion factors from these two reactions?

Answers

1. $\frac{1\,pound\,butter}{1\,pound\,flour}$ or $\frac{1\,pound\,sugar}{1\,pound\,eggs}$ are two conversion factors that can be constructed from the pound cake recipe. Other conversion factors are also possible.

3. $\frac{2\,molecules\,H_2}{1\,molecule\,O_2}$, $\frac{1\,molecule\,O_2}{2\,molecules\,H_2O}$, $\frac{2\,molecules\,H_2}{2\,molecules\,H_2O}$ and their reciprocals are the conversion factors that can be constructed.

5.

a. $2Na(s) + 2H_2O(\ell) \rightarrow 2NaOH(aq) + H_2(g)$
b. 166 molecules

7. 120 molecules

9. 107 molecules

11. 435,000 molecules

13.

a. $\frac{2\,molecules\,H_2}{1\,molecule\,O_2}$, $\frac{1\,molecule\,O_2}{2\,molecules\,H_2O}$, $\frac{2\,molecules\,H_2}{2\,molecules\,H_2O}$
b. $\frac{1\,molecules\,H_2}{1\,molecule\,O_2}$, $\frac{1\,molecule\,O_2}{2\,molecules\,H_2O_2}$, $\frac{1\,molecule\,H_2}{1\,molecule\,H_2O_2}$

Excercises (The Mole)

1. How many atoms are present in 4.55 mol of Fe?

2. How many atoms are present in 0.0665 mol of K?

3. How many molecules are present in 2.509 mol of H_2S?

4. How many molecules are present in 0.336 mol of acetylene (C_2H_2)?

5. How many moles are present in 3.55×10^{24} Pb atoms?

6. How many moles are present in 2.09×10^{22} Ti atoms?

7. How many moles are present in 1.00×10^{23} PF_3 molecules?

8. How many moles are present in 5.52×10^{25} penicillin molecules?

9. Determine the molar mass of each substance.
   a. Si
   b. SiH_4
   c. K_2O

10. Determine the molar mass of each substance.
   a. Cl_2
   b. $SeCl_2$
   c. $Ca(C_2H_3O_2)_2$

11. Determine the molar mass of each substance.
   a. Al
   b. Al_2O_3
   c. $CoCl_3$

12. Determine the molar mass of each substance.
   a. O_3
   b. NaI
   c. $C_{12}H_{22}O_{11}$

13. What is the mass of 4.44 mol of Rb?

14. What is the mass of 0.311 mol of Xe?

15. What is the mass of 12.34 mol of $Al_2(SO_4)_3$?

16. What is the mass of 0.0656 mol of $PbCl_2$?

17. How many moles are present in 45.6 g of CO?

18. How many moles are present in 0.00339 g of LiF?

19. How many moles are present in 1.223 g of SF_6?

20. How many moles are present in 48.8 g of $BaCO_3$?

21. How many moles are present in 54.8 mL of mercury if the density of mercury is 13.6 g/mL?

22. How many moles are present in 56.83 mL of O_2 if the density of O_2 is 0.00133 g/mL?

Answers

1. 2.74×10^{24} atoms

3. 1.511×10^{24} molecules

5. 5.90 mol

7. 0.166 mol

9.

   a. 28.086 g
   b. 32.118 g
   c. 94.195 g

11.

   a. 26.981 g
   b. 101.959 g
   c. 165.292 g

13. 379 g

15. 4,222 g

17. 1.63 mol

19. 0.008374 mol

21. 3.72 mol

Exercises (The Mole in Chemical Reactions)

1. Express in mole terms what this chemical equation means: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
2. Express in mole terms what this chemical equation means.

   $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$
3. How many molecules of each substance are involved in the equation in Exercise 1 if it is interpreted in terms of moles?
4. How many molecules of each substance are involved in the equation in Exercise 2 if it is interpreted in terms of moles?
5. For the chemical equation

   $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

   what equivalents can you write in terms of moles? Use the $\Leftrightarrow$ sign.
6. For the chemical equation

   $2Al + 3Cl_2 \rightarrow 2AlCl_3$

   what equivalents can you write in terms of moles? Use the $\Leftrightarrow$ sign.
7. Write the balanced chemical reaction for the combustion of C_5H_{12} (the products are CO_2 and H_2O) and determine how many moles of H_2O are formed when 5.8 mol of O_2 are reacted.
8. Write the balanced chemical reaction for the formation of $Fe_2(SO_4)_3$ from Fe_2O_3 and SO_3 and determine how many moles of $Fe_2(SO_4)_3$ are formed when 12.7 mol of SO_3 are reacted.
9. For the balanced chemical equation

$3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 4H_2O(\ell) + 2NO(g)$

how many moles of Cu^{2+} are formed when 55.7 mol of H^+ are reacted?

10. For the balanced chemical equation

$Al(s) + 3Ag^+(aq) \rightarrow Al^{3+}(aq) + 3Ag(s)$

how many moles of Ag are produced when 0.661 mol of Al are reacted?

11. For the balanced chemical reaction

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(\ell)$

how many moles of H_2O are produced when 0.669 mol of NH_3 react?

12. For the balanced chemical reaction

$4NaOH(aq) + 2S(s) + 3O_2(g) \rightarrow 2Na_2SO_4(aq) + 2H_2O(\ell)$

how many moles of Na_2SO_4 are formed when 1.22 mol of O_2 react?

13. For the balanced chemical reaction

$4KO_2(s) + 2CO_2(g) \rightarrow 2K_2CO_3(s) + 3O_2(g)$

determine the number of moles of both products formed when 6.88 mol of KO_2 react.

14. For the balanced chemical reaction

$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$

determine the number of moles of both products formed when 0.0552 mol of $AlCl_3$ react.

Answers

1. One mole of CH_4 reacts with 2 mol of O_2 to make 1 mol of CO_2 and 2 mol of H_2O.

3. 6.022×10^{23} molecules of CH_4, 1.2044×10^{24} molecules of O_2, 6.022×10^{23} molecules of CO_2, and 1.2044×10^{24} molecules of H_2O

5. 2 mol of C_2H_6 ⇔ 7 mol of O_2 ⇔ 4 mol of CO_2 ⇔ 6 mol of H_2O

7. $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$; 4.4 mol

9. 20.9 mol

11. 1.00 mol

13. 3.44 mol of K_2CO_3; 5.16 mol of O_2

Exercises (Mole-Mass and Mass-Mass Calculations)

1. What mass of CO_2 is produced by the combustion of 1.00 mol of CH_4?

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$

2. What mass of H_2O is produced by the combustion of 1.00 mol of CH_4?

   $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell)$

3. What mass of HgO is required to produce 0.692 mol of O_2?

   $2HgO(s) \rightarrow 2Hg(\ell) + O_2(g)$

4. What mass of $NaHCO_3$ is needed to produce 2.659 mol of CO_2?

   $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(\ell) + CO_2(g)$

5. How many moles of Al can be produced from 10.87 g of Ag?

   $Al(NO_3)\ 3(s) + 3Ag \rightarrow Al + 3AgNO_3$

6. How many moles of HCl can be produced from 0.226 g of $SOCl_2$?

   $SOCl_2(\ell) + H_2O(\ell) \rightarrow SO_2(g) + 2HCl(g)$

7. How many moles of O_2 are needed to prepare 1.00 g of $Ca(NO_3)_2$?

   $Ca(s) + N_2(g) + 3O_2(g) \rightarrow Ca(NO_3)\ 2(s)$

8. How many moles of C_2H_5OH are needed to generate 106.7 g of H_2O?

   $C_2H_5OH(\ell) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(\ell)$

9. What mass of O_2 can be generated by the decomposition of 100.0 g of $NaClO_3$?

   $2NaClO_3 \rightarrow 2NaCl(s) + 3O_2(g)$

10. What mass of Li_2O is needed to react with 1,060 g of CO_2?

    $Li_2O(aq) + CO_2(g) \rightarrow Li_2CO_3(aq)$

11. What mass of Fe_2O_3 must be reacted to generate 324 g of Al_2O_3?

    $Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(s) + Al_2O_3(s)$

12. What mass of Fe is generated when 100.0 g of Al are reacted?

    $Fe_2O_3(s) + 2Al(s) \rightarrow 2Fe(s) + Al_2O_3(s)$

13. What mass of MnO_2 is produced when 445 g of H_2O are reacted?

    $H_2O(\ell) + 2MnO_4^-(aq) + Br^-(aq) \rightarrow BrO_3^-(aq) + 2MnO_2(s) + 2OH^-(aq)$

14. What mass of $PbSO_4$ is produced when 29.6 g of H_2SO_4 are reacted?

    $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(s) + 2H_2O(\ell)$

15. If 83.9 g of ZnO are formed, what mass of Mn_2O_3 is formed with it?

    $Zn(s) + 2MnO_2(s) \rightarrow ZnO(s) + Mn_2O_3(s)$

16. If 14.7 g of NO_2 are reacted, what mass of H_2O is reacted with it?

    $3NO_2(g) + H_2O(\ell) \rightarrow 2HNO_3(aq) + NO(g)$

17. If 88.4 g of CH_2S are reacted, what mass of HF is produced?

    $CH_2S + 6F_2 \rightarrow CF_4 + 2HF + SF_6$

18. If 100.0 g of Cl_2 are needed, what mass of NaOCl must be reacted?

    $NaOCl + HCl \rightarrow NaOH + Cl_2$

Answers

1. 44.0 g

3. 3.00×10^2 g

5. . 0.0336 mol

7. 0.0183 mol

9. 45.1 g

11. 507 g

13. 4.30×10^3 g

15. 163 g

17. 76.7 g

Exercises (Yields)

1. What is the difference between the theoretical yield and the actual yield?
2. What is the difference between the actual yield and the percent yield?
3. A worker isolates 2.675 g of SiF_4 after reacting 2.339 g of SiO_2 with HF. What are the theoretical yield and the actual yield?

   $SiO_2(s) + 4HF(g) \rightarrow SiF_4(g) + 2H_2O(\ell)$

4. A worker synthesizes aspirin, $C_9H_8O_4$, according to this chemical equation. If 12.66 g of $C_7H_6O_3$ are reacted and 12.03 g of aspirin are isolated, what are the theoretical yield and the actual yield?

   $C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + HC_2H_3O_2$

5. A chemist decomposes 1.006 g of $NaHCO_3$ and obtains 0.0334 g of Na_2CO_3. What are the theoretical yield and the actual yield?

   $2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(\ell) + CO_2(g)$

6. A chemist combusts a 3.009 g sample of C_5H_{12} and obtains 3.774 g of H_2O. What are the theoretical yield and the actual yield?

   $C_5H_{12}(\ell) + 8O_2(g) \rightarrow 5CO_2 + 6H_2O(\ell)$

7. What is the percent yield in Exercise 3?
8. What is the percent yield in Exercise 4?
9. What is the percent yield in Exercise 5?
10. What is the percent yield in Exercise 6?

Answers

1. Theoretical yield is what you expect stoichiometrically from a chemical reaction; actual yield is what you actually get from a chemical reaction.

3. theoretical yield = 4.0052 g; actual yield = 2.675 g

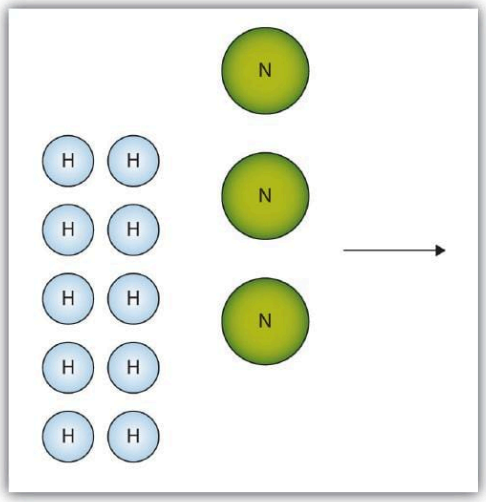
5. theoretical yield = 0.635 g; actual yield = 0.0334 g

7. 66.02%

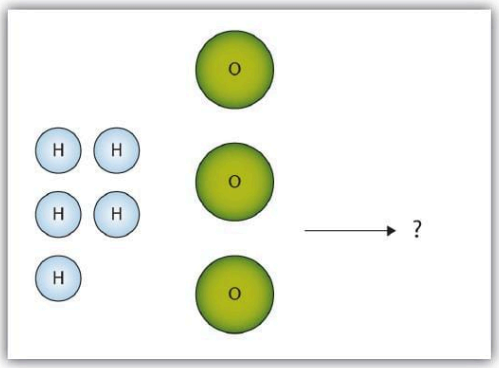
9. 5.26%

Exercises (Limiting Reagents)

1. The box below shows a group of nitrogen and hydrogen molecules that will react to produce ammonia, NH_3. What is the limiting reagent?



2. The box below shows a group of hydrogen and oxygen molecules that will react to produce water, H_2O. What is the limiting reagent?



3. Given the statement “20.0 g of methane is burned in excess oxygen,” is it obvious which reactant is the limiting reagent?

4. Given the statement “the metal is heated in the presence of excess hydrogen,” is it obvious which substance is the limiting reagent despite not specifying any quantity of reactant?

5. Acetylene (C_2H_2) is formed by reacting 7.08 g of C and 4.92 g of H_2.

   $2C(s) + H_2(g) \rightarrow C_2H_2(g)$

   What is the limiting reagent? How much of the other reactant is in excess?

6. Ethane (C_2H_6) is formed by reacting 7.08 g of C and 4.92 g of H_2.

   $2C(s) + 3H_2(g) \rightarrow C_2H_6(g)$

   What is the limiting reagent? How much of the other reactant is in excess?

7. Given the initial amounts listed, what is the limiting reagent, and how much of the other reactant is in excess?

$$\underset{35.6\ g}{P_4O_6(s)} + \underset{4.77\ g}{6H_2O(l)} \rightarrow 4H_3PO_4 \tag{5.E.1}$$

8. Given the initial amounts listed, what is the limiting reagent, and how much of the other reactant is in excess?

$$\underset{377\ g}{3NO_2(g)} + \underset{244\ g}{H_2O(l)} \rightarrow 2HNO_3(aq) + NO(g) \tag{5.E.2}$$

9. To form the precipitate $PbCl_2$, 2.88 g of NaCl and 7.21 g of $Pb(NO_3)_2$ are mixed in solution. How much precipitate is formed? How much of which reactant is in excess?

10. In a neutralization reaction, 18.06 g of KOH are reacted with 13.43 g of HNO_3. What mass of H_2O is produced, and what mass of which reactant is in excess?

Answers

1. Nitrogen is the limiting reagent.

3. Yes; methane is the limiting reagent.

5. C is the limiting reagent; 4.33 g of H_2 are left over.

7. H_2O is the limiting reagent; 25.9 g of P_4O_6 are left over.

9. 6.06 g of $PbCl_2$ are formed; 0.33 g of NaCl is left over.

Additional Exercises

1. How many molecules of O_2 will react with 6.022×10^{23} molecules of H_2 to make water? The reaction is $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$.

2. How many molecules of H_2 will react with 6.022×10^{23} molecules of N_2 to make ammonia? The reaction is $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$.

3. How many moles are present in 6.411 kg of CO_2? How many molecules is this?

4. How many moles are present in 2.998 mg of SCl_4? How many molecules is this?

5. What is the mass in milligrams of 7.22×10^{20} molecules of CO_2?

6. What is the mass in kilograms of 3.408×10^{25} molecules of SiS_2?

7. What is the mass in grams of 1 molecule of H_2O?

8. What is the mass in grams of 1 atom of Al?

9. What is the volume of 3.44 mol of Ga if the density of Ga is 6.08 g/mL?

10. What is the volume of 0.662 mol of He if the density of He is 0.1785 g/L?

11. For the chemical reaction

    $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(\ell)$

    assume that 13.4 g of C_4H_{10} reacts completely to products. The density of CO_2 is 1.96 g/L. What volume in liters of CO_2 is produced?

12. For the chemical reaction

    $2GaCl_3(s) + 3H_2(g) \rightarrow 2Ga(\ell) + 6HCl(g)$

    if 223 g of $GaCl_3$ reacts completely to products and the density of Ga is 6.08 g/mL, what volume in milliliters of Ga is produced?

13. Calculate the mass of each product when 100.0 g of CuCl react according to the reaction

    $2CuCl(aq) \rightarrow CuCl_2(aq) + Cu(s)$

    What do you notice about the sum of the masses of the products? What concept is being illustrated here?

14. Calculate the mass of each product when 500.0 g of $SnCl_2$ react according to the reaction

    $2SnCl_2(aq) \rightarrow SnCl_4(aq) + Sn(s)$

    What do you notice about the sum of the masses of the products? What concept is being illustrated here?

15. What mass of CO_2 is produced from the combustion of 1 gal of gasoline? The chemical formula of gasoline can be approximated as C_8H_{18}. Assume that there are 2,801 g of gasoline per gallon.

16. What mass of H_2O is produced from the combustion of 1 gal of gasoline? The chemical formula of gasoline can be approximated as C_8H_{18}. Assume that there are 2,801 g of gasoline per gallon.

17. A chemical reaction has a theoretical yield of 19.98 g and a percent yield of 88.40%. What is the actual yield?

18. A chemical reaction has an actual yield of 19.98 g and a percent yield of 88.40%. What is the theoretical yield?

19. Given the initial amounts listed, what is the limiting reagent, and how much of the other reactants are in excess?

$$\underset{35.0\ g}{P_4} + 3NaOH + \underset{9.33\ g}{3H_2O} \rightarrow 2Na_2HPO_4 + PH_3 \tag{5.E.3}$$

20. Given the initial amounts listed, what is the limiting reagent, and how much of the other reactants are in excess?

$$\underset{46.3\ g}{2NaCrO_2} + \underset{88.2\ g}{3NaBrO_4} + \underset{32.5\ g}{2NaOH} \rightarrow 3NaBrO_3 + 2Na_2CrO_4 + H_2O \tag{5.E.4}$$

21. Verify that it does not matter which product you use to predict the limiting reagent by using both products in this combustion reaction to determine the limiting reagent and the amount of the reactant in excess. Initial amounts of each reactant are given.

$$\underset{26.3\ g}{C_3H_8} + \underset{21.8\ g}{5O_2} \rightarrow 3CO_2(g) + 4H_2O(l) \tag{5.E.5}$$

22. Just in case you suspect Exercise 21 is rigged, do it for another chemical reaction and verify that it does not matter which product you use to predict the limiting reagent by using both products in this combustion reaction to determine the limiting reagent and the amount of the reactant in excess. Initial amounts of each reactant are given.

$$\underset{35.0\ g}{2P_4} + \underset{12.7\ g}{6NaOH} + \underset{9.33\ g}{6H_2O} \rightarrow 3Na_2HPO_4 + 5PH_3 \tag{5.E.6}$$

Answers

1. 1.2044 × 10^{24} molecules

3. 145.7 mol; 8.77 × 10^{25} molecules

5. 52.8 mg

7. 2.99 × 10^{-23} g

9. 39.4 mL

11. 20.7 L

13. 67.91 g of $CuCl_2$; 32.09 g of Cu. The two masses add to 100.0 g, the initial amount of starting material, demonstrating the law of conservation of matter.

15. 8,632 g

17. 17.66 g

19. The limiting reagent is NaOH; 21.9 g of P_4 and 3.61 g of H_2O are left over.

21. Both products predict that O_2 is the limiting reagent; 20.3 g of C_3H_8 are left over.

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6: CHEMISTRY OF EARTH

6.1: Spaceship Earth- Structure and Composition
6.2: Silicates and the Shapes of Things
6.3: Metals and Ores
6.4: Earth's Dwindling Resources

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6.1: Spaceship Earth- Structure and Composition

Learning Objective

- Describe the chemical composition and structure of Earth.

The layers found inside Earth are divided by composition into core, mantle, and crust or by mechanical properties into lithosphere and asthenosphere. Scientists use information from earthquakes and computer modeling to learn about Earth's interior.

THE EARTH'S LAYERS

The layers scientists recognize are pictured in Figure 6.1.1.

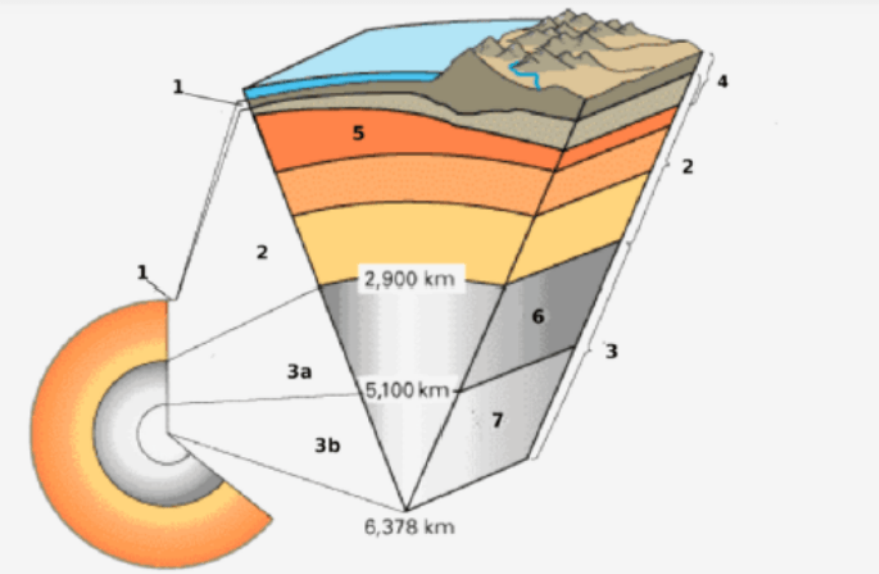


Figure 6.1.1 *A cross section of Earth showing the following layers: (1) crust (2) mantle (3a) outer core (3b) inner core (4) lithosphere (5) asthenosphere (6) outer core (7) inner core.*

Core, mantle, and crust are divisions based on chemical composition:

Crust: The Earth's surface is the crust. Generally speaking, the crust is predominately silicon oxide and aluminum oxide. Continental crust is thicker and less dense than oceanic crust. Earth's crust varies in thickness from less than 5 km (under mid-ocean spreading ridges) to more than 70 km (beneath the highest mountain range).

Mantle: The next layer down chemically is the mantle. The mantle has an ultramafic composition – it contains more iron, magnesium, less aluminum and somewhat less silicon than the crust. The mantle is roughly 2,900 km thick. In terms of volume, the mantle is the largest of earth's three chemical layers.

Core:The final layer is the core, which is mostly iron and nickel. The core is about 3,500 km thick.

Table 6.1.1 summarizes the chemical layers of the earth.

Table 6.1.1 Chemical Layers of Earth.

Crust	Mantle	Core
composition: high Si, Al, & O	*composition:* moderate Si, high Mg & Fe	*composition:* Fe & Ni

thickness: 5 to 70 km	*thickness:* 2,900 km	*thickness:* 3,500 km

Table 6.1.2 provides the elemental composition of the Earth's crust.

Table 6.1.2 The Elements of Earth's Crust. Source: Wikipedia

Most Abundant Elements of Earth's Crust	Approximate % by weight
O	46.6
Si	27.7
Al	8.1
Fe	5.0
Ca	3.6
Na	2.8
K	2.6
Mg	1.5

Lithosphere and asthenosphere are divisions based on mechanical properties:

1. The lithosphere is composed of both the crust and the portion of the upper mantle that behaves as a brittle, rigid solid.
2. The asthenosphere is partially molten upper mantle material that behaves plastically and can flow.

Compositional and Mechanical Layers of the Earth

https://www.khanacademy.org/science/cosmology-and-astronomy/earth-history-topic/plate-techtonics/v/compositional-and-mechanical-layers-of-the-earth

Video 6.1.1 *A comparison of the compositional and mechanical layers of the earth.*

Crust and Lithosphere

Earth's outer surface is its crust; a cold, thin, brittle outer shell made of rock. The crust is very thin, relative to the radius of the planet. There are two very different types of crust, each with its own distinctive physical and chemical properties, which are summarized in Table 6.1.3.

Table 6.1.3 The Physical and Chemical Properties of Oceanic and Continental Crust.

Crust	Thickness	Density	Composition	Rock types
Oceanic	5-12 km (3-8 mi)	3.0 g/cm^3	Mafic	Basalt and gabbr
Continental	Avg. 35 km (22 mi)	2.7 g/cm^3	Felsic	All types

Oceanic crust is composed of mafic magma that erupts on the seafloor to create basalt lava flows or cools deeper down to create the intrusive igneous rock gabbro (Figure 6.1.2).



Figure 6.1.2 Gabbro

Sediments, primarily muds and the shells of tiny sea creatures, coat the sea floor. Sediment is thickest near the shore where it comes off the continents in rivers and on wind currents.

Continental crust is made up of many different types of igneous, metamorphic, and sedimentary rocks. The average composition is granite, which is much less dense than the mafic rocks of the oceanic crust (Figure 6.1.3). Because it is thick and has relatively low density, continental crust rises higher on the mantle than oceanic crust, which sinks into the mantle to form basins. When filled with water, these basins form the planet's oceans.

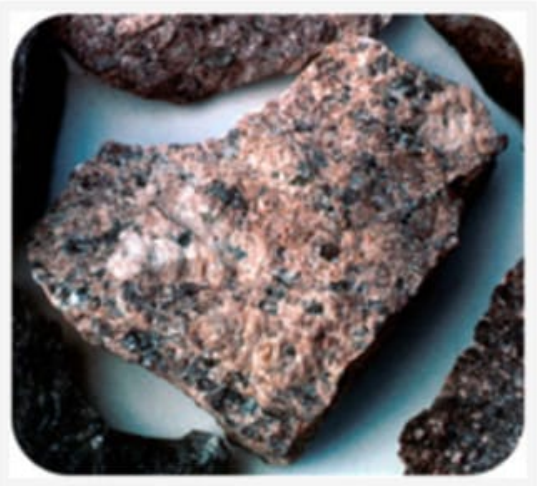


Figure 6.1.3 *The granite from Missouri is more than 1 billion years old.*

The **lithosphere** is the outermost mechanical layer, which behaves as a brittle, rigid solid. The lithosphere is about 100 kilometers thick. Look at Figure 6.1.1. Can you find where the crust and the lithosphere are located? How are they different from each other?

The definition of the lithosphere is based on how earth materials behave, so it includes the crust and the uppermost mantle, which are both brittle. Since it is rigid and brittle, when stresses act on the lithosphere, it breaks. This is what we experience as an earthquake.

Mantle

The two most important things about the mantle are: (1) it is made of solid rock, and (2) it is hot. Scientists know that the mantle is made of rock based on evidence from seismic waves, heat flow, and meteorites.

The properties fit the ultramafic rock peridotite, which is made of the iron- and magnesium-rich silicate minerals (Figure 6.1.4). Peridotite is rarely found at Earth's surface.

Scientists know that the mantle is extremely hot because of the heat flowing outward from it and because of its physical properties.



Figure 6.1.4 *Peridotite is formed of crystals of olivine (green) and pyroxene (black).*

Core

At the planet's center lies a dense metallic core. Scientists know that the core is metal because:

1. The density of Earth's surface layers is much less than the overall density of the planet, as calculated from the planet's rotation. If the surface layers are less dense than average, then the interior must be denser than average. Calculations indicate that the core is about 85% iron metal with nickel metal making up much of the remaining 15%.
2. Metallic meteorites are thought to be representative of the core. The 85% iron/15% nickel calculation above is also seen in metallic meteorites (Figure 6.1.5).



Figure 6.1.5 *An iron meteorite is the closest thing to the Earth's core that we can hold in our hands.*

If Earth's core were not metal, the planet would not have a magnetic field. Metals such as iron are magnetic, but rock, which makes up the mantle and crust, is not.

Summary

- The core, mantle, and crust are divisions of Earth based on chemical composition.
- The lithosphere and asthenosphere are divisions of Earth based on mechanical properties.
- The three most abundant elements of Earth's crust are oxygen (46.6% by weight), silicon (27.7%), and aluminum (8.1%).

Contributors

Libretext: Fundamentals of Geology (Schulte)

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Wikipedia

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6.2: Silicates and the Shapes of Things

Learning Objective

- Describe different types of silicate based minerals.

The silicates are the largest, the most interesting and the most complicated class of minerals than any other minerals. Approximately 30% of all minerals are silicates and some geologists estimate that 90% of the Earth's crust is made up of silicates, SiO_4^{4-} based material. Thus, oxygen and silicon are the two most abundant elements in the earth's crust.

The building block of all of these minerals is the **silica tetrahedron**, a combination of four oxygen atoms and one silicon atom. As we've seen, it's called a tetrahedron because planes drawn through the oxygen atoms form a shape with 4 surfaces (Figure 6.2.1). Since the silicon ion has a charge of 4 and each of the four oxygen ions has a charge of –2, the silica tetrahedron has a net charge of –4.

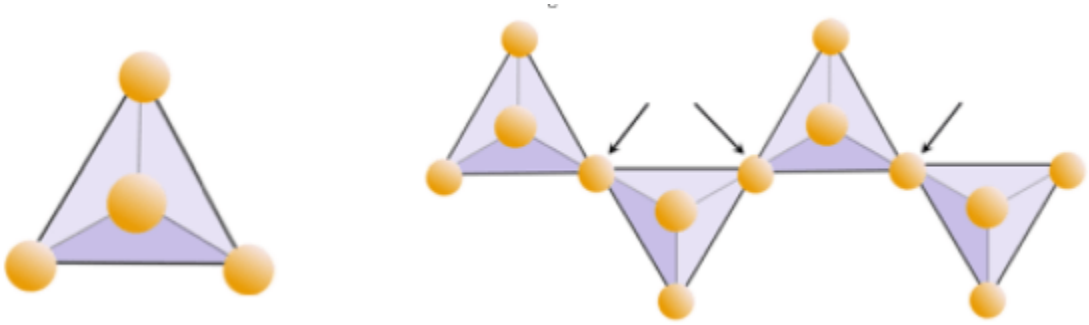
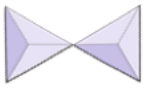
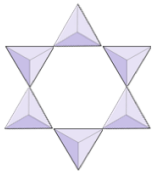
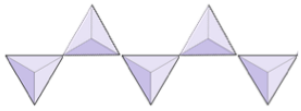
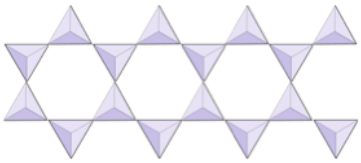
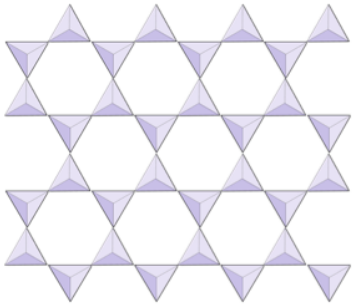


Figure 6.2.1 *A single silica tetrahedron (left) with four oxygen ions per silicon ion (SiO_4). Part of a single chain of tetrahedra (right), where the oxygen atoms at the adjoining corners are shared between two tetrahedra (arrows). For a very long chain the resulting ratio of silicon to oxygen is 1 to 3 (SiO_3).*

In silicate minerals, these tetrahedra are arranged and linked together in a variety of ways, from single units to complex frameworks (Table 6.2.1).

Table 6.2.1 Silicate Mineral Configurations (The triangles represent silica tetrahedra).

Tetrahedron Configuration Picture	Tetrahedron Configuration Name	Example Minerals
	Isolated (nesosilicates)	Olivine, garnet, zircon, kyanite
	Pairs (sorosilicates)	Epidote, zoisite
	Rings (cyclosilicates)	Tourmaline
	Single chains (inosilicates)	Pyroxenes, wollastonite
	Double chains (inosilicates)	Amphiboles
	Sheets (phyllosilicates)	Micas, clay minerals, serpentine, chlorite
3-dimensional structure	Framework (tectosilicates)	Feldspars, quartz, zeolite

In the extreme case, the tetrahedra are arranged in a regular, orderly fashion forming a three-dimensional network. **Quartz** is such a structure (Figure 6.2.2), and its formula is SiO_2. If silica in the molten state is cooled very slowly it crystallizes at the freezing point. But if molten silica is cooled more rapidly, the resulting solid is a disorderly arrangement which is called **glass**, often also called *quartz*.

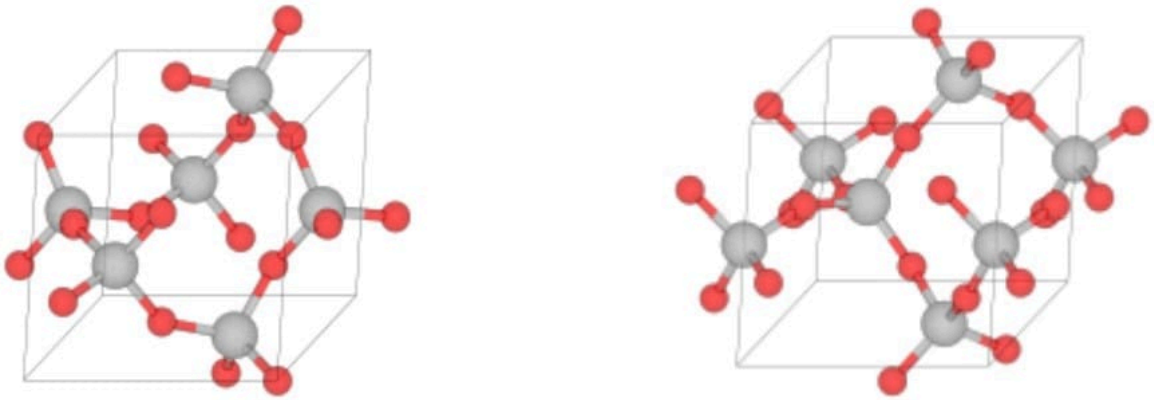


Figure 6.2.2 Crystal structure of alpha (left) and beta (right) quartz. The red balls are oxygen and gray are silicon.

Note:

Silicates are extremely important materials, both natural (such as granite, gravel, and garnet) and artificial (such as Portland cement, ceramics, glass, and waterglass), for all sorts of technological and artistic activities. Source: Wikipedia

Glasses

Glass is a non-crystalline, often transparent amorphous solid, that has widespread practical, technological, and decorative use in, for example, window panes, tableware, optics, and optoelectronics. The most familiar, and historically the oldest, types of manufactured glass are "silicate glasses" based on the chemical compound silica (silicon dioxide, or quartz), the primary constituent of sand. The term *glass*, in popular usage, is often used to refer only to this type of material, which is familiar from use as window glass and glass bottles. Of the many silica-based glasses that exist, ordinary glazing and container glass is formed from a specific type called soda-lime glass, composed of approximately 75% silicon dioxide (SiO_2), sodium oxide (Na_2O) from sodium carbonate (Na_2CO_3), calcium oxide (CaO), also called lime, and several minor additives.

Table 6.2.2 list the more common types of silicate glasses and their ingredients, properties, and applications. Figure 6.2.3 are examples of silicate glasses.

Table 6.2.2 Common Types of Silicate Glasses.

Type and Properties	Key Ingredients	Applications
Fused quartz has very low thermal expansion and excellent resistance to thermal shock, being able to survive immersion in water while red hot, resists high temperatures (1000–1500 °C) and chemical weathering, and is very hard.	Chemically pure silica (silicon dioxide)	Used for high-temperature applications such as furnace tubes, lighting tubes, melting crucibles, etc.
Soda-lime-silica glass is transparent, easily formed, and most suitable for window glass and tableware.[68] However, it has a high thermal expansion and poor resistance to heat.	(Na_2O) + lime (CaO) + magnesia (MgO) + alumina (Al_2O_3) account for over 75% of manufactured glass, containing about 70 to 74% silica by weight	Typically used for windows, bottles, light bulbs, and jars.[
Sodium borosilicate glasses have fairly low coefficients of thermal expansion (7740 Pyrex CTE is 3.25×10^{-6}/°C[69] as compared to about 9×10^{-6}/°C for a typical soda-lime glass[70]). They are, therefore, less subject to stress caused by thermal expansion and thus less vulnerable to cracking from thermal shock.	5–13% boron trioxide (B_2O_3)	Commonly used for e.g. labware (e.g. Pyrex, Duran), household cookware, and sealed beam car head lamps.[

Lead-oxide glass, crystal glass, lead glass has high density which results in a high electron density, and hence high refractive index, making the look of glassware more brilliant and causing noticeably more specular reflection and increased optical dispersion.[61][72] Lead glass has a high elasticity, making the glassware more workable and giving rise to a clear "ring" sound when struck.	silica + lead oxide (PbO) + potassium oxide (K_2O) + soda (Na_2O) + zinc oxide (ZnO) + alumina	Used for tableware, art objects, optical glass.
Aluminosilicate glass tends to be more difficult to melt and shape compared to borosilicate compositions, but has excellent thermal resistance and durability.	alumina and silica	Extensively used for fiberglass, used for making glass-reinforced plastics (boats, fishing rods, etc.), top-of-stove cookware, and halogen bulb glass.
Germanium-oxide glass is an extremely clear glass. Light loses only 5% of its intensity through 1 km of glass fiber.	alumina + germanium dioxide (GeO_2).	Used for fiber-optic waveguides in communication networks.
Colored glass is clear glass with oxides added .	Different oxide additives produce the different colors in glass: turquoise (copper(II) oxide),[121] purple (manganese dioxide), red (cadmium sulfide), blue (cobalt oxide) and green (Iron(II) oxide and chromium(III) oxide).	Used for tableware and decorative glassware.






a. b. c.

Figure 6.2.3 *Examples of silicate glasses: a. Pyrex measuring cup, b. red glass bottle with yellow glass overlay, and c. four-color Roman glass bowl, manufactured circa 1st century B.C..*

Asbestos

Asbestos is the name applied to six naturally occurring minerals that are mined from the earth. The different types of asbestos are:

- Amosite

- Chrysotile
- Tremolite
- ActinoliteUnlink
- Anthophyllite
- Crocidolite

Of these six, three are used more commonly. **Chrysotile** (white) is the most common, but it is not unusual to encounter**. Amosite** (brown / off-white), or **Crocidolite** (blue) as well. Asbestos are noncombustable fibrous material, and they have been used for terminal insulation material, brake linings, construction material, and filters. When mixed with cement, it reinforce the mechanical strength of concrete. It decomposes due to loss of water, and forms forsterite and silica at high temperature.

Ceramics

A **ceramic** is a solid material comprising an inorganic compound of metal, non-metal or ionic and covalent bonds. Common examples are earthenware, porcelain, and brick.

Varying crystallinity and electron composition in the ionic and covalent bonds cause most ceramic materials to be good thermal and electrical insulators (extensively researched in ceramic engineering). With such a large range of possible options for the composition/structure of a ceramic (e.g. nearly all of the elements, nearly all types of bonding, and all levels of crystallinity), the breadth of the subject is vast, and identifiable attributes (e.g. hardness, toughness, electrical conductivity, etc.) are difficult to specify for the group as a whole. General properties such as high melting temperature, high hardness, poor conductivity, high moduli of elasticity, chemical resistance and low ductility are the norm,[1] with known exceptions to each of these rules (e.g. piezoelectric ceramics, glass transition temperature, superconductive ceramics, etc.).

Cement and Concrete

A **cement** is a binder, a substance used for construction that sets, hardens, and adheres to other materials to bind them together. Cement is seldom used on its own, but rather to bind sand and gravel (aggregate) together. Cement mixed with fine aggregate produces mortar for masonry, or with sand and gravel, produces concrete. Cement is the most widely used material in existence and is only behind water as the planet's most-consumed resource.[1]

Cements used in construction are usually inorganic, often lime or calcium silicate based.

Concrete is a composite material composed of fine and coarse aggregate bonded together with a fluid cement (cement paste) that hardens over time—most frequently in the past a lime-based cement binder, such as lime putty, but sometimes with other hydraulic cements, such as a calcium aluminate cement or with Portland cement to form **Portland cement concrete** (for its visual resemblance to Portland stone).[2][3] Many other non-cementitious types of concrete exist with different methods of binding aggregate together, including asphalt concrete with a bitumen binder, which is frequently used for road surfaces, and polymer concretes that use polymers as a binder.

Summary

- The silicates are the largest, the most interesting and the most complicated class of minerals than any other minerals. Approximately 90% of the Earth's crust is made up of silicates, SiO_4^{4-} based material.
- Silicates are extremely important materials, both natural (such as granite, gravel, and garnet) and artificial (such as Portland cement, ceramics, glass, and waterglass), for all sorts of technological and artistic activities

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- Libretext: Inorganic Chemistry
- Wikipedia

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6.3: Metals and Ores

Learning Objectives

- Identify important metals and describe their extraction from their main ores.
- List different metals, their uses, and their alloys.
- Describe the environmental impact of metal production.

Most metals are found as types of rock in the Earth's crust. These ores contain sufficient minerals with important elements including metals that can be economically extracted from the rock. Metal ores are generally oxides, sulfides, silicates (Table 6.3.1) or "native" metals (such as native copper) that are not commonly concentrated in the Earth's crust, or "noble" metals (not usually forming compounds) such as gold (Figure 6.3.1). The ores must be processed to extract the metals of interest from the waste rock and from the ore minerals.



Figure 6.3.1: Three Common Ores. (left) Iron ore, (middle) Manganese ore – psilomelane, and (right) Lead ore – galena and anglesite.

Alloys

An **alloy** is a mixture composed of two or more elements, at least one of which is a metal. You are probably familiar with some alloys of copper (such as brass and bronze) and iron (steel) . Alloys can be one of two general types. In one type, called a **substitutional alloy**, the various atoms simply replace each other in the crystal structure. In another type, called an **interstitial alloy**, the smaller atoms such as carbon fit in between the larger atoms in the crystal packing arrangement.

Steels are a very important class of alloys. The many types of steels are primarily composed of iron, with various amounts of the elements carbon, chromium, manganese, nickel, molybdenum, and boron. Steels are widely used in building construction because of their strength, hardness, and resistance to corrosion. Most large modern structures like skyscrapers and stadiums are supported by a steel skeleton (see figure below).



Figure 6.3.2: The Willis Tower (formerly called the Sears Tower) in Chicago was once the tallest building in the world and is still the tallest in the Western Hemisphere. The use of steel columns makes it possible to build taller, stronger, and lighter buildings.

Copper,Brass, and Bronze

Copper is a chemical element with the symbol **Cu** (from Latin:) and atomic number 29. It is a soft, malleable, and ductile metal with very high thermal and electrical conductivity. A freshly exposed surface of pure copper has a pinkish-orange color. Copper is used as a conductor of heat and electricity and as a building material.

Copper is one of the few metals that can occur in nature in a directly usable metallic form (native metals). This led to very early human use in several regions, from c. 8000 BC. Thousands of years later, it was the first metal to be smelted from sulfide ores, c. 5000 BC, the first metal to be cast into a shape in a mold, c. 4000 BC and the first metal to be purposefully alloyed with another metal, tin, to create bronze, c. 3500 BC.

Most commercial ores are sulfides, especially chalcopyrite ($CuFeS_2$), bornite (Cu_5FeS_4) and, to a lesser extent, covellite (CuS) and chalcocite (Cu_2S).

Cu_2S, is converted into oxides:

$$2\ Cu_2S + 3\ O_2 \rightarrow 2\ Cu_2O + 2\ SO_2$$

The cuprous oxide is then converted to copper upon heating:

$$2\ Cu_2O \rightarrow 4\ Cu + O_2$$

Copper is a constituent of various metal alloys, such as sterling silver used in jewelry, cupronickel used to make marine hardware and coins, and constantan used in strain gauges and thermocouples for temperature measurement. Bronze, an alloy of copper and tin has been in use since ancient times. The Bronze Age saw the increased use of metals rather than stone for weapons, tools, and decorative objects. Brass, an alloy of copper and zinc, is widely used in musical instruments like the trumpet and trombone. Alloys are commonly used in manufactured items because the properties of these metal mixtures are often superior to a pure metal. Bronze is harder than copper and more easily cast. Brass is very malleable and its acoustic properties make it useful for musical instruments (Figure 6.3.1).



Figure 6.3.3 A bronze helmet and a trumpet made of brass. Copper alloys are widely used in the production of coinage; seen here are two examples - post-1964 American dimes, which are composed of the alloy cupronickel and a pre-1968 Canadian dime, which is composed of an alloy of 80 percent silver and 20 percent copper. (Source: Wikipedia)

Iron and Steel

The early application of iron to the manufacture of tools and weapons was possible because of the wide distribution of iron ores and the ease with which iron compounds in the ores could be reduced by carbon. Iron ore is reduced with coke in a blast furnace (Figure 6.3.1). The blast furnace is loaded with iron ores, usually hematite Fe_2O_3 or magnetite Fe_3O_4, together with coke (coal that has been separately baked to remove volatile components). Air pre-heated to 900 °C is blown through the mixture, in sufficient amount to turn the carbon into carbon monoxide:

$$2\ C + O_2 \rightarrow 2\ CO$$

This reaction raises the temperature to about 2000 °C The carbon monoxide reduces the iron ore to metallic iron[112]

$$Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$$

Some iron in the high-temperature lower region of the furnace reacts directly with the coke:[112]

$$2\ Fe_2O_3 + 3\ C \rightarrow 4\ Fe + 3\ CO_2$$

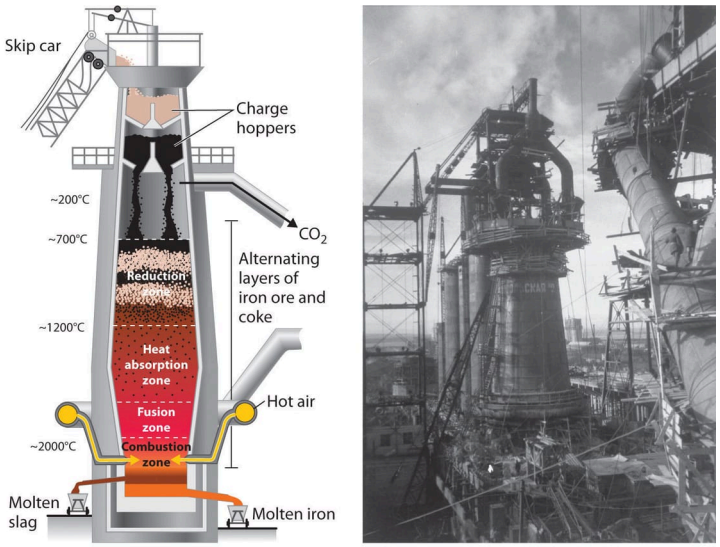


Figure 6.3.4: A Blast Furnace for Converting Iron Oxides to Iron Metal. (a) The furnace is charged with alternating layers of iron ore (largely Fe2O3) and a mixture of coke (C) and limestone (CaCO3). (b) This blast furnace in Magnitogorsk, Russia, was the largest in the world when it was built in 1931.

Much of the iron produced is refined and converted into steel. Steel is made from iron by removing impurities and adding substances such as manganese, chromium, nickel, tungsten, molybdenum, and vanadium to produce alloys with properties that make the material suitable for specific uses. Most steels also contain small but definite percentages of carbon (0.04%–2.5%). However, a large part of the carbon contained in iron must be removed in the manufacture of steel; otherwise, the excess carbon would make the iron brittle. However, there is not just one substance called steel - they are a family of alloys of iron with carbon or various metals.

Impurities in the iron from the Blast Furnace include carbon, sulfur, phosphorus and silicon, which have to be removed.

- **Removal of sulfur**: Sulfur has to be removed first in a separate process. Magnesium powder is blown through the molten iron and the sulfur reacts with it to form magnesium sulfide. This forms a slag on top of the iron and can be removed.

$$Mg + S \rightarrow MgS \tag{6.3.1}$$

- **Removal of carbon:** The still impure molten iron is mixed with scrap iron (from recycling) and oxygen is blown on to the mixture. The oxygen reacts with the remaining impurities to form various oxides. The carbon forms carbon monoxide. Since this is a gas it removes itself from the iron! This carbon monoxide can be cleaned and used as a fuel gas.
- **Removal of other elements:** Elements like phosphorus and silicon react with the oxygen to form acidic oxides. These are removed using quicklime (calcium oxide) which is added to the furnace during the oxygen blow. They react to form compounds such as calcium silicate or calcium phosphate which form a slag on top of the iron.

Cast iron has already been mentioned above. This section deals with the types of iron and steel which are produced as a result of the steel-making process.

- Wrought iron: If all the carbon is removed from the iron to give high purity iron, it is known as wrought iron. Wrought iron is quite soft and easily worked and has little structural strength. It was once used to make decorative gates and railings, but these days mild steel is normally used instead.
- Mild steel: Mild steel is iron containing up to about 0.25% of carbon. The presence of the carbon makes the steel stronger and harder than pure iron. The higher the percentage of carbon, the harder the steel becomes. Mild steel is used for lots of things - nails, wire, car bodies, ship building, girders and bridges amongst others.
- High carbon steel: High carbon steel contains up to about 1.5% of carbon. The presence of the extra carbon makes it very hard, but it also makes it more brittle. High carbon steel is used for cutting tools and masonry nails (nails designed to be driven into concrete blocks or brickwork without bending). High carbon steel tends to fracture rather than bend if mistreated.
- Special steels: These are iron alloyed with other metals (Table 6.3.1).

Table 6.3.1 *Special Steels*

	Iron mixed with	Special properties	Uses include
stainless steel	chromium and nickel	resists corrosion	cutlery, cooking utensils, kitchen sinks, industrial equipment for food and drink processing
titanium steel	titanium	withstands high temperatures	gas turbines, spacecraft
manganese steel	manganese	very hard	rock-breaking machinery, some railway track (e.g. points), military helmets

Steelmaking Animation

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 6.3.1 The process of steelmaking.

Aluminum

Aluminum is too high in the electrochemical series (reactivity series) to extract it from its ore using carbon reduction. The temperatures needed are too high to be economic. Instead, it is extracted by electrolysis. The ore is first converted into pure aluminum oxide by the Bayer Process, and this is then electrolyzed in solution in molten cryolite - another aluminum compound. The aluminum oxide has too high a melting point to electrolyse on its own. The usual aluminum ore is bauxite. Bauxite is essentially an impure aluminum oxide. The major impurities include iron oxides, silicon dioxide and titanium dioxide.

Crushed bauxite is treated with moderately concentrated sodium hydroxide solution. The concentration, temperature and pressure used depend on the source of the bauxite and exactly what form of aluminum oxide it contains. Temperatures are typically from 140°C to 240°C; pressures can be up to about 35 atmospheres. With hot concentrated sodium hydroxide solution, aluminum oxide reacts to give a solution of sodium tetrahydroxoaluminate.

$$Al_2O_3 + 2NaOH + 3H_2O \longrightarrow 2NaAl(OH)_4$$

The sodium tetrahydroxoaluminate solution is cooled, and "seeded" with some previously produced aluminum hydroxide. This provides something for the new aluminum hydroxide to precipitate around.

$$NaAl(OH)_4 \longrightarrow Al(OH)_3 + NaOH$$

Aluminum oxide (sometimes known as alumina) is made by heating the aluminum hydroxide to a temperature of about 1100 - 1200°C.

$$2Al(OH)_3 \longrightarrow Al_2O_3 + 3H_2O$$

The aluminum oxide is electrolyzed in solution in molten cryolite, Na_3AlF_6. Cryolite is another aluminum ore, but is rare and expensive, and most is now made chemically.

Uses of Aluminum

Aluminum is usually alloyed with other elements such as silicon, copper or magnesium. Pure aluminum isn't very strong, and alloying it adds to it strength. Aluminum is especially useful because it

- has a low density;

- is strong when alloyed;
- is a good conductor of electricity;
- has a good appearance;
- resists corrosion because of the strong thin layer of aluminum oxide on its surface. This layer can be strengthened further by anodizing the aluminum.

Anodizing essentially involves etching the aluminum with sodium hydroxide solution to remove the existing oxide layer, and then making the aluminum article the anode in an electrolysis of dilute sulfuric acid. The oxygen given of at the anode reacts with the aluminum surface, to build up a film of oxide up to about 0.02 mm thick. As well as increasing the corrosion resistance of the aluminum, this film is porous at this stage and will also take up dyes. (It is further treated to make it completely non-porous afterwards.) That means that you can make aluminum articles with the colour built into the surface.

Some uses include:

Aluminum is used for	because
aircraft	light, strong, resists corrosion
other transport such as ships' superstructures, container vehicle bodies, tube trains (metro trains)	light, strong, resists corrosion
overhead power cables (with a steel core to strengthen them)	light, resists corrosion, good conductor of electricity
saucepans	light, resists corrosion, good appearance, good conductor of heat

Recycling

Aluminum is an infinitely recyclable material, and it takes up to 95 percent less energy to recycle it than to produce primary aluminum, which also limits emissions, including greenhouse gases. Today, about 75 percent of all aluminum produced in history, nearly a billion tons, is still in use.[6]

The recycling of aluminum generally produces significant cost savings over the production of new aluminum, even when the cost of collection, separation and recycling are taken into account.[7] Over the long term, even larger national savings are made when the reduction in the capital costs associated with landfills, mines, and international shipping of raw aluminum are considered.

Environmental Impact of Steel and Aluminum Production

The impact of steel and aluminum production on the environment can be traced back from the mining of the ores to the production of the final commercial products (i.e. steel and aluminum). The main sources of emissions during the different phases of manufacture include the products of combustion such as nitrous oxide, carbon dioxide, carbon monoxide, and sulfur dioxide and fugitive dust from the operation of equipment. The effect of the different emissions on air quality (i.e. smog formation, greenhouse effect, acid rain etc.) will be discussed in more detail in Chapter 13.

Sulfuric acid is created when water and oxygen interact with sulfur bearing minerals and chemicals in rocks. Many metals become mobile as water becomes more acidic and at high concentrations these metals become toxic to most life forms. There is also production of enormous amounts of wastewater contaminants, hazardous wastes, and solid wastes.

Summary

- Metal ores contain sufficient minerals with important elements including metals that can be economically extracted from the rock. The ores must be processed to extract the metals of interest from the waste rock and from the ore minerals.
- Alloys are mixtures of materials, at least one of which is a metal.
- Bronze alloys were widely used in weapons.
- Brass alloys have long been employed in musical instruments.
- Steel alloys are strong and durable.

- Aluminum alloys are widely used due to its durability, resistance to corrosion, and good conductivity.

Contributors and Attributions

-
- TextMap: General Chemistry: General Chemistry (Petrucci et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Wikipedia

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6.4: Earth's Dwindling Resources

Learning Objectives

- Describe the composition and classification of solid waste.
- List the 3R's of Garbage and its benefits.

Land Pollution: Solid Wastes

Municipal solid waste (**MSW**), commonly known as **trash** or **garbage** in the United States and **rubbish** in Britain, is a waste type consisting of everyday items that are discarded by the public. "Garbage" can also refer specifically to food waste, as in a garbage disposal; the two are sometimes collected separately. In the European Union, the semantic definition is 'mixed municipal waste,' given waste code 20 03 01 in the European Waste Catalog. Although the waste may originate from a number of sources that has nothing to do with a municipality, the traditional role of municipalities in collecting and managing these kinds of waste have produced the particular etymology 'municipal.'

The composition of municipal solid waste varies greatly from municipality to municipality,[1] and it changes significantly with time. In municipalities which have a well developed waste recycling system, the waste stream mainly consists of intractable wastes such as plastic film and non-recyclable packaging materials. At the start of the 20th century, the majority of domestic waste (53%) in the UK consisted of coal ash from open fires.[2] In developed areas without significant recycling activity it predominantly includes food wastes, market wastes, yard wastes, plastic containers and product packaging materials, and other miscellaneous solid wastes from residential, commercial, institutional, and industrial sources.[3] Most definitions of municipal solid waste do not include industrial wastes, agricultural wastes, medical waste, radioactive waste or sewage sludge.[4] Waste collection is performed by the municipality within a given area. The term *residual waste* relates to waste left from household sources containing materials that have not been separated out or sent for processing.[5] Waste can be classified in several ways but the following list represents a typical classification:

- Biodegradable waste: food and kitchen waste, green waste, paper (most can be recycled, although some difficult to compost plant material may be excluded[6])
- Recyclable materials: paper, cardboard, glass, bottles, jars, tin cans, aluminum cans, aluminium foil, metals, certain plastics, textiles, clothing, tires, batteries, etc.
- Inert waste: construction and demolition waste, dirt, rocks, debris
- Electrical and electronic waste (WEEE) - electrical appliances, light bulbs, washing machines, TVs, computers, screens, mobile phones, alarm clocks, watches, etc.
- Composite wastes: waste clothing, Tetra Pack food and drink cartons, waste plastics such as toys and plastic garden furniture
- Hazardous waste including most paints, chemicals, tires, batteries, light bulbs, electrical appliances, fluorescent lamps, aerosol spray cans, and fertilizers
- Toxic waste including pesticides, herbicides, and fungicides
- Biomedical waste, expired pharmaceutical drugs, etc.

For example, typical municipal solid waste in China is composed of 55.9% food residue, 8.5% paper, 11.2% plastics, 3.2% textiles, 2.9% wood waste, 0.8% rubber, and 18.4% non-combustibles. The breakdown of municipal waste by material is shown below (Figure 6.4.1).

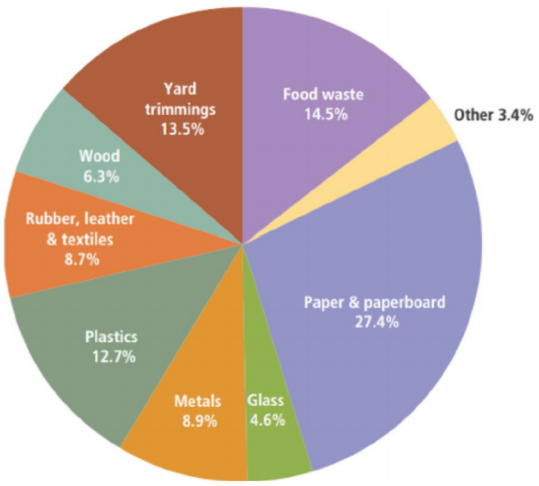


Figure 6.4.1 Total municipal waste generation by material. Source: US EPA

The Three R's of Garbage: Reduce, Reuse and Recyle

Reducing and Reusing Basics

The most effective way to reduce waste is to not create it in the first place. Making a new product requires a lot of materials and energy - raw materials must be extracted from the earth, and the product must be fabricated then transported to wherever it will be sold. As a result, reduction and reuse are the most effective ways you can save natural resources, protect the environment and save money.

Benefits of Reducing and Reusing

- Prevents pollution caused by reducing the need to harvest new raw materials
- Saves energy
- Reduces greenhouse gas emissions that contribute to global climate change
- Helps sustain the environment for future generations
- Saves money
- Reduces the amount of waste that will need to be recycled or sent to landfills and incinerators
- Allows products to be used to their fullest extent

Ideas on How to Reduce and Reuse

- Buy used. You can find everything from clothes to building materials at specialized reuse centers and consignment shops. Often, used items are less expensive and just as good as new.
- Look for products that use less packaging. When manufacturers make their products with less packaging, they use less raw material. This reduces waste and costs. These extra savings can be passed along to the consumer. Buying in bulk, for example, can reduce packaging and save money.
- Buy reusable over disposable items. Look for items that can be reused; the little things can add up. For example, you can bring your own silverware and cup to work, rather than using disposable items.
- Maintain and repair products, like clothing, tires and appliances, so that they won't have to be thrown out and replaced as frequently.
- Borrow, rent or share items that are used infrequently, like party decorations, tools or furniture.

Recycling Basics

Recycling is the process of collecting and processing materials that would otherwise be thrown away as trash and turning them into new products. Recycling can benefit your community and the environment.

Benefits of Recycling

- Reduces the amount of waste sent to landfills and incinerators
- Conserves natural resources such as timber, water and minerals
- Increases economic security by tapping a domestic source of materials
- Prevents pollution by reducing the need to collect new raw materials
- Saves energy
- Supports American manufacturing and conserves valuable resources
- Helps create jobs in the recycling and manufacturing industries in the United States

Recycling Creates Jobs

EPA released significant findings on the economic benefits of the recycling industry with an update to the national Recycling Economic Information (REI) Study in 2016. This study analyzes the numbers of jobs, wages and tax revenues attributed to recycling. The study found that in a single year, recycling and reuse activities in the United States accounted for:

- 757,000 jobs
- $36.6 billion in wages; and
- $6.7 billion in tax revenues.

This equates to 1.57 jobs, $76,000 in wages, and $14,101 in tax revenues for every 1,000 tons of materials recycled.

Summary

- Municipal solid waste (MSW), (i.e. trash or garbage in the United States and rubbish in Britain), is a type of waste consisting of everyday items that are discarded by the public.
- Most definitions of municipal solid waste do not include industrial wastes, agricultural wastes, medical waste, radioactive waste or sewage sludge.
- Typical classification of waste includes biodegradable wastes, recyclable materials, inert waste, electrical and electronic waste, composite wastes, hazardous wastes, toxic waste, and biomedical wastes.
- Reducing, reusing, and recycling solid wastes reduces the amount of waste in landfills and incinerators and promotes safer and cleaner environment.

Contributors and Attributions

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US EPA (Environmental Protection Agency)

Wikipedia

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7: AIR

7.1: Earth's Atmosphere- Divisions and Composition

7.2: Chemistry of the Atmosphere

7.3: Outdoor Air Pollution

7.4: Automobile Emissions

7.5: Photochemical Smog- Making Haze While the Sun Shines

7.6: Acid Rain- Air Pollution Water Pollution

7.7: Indoor Air Pollution

7.8: Stratospheric Ozone- Earth's Vital Shield

7.9: Carbon Dioxide and Climate Change

7.10: Who Pollutes? Who Pays?

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7.1: Earth's Atmosphere- Divisions and Composition

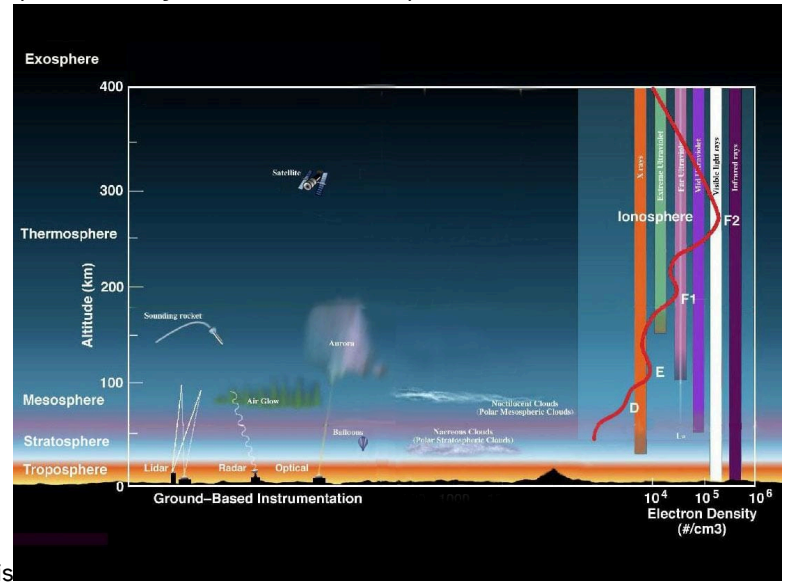
Learning Objectives

- Describe the different layers of the atmosphere.
- Know the composition of air.

The Earth's atmosphere is composed of several layers (Figure 7.1.1). The lowest layer, the troposphere, extends from the Earth's surface up to about 6 miles or 10 kilometers (km) in altitude. Virtually all human activities occur in the troposphere. Mt. Everest, the tallest mountain on the planet, is only about 5.6 miles (9 km) high. The next layer, the stratosphere, continues from 6 miles (10 km) to about 31 miles (50 km). Most commercial airplanes fly in the lower part of the stratosphere.

Most atmospheric ozone is concentrated in a layer in the stratosphere, about 9 to 18 miles (15 to 30 km) above the Earth's surface (see the figure below). Ozone is a molecule that contains three oxygen atoms. At any given time, ozone molecules are constantly formed and destroyed in the stratosphere. The total amount has remained relatively stable during the decades that it has been measured.

The ozone layer in the stratosphere absorbs a portion of the radiation from the sun, preventing it from reaching the planet's surface. Most importantly, it absorbs the portion of UV light called UVB. UVB has been linked to many harmful effects, including skin cancers, cataracts, xand harm to some crops and marine life.

The mesosphere starts just above the stratosphere and extends to 85 kilometers (53 miles) high. Meteors burn up in this layer
The thermosphere starts just above the mesosphere and extends to 600 kilometers (372 miles) high. Aurora and satellites occur in this 

layer. The ionosphere is an abundant layer of electrons and ionized atoms and molecules that stretches from about 48 kilometers (30 miles) above the surface to the edge of space at about 965 km (600 mi), overlapping into the mesosphere and thermosphere. This dynamic region grows and shrinks based on solar conditions and divides further into the sub-regions: D, E and F; based on what wavelength of solar vvradiation is absorbed. The ionosphere is a critical link in the chain of Sun-Earth interactions. This region is what makes radio communications possible. The exosphere is is the upper limit of the atmosphere. It extends from the top of the thermosphere up to 10,000 km (6,200 mi).

Figure 7.1.1 *The Earth's atmosphere* ***Credit:*** *NASA/Goddard*

Composition of the Atmosphere

Except for water vapor, whose atmospheric abundance varies from practically zero up to 4%, the fractions of the major

atmospheric components N_2, O_2, and Ar are remarkably uniform below about 100 km (Table 7.1.1). At greater heights, diffusion becomes the principal transport process, and the lighter gases become relatively more abundant. In addition, photochemical processes result in the formation of new species whose high reactivities would preclude their existence in significant concentrations at the higher pressures found at lower elevations.

The atmospheric gases fall into three abundance categories: major, minor, and trace. Nitrogen, the most abundant component, has accumulated over time as a result of its geochemical inertness; a very small fraction of it passes into the other phases as a result of biological activity and natural fixation by lightning. It is believed that denitrifying bacteria in marine sediments may provide the major route for the return of N_2 to the atmosphere. Oxygen is almost entirely of biological origin, and cycles through the hydrosphere, the biosphere, and sedimentary rocks. Argon consists mainly of Ar^{40} which is a decay product of K^{40} in the mantle and crust.

Table 7.1.1 The Major Components of Clean Dry Air in the Lower Atmosphere

nitrogen	N_2	78.08 %
oxygen	O_2	20.95 %
argon	Ar	0.93 %

The most abundant of the minor gases aside from water vapor is carbon dioxide (Table 7.1.2). Next in abundance are neon and helium. Helium is a decay product of radioactive elements in the earth, but neon and the other inert gases are primordial, and have probably been present in their present relative abundances since the earth's formation. Two of the minor gases, ozone and carbon monoxide, have abundances that vary with time and location. A variable abundance implies an imbalance between the rates of formation and removal. In the case of carbon monoxide, whose major source is anthropogenic (a small amount is produced by biological action), the variance is probably due largely to localized differences in fuel consumption, particularly in internal combustion engines. The nature of the carbon monoxide sink (removal mechanism) is not entirely clear; it may be partly microbial.

Table 7.1.2 The Minor Components of Air in the Lower Atmosphere.

water	H_2O	0-4 %
carbon dioxide	CO_2	325 ppm
neon	Ne	18 ppm
helium	He	5 ppm
methane	CH_4	2 ppm
krypton	Kr	1 ppm
hydrogen	H_2	0.5 ppm
nitrous oxide	N_2O	0.3 ppm
carbon monoxide	CO	0.05-0.2 ppm
ozone	O_3	0.02 - 10 ppm
xenon	Xe	0.08 ppm

Summary

- The layers of the atmosphere include the trophosphere, stratosphere, mesosphere, thermosphere, and exosphere.
- The major components of dry clean air are N_2, O_2, and Ar.
- The most abundant of the minor gases are water vapor and CO_2

Contributors and Attributions

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7.2: Chemistry of the Atmosphere

Learning Objectives

- Describe the nitrogen cycle.
- Describe the oxygen cycle.
- Describe the conditions for temperature inversion.

The mix of gases in the atmosphere forms a complex system organized into layers that together support life on Earth. Although there are numerous gases, as shown in Table 13.1.1, the top four gases make up 99.998 % of the volume of *clean dry air* (unpolluted air that does not contain water vapor). Of this dry composition of the atmosphere **nitrogen**, by far, is the most common (78%). Nitrogen dilutes oxygen and prevents rapid or instantaneous burning at the Earth's surface, as oxygen gas is a necessary reactant of the combustion process. Nitrogen is also needed and used by living things to make proteins, though as nitrogen gas, N_2, it is unavailable to most living things. **Oxygen** is used by all living things to make molecules that are essential for life. It is also essential for aerobic respiration as well as combustion or burning.

Nitrogen Cycle

All life requires nitrogen-compounds, e.g., proteins and nucleic acids. Air, which is 79% nitrogen gas (N_2), is the major reservoir of nitrogen. But most organisms cannot use nitrogen in this form. Figure 7.2.1 illustrates the entire nitrogen cycle. Plants must secure their nitrogen in "fixed" form, i.e., incorporated in compounds such as: nitrate ions (NO_3^-), ammonium ions (NH_4^+) and urea $(NH_2)_2CO$. Animals secure their nitrogen (and all other) compounds from plants (or animals that have fed on plants).

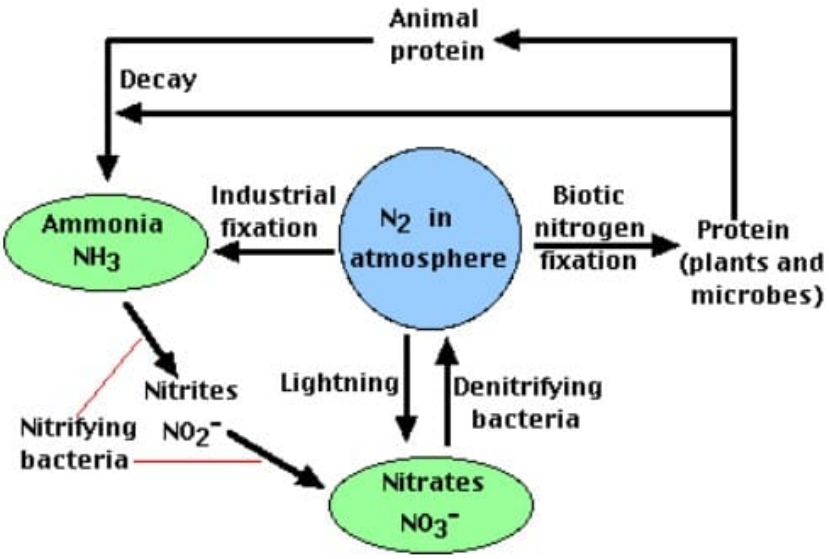


Figure 7.2.1 Nitrogen cycle

Four processes participate in the cycling of nitrogen through the biosphere: (1) nitrogen fixation, (2) decay, (3) nitrification, and (4) denitrification. Microorganisms play major roles in all four of these.

Nitrogen Fixation

The nitrogen molecule (N_2) is quite inert. To break it apart so that its atoms can combine with other atoms requires the input of substantial amounts of energy. Three processes are responsible for most of the nitrogen fixation in the biosphere:

- **atmospheric fixation** by lightning
- **biological fixation** by certain microbes alone or in a symbiotic relationship with some plants and animals
- **industrial fixation**

Atmospheric Fixation

The enormous energy of lightning breaks nitrogen molecules and enables their atoms to combine with oxygen in the air forming nitrogen oxides. These dissolve in rain, forming nitrates, that are carried to the earth. Atmospheric nitrogen fixation probably contributes some 5– 8% of the total nitrogen fixed.

Industrial Fixation

Under great pressure, at a temperature of 600°C, and with the use of a catalyst, atmospheric nitrogen and hydrogen (usually derived from natural gas or petroleum) can be combined to form ammonia (NH3). Ammonia can be used directly as fertilizer, but most of its is further processed to urea and ammonium nitrate (NH4NO3).

Biological Fixation

The ability to fix nitrogen in the soil (Figure $7.2.2$) only in certain bacteria and archaea.

- Some live in a symbiotic relationship with plants of the legume family (e.g., soybeans, alfalfa).
- Some establish symbiotic relationships with plants other than legumes (e.g., alders).
- Some establish symbiotic relationships with animals, e.g., termites and "shipworms" (wood-eating bivalves).
- Some nitrogen-fixing bacteria live free in the soil.
- Nitrogen-fixing cyanobacteria are essential to maintaining the fertility of semi-aquatic environments like rice paddies.

Biological nitrogen fixation requires a complex set of enzymes and a huge expenditure of ATP. Although the first stable product of the process is ammonia, this is quickly incorporated into protein and other organic nitrogen compounds.

Decay

The proteins made by plants enter and pass through food webs just as carbohydrates do. At each trophic level, their metabolism produces organic nitrogen compounds that return to the environment, chiefly in excretions. The final beneficiaries of these materials are microorganisms of decay. They break down the molecules in excretions and dead organisms into **ammonia**.

Nitrification

Ammonia can be taken up directly by plants — usually through their roots. However, most of the ammonia produced by decay is converted into **nitrates**. Until recently this was thought always to be accomplished in two steps:

1. Bacteria of the genus *Nitrosomonas* oxidize $\mathrm{NH_3}$ to **nitrites** ($\mathrm{NO_2^-}$).
2. Bacteria of the genus *Nitrobacter* oxidize the nitrites to **nitrates** ($\mathrm{NO_3^-}$).

These two groups of autotrophic bacteria are called **nitrifying bacteria**. Through their activities (which supply them with all their energy needs), nitrogen is made available to the roots of plants. However, in 2015, two groups reported finding that bacteria in the genus *Nitrospira* were able to carry out both steps: ammonia to nitrite and nitrite to nitrate. This ability is called "comammox" (for complete ammonia oxidation).

In addition, both soil and the ocean contain **archaeal** microbes, assigned to the Crenarchaeota, that convert ammonia to nitrites. They are more abundant than the nitrifying bacteria and may turn out to play an important role in the nitrogen cycle.

Many legumes, in addition to fixing atmospheric nitrogen, also perform nitrification - converting some of their organic nitrogen to nitrites and nitrates. These reach the soil when they shed their leaves.

Denitrification

The three processes above remove nitrogen from the atmosphere and pass it through ecosystems. Denitrification reduces nitrates and nitrites to nitrogen gas, thus replenishing the atmosphere. In the process several intermediates are formed:

- nitric oxide (NO)
- nitrous oxide (N_2O)(a greenhouse gas 300 times as potent as CO_2)
- nitrous acid (HONO)

Once again, bacteria are the agents. They live deep in soil and in aquatic sediments where conditions are anaerobic. They use nitrates as an alternative to oxygen for the final electron acceptor in their respiration.

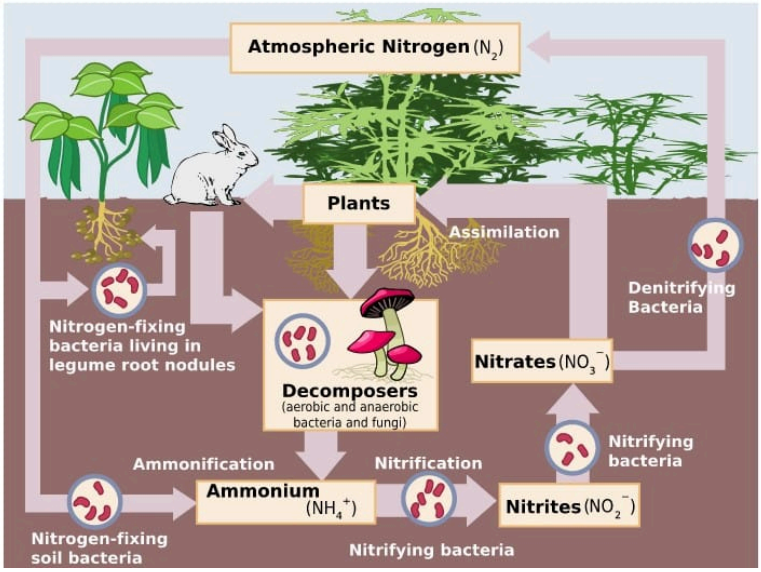


Figure 7.2.2 The role of soil bacteria in the Nitrogen cycle: Nitrogen transitions between various biologically useful forms.

Anammox (anaerobic ammonia oxidation)

Under anaerobic conditions in marine and freshwater sediments, other species of bacteria are able to oxidize ammonia (with NO_2^-) forming nitrogen gas.

$$NH_4^+ + NO_2^- \longrightarrow N_2 + 2\,H_2O$$

The anammox reaction may account for as much as 50% of the denitrification occurring in the oceans. All of these processes participate in closing the nitrogen cycle.

Are the denitrifiers keeping up?

Agriculture may now be responsible for one-half of the nitrogen fixation on earth through the use of fertilizers produced by industrial fixation and the the growing of legumes like soybeans and alfalfa. This is a remarkable influence on a natural cycle. Are the denitrifiers keeping up the nitrogen cycle in balance? Probably not. Certainly, there are examples of nitrogen enrichment in ecosystems. One troubling example: the "blooms" of algae in lakes and rivers as nitrogen fertilizers leach from the soil of adjacent farms (and lawns). The accumulation of dissolved nutrients in a body of water is called **eutrophication**.

Oxygen Cycle

Oxygen is the most abundant element on the earth's crust. The earth's surface is composed of the crust, atmosphere, and hydrosphere. About 50% of the mass of the earth's crust consists of oxygen (combined with other elements, principally silicon). Oxygen occurs as O_2 molecules and, to a limited extent, as O_3 (ozone) molecules in air. It forms about 20% of the mass of the

air. About 89% of water by mass consists of combined oxygen. In combination with carbon, hydrogen, and nitrogen, oxygen is a large part of plants and animals.

Oxygen is a colorless, odorless, and tasteless gas at ordinary temperatures. It is slightly denser than air. Although it is only slightly soluble in water (49 mL of gas dissolves in 1 L at STP), oxygen's solubility is very important to aquatic life.

Oxygen is essential in combustion processes such as the burning of fuels. Plants and animals use the oxygen from the air in respiration (*Figure* 7.2.4*)*. The main way free oxygen is lost from the atmosphere is via and , mechanisms in which life and consume oxygen and release carbon dioxide.

The respiration process is represented as:

$$6\,O_2{}^{+}C_6H_{12}O_6 \longrightarrow 6\,CO_2{}^{+}6H_2O$$

Green plants continually replenish the oxygen in the atmosphere by a process called photosynthesis (*Figure* 7.2.3*)* . The products of photosynthesis may vary, but, in general, the process converts carbon dioxide and water into glucose (a sugar) and oxygen using the energy of light:

$$6\,CO_2(g) + 6\,H_2O(l) \xrightarrow[\text{light}]{\text{chlorophyll}} C_6H_{12}O_6(aq) + 6\,O_2(g) \tag{7.2.1}$$

$$\text{carbon dioxide} \qquad \text{water} \qquad \text{glucose} \qquad \text{oxygen} \tag{7.2.2}$$

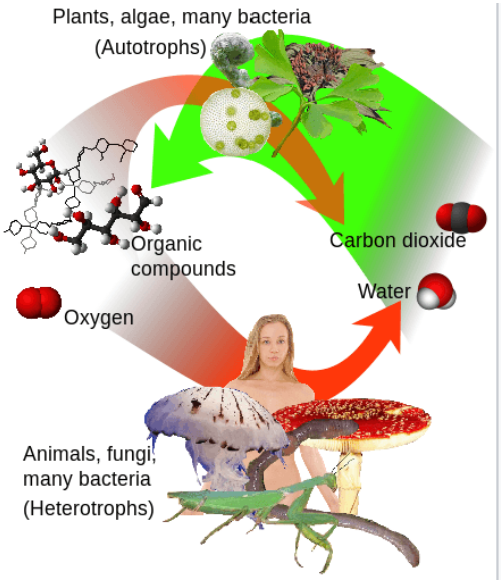


Figure 7.2.3

Overview of and photosynthesis (green) and respiration (red).

Water (at right), together with carbon dioxide (CO_2), form oxygen and organic compounds (at left),

which can be respired to water and (CO_2). Source: Wikipedia

Thus, the oxygen that became carbon dioxide and water by the metabolic processes in plants and animals returns to the atmosphere by photosynthesis. Photosynthesizing organisms include the plant life of the land areas as well as the of the oceans. The tiny marine was discovered in 1986 and accounts for more than half of the photosynthesis of the open ocean.

Oxygen is a key reactant in various oxidation reactions mentioned in section 8.5. Atmospheric free oxygen is also consumed by chemical weathering and surface reactions. An example of surface weathering is formation of rust:

$$\mathbf{4\,FeO + O_2 \longrightarrow 2\,Fe_2O_3}$$

Ozone forms naturally in the upper atmosphere by the action of ultraviolet light from the sun on the oxygen there. Most

atmospheric ozone occurs in the stratosphere, a layer of the atmosphere extending from about 10 to 50 kilometers above the earth's surface. This ozone acts as a barrier to harmful ultraviolet light from the sun by absorbing it via a chemical decomposition reaction:

$$\mathrm{O_3}(g) \xrightarrow{\text{ultraviolet light}} \mathrm{O}(g) + \mathrm{O_2}(g)$$

Temperature Inversions

In meteorology, an **inversion**, also known as a **temperature inversion**, is a deviation from the normal change of an amospheric property with altitude. It almost always refers to an inversion of the thermal lapse rate. Normally, air temperature decreases with an increase in altitude. During an inversion, warmer air is held above cooler air; the normal temperature profile with altitude is inverted.

An inversion traps air pollution, such as smog, close to the ground. An inversion can also suppress convection by acting as a "cap". If this cap is broken for any of several reasons, convection of any moisture present can then erupt into violent thunderstorms Temperature inversion can notoriously result in freezing rain in cold climates.

Summary

- The different forms of nitrogen that can be used in metabolism are produced through the process of nitrogen fixation.
- Bacteria in the soil carry out a process known as denitrification which converts nitrates back to nitrogen gas.
- Oxygen is produced mainly through photosynthesis. An additional source of atmospheric free oxygen comes from , whereby high-energy radiation breaks down atmospheric water and nitrous oxide.
- Oxygen is utilized during cellular respiration and decay. As well as during chemical weathering and various oxidation reactions.
- Temperature inversion is when a layer of cool air is trapped under a layer of warmer air. The cool dense air can trap and accumulate air pollutants.

Contributors and Attributions

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- Wikipedia

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7.3: Outdoor Air Pollution

Learning Objectives

- Define air pollution.
- Know the sources and effects of major outdoor pollutants

Air pollution refers to the introduction, into the atmosphere, of substances that have harmful effects on humans, other living organisms, and the environment either as solid particles, liquid droplets or gases. Air pollution can result from natural processes such as dust storms, forest fires, and volcanic eruptions, or from human activities such as biomass burning, vehicular emissions, mining, agriculture, and industrial processes. Improved technology and government policies have helped reduce most types of outdoor air pollution in many industrialized countries including the United States, in recent decades. However, outdoor air quality is still a problem in less industrialized nations, especially in megacities of rapidly industrializing nations such as China and India.

Outdoor pollutants can come from **stationary (point)** sources or **mobile (nonpoint)** sources (Figure 7.3.1). Stationary sources have a fixed location, for example power plant smokestacks, burning, construction sites, farmlands and surface mines among others. Mobile sources of air pollutants move from place to place while emitting pollutants. Examples of mobile sources include vehicles, aircrafts, ships, and trains.

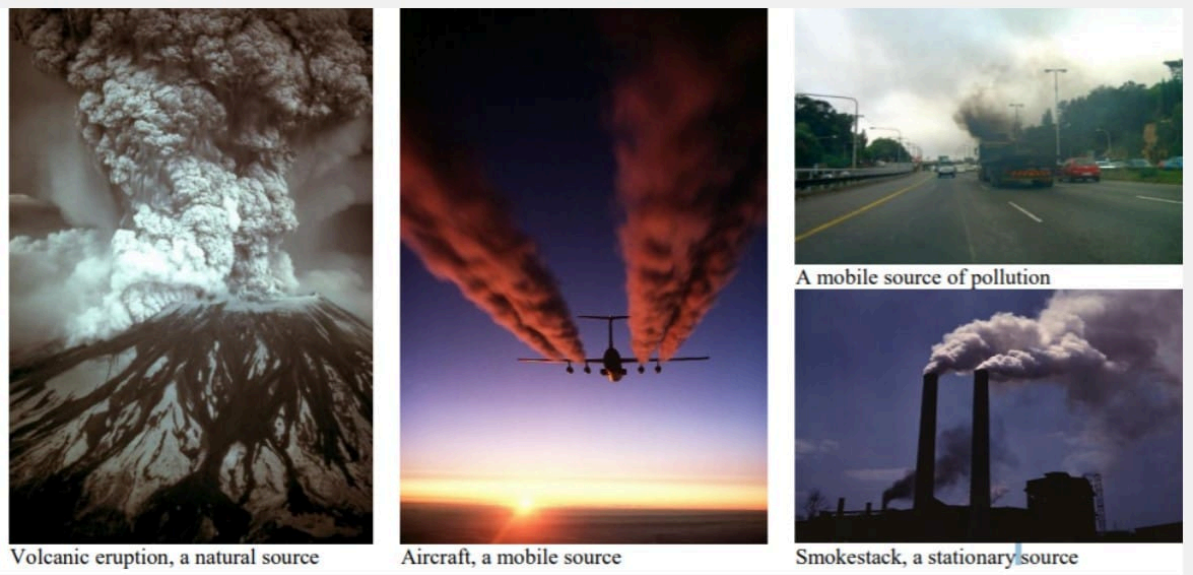


Figure 7.3.1: Images showing various sources of air pollution, including natural and anthropogenic, stationary and mobile. Source: All images obtained from Wiki Commons commons.wikimedia.org/wiki/Air_pollution

Pollutants are categorized into two major types based on how they originated namely primary and secondary pollutants. **Primary pollutants** are those released directly from the source into the air in a harmful form. The primary pollutants that account for nearly all air pollution problems are carbon monoxide (58%), volatile organic compounds (VOCs, 11%), nitrogen oxides (15%), sulfur dioxides (13%), and particulate material (3%). **Secondary pollutants** are produced through reactions between primary pollutants and normal atmospheric compounds. For example, ground-level ozone forms over urban areas through reactions, powered by sunlight, between primary pollutants (oxides of nitrogen) and other atmospheric gases such as VOCs.

Criteria pollutants

Under the Clean Air Act, the Environmental Protection Agency (EPA) establishes air quality standards to protect public health and the environment. EPA has set national air quality standards for six common air pollutants namely: 1) carbon monoxide; 2) ground-level ozone; 3) nitrogen dioxide; 4) Sulfur dioxide; 5) lead; and 6) particulate matter (also known as particle pollution). Of the six pollutants, particle pollution and ground-level ozone are the most widespread health threats. EPA calls these pollutants "criteria"

air pollutants because it regulates them by developing human health-based and/or environmentally-based criteria (science-based guidelines) for setting permissible levels. The set of limits based on human health is called primary standards. Another set of limits intended to prevent environmental and property damage is called secondary standards.

1. **Carbon Monoxide (CO)**: is a colorless, odorless gas emitted from combustion processes, specifically, the incomplete combustion of fuel. Nationally and, particularly in urban areas, the majority of CO emissions to ambient air come from mobile sources. CO can cause harmful health effects by reducing oxygen delivery to the body's organs (like the heart and brain) and tissues. At extremely high levels, CO can cause death.
2. **Ground-level ozone (O_3)**: is a colorless gas with a slightly sweet odor that is not emitted directly into the air, but is created by the interaction of sunlight, heat, oxides of nitrogen (NO_x) and volatile organic compounds (VOCs). Ozone is likely to reach unhealthy levels on hot sunny days in urban environments. Emissions from industrial facilities and electric utilities, motor vehicle exhaust, gasoline vapors, and chemical solvents are some of the major sources of NO_x and VOCs.
3. **Nitrogen dioxide (NO_2)**: is one of a group of highly reactive gasses known as "oxides of nitrogen," or "nitrogen oxides (NO_x)." Other nitrogen oxides include nitrous acid and nitric acid. NO_2 is a yellowish-brown to reddish-brown foul-smelling gas that is a major contributor to smog and acid rain. Nitrogen oxides result when atmospheric nitrogen and oxygen react at the high temperatures created by combustion engines. Most emissions in the U.S. result from combustion in vehicle engines, electrical utility, and industrial combustion.
4. **Sulfur dioxide (SO_2)**: Sulfur dioxide is one of a group of highly reactive gasses known as “oxides of sulfur.” The largest sources of SO_2 emissions are from fossil fuel combustion at power plants (73%) and other industrial facilities (20%). Smaller sources of SO_2 emissions include industrial processes such as extracting metals from their ores, and the burning of high sulfur containing fuels by locomotives, large ships, and non-road equipment.
5. **Lead (Pb)**: is a metal found naturally in the environment as well as in manufactured products. The major sources of lead emissions have historically been from fuels in motor vehicles (such as cars and trucks) and industrial sources. As a result of EPA's regulatory efforts to remove lead from gasoline, emissions of lead from the transportation sector dramatically declined by 95 percent between 1980 and 1999, and levels of lead in the air decreased by 94 percent during that time period. The major sources of lead emissions today are ore and metal processing and piston-engine aircraft operating on leaded aviation gasoline. Today, the highest levels of lead in air are usually found near lead smelters.
6. **Particulate material (PM)**, sometimes known simply as “**particulates**” refers to solid particles and liquid droplets suspended in the air we breathe. Particulate pollution is made up of a variety of components, including acids (nitrates and sulfates), organic chemicals, metals, soil or dust particles, and allergens (pollen and mold spores). The size of the particles in directly linked to their potential for causing health problems. Particles that are 10 micrometers in diameter or smaller generally pass through the throat and nose and enter the lungs. EPA groups these into two types: "**inhalable coarse particles**," with diameters larger than 2.5 micrometers and smaller than 10 micrometers and "**fine particles**," with diameters that are 2.5 micrometers and smaller. How small is 2.5 micrometers? Think about a single hair from your head. The average human hair is about 70 micrometers in diameter – making it 30 times larger than the largest fine particle (Figure 7.3.2). Our respiratory systems are equipped to filter larger particles out of the air once it is inhaled. However, the lungs are vulnerable to both coarse particles (PM_{10}), and fine particles ($PM_{2.5}$). These can slip past the respiratory system's natural defenses and get deep into the lungs and some may even get into the bloodstream. Coarse particles come from road dust while fine particles come from combustion processes.

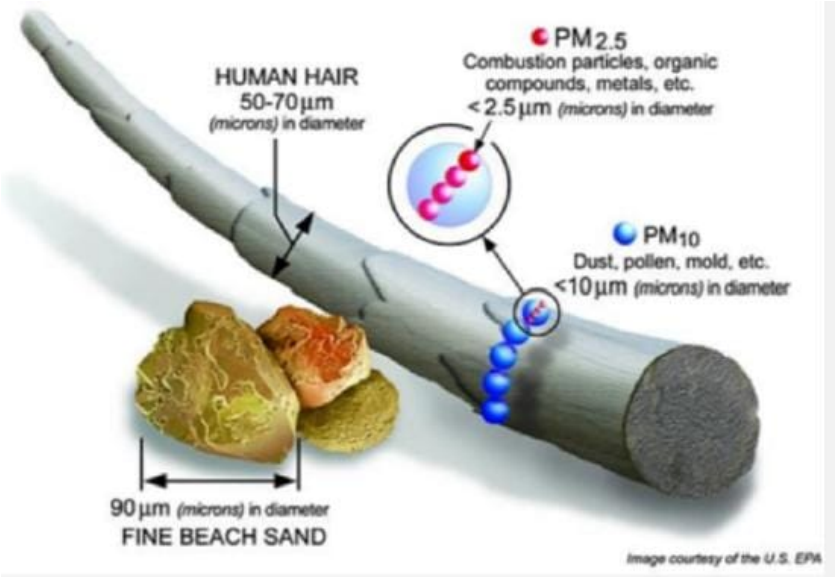


Figure 7.3.2 Graphic showing size comparison of particulate material (PM_{10} and $PM_{2.5}$) compared to fine beach sand and human hair. Source: US EPA http://www.epa.gov/air/particlepollution/basic.html

Volatile Organic Compounds

Volatile organic compounds (VOCs) are carbon-containing chemicals emitted as gases from natural and human-made sources. Natural sources of VOCs include plants, the largest source, and bacteria in the guts of termites and ruminant animals. These compounds are generally oxidized to carbon monoxide and carbon dioxide in the atmosphere. VOCs are of great concern because they are precursors for the formation of ozone, a secondary air pollutant.

A large number of synthetic organic chemicals such as benzene, toluene, formaldehyde, vinyl chloride, chloroform, and phenols are widely used as ingredients in countless household products. Paints, paint strippers, varnishes, many cleaning, disinfecting, cosmetic, degreasing, and hobby products all contain VOCs. Fuels are also made up of organic chemicals. All of these products can release organic compounds while you are using them, and, to some degree, when they are stored. The “new car smell” characteristic of new cars is from a complex mix of dozens of VOC. Also, concentrations of many VOCs are consistently higher indoors (up to ten times higher) than outdoors. They are often held responsible for sick building syndrome, an illness resulting from indoor pollution in which the specific cause is not identified.

Smog

Smog is a mixture of air pollutants (sulfur dioxide, nitrogen oxides, ozone, and particulates) that often form over urban areas as a result of fossil fuel combustion. The term was coined from the terms “smoke” and “fog” referring to a brownish haze that pollutes the air, greatly reducing visibility and making it difficult for some people to breathe (Figure 7.3.3 and 7.3.4). There are two main types of smog: industrial and photochemical smog. **Industrial smog** is produced primarily by the burning of fossil fuels which produces carbon dioxide (from complete combustion), carbon monoxides (from partial combustion), sulfur, and mercury. The sulfur reacts with other chemicals in the atmosphere producing several sulfur compounds including sulfur dioxide. These compounds along with particulate material make up industrial smog. **Photochemical smog** is formed when sunlight drives chemical reactions between primary pollutants from automobiles and normal atmospheric compounds. The product is a mix of over 100 different chemicals with the most abundant being ground-level ozone. This type of smog will be discussed in more detail in section 13.5.



Figure 7.3.3: Smog over Almaty city, Kazakhstan. Photo by Igors Jefimovs. Source: Wikicommons commons.wikimedia.org/wiki/C...dia/File:Smog_ over_Almaty.jpg



Figure 7.3.4: Smog over Santiago in Chile. Source: German Wikipedia commons.wikimedia.org/wiki/C...#/media/File:S antiago30std.jpg

Industrial Smog

Industrial smog or London-type smog is mainly a product of burning large amounts of high sulfur coal. Clean air laws passed in 1956 have greatly reduced smog formation in the United Kingdom; however, in other parts of the world London-type smog is still very prevalent. The main constituent of London-type smog is soot; however, these smogs also contain large quantities of fly ash, sulfur dioxide, sodium chloride and calcium sulfate particles. If concentrations are high enough, sulfur dioxide can react with atmospheric hydroxide to produce sulfuric acid, which will precipitate as acid rain.

$$SO_2 + OH \rightarrow HOSO_2 \tag{7.3.1}$$

$$HOSO_2 + O_2 \rightarrow HO_2 + SO_3 \tag{7.3.2}$$

$$SO_3 + H_2O \rightarrow H_2SO_4 \tag{7.3.3}$$

Concerns about the harmful effects of acid rain have led to strong pressure on industry to minimize the release of SO_2 and NO. For example, coal-burning power plants now use SO_2 "scrubbers," which trap SO_2 by its reaction with lime (CaO) to produce calcium sulfite dihydrate ($CaSO_3 \cdot 2H_2O$; Figure 7.3.5).

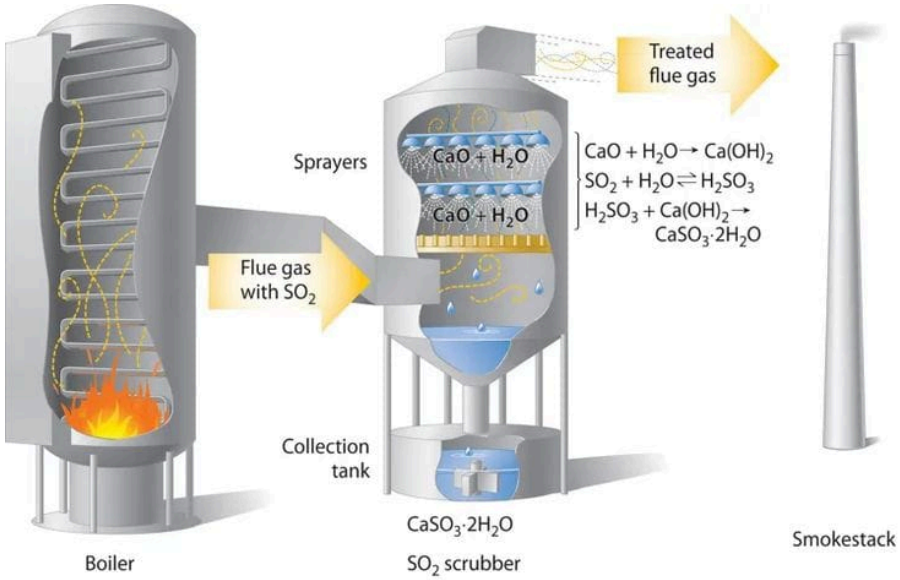


Figure 7.3.5: Schematic Diagram of a Wet Scrubber System. In coal-burning power plants, SO_2 can be removed ("scrubbed") from exhaust gases by its reaction with a lime (CaO) and water spray to produce calcium sulfite dihydrate ($CaSO_3 \cdot 2H_2O$). Removing SO_2 from the gases prevents its conversion to SO_3 and subsequent reaction with rainwater (acid rain). Scrubbing systems are now commonly used to minimize the environmental effects of large-scale fossil fuel combustion.

Toxic pollutants

Toxic air pollutants, also known as hazardous air pollutants, are those pollutants that are known or suspected to cause cancer or other serious health effects, such as reproductive effects or birth defects, or adverse environmental effects. Examples of toxic air pollutants include **benzene**, which is found in gasoline; **perchloroethylene,** which is emitted from some dry cleaning facilities; **methylene chloride**, which is used as a solvent and paint stripper by a number of industries; and others such as dioxin, asbestos, toluene, and metals such as cadmium, mercury, chromium, and lead compounds.

Most air toxics originate from human-made sources, including mobile sources (e.g., cars, trucks, buses) and stationary sources (e.g., factories, refineries, power plants), as well as indoor sources (e.g., some building materials and cleaning solvents). Some air toxics are also released from natural sources such as volcanic eruptions and forest fires. Exposure to air toxics is mainly through breathing but some toxic air pollutants such as mercury can deposit onto soils or surface waters, where they are taken up by plants and ingested by animals and are eventually magnified up through the food chain. Like humans, animals may experience health problems if exposed to sufficiently high quantities of air toxics over time.

Effects of Air Pollution on Human Health

The World Health Organization (WHO) and other international agencies recognize air pollution as a major threat to human health. Numerous scientific studies have linked air pollution to a variety of health problems (Table 7.3.1) including: aggravation of respiratory and cardiovascular diseases; decreased lung function; increased frequency and severity of respiratory symptoms such as difficulty breathing and coughing; increased susceptibility to respiratory infections; effects on the nervous system, including the brain, such as IQ loss and impacts on learning, memory, and behavior; cancer; and premature death. Immediate effects of air pollution may show up after a single exposure or repeated exposures. Other health effects may show up either years after exposure has occurred or only after long or repeated periods of exposure.

Immediate effects of air pollution include irritation of the eyes, nose, and throat, headaches, dizziness, and fatigue. Such immediate effects are usually short-term and treatable. Sometimes the treatment is simply eliminating the person's exposure to the source of the pollution, if it can be identified. Symptoms of some diseases, including asthma, hypersensitivity pneumonitis, and humidifier fever , may also show up soon after exposure to some indoor air pollutants.

Table 7.3.1: Sources and health effects of criteria pollutants

Pollutant	Sources	Health Effects
Ground-level Ozone (O_3)	Secondary pollutant typically formed by chemical reaction of volatile organic compounds (VOCs) and NO_x in the presence of sunlight.	Decreases lung function and causes respiratory symptoms, such as coughing and shortness of breath; aggravates asthma and other lung diseases leading to increased medication use, hospital admissions, emergency department (ED) visits, and premature mortality.
Particulate Matter (PM)	Emitted or formed through chemical reactions; fuel combustion (e.g., burning coal, wood, diesel); industrial processes; agriculture (plowing, field burning); and unpaved roads.	Short-term exposures can aggravate heart or lung diseases leading to respiratory symptoms, increased medication use, hospital admissions, ED visits, and premature mortality; long-term exposures can lead to the development of heart or lung disease and premature mortality.
Lead	Smelters (metal refineries) and other metal industries; combustion of leaded gasoline in piston engine aircraft; waste incinerators; and battery manufacturing.	Damages the developing nervous system, resulting in IQ loss and impacts on learning, memory, and behavior in children. Cardiovascular and renal effects in adults and early effects related to anemia.
Oxides of Nitrogen (NO_x)	Fuel combustion (e.g., electric utilities, industrial boilers, and vehicles) and wood burning.	Aggravate lung diseases leading to respiratory symptoms, hospital admissions, and ED visits; increased susceptibility to respiratory infection.
Carbon Monoxide (CO)	Fuel combustion (especially vehicles), industrial processes, fires, waste combustion, and residential wood burning.	Reduces the amount of oxygen reaching the body's organs and tissues; aggravates heart disease, resulting in chest pain and other symptoms leading to hospital admissions and ED visits.
Sulfur Dioxide (SO_2)	Fuel combustion (especially high-sulfur coal); electric utilities and industrial processes; and natural sources such as volcanoes.	Aggravates asthma and increased respiratory symptoms. Contributes to particle formation with associated health effects.

Source: www.epa.gov

The likelihood of immediate reactions to air pollutants depends on several factors. Age and preexisting medical conditions are two important influences. Some sensitive individuals appear to be at greater risk for air pollution-related health effects, for example, those with pre-existing heart and lung diseases (e.g., heart failure/ischemic heart disease, asthma, emphysema, and chronic bronchitis), diabetics, older adults, and children. In other cases, whether a person reacts to a pollutant depends on individual sensitivity, which varies tremendously from person to person. Some people can become sensitized to biological pollutants after repeated exposures, and it appears that some people can become sensitized to chemical pollutants as well.

Summary

- Air pollution refers to the introduction, into the atmosphere, of substances that have harmful effects on humans, other living organisms, and the environment either as solid particles, liquid droplets or gases.
- Major sources of air pollution are gas emissions from fossil fueled vehicles and their reaction products (CO, NO_x, Ozone) as well as particulate matter, SO_2, and VOCs.
- Examples of toxic air pollutants include benzene, which is found in gasoline; perchloroethylene, which is emitted from some dry cleaning facilities; methylene chloride, which is used as a solvent and paint stripper by a number of industries; and others such as dioxin, asbestos, toluene, and metals such as cadmium, mercury, chromium, and lead compounds.
- Smog is a mixture of air pollutants (sulfur dioxide, nitrogen oxides, ozone, and particulates) that often form over urban areas as a result of fossil fuel combustion. The term was coined from the terms "smoke" and "fog" referring to a brownish haze that pollutes the air, greatly reducing visibility and making it difficult for some people to breathe.
- Numerous scientific studies have linked air pollution to a variety of health problems.

Contributors and Attributions

- Libretext: Introduction to Environmental Science (Zendher et al.)
- TextMap: General CHemistry (Averill and Eldredge)

- Marisa Alviar-Agnew (Sacramento City College)
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7.4: Automobile Emissions

Learning Objective

- Identify the different pollutant gases in automobile emissions.

Exhaust gas or flue gas is emitted as a result of the combustion of fuels such as natural gas, gasoline, petrol, biodiesel blends, diesel fuel, fuel oil, or coal. It is a major component of motor vehicle emissions (and from stationary internal combustion engines), which can also include:

- Crankcase blow-by
- Evaporation of unused gasoline

Motor vehicle emissions contribute to air pollution and are a major ingredient in the creation of smog in some large cities. A 2013 study by MIT indicates that 53,000 early deaths occur per year in the United States alone because of vehicle emissions.

Carbon monoxide and nitrogen oxides are the two main pollutant gases from automobile emissions. Ozone is a result of the reaction between nitrogen oxides and volatile organic compounds (VOCs).

Carbon Monoxide

Carbon Monoxide (CO) is a colorless, odorless gas that can be harmful when inhaled in large amounts. CO is released when something is burned. The greatest sources of CO to outdoor air are cars, trucks and other vehicles or machinery that burn fossil fuels. A variety of items in your home such as unvented kerosene and gas space heaters, leaking chimneys and furnaces, and gas stoves also release CO and can affect air quality indoors.

Harmful effects

Breathing air with a high concentration of CO reduces the amount of oxygen that can be transported in the blood stream to critical organs like the heart and brain.

At very high levels, which are possible indoors or in other enclosed environments, CO can cause dizziness, confusion, unconsciousness and death.

Very high levels of CO are not likely to occur outdoors. However, when CO levels are elevated outdoors, they can be of particular concern for people with some types of heart disease. These people already have a reduced ability for getting oxygenated blood to their hearts in situations where the heart needs more oxygen than usual. They are especially vulnerable to the effects of CO when exercising or under increased stress. In these situations, short-term exposure to elevated CO may result in reduced oxygen to the heart accompanied by chest pain also known as angina.

Nitrogen Oxides

Nitrogen dioxide (NO_2) is one of a group of highly reactive gasses known as "oxides of nitrogen," or "nitrogen oxides (NO_X)." Other nitrogen oxides include nitrous acid and nitric acid. NO_2 is a yellowish-brown to reddish-brown foul-smelling gas that is a major contributor to smog and acid rain. Nitrogen oxides result when atmospheric nitrogen and oxygen react at the high temperatures created by combustion engines. Most emissions in the U.S. result from combustion in vehicle engines, electrical utility, and industrial combustion.

NO_2 primarily gets in the air from the burning of fuel. NO_2 forms from emissions from cars, trucks and buses, power plants, and off-road equipment.

Health effects

Breathing air with a high concentration of NO_2 can irritate airways in the human respiratory system. Such exposures over short periods can aggravate respiratory diseases, particularly asthma, leading to respiratory symptoms (such as coughing, wheezing or difficulty breathing), hospital admissions and visits to emergency rooms. Longer exposures to elevated concentrations of NO_2 may contribute to the development of asthma and potentially increase susceptibility to respiratory infections. People with asthma, as well as children and the elderly are generally at greater risk for the health effects of NO_2.

NO_2 along with other NO_x reacts with other chemicals in the air to form both particulate matter and ozone. Both of these are also harmful when inhaled due to effects on the respiratory system.

Environmental effects

NO_2 and other NO_x interact with water, oxygen and other chemicals in the atmosphere to form acid rain. Acid rain harms sensitive ecosystems such as lakes and forests.

The nitrate particles that result from NO_x make the air hazy and difficult to see though. This affects the many national parks that we visit for the view. NO_x in the atmosphere contributes to nutrient pollution in coastal waters.

Ozone

Ground-level ozone (O_3) is a colorless gas with a slightly sweet odor that is not emitted directly into the air, but is created by the interaction of sunlight, heat, oxides of nitrogen (NO_x) and volatile organic compounds (VOCs). Ozone is likely to reach unhealthy levels on hot sunny days in urban environments. Emissions from industrial facilities and electric utilities, motor vehicle exhaust, gasoline vapors, and chemical solvents are some of the major sources of NO_x and VOCs.

Ozone is a gas composed of three atoms of oxygen (O_3). Ozone occurs both in the Earth's upper atmosphere and at ground level. Ozone can be good or bad, depending on where it is found.

Called stratospheric ozone, good ozone occurs naturally in the upper atmosphere, where it forms a protective layer that shields us from the sun's harmful ultraviolet rays. This beneficial ozone has been partially destroyed by manmade chemicals, causing what is sometimes called a "hole in the ozone." The good news is, this hole is diminishing.

Ozone at ground level is a harmful air pollutant, because of its effects on people and the environment, and it is the main ingredient in “smog."

How does ground-level ozone form?

Tropospheric, or ground level ozone, is not emitted directly into the air, but is created by chemical reactions between oxides of nitrogen (NOx) and volatile organic compounds (VOC). This happens when pollutants emitted by cars, power plants, industrial boilers, refineries, chemical plants, and other sources chemically react in the presence of sunlight.

Ozone is most likely to reach unhealthy levels on hot sunny days in urban environments, but can still reach high levels during colder months. Ozone can also be transported long distances by wind, so even rural areas can experience high ozone levels.

This graphic depicts the formation of ozone from car and factory emissions

Health effects

Ozone in the air we breathe can harm our health especially on hot sunny days when ozone can reach unhealthy levels. People most at risk from breathing air containing ozone include people with asthma, children, older adults, and people who are active outdoors, especially outdoor workers.

Environmental effects

Ozone affects sensitive vegetation and ecosystems, including forests, parks, wildlife refuges and wilderness areas. In particular, ozone harms sensitive vegetation during the growing season.

Summary

- Carbon monoxide and nitrogen oxides are the two main compounds from automobile emissions.
- Ozone but is created by chemical reactions between oxides of nitrogen (NOx) and volatile organic compounds (VOC).
- Different levels of exposure to automobile emissions and ozone can lead to various health problems.

Contributor

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7.5: Photochemical Smog- Making Haze While the Sun Shines

Learning Objectives

- Describe photochemical smog.
- List different means to address photochemical smog.

Photochemical smog is a type of air pollution due to the reaction of solar radiation with airborne pollutant mixtures of nitrogen oxides (NOx) and volatile organic compounds (hydrocarbons). Smog is a byproduct of modern industrialization. Due to industry and the number of motor vehicles, this is more of a problem in large cities that have a warm, sunny and dry climate.

- **Oxidation**: Photochemical smog is also referred to as oxidizing smog. Oxidation reactions have been defined several ways. In terms of oxygen transfer, oxidation is a gain of oxygen. Oxidation can also be defined as a loss of hydrogen. The most important use of oxidation is described in terms of electron transfer. Oxidation can be described as an increase in oxidation number or loss of electrons. Oxidation numbers represents a distribution of charge. In other words, oxidation numbers represent the charge of the atom if the compound was composed of ions.
- **Reduction**: Reduction can involve the gain of hydrogen or loss of oxygen. Reduction can refer to the gain of electrons, which results in a decrease in oxidation number.

Formation of Photochemical Smog

The different reactions involved in the formation of photochemical smog are given below.

Step 1: People begin driving in the morning, nitrogen is burned or oxidized

$$N_2 + O_2 \rightarrow 2NO$$

- Oxidation number of N_2 is 0. The nitrogen in NO has acquired an oxidation number of +2.

Step 2: After a few hours, NO combines with O_2, in another oxidation reaction

$$2NO + O_2 \rightarrow 2NO_2$$

- The nitrogen in NO has an oxidation number of +2. The nitrogen in NO_2 has an oxidation number of +4.

Step 3: Nitrogen dioxide absorbs light energy, resulting in a reduction reaction

$$NO_2 \rightarrow NO + O$$

- The nitrogen in NO_2 has an oxidation number of +4 and the nitrogen in NO is +2.

Step 4: In sunlight, atomic oxygen combines with oxygen gas to form ozone

$$O + O_2 \rightarrow O_3$$

Step 5: Reaction is temperature and sunlight dependent

$$O_3 + NO \rightleftharpoons NO_2 + O_2$$

Alternative Reactions

NO and NO_2 can also react with the hydrocarbons instead of ozone to form other volatile compounds known as PAN (peroxyacetyl nitrate) as shown in Figure . The accumulation of ozone and volatile organic compounds along with the energy from the sun forms the brown, photochemical smog seen on hot, sunny days. Panoramic view of Santiago covered by a layer of smog on May 10,

2006. The Metropolitan Region of Santiago facing the driest autumn last 28 years due to lack of rainfall, which coupled with poor air circulation, causes an increase in smog.

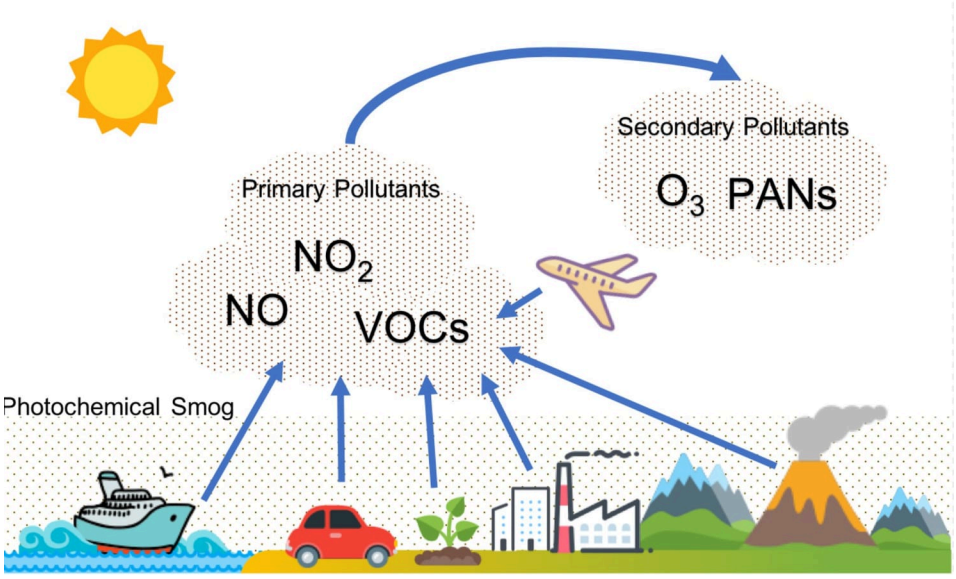


Figure 7.5.1 The photochemical smog formation diagram. Source: Wikipedia

Health Hazards

Because ozone is highly reactive, it has the ability to oxidize and destroy lung tissue. Short term exposures to elevated levels of ozone (above .75 ppm) have been linked to a host of respiratory irritations including coughing, wheezing, substernal soreness, pharyngitis, and dyspnea. Prolonged exposure to the molecule has been proven to cause a permanent reduction in lung function, as well as elevate the risk of developing asthma. Sulfur dioxide is a common component of London smog. Epidemiological studies have linked short term sulfur dioxide exposure to respiratory irritations including coughing, wheezing, and pharyngitis.

Other Harmful Effects of Smog

Plants are harmed by exposure to nitrogen oxides, ozone, and peroxyacetyl nitrate (PAN, see above), all oxidants present in a smoggy atmosphere. PAN is the most harmful of these constituents, damaging younger plant leaves, especially. Ozone exposure causes formation of yellow spots on leaves, a condition called chlorotic stippling. Some plant species, including sword-leaf lettuce, black nightshade, quickweed, and double-fortune tomato, are extremely susceptible to damage by oxidant species in smog and are used as bioindicators of the presence of smog. Costs of crop and orchard damage by smog run into millions of dollars per year in areas prone to this kind of air pollution, such as southern California.

Materials that are adversely affected by smog are generally those that are attacked by oxidants. The best example of such a material is rubber, especially natural rubber, which is attacked by ozone. Indeed, the hardening and cracking of natural rubber has been used as a test for atmospheric ozone.

Visibility-reducing atmospheric aerosol particles are the most common manifestation of the harm done to atmospheric quality by smog. The smog-forming process occurs by the oxidation of organic materials in the atmosphere, and carbon-containing organic materials are the most common constituents of the aerosol particles in an atmosphere afflicted by smog. Conifer trees(pine and cypress) and citrus trees are major contributors to the organic hydrocarbons that are precursors to organic particle formation in smog.

Controlling Photochemical Smog

Every new vehicle sold in the United States must include a catalytic converter to reduce photochemical emissions. Catalytic converters force CO and incompletely combusted hydrocarbons to react with a metal catalyst, typically platinum, to produce CO_2 and H_2O. Additionally, catalytic converters reduce nitrogen oxides from exhaust gases into O_2 and N_2, eliminating the cycle of

ozone formation. Many scientists have suggested that pumping gas at night could reduce photochemical ozone formation by limiting the amount of exposure VOCs have with sunlight.

Preventing Smog with Green Chemistry

Smog is basically a chemical problem, which would indicate that it should be amenable to chemical solutions. Indeed, the practice of green chemistry and the application of the principles of industrial ecology can help to reduce smog. This is due in large part to the fact that a basic premise of green chemistry is to avoid the generation and release of chemical species with the potential to harm the environment. The best way to prevent smog formation is to avoid the release of nitrogen oxides and organic vapors that enable smog to form. At an even more fundamental level, measures can be taken to avoid the use of technologies likely to release such substances, for example, by using alternatives to polluting automobiles for transportation.

The evolution of automotive pollution control devices to reduce smog provides an example of how green chemistry can be used to reduce pollution. The first measures taken to reduce hydrocarbon and nitrogen oxide emissions from automobiles were very much command-and-control and “end-of-pipe” measures. These primitive measures implemented in the early 1970s did reduce emissions, but with a steep penalty in fuel consumption and in driving performance of vehicles. However, over the last three decades, the internal combustion automobile engine has evolved into a highly sophisticated computer-controlled machine that generally performs well, emits few air pollutants, and is highly efficient. (And it would be much more efficient if those drivers who feel that they must drive “sport utility” behemoths would switch to vehicles of a more sensible size.) This change has required an integrated approach involving reformulation of gasoline. The first major change was elimination from gasoline of tetraethyllead, an organometallic compound that poisoned automotive exhaust catalysts (and certainly was not good for people). Gasoline was also reformulated to eliminate excessively volatile hydrocarbons and unsaturated hydrocarbons (those with double bonds between carbon atoms) that are especially reactive in forming photochemical smog.

An even more drastic approach to eliminating smog-forming emissions is the use of electric automobiles that do not burn gasoline. These vehicles certainly do not pollute as they are being driven, but they suffer from the probably unsolvable problem of a very limited range between charges and the need for relatively heavy batteries. However, hybrid automobiles using a small gasoline or diesel engine that provides electricity to drive electric motors propelling the automobile and to recharge relatively smaller batteries can largely remedy the emission and fuel economy problems with automobiles. The internal combustion engine on these vehicles runs only as it is needed to provide power and, in so doing, can run at a relatively uniform speed that provides maximum economy with minimum emissions.

Another approach that is being used on vehicles as large as buses that have convenient and frequent access to refueling stations is the use of fuel cells that can generate electricity directly from the catalytic combination of elemental hydrogen and oxygen, producing only harmless water as a product . There are also catalytic process that can generate hydrogen from liquid fuels, such as methanol, so that vehicles carrying such a fuel can be powered by electricity generated in fuel cells.

Green chemistry can be applied to devices and processes other than automobiles to reduce smog-forming emissions. This is especially true in the area of organic solvents used for parts cleaning and other industrial operations, vapors of which are often released to the atmosphere. The substitution of water with proper additives or even the use of supercritical carbon dioxide fluid can eliminate such emissions.

Summary

- Photochemical smog is a mixture of pollutants that are formed (mostly during the hot summer months) when nitrogen oxides and volatile organic compounds (VOCs) react to sunlight, creating a brown haze above cities.
- Photochemical smog is formed from the reactions of natural and man-made emissions of nitrogen oxides and VOCs.
- Smog is a serious problem in many cities and continues to harm human health and are especially harmful for senior citizens, children, and people with heart and lung conditions such as emphysema, bronchitis, and asthma.
- Catalytic converters in gas powered vehicles help reduce photochemical emissions.
- The practice of green chemistry and the application of the principles of industrial ecology can help to reduce smog.

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7.6: Acid Rain- Air Pollution Water Pollution

Learning Objective

- Describe acid rain and its effects.

Acid rain is a term referring to a mixture of wet and dry deposition (deposited material) from the atmosphere (Figure 7.6.1) containing higher than normal amounts of nitric and sulfuric acids. The precursors, or chemical forerunners, of acid rain formation result from both natural sources, such as volcanoes and decaying vegetation, and man-made sources, primarily emissions of sulfur dioxide (SO_2) and nitrogen oxides (NO_x) resulting from fossil fuel combustion. Acid rain occurs when these gases react in the atmosphere with water, oxygen, and other chemicals to form various acidic compounds. The result is a mild solution of sulfuric acid and nitric acid.

$$SO_2 + HOH \rightarrow H_2SO_3 \tag{7.6.1}$$

$$2NO_2 + HOH \rightarrow HNO_2 + HNO_3 \tag{7.6.2}$$

When sulfur dioxide and nitrogen oxides are released from power plants and other sources, *prevailing winds* blow these compounds across state and national borders, sometimes over hundreds of miles.

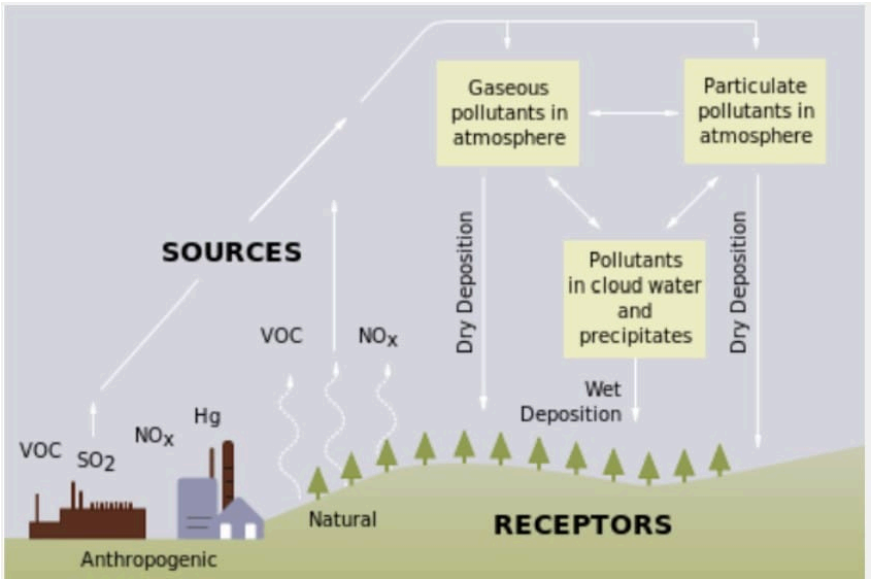


Figure 7.6.1 Processes involved in acid deposition.

Acid rain is measured using a scale called “pH.” The lower a substance’s pH, the more acidic it is. Pure water has a pH of 7.0. However, normal rain is slightly acidic because carbon dioxide (CO_2) dissolves into it forming weak carbonic acid, giving the resulting mixture a pH of approximately 5.6 at typical atmospheric concentrations of CO_2. As of 2000, the most acidic rain falling in the U.S. has a pH of about 4.3.

Effects of Acid Rain

Acid rain causes **acidification** of lakes and streams and contributes to the damage of trees at high elevations (for example, red spruce trees above 2,000 feet) and many sensitive forest soils. In addition, acid rain accelerates the decay of building materials and paints, including irreplaceable buildings, statues, and sculptures that are part of our nation’s cultural heritage. Prior to falling to the earth, sulfur dioxide (SO_2) and nitrogen oxide (NO_x) gases and their particulate matter derivatives—sulfates and nitrates—contribute to visibility degradation and harm public health.

The **ecological** effects of acid rain are most clearly seen in the aquatic, or water, environments, such as streams, lakes, and marshes. Most lakes and streams have a pH between 6 and 8, although some lakes are naturally acidic even without the effects of acid rain. Acid rain primarily affects sensitive bodies of water, which are located in watersheds whose soils have a limited ability to neutralize acidic compounds (called “buffering capacity”). Lakes and streams become acidic (i.e., the pH value goes down) when the water itself and its surrounding soil cannot buffer the acid rain enough to neutralize it. In areas where buffering capacity is low, acid rain releases aluminum from soils into lakes and streams; aluminum is highly toxic to many species of aquatic organisms. Acid rain causes slower growth, injury, or death of forests as shown in Figure 7.6.2. Of course, acid rain is not the only cause of such conditions. Other factors contribute to the overall stress of these areas, including air pollutants, insects, disease, drought, or very cold weather. In most cases, in fact, the impacts of acid rain on trees are due to the combined effects of acid rain and these other environmental stressors.



Figure 7.6.2. Acid rain damage to a forest in the Czech Republic.

Acid rain and the dry deposition of acidic particles contribute to the corrosion of **metals**(such as bronze).

The damage that acid rain does to limestone and marble buildings and sculptures is due to a classic acid–base reaction. Marble and limestone both consist of calcium carbonate ($CaCO_3$), a salt derived from the weak acid H_2CO_3. The reaction of a strong acid with a salt of a weak acid goes to completion. Thus we can write the reaction of limestone or marble with dilute sulfuric acid as follows:

$$CaCO_{3(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + H_2O_{(l)} + CO_{2(g)} \tag{7.6.3}$$

Because $CaSO_4$ is sparingly soluble in water, the net result of this reaction is to dissolve the marble or limestone.

These effects significantly reduce the societal value of buildings, bridges, cultural objects (such as statues, monuments, and tombstones), and cars (Figure 7.6.3).



Figure 7.6.3. A gargoyle that has been damaged by acid rain.

Sulfates and nitrates that form in the atmosphere from sulfur dioxide (SO_2) and nitrogen oxides (NO_x) emissions contribute to **visibility impairment,** meaning we cannot see as far or as clearly through the air. The pollutants that cause acid rain—sulfur dioxide (SO_2) and nitrogen oxides (NO_x)—damage **human health**. These gases interact in the atmosphere to form fine sulfate and nitrate particles that can be transported long distances by winds and inhaled deep into people's lungs. Fine particles can also penetrate indoors. Many scientific studies have identified a relationship between elevated levels of fine particles and increased illness and premature death from heart and lung disorders, such as asthma and bronchitis.

Summary

- Acid rain is a term referring to a mixture of wet and dry deposition (deposited material) from the atmosphere containing higher than normal amounts of nitric and sulfuric acids.
- Acidic rain water in the soil, streams, lakes, and marshes (and other bodies of water) can be harmful to trees, plants, animals, especially aquatic plants and animals.
- Acid rain and the dry deposition of acidic particles contribute to the corrosion of **metals**(such as bronze) and the deterioration of paint and stone (such as marble and limestone).

Contributors and Attributions

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7.7: Indoor Air Pollution

Learning Objectives

- Name the different indoor pollutants and their sources.
- Know the harmful effects of indoor pollutants.

In both developed and developing nations, indoor air pollution poses a greater health risk than outdoor air pollution. According to the World Health Organization (WHO) and other agencies such as the Environmental Protection Agency (EPA), indoor air generally contains higher concentrations of toxic pollutants than outdoor air. Additionally, people generally spend more time indoors than outdoors, hence, the health effects from indoor air pollution in workplaces, schools, and homes are far greater than outdoor. Indoor pollution sources that release gases or particles into the air are the primary cause of indoor air quality problems in homes. Inadequate ventilation can increase indoor pollutant levels by not bringing in enough outdoor air to dilute emissions from indoor sources and by not carrying indoor air pollutants out of the home.

Outdoor air enters and leaves a building by infiltration, natural ventilation, and mechanical ventilation. In infiltration, outdoor air flows into the house through openings, joints, and cracks in walls, floors, and ceilings, and around windows and doors. In natural ventilation, air moves through opened windows and doors. Air movement associated with infiltration and natural ventilation is caused by air temperature differences between indoors and outdoors and by wind. Finally, there are a number of mechanical ventilation devices, from outdoor-vented fans that intermittently remove air from a single room, such as bathrooms and kitchen, to air handling systems that use fans and duct work to continuously remove indoor air and distribute filtered and conditioned outdoor air to strategic points throughout the house. The rate at which outdoor air replaces indoor air is described as the air exchange rate. When there is little infiltration, natural ventilation, or mechanical ventilation, the air exchange rate is low and pollutant levels can increase. High temperature and humidity levels can also increase concentrations of some pollutants.

There are many sources of indoor air pollution in any home (Figure 7.7.1). These include combustion sources such as oil, gas, kerosene, coal, wood, and tobacco products; building materials and furnishings as diverse as deteriorated, asbestos-containing insulation, wet or damp carpet, and cabinetry or furniture made of certain pressed wood products; products for household cleaning and maintenance, personal care, or hobbies; central heating and cooling systems and humidification devices. Pollutants causing indoor air pollution can also originate from outside sources such as radon, pesticides, and outdoor air pollution. Radon is a naturally occurring radioactive gas produced from the decay of uranium in rock. If a building/home is constructed in an area with uranium containing rock, the gas can seep through the foundations and accumulate in basements. Exposure to radon can cause lung cancer.

The relative importance of any single source depends on how much of a given pollutant it emits and how hazardous those emissions are. In some cases, factors such as how old the source is and whether it is properly maintained are significant. For example, an improperly adjusted gas stove can emit significantly more carbon monoxide than one that is properly adjusted. Some sources, such as building materials, furnishings, and household products like air fresheners, release pollutants more or less continuously. Other sources, related to activities carried out in the home, release pollutants intermittently. These include smoking, the use of unvented or malfunctioning stoves, furnaces, or space heaters, the use of solvents in cleaning and hobby activities, the use of paint strippers in redecorating activities, and the use of cleaning products and pesticides in house-keeping. High pollutant concentrations can remain in the air for long periods after some of these activities.

Risks from indoor air pollution differ between less industrialized and industrialized nations. Indoor pollution has a greater impact in less industrialized nations where many people use cheaper sources of fuel such as wood, charcoal, and crop waste among others for cooking and heating, often with little or no ventilation. The most significant indoor pollutant, therefore, is soot and carbon monoxide. In industrialized nations, the primary indoor air health risks are cigarette smoke and radon.

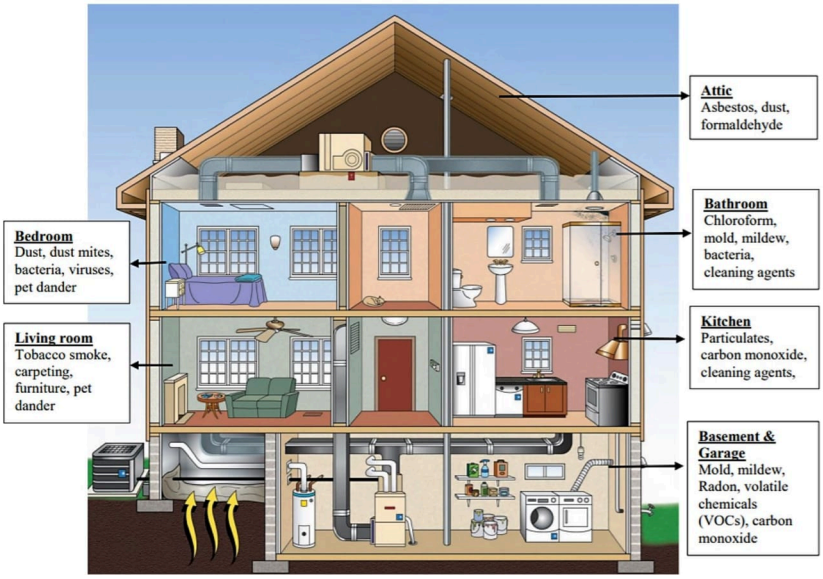


Figure 7.7.1: Sources of indoor air pollution. Image source: EPA http://www.epa.gov/iaq/pdfs/careforyourair.pdf

Combustion Pollutants

Combustion pollutants are gases or particles that come from burning materials. In homes, the major source of combustion pollutants are improperly vented or unvented fuel-burning appliances such as space heaters, wood stoves, gas stoves, water heaters, dryers, and fireplaces. The types and amounts of pollutants produced depends on the type of appliance, how well the appliance is installed, maintained and vented and the kind of fuel it uses. Common combustion pollutants include:

Carbon monoxide (CO) which is a colorless, odorless gas that interferes with the delivery of oxygen throughout the body. Carbon monoxide causes headaches, dizziness, weakness, nausea and even death. Average levels in homes without gas stoves vary from 0.5 to 5 parts per million (ppm). Levels near properly adjusted gas stoves are often 5 to 15 ppm and those near poorly adjusted stoves may be 30 ppm or higher.

Nitrogen dioxide (NO_2) which is a colorless, odorless gas that causes eye, nose and throat irritation, shortness of breath, and an increased risk of respiratory infection. Average level in homes without combustion appliances is about half that of outdoors. In homes with gas stoves, kerosene heaters or un-vented gas space heaters, indoor levels often exceed outdoor levels.

Particulate Matter (PM) which refers to fine particles that forms in smoke when wood or other organic matter burns.

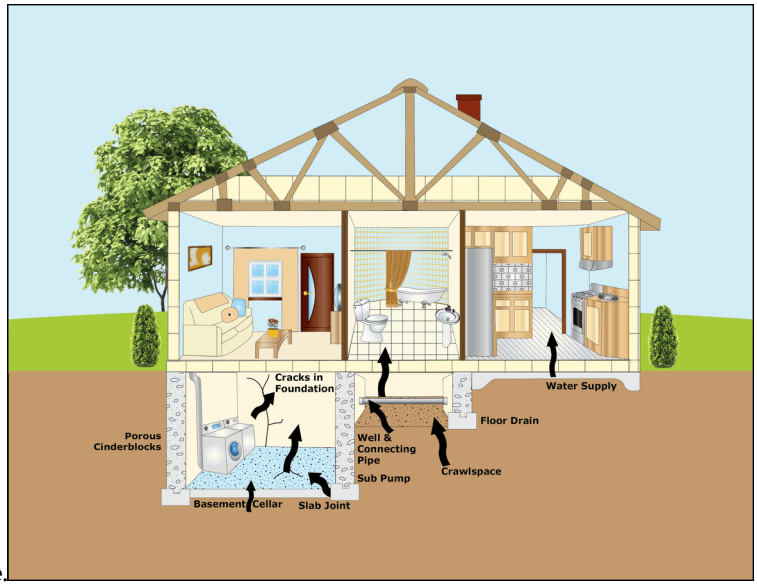
Other pollutants from wood smoke In addition to particle pollution, wood smoke contains several toxic harmful air pollutants including benzene, formaldehyde, acrolein, and polycyclic aromatic hydrocarbons (PAHs).

Secondhand Smoke

Secondhand smoke is a mixture of the smoke given off by the burning of tobacco products, such as cigarettes, cigars or pipes and the smoke exhaled by smokers. Secondhand smoke is also called environmental tobacco smoke (ETS). Exposure to secondhand smoke is sometimes called involuntary or passive smoking. Secondhand smoke, classified by EPA as a Group A carcinogen, contains more than 7,000 substances. Secondhand smoke exposure commonly occurs indoors, particularly in homes and cars. Secondhand smoke can move between rooms of a home and between apartment units. Opening a window or increasing ventilation in a home or car is not protective from secondhand smoke.

Radon

Radon is a radioactive gas that results from the natural decay of uranium and radiumfound in nearly all rocks and soils. Elevated



radon levels have been found in every state. Radon is in the atmosphere and can also be found in ground water. The national average for radon in outdoor air is 0.4 picocuries per liter (pCi/L), while the average for indoor air is 1.3 pCi/L.

Any building can have high levels of radon, including new and old homes, well-sealed and drafty homes, office buildings and schools, and homes with or without basements. Radon gas can get into buildings through cracks in solid floors and walls, construction joints, gaps in suspended floors and around service pipes, cavities inside walls, the water supply and building materials. Testing is the only way to know if your home has elevated radon levels.EPA recommends fixing your home when the radon level is at or above 4 pCi/L.

Figure 7.7.2***:*** *Sources of Radon. Image source: EPA*https://www.epa.gov/radiation/radionuclide-basics-radon

EPA estimates that about 21,000 lung cancer deaths each year in the U.S. are radon-related. Exposure to radon is the second leading cause of lung cancer after smoking. For most people, radon is the single greatest environmental source of radiation exposure. EPA recommends that all homes and schools be tested for radon. For smokers, the risk of lung cancer is heightened due to the combined effects of radon and smoking.

Other Indoor Pollutants

Molds are usually not a problem indoors, unless mold spores land on a wet or damp spot and begin growing. Molds have the potential to cause health problems. Molds produce allergens (substances that can cause allergic reactions) and irritants. Inhaling or touching mold or mold spores may cause allergic reactions in sensitive individuals. Allergic responses include hay fever-type symptoms, such as sneezing, runny nose, red eyes, and skin rash.

Allergic reactions to mold are common. They can be immediate or delayed. Molds can also cause asthma attacks in people with asthma who are allergic to mold. In addition, mold exposure can irritate the eyes, skin, nose, throat, and lungs of both mold-allergic and non-allergic people. Symptoms other than the allergic and irritant types are not commonly reported as a result of inhaling mold.

Volatile organic compounds (VOCs) are emitted as gases from certain solids or liquids. VOCs include a variety of chemicals, some of which may have short- and long-term adverse health effects. Concentrations of many VOCs are consistently higher indoors (up to ten times higher) than outdoors. VOCs are emitted by a wide array of products numbering in the thousands.

Organic chemicals are widely used as ingredients in household products. Paints, varnishes and wax all contain organic solvents, as do many cleaning, disinfecting, cosmetic, degreasing and hobby products. Fuels are made up of organic chemicals. All of these products can release organic compounds while you are using them, and, to some degree, when they are stored.

EPA's Office of Research and Development's "Total Exposure Assessment Methodology (TEAM) Study" (Volumes I through IV, completed in 1985) found levels of about a dozen common organic pollutants to be 2 to 5 times higher inside homes than outside, regardless of whether the homes were located in rural or highly industrial areas. TEAM studies indicated that while people are

using products containing organic chemicals, they can expose themselves and others to very high pollutant levels, and elevated concentrations can persist in the air long after the activity is completed.

VOCs are emitted by a wide array of products used in homes including paints and lacquers, paint strippers, cleaning supplies, varnishes and waxes, pesticides, building materials and furnishings, office equipment, moth repellents, air fresheners, and dry-cleaned clothing. VOCs evaporate into the air when these products are used or sometimes even when they are stored.

Volatile organic compounds irritate the eyes, nose and throat, and cause headaches, nausea, and damage to the liver, kidneys and central nervous system. Some of them can cause cancer.

Asbestos is a mineral fiber that occurs in rock and soil. Because of its fiber strength and heat resistance asbestos has been used in a variety of building construction materials for insulation and as a fire retardant. Asbestos has also been used in a wide range of manufactured goods, mostly in building materials (roofing shingles, ceiling and floor tiles, paper products, and asbestos cement products), friction products (automobile clutch, brake, and transmission parts), heat-resistant fabrics, packaging, gaskets, and coatings.

Asbestos fibers may be released into the air by the disturbance of asbestos-containing material during product use, demolition work, building or home maintenance, repair, and remodeling. In general, exposure may occur only when the asbestos-containing material is disturbed or damaged in some way to release particles and fibers into the air.

Three of the major health effects associated with asbestos exposure are:

· lung cancer

· mesothelioma, a rare form of cancer that is found in the thin lining of the lung, chest and the abdomen and heart

· asbestosis, a serious progressive, long-term, non-cancer disease of the lungs

Improving Your Indoor Air Quality

Take steps to help improve your air quality and reduce your IAQ-related health risks at little or no cost by:

Controlling the sources of pollution: Usually the most effective way to improve indoor air is to eliminate individual sources or reduce their emissions.

Ventilating: Increasing the amount of fresh air brought indoors helps reduce pollutants inside. When weather permits, open windows and doors, or run an air conditioner with the vent control open. Bathroom and kitchen fans that exhaust to the outdoors also increase ventilation and help remove pollutants.

Always ventilate and follow manufacturers' instructions when you use products or appliances that may release pollutants into the indoor air.

Changing filters regularly: Central heaters and air conditioners have filters to trap dust and other pollutants in the air. Make sure to change or clean the filters regularly, following the instructions on the package.

Adjusting humidity: The humidity inside can affect the concentrations of some indoor air pollutants. For example, high humidity keeps the air moist and increases the likelihood of mold.

Keep indoor humidity between 30 and 50 percent. Use a moisture or humidity gauge, available at most hardware stores, to see if the humidity in your home is at a good level. To increase humidity, use a vaporizer or humidifier. To decrease humidity, open the windows if it is not humid outdoors. If it is warm, turn on the air conditioner or adjust the humidity setting on the humid

Get a quick glimpse of some of the most important ways to protect the air in your home by touring the Indoor Air Quality (IAQ) House. Room-by-room, you'll learn about the key pollutants and how to address them.

Interactive Version

https://www.epa.gov/indoor-air-quality-iaq/interactive-tour-indoor-air-quality-demo-house#mainHouse

Text Version

https://www.epa.gov/indoor-air-quality-iaq/text-version-indoor-air-quality-house-tour

Summary

- Indoor pollutants include combustion pollutants (from burning of oil, gas, kerosene, coal, wood, and tobacco products); building materials and furnishings products for household cleaning and maintenance, personal care, or hobbies; central heating and cooling systems and humidification devices.
- Pollutants causing indoor air pollution can also originate from outside sources such as radon, pesticides, and outdoor air pollution.
- Indoor pollutants can cause serious health problems to sensitive groups.
- Several suggestions were provided to improve indoor air quality.

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7.8: Stratospheric Ozone- Earth's Vital Shield

Learning Objectives

- Describe the depletion of the ozone layer.
- Explain how chlorine and bromine atoms react with ozone that leads to the depletion of the ozone layer.

The earth's stratospheric ozone layer plays a critical role in absorbing ultraviolet radiation emitted by the sun. In the last thirty years, it has been discovered that stratospheric ozone is depleting as a result of anthropogenic pollutants. There are a number of chemical reactions that can deplete stratospheric ozone; however, some of the most significant of these involves the catalytic destruction of ozone by halogen radicals such as chlorine and bromine.

Introduction

The atmosphere of the Earth is divided into five layers. In order of closest and thickest to farthest and thinnest the layers are listed as follows: troposphere, stratosphere, mesosphere, thermosphere and exosphere. The majority of the ozone in the atmosphere resides in the stratosphere, which extends from six miles above the Earth's surface to 31 miles. Humans rely heavily on the absorption of ultraviolet B rays by the ozone layer because UV-B radiation causes skin cancer and can lead to genetic damage. The ozone layer has historically protected the Earth from the harmful UV rays, although in recent decades this protection has diminished due to stratospheric ozone depletion.

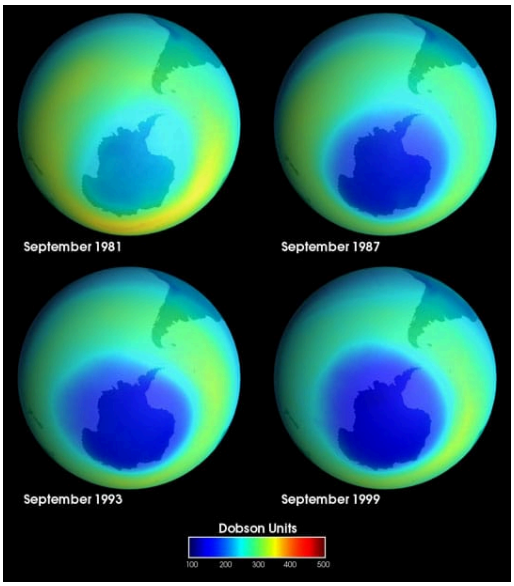


Figure courtesy of NASA.

Ozone depletion is largely a result of man-made substances. Humans have introduced gases and chemicals into the atmosphere that have rapidly depleted the ozone layer in the last century. This depletion makes humans more vulnerable to the UV-B rays which are known to cause skin cancer as well as other genetic deformities. The possibility of ozone depletion was first introduced by scientists in the late 1960's as dreams of super sonic transport began to become a reality. Scientists had long been aware that nitric oxide (NO) can catalytically react with ozone (O_3) to produce O_2 molecules; however, NO molecules produced at ground level have a half life far too short to make it into the stratosphere. It was not until the advent of commercial super sonic jets (which fly in the stratosphere and at an altitude much higher then conventional jets) that the potential for NO to react with stratospheric ozone became a possibility. The threat of ozone depletion from commercial super sonic transport was so great that it is often cited as the main reason why the US federal government pulled support for its development in 1971. Fear of ozone depletion was abated until 1974 when Sherwood Rowland and Mario Molina discovered that chlorofluorocarbons could be photolyzed by high energy photons in the stratosphere. They discovered that this process could releasing chlorine radicals that would catalytically

react with O_3 and destroy the molecule. This process is called the Rowland-Molina theory of O_3 depletion.

The Chapman Cycle

The stratosphere is in a constant cycle with oxygen molecules and their interaction with ultraviolet rays. This process is considered a cycle because of its constant conversion between different molecules of oxygen. The ozone layer is created when ultraviolet rays react with oxygen molecules (O_2) to create ozone (O_3) and atomic oxygen (O). This process is called the *Chapman cycle*.

Step 1: An oxygen molecules is photolyzed by solar radiation, creating two oxygen radicals:

$$h\nu + O_2 \rightarrow 2O^{\cdot}$$

Step 2: Oxygen radicals then react with molecular oxygen to produce ozone:

$$O_2 + O^{\cdot} \rightarrow O_3$$

Step 3: Ozone then reacts with an additional oxygen radical to form molecular oxygen:

$$O_3 + O^{\cdot} \rightarrow 2O_2$$

Step 4: Ozone can also be recycled into molecular oxygen by reacting with a photon:

$$O_3 + h\nu \rightarrow O_2 + O^{\cdot}$$

It is important to keep in mind that ozone is constantly being created and destroyed by the Chapman cycle and that these reactions are natural processes, which have been taking place for millions of years. Because of this, the thickness the ozone layer at any particular time can vary greatly. It is also important to know that O_2 is constantly being introduced into the atmosphere through photosynthesis, so the ozone layer has the capability of regenerating itself.

Chemistry of Ozone Depletion

CFC molecules are made up of chlorine, fluorine and carbon atoms and are extremely stable. This extreme stability allows CFC's to slowly make their way into the stratosphere (most molecules decompose before they can cross into the stratosphere from the troposphere). This prolonged life in the atmosphere allows them to reach great altitudes where photons are more energetic. When the CFC's come into contact with these high energy photons, their individual components are freed from the whole. The following reaction displays how Cl atoms have an ozone destroying cycle:

$$Cl + O_3 \rightarrow ClO + O_2 \qquad (\text{step } 1)$$

$$ClO + O^{\cdot} \rightarrow Cl + O_2 \qquad (\text{step } 2)$$

$$O_3 + O^{\cdot} \rightarrow 2O_2 \qquad (\text{Overall reaction})$$

Chlorine is able to destroy so much of the ozone because it acts as a catalyst. Chlorine initiates the breakdown of ozone and combines with a freed oxygen to create two oxygen molecules. After each reaction, chlorine begins the destructive cycle again with another ozone molecule. One chlorine atom can thereby destroy thousands of ozone molecules. Because ozone molecules are being broken down they are unable to absorb any ultraviolet light so we experience more intense UV radiation at the earths surface.

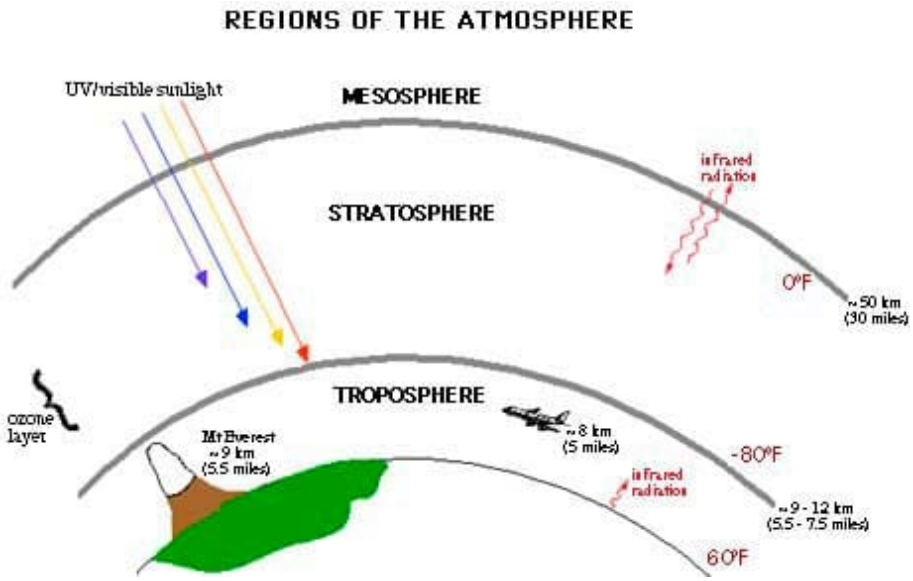


Figure courtesy of NOAA.

From 1985 to 1988, researchers studying atmospheric properties over the south pole continually noticed significantly reduced concentrations of ozone directly over the continent of Antarctica. For three years it was assumed that the ozone data was incorrect and was due to some type of instrument malfunction. In 1988, researchers finally realized their error and concluded that an enormous hole in the ozone layer had indeed developed over Antarctica. Examination of NASA satellite data later showed that the hole had begun to develop in the mid 1970's.

The ozone hole over Antarctica is formed by a slew of unique atmospheric conditions over the continent that combine to create an ideal environment for ozone destruction.

- Because Antarctica is surrounded by water, winds over the continent blow in a unique clockwise direction creating a so called "polar vortex" that effectively contains a single static air mass over the continent. As a result, air over Antarctica does not mix with air in the rest of the earth's atmosphere.
- Antarctica has the coldest winter temperatures on earth, often reaching -110 F. These chilling temperatures result in the formation of polar stratospheric clouds (PSC's) which are a conglomeration of frozen H_2O and HNO_3. Due to their extremely cold temperatures, PSC's form an electrostatic attraction with CFC molecules as well as other halogenated compounds

As spring comes to Antarctica, the PSC's melt in the stratosphere and release all of the halogenated compounds that were previously absorbed to the cloud. In the antarctic summer, high energy photons are able to photolyze the halogenated compounds, freeing halogen radicals that then catalytically destroy O_3. Because Antarctica is constantly surrounded by a polar vortex, radical halogens are not able to be diluted over the entire globe. The ozone hole develops as result of this process.

Resent research suggests that the strength of the polar vortex from any given year is directly correlated to the size of the ozone hole. In years with a strong polar vortex, the ozone hole is seen to expand in diameter, whereas in years with a weaker polar vortex, the ozone hole is noted to shrink

Ozone Depleting Substances

The following substances are listed as ozone depleting substances under Title VI of the United State Clean Air Act:

Table 7.8.1: Ozone Depleting Substances And Their Ozone-Depletion Potential. Taken directly from the Clean Air Act, as of June 2010.

Substance	Ozone- depletion potential
chlorofluorocarbon-11 (CFC–11)	1.0
chlorofluorocarbon-12 (CFC–12)	1.0

Substance	Ozone- depletion potential
chlorofluorocarbon-13 (CFC–13)	1.0
chlorofluorocarbon-111 (CFC–111)	1.0
chlorofluorocarbon-112 (CFC–112)	1.0
chlorofluorocarbon-113 (CFC–113)	0.8
chlorofluorocarbon-114 (CFC–114)	1.0
chlorofluorocarbon-115 (CFC–115)	0.6
chlorofluorocarbon-211 (CFC–211)	1.0
chlorofluorocarbon-212 (CFC–212)	1.0
chlorofluorocarbon-213 (CFC–213)	1.0
chlorofluorocarbon-214 (CFC–214)	1.0
chlorofluorocarbon-215 (CFC–215)	1.0
chlorofluorocarbon-216 (CFC–216)	1.0
chlorofluorocarbon-217 (CFC–217)	1.0
halon-1211	3.0
halon-1301	10.0
halon-2402	6.0
carbon tetrachloride	1.1
methyl chloroform	0.1
hydrochlorofluorocarbon-22 (HCFC–22)	0.05
hydrochlorofluorocarbon-123 (HCFC–123)	0.02
hydrochlorofluorocarbon-124 (HCFC–124)	0.02
hydrochlorofluorocarbon-141(b) (HCFC–141(b))	0.1
hydrochlorofluorocarbon-142(b) (HCFC–142(b))	0.06

Summary

- Any disruption of the balance that results in a higher rate of ozone destruction than ozone creation would result in depletion of ozone.
- Some compounds that release chlorine or bromine when they are exposed to intense UV light in the stratosphere contribute to ozone depletion, and are called ozone-depleting substances (ODS)
- The Montreal Protocol is an international agreement that committed all parties (signatory nations) to a schedule for phasing out the production and use of CFCs and other substances known to be harmful to the ozone layer.

Contributors and Attributions

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7.9: Carbon Dioxide and Climate Change

Learning Objectives

- Define global warming, climate change, the greenhouse effect, and greenhouse gases.
- Discuss strategies for reducing the intensity of the human influence on the greenhouse effect.

Global warming refers to the increase in the average temperature of the Earth's atmosphere due to elevated greenhouse gas concentrations, heightening the greenhouse effect. **Climate change** includes both global warming driven by human-induced emissions of greenhouse gases and the resulting large-scale shifts in weather patterns. Though there have been previous periods of climatic change, since the mid-20th century humans have had an unprecedented impact on Earth's climate system and caused change on a global scale.

The largest driver of warming is the emission of gases that create a greenhouse effect, of which more than 90% are carbon dioxide (CO_2) and methane. Fossil fuel burning (coal, oil, and natural gas) for energy consumption is the main source of these emissions, with additional contributions from agriculture, deforestation, and manufacturing. The human cause of climate change is not disputed by any scientific body of national or international standing. Temperature rise is accelerated or tempered by climate feedbacks, such as loss of sunlight-reflecting snow and ice cover, increased water vapour (a greenhouse gas itself), and changes to land and ocean carbon sinks.

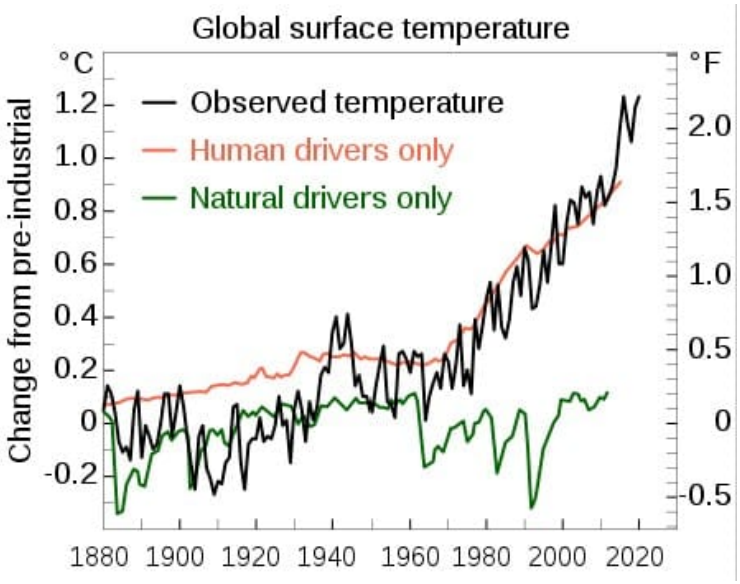


Figure 7.9.1 Observed temperature from NASA versus the 1850-1900 average as a pre-industrial baseline.

The **greenhouse effect** is the process by which radiation from a planet's atmosphere warms the planet's surface to a temperature above what it would be without this atmosphere (Figure 7.9.2). Radiatively active gases (i.e. greenhouse gases) in a planet's atmosphere radiate energy in all directions. Part of this radiation is directed towards the surface, thus warming it. The intensity of downward radiation – that is, the strength of the greenhouse effect – depends on the amount of greenhouse gases that the atmosphere contains. The temperature rises until the intensity of upward radiation from the surface, thus cooling it, balances the downward flow of energy.

Earth's natural greenhouse effect is critical to supporting life and initially was a precursor to life moving out of the ocean onto land. Human activities, mainly the burning of fossil fuels and clear cutting of forests, have increased the greenhouse effect and caused global warming.

The term *greenhouse effect* is a slight misnomer in the sense that physical greenhouses warm via a different mechanism. The greenhouse effect as an atmospheric mechanism functions through radiative heat loss while a traditional greenhouse as a built

structure blocks convective heat loss. The result, however, is an increase in temperature in both cases.

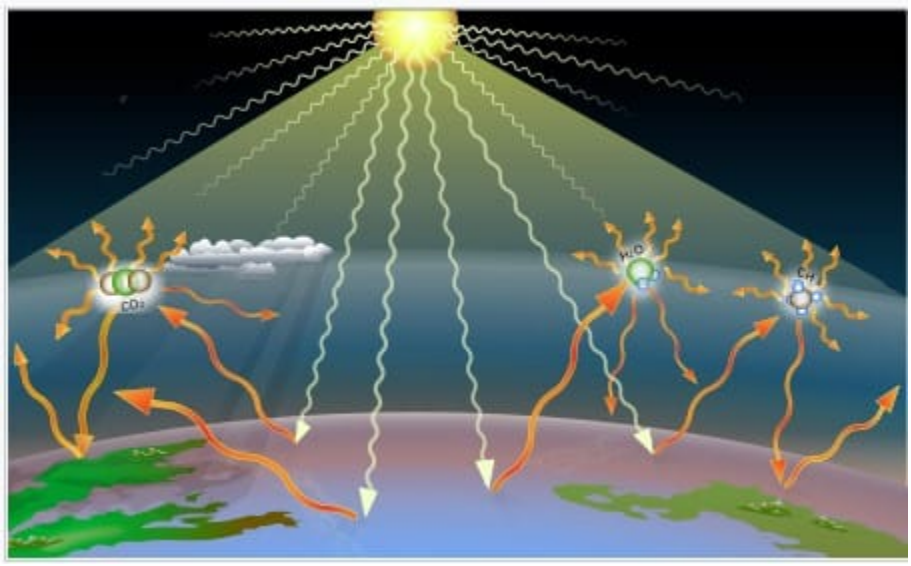


Figure 7.9.2 The greenhouse effect of solar radiation on the Earth's surface caused by emission of greenhouse gases. Source: Wikipedia

Greenhouse Gases and Global Warming

A **greenhouse gas** (sometimes abbreviated **GHG**) is a gas that absorbs and emits radiant energy within the thermal infrared range, causing the greenhouse effect. It's important to realize that **water vapor** (H_2O) is also a greenhouse gas. While humans have little direct impact on water vapor concentrations in the atmosphere, it is still an essential component of the natural greenhouse effect that occurs in our atmosphere. The four major categories of greenhouse gases that have been impacted by humans the most will be discussed in detail below. See Table 7.9.1 for a numeric comparison of these greenhouse gases.

- Carbon dioxide, CO_2
- Methane, CH_4
- Nitrous oxide, N_2O
- Synthetic **fluorinated gases**, including **hydrofluorocarbons** (HFCs), **perfluorocarbons** (PFCs), and **sulfur hexafluoride** (SF_6)

Carbon dioxide (CO_2) is the greenhouse gas responsible for most of the human-caused climate change in our atmosphere. It has the highest concentration in the atmosphere of any of the greenhouse gases that we'll discuss here. Remember that CO_2 is a direct product of both combustion and cellular respiration, causing it to be produced in great quantities both naturally and anthropogenically. Any time biomass or fossil fuels are burned, CO_2 is released. Major anthropogenic sources include: electricity production from coal-fired and natural gas power plants, transportation, and industry (Chapter 4). To get an idea of how CO_2 concentration has changed over time, watch this video compiled by the National Oceanic and Atmospheric Administration (NOAA): http://www.esrl.noaa.gov/gmd/ccgg/trends/history.html. This video contains atmospheric CO_2 concentrations measured directly, dating back to 1958, as well as atmospheric CO_2 concentrations measured indirectly from ice core data, dating back to 800,000 BCE. By 1990, a quantity of over seven billion tons of carbon (equivalent to 26 billion tons of carbon dioxide when the weight of the oxygen atoms are also considered) was being emitted into the atmosphere every year, much of it from industrialized nations. Similar to the action of the naturally existing greenhouse gases, any additional greenhouse gases leads to an increase in the surface temperature of the Earth.

While CO_2 is produced by aerobic cellular respiration, gases such as CH_4 and N_2O are often the products of anaerobic metabolisms. Agriculture is a major contributor to CH_4 emissions. In addition to anaerobic bacteria, methane is also a significant component of natural gas, and is commonly emitted through the mining and use of natural gas and petroleum, in addition to coal

mining. Finally, landfills contribute significantly to CH_4 emissions, as the waste put into the landfill largely undergoes anaerobic decomposition as it is buried under many layers of trash and soil. Natural sources of CH_4 include swamps and wetlands, and volcanoes.

The vast majority of N_2O production by humans comes from agricultural land management. While some N_2O is naturally emitted to the atmosphere from soil as part of the nitrogen cycle, human changes in land management, largely due to agricultural practices, have greatly increased N_2O emissions. Some N_2O is also emitted from transportation and industry.

Due to their relatively high concentrations in the atmosphere compared to synthetic gases, CO_2, CH_4, and N_2O, are responsible for most of the human-caused global climate change over the past century. Figure 7.9.2 shows the increases in all three gases since 1750. Ice core data shows us that the atmospheric CO_2 concentration never exceeded 300 ppm before the industrial revolution. As of early 2015, the current atmospheric CO_2 concentration is 400 ppm.

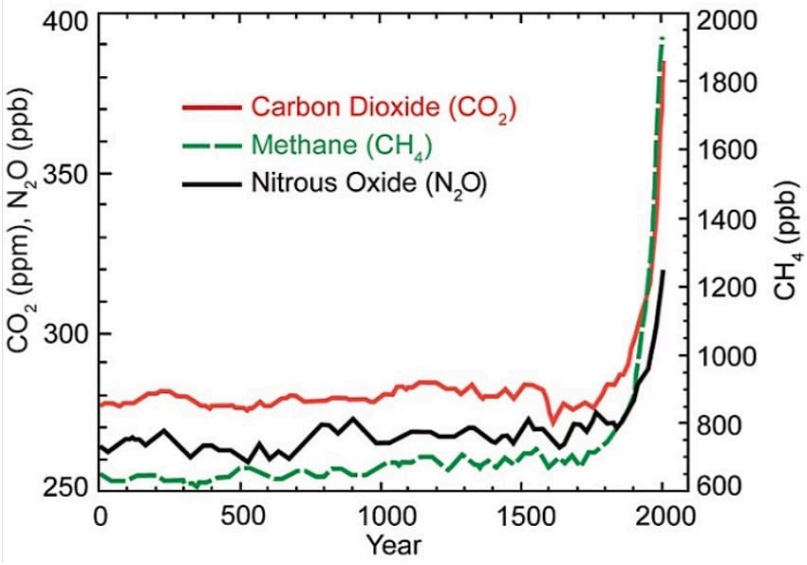


Figure 7.9.2: Increase in greenhouse gas concentrations in the atmosphere over the last 2,000 years. Increases in concentrations of these gases since 1750 are due to human activities in the industrial era. Concentration units are parts per million (ppm) or parts per billion (ppb), indicating the number of molecules of the greenhouse gas per million or billion molecules of air. Source: USGCRP (2009)

One class of greenhouse gas chemicals that has no natural sources is the fluorinated gases. These include HFCs, PFCs, and SF6, among others. Because these are synthetic chemicals that are only created by humans, these gases were essentially non-existent before the industrial revolution. These synthetic gases are used for a wide variety of applications, from refrigerants to semiconductor manufacturing, and propellants to fire retardants. They tend to have a long lifetime in the atmosphere, as seen in Table 7.9.1. Some of these chemicals, as well as the older **chlorofluorocarbons** (CFCs), have been phased out by international environmental legislation under the Montreal Protocol. Due to their long lifespan, many of these now banned CFCs remain in the atmosphere. Newer chemical replacements, such as HFCs, provide many of the same industrial applications, but unfortunately have their own environmental consequences.

Just as greenhouse gases differ in their sources and their residence time in the atmosphere, they also differ in their ability to produce the greenhouse effect. This is measured by the **global warming potential**, or GWP, of each greenhouse gas. The GWP of a greenhouse gas is based on its ability to absorb and scatter energy, as well as its lifetime in the atmosphere. Since CO_2 is the most prevalent greenhouse gas, all other greenhouse gases are measured relative to it. As the reference point, CO_2 always has a GWP of 1. Note the very high GWP values of the synthetic fluorinated gases in Table 7.9.1. This is largely due to their very long residence time in the atmosphere. Also note the higher GWP values for CH_4 and N_2O compared to CO_2.

Table 7.9.1. Comparision of Greenhouse Gases.

Greenhouse gas	Chemical formula or abbreviation	Lifetime in atmosphere	Global warming potential (100-year)
Carbon dioxide	CO_2	Variable	1
Methane	CH_4	12 years	28-36

Nitrous oxide	N_2O	114 years	298
Hydrofluorocarbons	Abbreviation: HFCs	1-270 years	12-14,800
Perfluorocarbons	Abbreviation: PFCs	2,600-50,000 years	7,390
Sulfur hexafluoride	SF6	3,200 years	22,800

Other climate influencers

In addition to greenhouse gases, other manmade changes may be forcing climate change. Increases in near-surface ozone from internal combustion engines, aerosols such as carbon black, mineral dust and aviation-induced exhaust are acting to raise the surface temperature. This primarily occurs due to a decrease in the **albedo** of light-colored surfaces by the darker-colored carbon black, soot, dust, or particulate matter. As you know, it is more comfortable to wear a white shirt on a hot summer day than a black shirt. Why is this? Because the lighter-colored material bounces more solar radiation back toward space than the darker-colored material does, allowing it to stay cooler. The darker-colored material absorbs more solar radiation, increasing its temperature. Just as the white shirt has a higher albedo than the black shirt, light-colored objects in nature (such as snow) have a higher albedo than dark-colored objects (such as soot or dust). As humans increase the amount of carbon black, soot, dust, and particulates in the atmosphere, we decrease the albedo of light-colored surfaces, causing them to absorb more solar radiation and become warmer than they would without human influence. An example of this can be seen in the snow on Figure 7.9.3.



Figure 7.9.3: A photograph of the extreme dust deposition from the deserts of the Colorado Plateau onto the Colorado Rockies snowpack in 2009. Taken from the high point of the Senator Beck Basin in the San Juan Mountains, it captures the extent of the impact of darkening in which the snow albedo dropped to about 30%, more than doubling the absorption of sunlight. Credit: S. McKenzie Skiles, Snow Optics Laboratory, NASA/JPL

Consequences of Climate Change

We will only discuss some of the consequences of climate change in this section, including changes in temperature, precipitation, ocean level, and ocean acidity. There are many more changes that have been seen, and are projected to continue in the future. These include: changes in the amount and distribution of ice and snow, changes in seasonality, ecosystem shift, and habitat changes of plant and animal populations, in addition to others. For more information about these consequence of climate change, visit this site: http://www.epa.gov/climatechange/sci...ors/index.html.

Temperature and precipitation

Temperature and **precipitation** are the two most direct impacts on the Earth's climate due to climate change. By now, you should already understand why an increase in greenhouse gas levels in the atmosphere causes an increase in temperature. But why does

it also impact precipitation patterns? As you already know, water vapor is an important component of the Earth's atmosphere . As the air in the troposphere warms and cools, the amount of water vapor that it holds changes dramatically. Here in Georgia, we have very hot and humid summers. The high summer humidity in this region is possible due to the increased capability warm air has to hold water vapor. Simply put, warmer air can hold more water than cooler air. As air cools, its ability to hold water vapor decreases, and any excess water will leave the air as liquid water. A great example of this is the formation of dew on surfaces overnight. During the day, the temperature is warmer than it is at night, and the air has a relatively high holding capacity for water vapor. When the sun sets, the air cools, decreasing its capacity to hold water vapor. That extra water must go somewhere, and it does that by accumulating on surfaces. Similarly, when warm and cool air fronts collide, the chances for rain and thunderstorms increase. Furthermore, an increase in temperature enhances evaporation occurring at the Earth's surface. This increased evaporation leads to greater concentrations of water vapor in the atmosphere which can lead to increased precipitation.

The change in temperature that we have already seen in the Earth's average atmospheric temperature is relatively small (about1.2 °C, according to Figure 7.5.1). However, as with many of the aspects of climate change, the potential for greater changes increases dramatically as time progresses in the future. This can be seen in Figure 7.9.4, which displays a model of the predicted temperature increase. Notice that these changes occur relatively rapidly, and are not uniform across the globe. What might be some of the reasons for this?

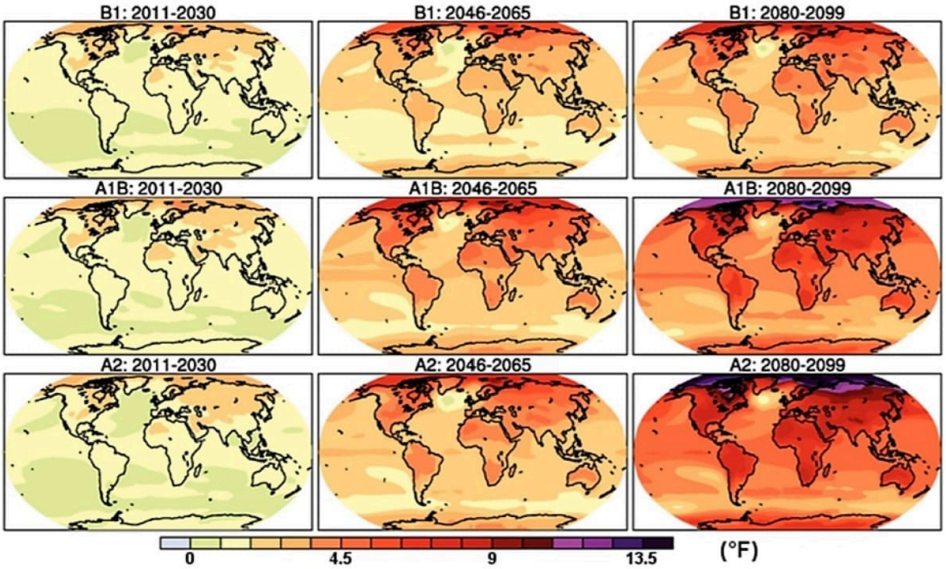


Figure 7.9.4: Projected changes in global average temperatures under three emissions scenarios (rows) for three different time periods (columns). Changes in temperatures are relative to 1961- 1990 averages. The scenarios come from the IPCC Special Report on Emissions Scenarios: B1 is a low emissions scenario, A1B is a medium-high emissions scenario, and A2 is a high emissions scenario. Source: IPCC Working Group I: The Physical Science Basis, 2007.

Changes in precipitation occur due to a variety of factors, including changes in atmospheric water vapor content due to changing temperature, as discussed above. Also at play is the heightened **evaporation** rate of water on Earth's surface under warmer temperatures. More evaporation leads to more precipitation. Finally, shifts in wind patterns impact the distribution of precipitation events. As you can see in Figure 7.9.5, there are some areas of the globe that are expected to have an increase in precipitation, while others are expected to have a dramatic decrease. Some major population centers projected to have a moderate to severe precipitation increase include (population estimates of the metropolitan area given in parentheses): New York, United States (20.1 million); Bogotá, Colombia (12.1 m.); and Manila, Philippines (11.9 m.). What sort of challenges might these cities face in the future as they deal with this change in their climate?

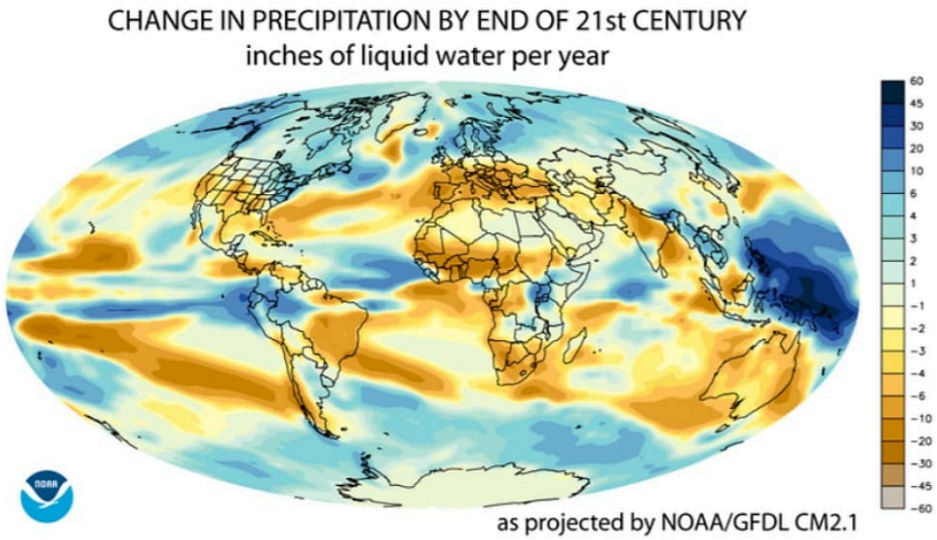


Figure 7.9.5 The plotted precipitation differences were computed as the difference between the 2081 to 2100 20-year averages minus the 1951 to 2000 50-year average. Blue areas project increases in precipitation; brown areas project decreases.

In contrast, many more major metropolitan areas are projected to have a moderate to severe precipitation decrease (droughts) by the end of the 21st century. These include Delhi, India (21.8 m.); Lagos, Nigeria (21 m.); São Paulo, Brazil (20.9 m.); Kolkata, India (14.6 m.); Istanbul, Turkey (14.4 m.); Los Angeles, United States (13.3 m.); Rio de Janeiro, Brazil (12 m.); Paris, France (12 m.); and Lahore, Pakistan (11.3 m.). The largest challenge that these areas are likely to face is a dwindling water supply for drinking and agriculture. See Chapter 8 for more detail on challenges faced by societies to supply clean, reliable water to their populations and farms.

Additional challenges may be felt by all areas of the world with regard to changes in the seasonality or timing of precipitation, as well the form in which precipitation falls (e.g., mist or downpour; rain, ice, or snow). All of these factors affect the availability of soil water for plants, the flow of rivers and streams, and the overall accessibility of water worldwide. Furthermore, scientists predict an increase in the number and severity of storms as climate change progresses. For a full discussion of the potential impacts of this, see the assigned article.

Sea level rise While we know that water continuously cycles around the world, and that the overall quantity of water on Earth will not change due to global climate change, the distribution of this water is changing. In particular, oceans are increasing in volume while land ice stores (such as **glaciers**) are decreasing. This contributes to an increase in sea level worldwide (Figure 7.9.6).

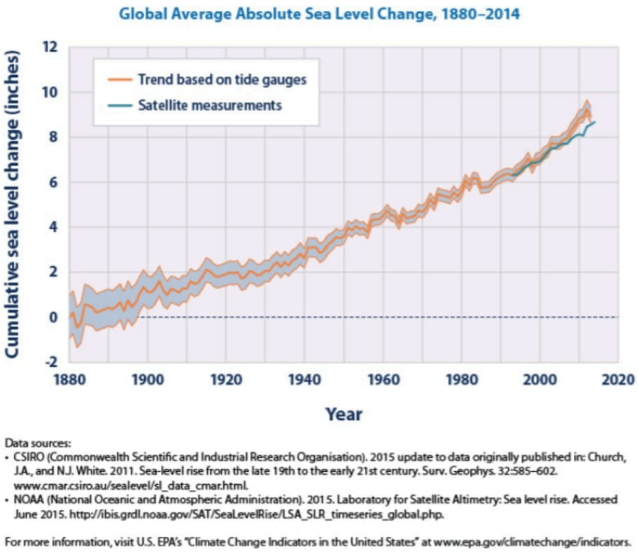


Figure 7.9.6: *This graph shows average absolute cumulative changes in sea level for the world's oceans since 1880, based on a combination of tide gauge measurements and recent satellite measurements. The shaded band shows the likely range of values, based on the number of measurements collected and the precision of the methods used.*

From the data in Figure 7.9.6, we see that sea level has increased at an average of 0.06 inches (0.15 cm) per year over the time period shown above. Most of this rise, however, has occurred within the most recent decades. The rate of increase has gone up to between 0.11 to 0.14 inches (0.28 to 0.36 cm) per year since 1993. There are two forces causing sea level to rise, both caused by climate change. First, the increased global temperature has caused increased ice melting in many regions of the globe. Melting **land ice** (such as the glacier shown in Figure 7.9.7) contributes to sea level rise because water that used to be stored in ice sitting on top of land becomes running water which reaches the ocean through **runoff.** We also observe **sea ice** melting (see http://www.epa.gov/climatechange/science/indicators/index.html for data and figures). Sea ice, such as the ice that covers the arctic regions of the Northern Hemisphere, has no land underneath it. When it melts, the water stays in the same locations, and the overall sea level does not change.



Figure 7.9.7: Glacier Retreat

The second factor that influences sea level rise is a phenomenon called thermal expansion. Due to the physical properties of water, as water warms, its density decreases. A less dense substance will have fewer molecules in a given area than a more dense substance (see Chapter 1 supplemental material). This means that as the overall temperature of the oceans increases due to global climate change, the same amount of water molecules will now occupy a slightly larger volume. This may not seem significant, but considering the 1.3 billion trillion liters (264 billion gallons) of water in the ocean, even a small change in density

can have large effects on sea level as a whole.

Scientists have already documented sea level rise in some areas of the world, including one familiar to most of us: the Southeastern United States. Figure 7.9.8 depicts the measured land area lost due to increasing sea level since 1996. Note that the Southeast (defined here as the Atlantic coast of North Carolina south to Florida) is particularly susceptible to land area loss due to the gently sloping nature of our coastline. Moving northward into the Mid-Atlantic States (defined here as Virginia north to Long Island, New York), coastal habitats tend to have a steeper geography, which protects against some losses.

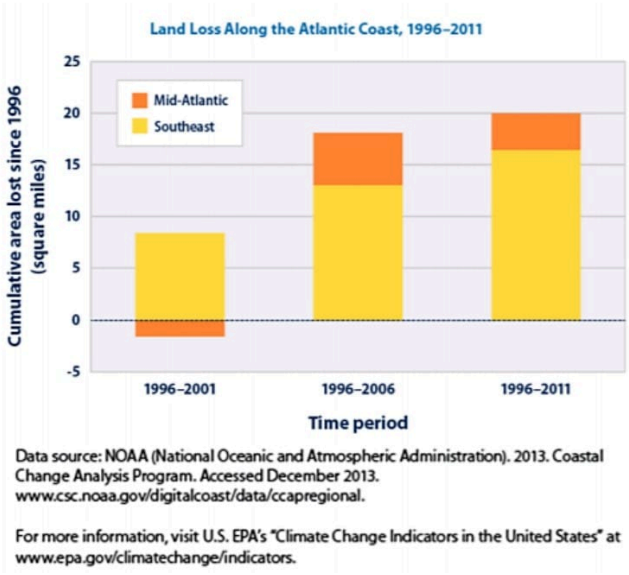


Figure 7.9.8: This graph shows the net amount of land converted to open water along the Atlantic coast during three time periods: 1996–2001, 1996–2006, and 1996–2011. The results are divided into two regions: the Southeast and the Mid-Atlantic. Negative numbers show where land loss is outpaced by the accumulation of new land.

While the ecological effects of sea level rise remain in the United States, we don't project any catastrophic loss of life, property, or livelihood for some time. This is, in part, due to large investments that we have made in infrastructure to protect our cities and farmlands. This is not the case in many areas of the world. For a discussion of the impacts of sea level rise on less industrialized nations of Bangladesh, Maldives, Kiribati, and Fiji, review the required article reading.

Ocean acidification

Dissolved CO_2 is essential for many organisms, including shell-building animals and other organisms that form a hard coating on their exterior (e.g., shellfish, corals, Haptophyte algae). This hard coating is built out of aragonite, a mineral form of the molecule calcium carbonate, $CaCO_3$. These organisms rely on the formation of carbonate ions (see Chapter 1 supplemental material for information on ions), CO_3^{2-}, from dissolved CO_2, through a natural, chemical reaction that occurs. This takes place through a chain-reaction equation, where bicarbonate (HCO_3^-) is formed as an intermediate, and hydrogen ions (H^+) are generated (equations 7.9.1 and 7.9.2).

$$CO_2 + H_2O \leftrightarrow H^+ + HCO_3^-$$

$$HCO_3^- \leftrightarrow H^+ + CO_3^{2-}$$

To have a better visualization of this process, follow along with the interactive graphic at: http://www.whoi.edu/home/oceanus_ima...ification.html.

As you can see, both equations 7.9.1 and 7.9.2 each produce one H^+. This is significant to water chemistry because an increase in H+ concentration means a decrease in the pH of the water. You can see in Figure 7.9.9 that a lower pH means that the liquid is more acidic. As shown in the interactive graphic, an increase in CO_2 in the atmosphere causes additional CO_2 to be dissolved in the ocean. This means that more CO_2 in the atmosphere leads to more acidic ocean environments.

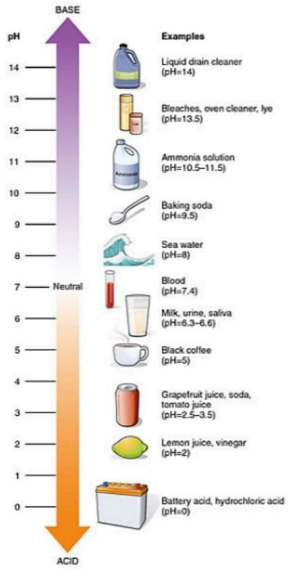


Figure 7.9.9: The pH scale and relative acidity. Illustration from Anatomy & Physiology, Connexions Web site. http://cnx.org/content/col11496/1.6/, Jun 19, 2013.

Unfortunately for shell-building animals, the buildup of H^+ in the more acidic ocean environment blocks the absorption of calcium and CO_3^{2-}, and makes the formation of aragonite more difficult. An aragonite deficit is already being documented in many of the world's oceans, as shown in Figure 7.9.10.

The increasing acidity of the world's oceans is resulting in habitat changes across the globe. This is only expected to worsen as atmospheric CO_2 levels continue to increase. Many organisms, including the corals that are the foundation species of the beautiful coral reefs, are very sensitive to changes in ocean pH. Scientists have documented cases of ecosystem destruction through coral bleaching, caused by the effects of climate change including ocean acidification and increased temperature. For more information, visit the NOAA Coral Reef Conservation Program website: coralreef.noaa.gov/threats/climate/.

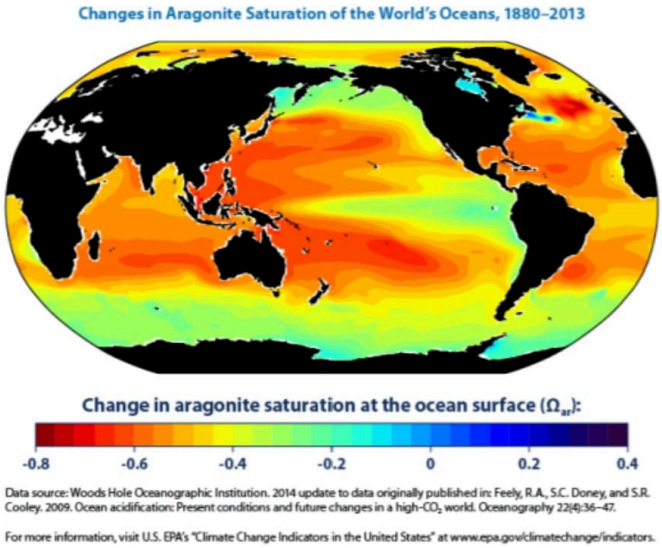


Figure 7.9.10: This map shows changes in the aragonite saturation level of ocean surface waters between the 1880s and the most recent decade (2004–2013). Aragonite is a form of calcium carbonate that many marine animals use to build their skeletons and shells. A negative change represents a decrease in saturation.

Climate Strategies

While the situation surrounding global climate change is in serious need of our attention, it is important to realize that many scientists, leaders, and concerned citizens are making solutions to climate change part of their life's work. The two solutions to the problems caused by climate change are **mitigation** and **adaptation**, and we will likely need a combination of both in order to prosper in the future.

Adaptation strategies

We know that climate change is already occurring, as we can see and feel the effects of it. For this reason, it is essential to also adapt to our changing environment. This means that we must change our behaviors in response to the changing environment around us. Some adaptation strategies are discussed in the required article reading.

Adaption strategies will vary greatly by region, depending on the largest specific impacts in that area. For example, in the city of Delhi, India, a dramatic decrease in rainfall is projected over the next century (Figure 7.6.2). This city will likely need to implement policies and practices relating to conservation of water, for example: rainwater harvesting, water re-use, and increased irrigation efficiency. Rain-limited cities near oceans, such as Los Angeles, California may choose to use desalination to provide drinking water to their citizens. **Desalination** involves taking the salt out of seawater to make it potable (Chapter 14).

Cities with low elevations near oceans may need to implement adaptation strategies to rising sea levels, from seawalls and levees to relocation of citizens. One adaptations strategy gaining use is the creation or conservation of **wetlands**, which provide natural protection against storm surges and flooding.

Mitigation strategies

In general, a strategy to mitigate climate change is one that reduces the amount of greenhouse gases in the atmosphere or prevents additional emissions. Mitigations strategies attempt to "fix" the problems caused by climate change. Governmental regulations regarding fuel efficiency of vehicles is one example of an institutionalized mitigation strategy already in place in the United States and in many other countries around the world. Unlike some other countries, there are no **carbon taxes** or charges on burning fossil fuels in the United States. This is another governmental mitigation strategy that has been shown to be effective in many countries including India, Japan, France, Costa Rica, Canada, and the United Kingdom.

In addition to government measures and incentives, technology can also be harnessed to mitigate climate change. One strategy for this is the use of **carbon capture and sequestration** (CCS). Through CCS, 80-90% of the CO_2 that would have been emitted to the atmosphere from sources such as a coal-fired power plant is instead captured and then stored deep beneath the Earth's surface. The CO_2 is often injected and sequestered hundreds of miles underground into porous rock formations sealed below an impermeable layer, where it is stored permanently (Figure 7.9.11).

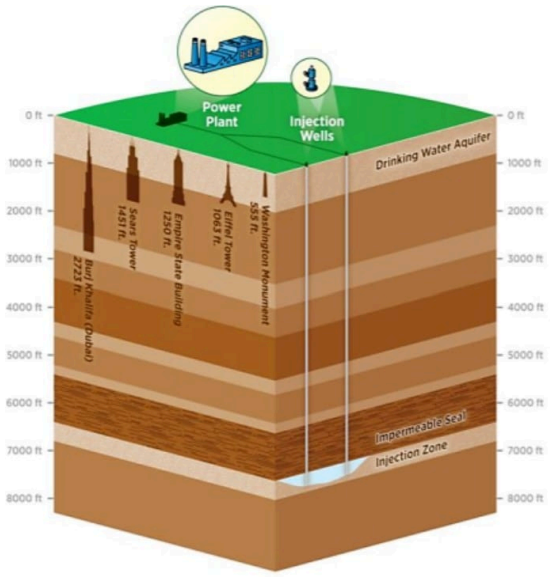


Figure 7.9.11: Carbon capture and sequestration schematic with landmarks shown to scale for depth reference. Source: US EPA.

Scientists are also looking into the use of soils and vegetation for carbon storage potential. Proper management of soil and forest ecosystems has been shown to create additional carbon sinks for atmospheric carbon, reducing the overall atmospheric CO_2 burden. Increasing soil carbon further benefits communities by providing better-quality soil for agriculture and cultivation.

Technologies related to alternative energy sources (Chapter 15) mitigate climate change by providing people with energy not derived from the combustion of fossil fuels. Finally, simple activities such as energy conservation, choosing to walk or bike instead of driving, and disposing of waste properly are activities that, when done by large numbers of people, actively mitigate climate change by preventing carbon emissions.

Take a moment to identify ways that you personally can be involved in the mitigation of or adaptation to climate change. What changes can you make in your own life to prevent excess carbon emissions? Similar to your ecological footprint, which you should have already calculated in lab, you can also calculate your carbon footprint. Use the EPA's carbon footprint calculator to do so, and investigate the Reduce Your Emissions section to find ways to decrease your carbon footprint.

Recommendations from the US Environmental Protection Agency

Figure 7.9.12 list various technologies and approaches that companies and individuals can adapt to reduce greenhouse gases. Technologies related to alternative energy sources mitigate climate change by providing people with energy not derived from the combustion of fossil fuels. Finally, simple activities such as energy conservation, choosing to walk or bike instead of driving, and disposing of waste properly are activities that, when done by large numbers of people, actively mitigate climate change by preventing carbon emissions.

Solutions for Transportation Air Pollution

Sources of air pollution cause: PM = soot = lung problems; CO2 = greenhouse gas = climate change; CO, NOx, SOx, & VOC = smog = asthma and poor air quality/ visibility. Solutions provide emission reductions that equal human and environmental health gains.

Figure 7.9.6 *Sources and solutions for transportation air pollution. Image: EPA*

Solutions for transportation air pollution, emission reductions, can lead to cleaner air and better health.

- Catalytic converters in conjunction with unleaded gasoline and low sulfur levels significantly reduce hydrocarbon and nitrogen oxide emissions.
- Fuel standards reduce exposure to pollutants like lead and benzene. Renewable fuels reduce CO_2 emissions.
- Engine technologies like computer controls, variable valve timing, multi-valve engines, turbo charging and gasoline direct injection improve fuel economy and reduce CO_2 emissions.
- Transmission technologies like 7+ speeds, dual clutch transmissions, (DCTs), and continuously variable transmissions (CVTs) improve fuel economy and reduce CO_2 emissions.
- Diesel filters reduce particulate matter from on road and off road diesel engines. Alternative vehicle technologies like plug-in electric vehicles and fuel cells equals zero tailpipe emissions.
- Better transportation planning for passengers and freight reduce emissions and fuel use.

Summary

- Gases that trap heat in the atmosphere are called greenhouse gases.
- The greenhouse effect is the process by which radiation from a planet's atmosphere warms the planet's surface to a temperature above what it would be without this atmosphere.
- Global warming refers to the increase in the average temperature of the Earth’s atmosphere due to elevated greenhouse gas concentrations, heightening the greenhouse effect.
- Climate change includes both global warming driven by human-induced emissions of greenhouse gasesand the resulting large-scale shifts in weather patterns.
- Greenhouse gases differ in their ability to produce the greenhouse effect as measured by the global warming potential, or GWP. A molecule of CH_4 is about 28 times more effective than one of CO_2 at absorbing infrared radiation, while N_2O is 298 times more effective . The synthetic fluorinated gases are evn much higher (in the thousands). This is largely due to their very long residence time in the atmosphere.
- Different technologies and approaches recommended by the US EPA to to reduce human greenhouse emmissions were provided.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)

- Libretext: Introduction to Environmental Science (Zendher et al.)
- US EPA
- Wikipedia

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7.10: Who Pollutes? Who Pays?

Learning Objectives

- Know the EPA standards for criteria air pollutants.
- Know the sources, health, and environmental effects of criteria air pollutants.

The Clean Air Act, which was last amended in 1990, requires EPA to set National Ambient Air Quality Standards (40 CFR part 50) for pollutants considered harmful to public health and the environment. The Clean Air Act identifies two types of national ambient air quality standards. ***Primary standards*** provide public health protection, including protecting the health of "sensitive" populations such as asthmatics, children, and the elderly. ***Secondary standards*** provide public welfare protection, including protection against decreased visibility and damage to animals, crops, vegetation, and buildings.

The EPA has set National Ambient Air Quality Standards for six principal pollutants, which are called "criteria" air pollutants. Periodically, the standards are reviewed and may be revised. The current standards are listed in Table 7.10.1. Units of measure for the standards are parts per million (ppm) by volume, parts per billion (ppb) by volume, and micrograms per cubic meter of air ($\mu g/m^3$).

Table 7.10.1 National Ambient Air Quality Standards for Criteria Air Pollutants.

Pollutant [links to historical tables of NAAQS reviews]		Primary/ Secondary	Averaging Time	Level	Form
Carbon Monoxide (CO)		primary	8 hours	9 ppm	Not to be exceeded more than once per year
			1 hour	35 ppm	
Lead (Pb)		primary and secondary	Rolling 3 month average	0.15 $\mu g/m^3$ (1)	Not to be exceeded
Nitrogen Dioxide (NO_2)		primary	1 hour	100 ppb	98th percentile of 1-hour daily maximum concentrations, averaged over 3 years
		primary and secondary	1 year	53 ppb (2)	Annual Mean
Ozone (O_3)		primary and secondary	8 hours	0.070 ppm (3)	Annual fourth-highest daily maximum 8-hour concentration, averaged over 3 years
Particle Pollution (PM)	$PM_{2.5}$	primary	1 year	12.0 $\mu g/m^3$	annual mean, averaged over 3 years
		secondary	1 year	15.0 $\mu g/m^3$	annual mean, averaged over 3 years
		primary and secondary	24 hours	35 $\mu g/m^3$	98th percentile, averaged over 3 years
	PM_{10}	primary and secondary	24 hours	150 $\mu g/m^3$	Not to be exceeded more than once per year on average over 3 years
Sulfur Dioxide (SO_2)		primary	1 hour	75 ppb (4)	99th percentile of 1-hour daily maximum concentrations, averaged over 3 years

Pollutant [links to historical tables of NAAQS reviews]	Primary/ Secondary	Averaging Time	Level	Form
	secondary	3 hours	0.5 ppm	Not to be exceeded more than once per year

(1) In areas designated nonattainment for the Pb standards prior to the promulgation of the current (2008) standards, and for which implementation plans to attain or maintain the current (2008) standards have not been submitted and approved, the previous standards (1.5 µg/m3 as a calendar quarter average) also remain in effect.

(2) The level of the annual NO2 standard is 0.053 ppm. It is shown here in terms of ppb for the purposes of clearer comparison to the 1-hour standard level.

(3) Final rule signed October 1, 2015, and effective December 28, 2015. The previous (2008) O3 standards additionally remain in effect in some areas. Revocation of the previous (2008) O3 standards and transitioning to the current (2015) standards will be addressed in the implementation rule for the current standards.

(4) The previous SO2 standards (0.14 ppm 24-hour and 0.03 ppm annual) will additionally remain in effect in certain areas: (1) any area for which it is not yet 1 year since the effective date of designation under the current (2010) standards, and (2)any area for which an implementation plan providing for attainment of the current (2010) standard has not been submitted and approved and which is designated nonattainment under the previous SO2 standards or is not meeting the requirements of a SIP call under the previous SO2 standards (40 CFR 50.4(3)). A SIP call is an EPA action requiring a state to resubmit all or part of its State Implementation Plan to demonstrate attainment of the required NAAQS.

Table 7.10.2 list the sources and harmful effects of criteria pollutants.

Table 7.10.2 Sources, Health and Environmental Effects of Criteria Pollutants.

Pollutant	Sources	Health Effects	Environmental Effects
Ground-level Ozone (O_3)	Secondary pollutant typically formed by chemical reaction of volatile organic compounds (VOCs) and NO_x in the presence of sunlight.	Decreases lung function and causes respiratory symptoms, such as coughing and shortness of breath; aggravates asthma and other lung diseases leading to increased medication use, hospital admissions, emergency department (ED) visits, and premature mortality.	Ozone affects sensitive vegetation and ecosystems, including forests, parks, wildlife refuges and wilderness areas. In particular, ozone harms sensitive vegetation during the growing season.

Particulate Matter (PM)	Emitted or formed through chemical reactions; fuel combustion (e.g., burning coal, wood, diesel); industrial processes; agriculture (plowing, field burning); and unpaved roads.	Short-term exposures can aggravate heart or lung diseases leading to respiratory symptoms, increased medication use, hospital admissions, ED visits, and premature mortality; long-term exposures can lead to the development of heart or lung disease and premature mortality.	Visibility impairment Fine particles ($PM_{2.5}$) are the main cause of reduced visibility (haze) in parts of the United States, including many of our treasured national parks and wilderness areas. Environmental damage Particles can be carried over long distances by wind and then settle on ground or water. Depending on their chemical composition, the effects of this settling may include: · making lakes and streams acidic · changing the nutrient balance in coastal waters and large river basins · depleting the nutrients in soil · damaging sensitive forests and farm crops · affecting the diversity of ecosystems · contributing to acid rain effects. Materials damage PM can stain and damage stone and other materials, including culturally important objects such as statues and monuments. Some of these effects are related to acid rain effects on materials.
Lead	Smelters (metal refineries) and other metal industries; combustion of leaded gasoline in piston engine aircraft; waste incinerators; and battery manufacturing.	Damages the developing nervous system, resulting in IQ loss and impacts on learning, memory, and behavior in children. Cardiovascular and renal effects in adults and early effects related to anemia.	Lead is persistent in the environment and can be added to soils and sediments through deposition from sources of lead air pollution. Other sources of lead to ecosystems include direct discharge of waste streams to water bodies and mining. Elevated lead in the environment can result in decreased growth and reproductive rates in plants and animals, and neurological effects in vertebrates.
Oxides of Nitrogen (NO_x)	Fuel combustion (e.g., electric utilities, industrial boilers, and vehicles) and wood burning.	Aggravate lung diseases leading to respiratory symptoms, hospital admissions, and ED visits; increased susceptibility to respiratory infection.	NO_2 and other NO_x interact with water, oxygen and other chemicals in the atmosphere to form acid rain. Acid rain harms sensitive ecosystems such as lakes and forests. NO_x in the atmosphere contributes to nutrient pollution in coastal waters.

Carbon Monoxide (CO)	Fuel combustion (especially vehicles), industrial processes, fires, waste combustion, and residential wood burning.	Reduces the amount of oxygen reaching the body's organs and tissues; aggravates heart disease, resulting in chest pain and other symptoms leading to hospital admissions and ED visits.	
Sulfur Dioxide (SO_2)	Fuel combustion (especially high-sulfur coal); electric utilities and industrial processes; and natural sources such as volcanoes.	Aggravates asthma and increased respiratory symptoms. Contributes to particle formation with associated health effects.	At high concentrations, gaseous SOx can harm trees and plants by damaging foliage and decreasing growth. SO_2 and other sulfur oxides can contribute to acid rain which can harm sensitive ecosystems.

Source: www.epa.gov

The likelihood of immediate reactions to air pollutants depends on several factors. Age and preexisting medical conditions are two important influences. Some sensitive individuals appear to be at greater risk for air pollution-related health effects, for example, those with pre-existing heart and lung diseases (e.g., heart failure/ischemic heart disease, asthma, emphysema, and chronic bronchitis), diabetics, older adults, and children. In other cases, whether a person reacts to a pollutant depends on individual sensitivity, which varies tremendously from person to person. Some people can become sensitized to biological pollutants after repeated exposures, and it appears that some people can become sensitized to chemical pollutants as well.

Paying the Price

Pollution has a cost. Manufacturing activities that cause air pollution impose health and clean-up costs on the whole of society, whereas the neighbors of an individual who chooses to fire-proof his home may benefit from a reduced risk of a fire spreading to their own homes. A manufacturing activity that causes air pollution is an example of a negative externality in production. A negative externality in production occurs "when a firm's production reduces the well-being of others who are not compensated by the firm." For example, if a laundry firm exists near a polluting steel manufacturing firm, there will be increased costs for the laundry firm because of the dirt and smoke produced by the steel manufacturing firm. If external costs exist, such as those created by pollution, the manufacturer will choose to produce more of the product than would be produced if the manufacturer were required to pay all associated environmental costs. Because responsibility or consequence for self-directed action lies partly outside the self, an element of externalization is involved. If there are external benefits, such as in public safety, less of the good may be produced than would be the case if the producer were to receive payment for the external benefits to others. However, goods and services that involve negative externalities in production, such as those that produce pollution, tend to be over-produced and underpriced since the externality is not being priced into the market.

Pollution can also create costs for the firms producing the pollution. Sometimes firms choose, or are forced by regulation, to reduce the amount of pollution that they are producing. The associated costs of doing this are called abatement costs, or marginal abatement costs if measured by each additional unit. In 2005 pollution abatement capital expenditures and operating costs in the US amounted to nearly $27 billion.

Summary

- The EPA has set National Ambient Air Quality Standards for six principal pollutants, which are called "criteria" air pollutants.
- The six criteria pollutants are carbon monoxide, lead, nitrogen dioxide, ozone, particulate matter, and sulfur dioxide.
- The key air pollutants can cause various environmental and health problems that could affect the respiratory, nervous, and cardiovascular systems.

Contributors and Attributions

- US EPA

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8: WATER

8.1: Water - Some Unique Properties

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8.4: Groundwater Contamination to Tainted Tap Water

8.5: Water Who Uses It and How Much?

8.6: Making Water Fit to Drink

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8.1: Water - Some Unique Properties

LEARNING OBJECTIVE

Describe the different properties of water as it relates to its polarity and ability to form hydrogen bonds.

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 is a simple molecule consisting of one oxygen atom bonded to two different hydrogen atoms (Figure $8.1.1$). Because of the higher electronegativity of the oxygen atom, the bonds are polar covalent (**polar bonds**). The oxygen atom attracts the shared electrons of the covalent bonds to a significantly greater extent than the hydrogen atoms. As a result, the oxygen atom requires a partial negative charge $(\delta-)$, while the hydrogen atoms each acquire a partial positive charge $(\delta+)$. The molecule adopts a bent structure because of the two lone pairs of electrons on the oxygen atom. The $\mathrm{H-O-H}$ bond angle is about 105°, slightly smaller than the ideal 109.5° of an sp^3 hybridized atomic orbital.The bent shape of the water molecule is critical because the polar $\mathrm{O-H}$ bonds do not cancel one another and the molecule as a whole is polar.

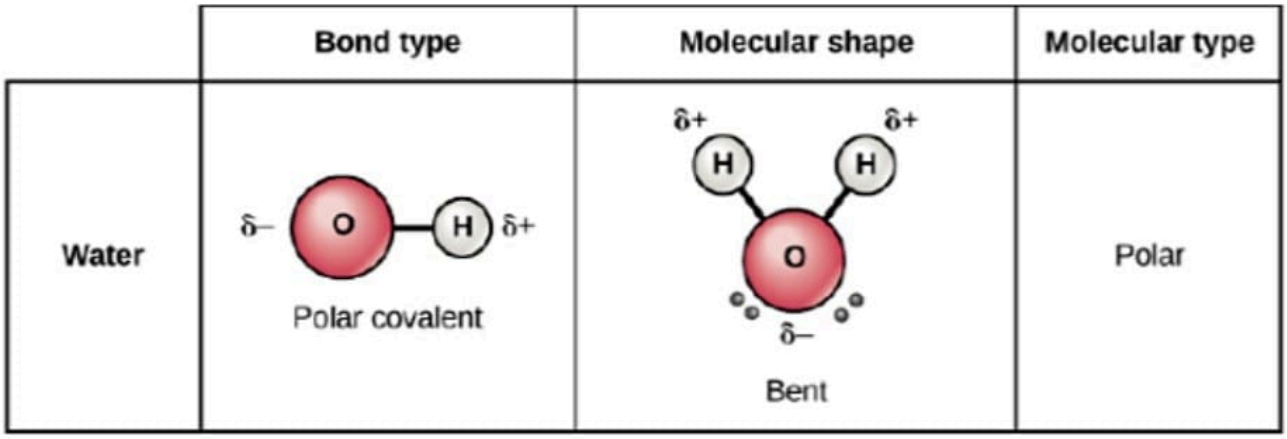


Figure $8.1.1$ *Polarity of the water molecule due to the uneven distribution of electrons in its covalent bond. C (From OpenStax Concepts of Biology text)*

Hydrogen Bonds

Due to water's polarity, each water molecule attracts other water molecules as oppositely charged ends of the molecules attract each other. When this happens, a weak interaction occurs between the *positive* hydrogen end from one molecule and the *negative oxygen* end of another molecule. This interaction is called a **hydrogen bond**. This hydrogen bonding contributes to the following water's unique properties.

1. Water is the universal solvent

2. Exists in nature as a solid, liquid, and gas

3. The density of ice is less than liquid water

4. Water has a high heat capacity

5. Water has a high heat of vaporization

6. Water exists as a liquid at room temperature It is important to note here that even we are only focusing on water in this text book, hydrogen bonding also occurs in other substances that have polar molecules.

Density of Water

Liquid water is a fluid. The hydrogen bonds in liquid water constantly break and reform as the water molecules tumble past one another. As water cools, its molecular motion slows and the molecules move gradually closer to one another. The **density** of any liquid increases as its temperature decreases. For most liquids, this continues as the liquid freezes and the solid state is denser than the liquid state. However, water behaves differently (Table $8.1.1$) . It actually reaches its highest density at about $4^{\circ}\mathrm{C}$.

Table $8.1.1$ *Density of Water and Ice*

Temperature $\left(^{\circ}\mathrm{C}\right)$	Density $\left(\mathrm{g/cm^3}\right)$
100 (liquid)	0.9584
50	0.9881
25	0.9971
10	0.9997
4	1.0000
0 (liquid)	0.9998
0 (solid)	0.9168

Between $4^{\circ}\mathrm{C}$ and $0^{\circ}\mathrm{C}$, the density gradually decreases as the hydrogen bonds begin to form a network characterized by a generally **hexagonal** structure with open spaces in the middle of the hexagons (Figure $8.1.2$).

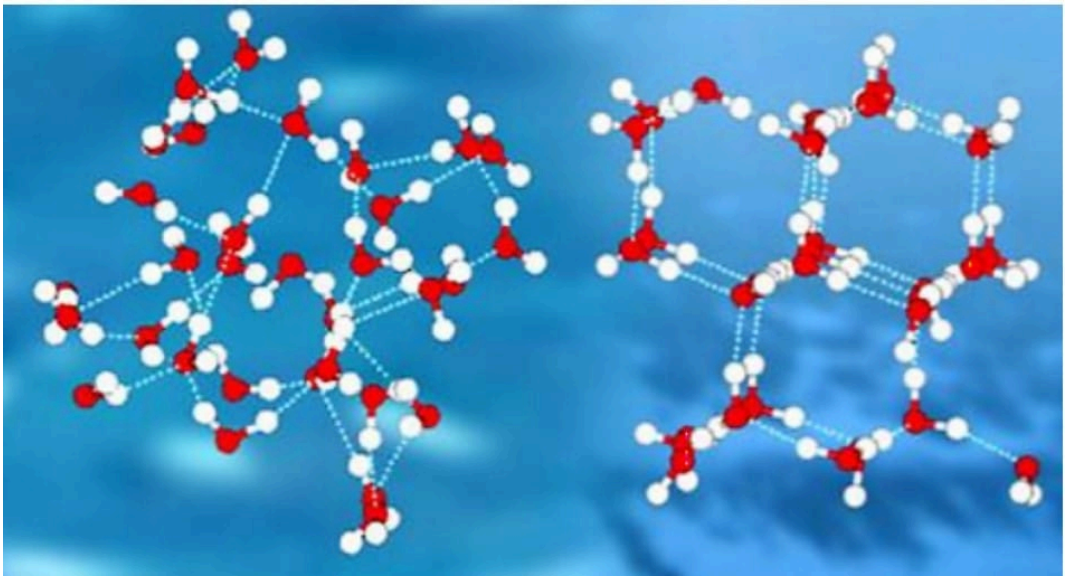


Figure $8.1.2$ *Hydrogen bonding makes ice less dense than liquid water. The lattice structure water is more condensed (left structure) than that of ice (right structure). The lattice structure of ice makes it less dense than freely flowing molecules of liquid water, enabling ice to float on liquid water. (Image credit: Lynn Yarris, http://www2.lbl.gov/Science-Articles...ter-solid.html)*

Ice is less dense than liquid water and so it floats. Ponds or lakes begin to freeze at the surface, closer to the cold air. A layer of ice forms, but does not sink as it would if water did not have this unique structure dictated by its shape, polarity, and hydrogen bonding. If the ice were to sink as it froze, entire lakes would freeze solid. Since the ice does not sink, liquid water remains under the ice all winter long. This is important, as fish and other organisms are capable of surviving through winter. Ice is one of only a very few solids that is less dense than its liquid form.

Solvation Ability of Water

Water typically dissolves many ionic compounds and polar molecules. Nonpolar molecules such as those found in grease or oil do not dissolve in water. We will first examine the process that occurs when an ionic compound such as table salt (sodium chloride) dissolves in water.

Water is attracted to the sodium chloride crystal because water is polar and has both a positive and a negative end. The positively charged sodium ions in the crystal attract the oxygen end of the water molecules because they are partially negative. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules because they are partially positive. The action of the polar water molecules takes the crystal lattice apart (see Figure $8.1.3$) .

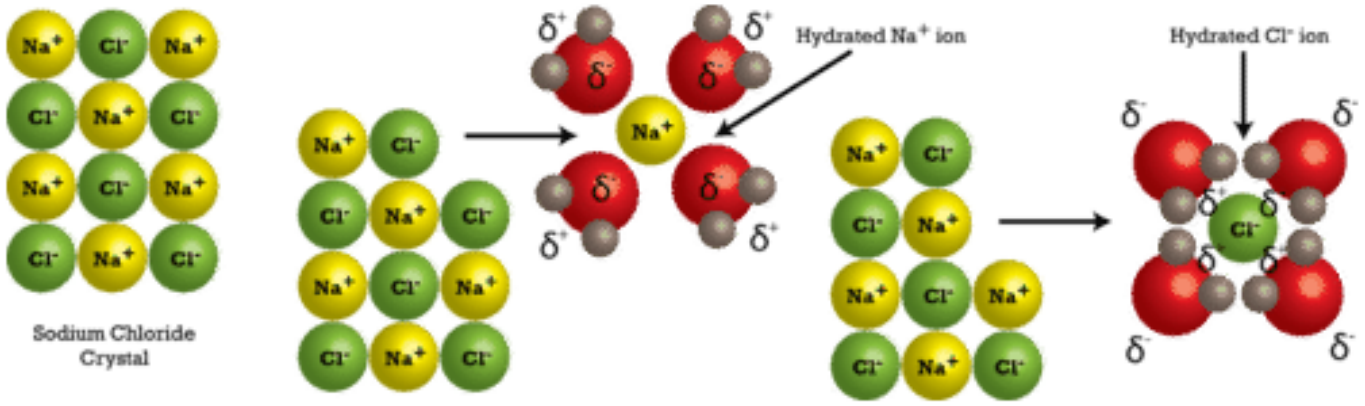


Figure $8.1.3$ *The dissolving of sodium chloride in water.*

After coming apart from the crystal, the individual ions are then surrounded by solvent particles in a process called **solvation**. Note that the individual $\mathrm{Na^+}$ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. **Hydration** is the process of solute particles being surrounded by water molecules arranged in a specific manner. Hydration helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and forming a precipitate.

Table sugar is sucrose $(\mathrm{C_{12}H_{22}O_{11}})$ and is an example of a molecular compound. Solid sugar consists of individual sugar molecules held together by intermolecular attractive forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the attractive forces, but does not break the covalent bonds between the carbon, hydrogen, and oxygen atoms. Dissolved sugar molecules are also hydrated, but without as distinct an orientation to the water molecules as in the case of the ions. The sugar molecules contain many $-\mathrm{OH}$ groups that can form hydrogen bonds with the water molecules, helping form the sucrose solution.

The Amphoteric Nature of Water

Water is amphoteric: it has the ability to act as either an acid or a base in chemical reactions. According to the Brønsted-Lowry definition, an acid is a proton (H^+) donor and a base is a proton acceptor. When reacting with a stronger acid, water acts as a base; when reacting with a stronger base, it acts as an acid. For instance, water receives an H^+ ion from HCl when hydrochloric acid is formed:

$$\mathrm{HCl(acid) + H_2O(base) \diamond H_3O^+ + Cl^-}$$

In the reaction with ammonia, NH_3, water donates a H^+ ion, and is thus acting as an acid:

$$\mathrm{NH_3(base) + H_2O(acid) \diamond NH^+_4 + OH^-}$$

High Heat Capacity and Specific Heat of Water

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water in the same sun will not become nearly as hot. Table $8.1.2$ list specific heats of various substances compared to water.Water has the highest **specific heat capacity** of any liquid. Water's high heat capacity is a property caused by hydrogen bonding among the water molecules. **Specific heat** is defined as the amount of heat one gram of a substance must absorb or lose to change its temperature by one degree Celsius. For water, this amount is **1 cal/g°C**. The units for specific heat can either be in the SI units of joules per gram per degree $(\mathrm{J/g^{\circ}C})$ or calories per gram per degree $(\mathrm{cal/g^{\circ}C})$. This text will use $\mathrm{J/g^{\circ}C}$ for specific heat.



Figure $8.1.4$ *This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmer water is pumped out of the plant and back into the lake.*

It takes water a long time to heat up and a long time to cool down. In fact, the specific heat capacity of water is about five times more than that of sand. This explains why land cools faster than the sea. Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days. Water is used as a coolant for machinery because it is able to absorb large quantities of heat (Figure $8.1.4$). Due to its high heat capacity, warm-blooded animals use water to disperse heat more evenly and maintain temperature in their bodies: it acts in a similar manner to a car's cooling system, transporting heat from warm places to cool places, causing the body to maintain a more even temperature.

Table $8.1.2$ *Specific Heats of Some Common Substances*

Substance	Specific Heat $(\mathrm{J/g^{\circ}C})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449

Substance	Specific Heat (J/g^oC)
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

High Heat of Vaporization of Water

Water in its liquid form has an unusually high boiling point temperature, a value close to 100°C. As a result of the network of hydrogen bonding present between water molecules, a high input of energy is required to transform one gram of liquid water into water vapor, an energy requirement called the **heat of vaporization.** Water has a heat of vaporization value of **40.65 kJ/mol**. A considerable amount of heat energy (586 calories) is required to accomplish this change in water. This process occurs on the surface of water. As liquid water heats up, hydrogen bonding makes it difficult to separate the water molecules from each other, which is required for it to enter its gaseous phase (steam). As a result, water acts as a heat sink, or heat reservoir, and requires much more heat to boil than does a liquid such as ethanol (grain alcohol), whose hydrogen bonding with other ethanol molecules is weaker than water's hydrogen bonding.

The fact that hydrogen bonds need to be broken for water to evaporate means that a substantial amount of energy is used in the process. As the water evaporates, energy is taken up by the process, cooling the environment where the evaporation is taking place. In many living organisms, including humans, the evaporation of sweat, which is 90 percent water, allows the organism to cool so that homeostasis of body temperature can be maintained.

Summary

- The polarity of water and its ability to hydrogen bond contributes to its unique properties.
- Ionic solute molecules are hydrated (surrounded by solvent molecules in a specific orientation).
- Ice is less dense than liquid water due to spaces in the intermolecular structure of ice not present in water.
- Heat capacity is the amount of heat required to raise the temperature of an object by 1^oC).
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by 1^oC.
- The dissociation of liquid water molecules, which changes the substance to a gas, requires a lot of energy.

Contributors and Attributions

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- Boundless: The Chemical Foundation of Life
- Libretext: Introduction to Environmental Science (Zendher et al.)
- TextMap: Introductory Chemistry (Tro)

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8.2: Water in Nature

Learning Objectives

- Know the distribution of earth's water.
- Describe the water (hydrologic) cycle.
- Know the different natural water contaminants

Water 's influence on the world around us is affected by its unique properties some of which where discussed in the previous section. 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.

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 8.2.1). 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**.

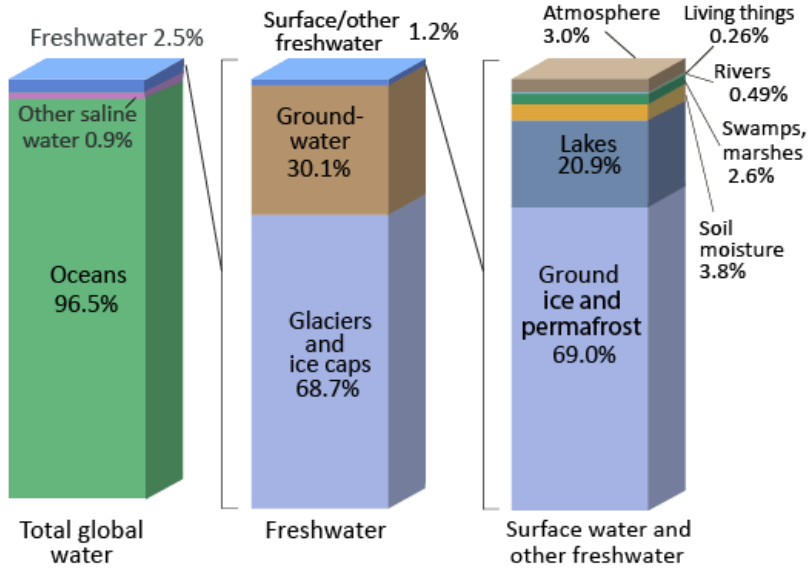


Figure 8.2.1 The distribution of Earth’s water. In the first bar, notice how only 2.5% of Earth's water is freshwater - the amount needed for life to survive. The middle bar shows the breakdown of freshwater. Almost all of it is locked up in ice and in the ground. Only a little more than 1.2% of all 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 Water (Hydrologic) Cycle and Natural Contaminants

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 **water cycle** (Figure 8.2.2).

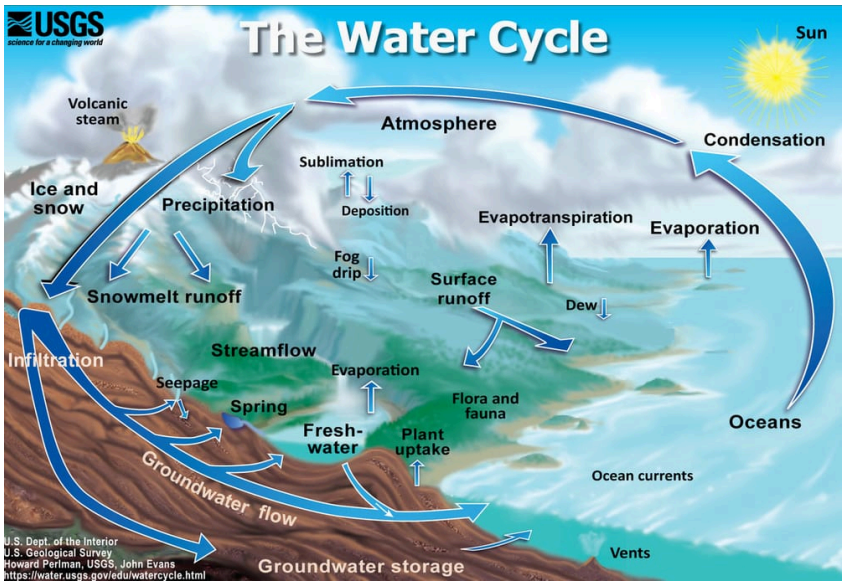


Figure 8.2.2 Because it is a cycle, the water cycle has no beginning and no end. Earth's water is always in movement, and the natural water cycle, also known as the hydrologic cycle, describes the continuous movement of water on, above, and below the surface of the Earth. Water is always changing states between liquid, vapor, and ice, with these processes happening in the blink of an eye and over millions of years.

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** and then precipitation. **Precipitation** can be rain, sleet, hail, or snow. At the surface, water or melted water, may eventually evaporate and reenter the atmosphere. A significant amount of water infiltrates into the ground. Water may seep through dirt and rock below the soil through pores infiltrating the ground to go into Earth's groundwater system (Figure Extra close brace or missing open brace). Groundwater enters **aquifers** which are bodies of rock or sediment that store (and yield) large amounts of usable water in their pores. Alternatively, the water may come to the surface through springs or find its way back to the oceans.

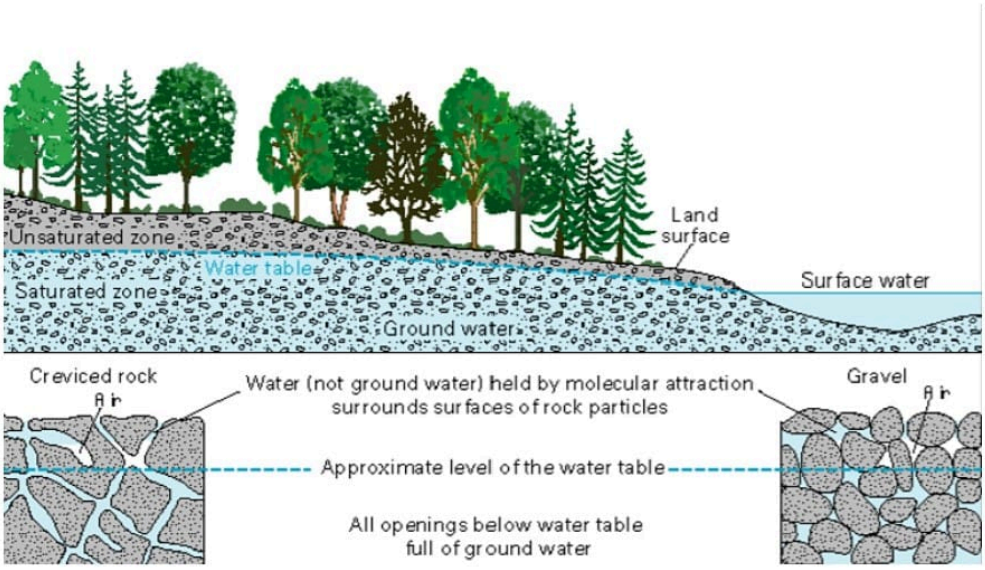


Figure Extra close brace or missing open brace *Model of groundwater system showing the different components of an unconfined groundwater system: http://water.usgs.gov/edu/earthgwaquifer.html*

Substances Present in Natural Waters

The amounts of water present in the atmosphere and on land (as surface runoff, lakes and streams) is great enough to make it a significant agent in transporting substances between the lithosphere (the rigid surface of the earth) and the oceans. Water

interacts with both the atmosphere and the lithosphere (the rigid surface of the earth) acquiring solutes from each, and thus provides the major chemical link between these two realms. The various transformations undergone by water through the different stages of the hydrologic cycle act to transport both dissolved and particulate substances between different geographic locations.

The Safe Drinking Water Act defines the term "**contaminant**" as meaning any physical, chemical, biological, or radiological substance or matter in water. Therefore, the law defines "contaminant" very broadly as being anything other than water molecules.

Gases

The carbon dioxide in natural water creates an interesting phenomenon. Rainwater is saturated with CO_2, and it dissolves limestones. When CO_2 is lost due to temperature changes or escaping from water drops, the reverse reaction takes place. The solid formed, however, may be a less stable phase called aragonite, which has the same chemical formula as, but a different crystal structure than that of calcite.

The rain dissolves calcium carbonate by the two reactions shown above. The water carries the ions with it, sips through the crack of the rocks. When it reached the ceiling of a cave, the drop dangles there for a long time before fallen. During this time, the carbon dioxide escapes and the pH of the water increases. Calcium carbonate crystals begin to appear. Calcite, aragonite, stalactite, and stalagmite are four common solids found in the formation of caves.

Dissolved Minerals

Table 8.2.1 list the major ions present in seawater. The composition does vary, depending on region, depth, latitude, and water temperature. Waters at the river mouths contain less salt. If the ions are utilized by living organism, its contents vary according to the populations of organisms.

Dust particles and ions present in the air are nucleation center of water drops. Thus, waters from rain and snow also contain such ions: Ca^{2+}, Mg^{2+}, Na^{+}, K^{+}, NH_4^{+}. These cations are balanced by anions, HCO_3^{-}, SO_4^{-}, NO_2^{-}, Cl^{-}, and NO_3^{-}. The pH of rain is between 5.5 and 5.6. Rain and snow waters eventually become river or lake waters. When the rain or snow waters fall, they interact with vegetation, top soil, bed rock, river bed and lake bed, dissolving whatever is soluble. Bacteria, algae, and water insects also thrive. Solubilities of inorganic salts are governed by the kinetics and equilibria of dissolution. The most common ions in lake and river waters are the same as those present in rainwater, but at higher concentrations. The pH of these waters depends on the river bed and lake bed. Natural waters contain dissolved minerals. Waters containing Ca^{2+} and Mg^{2+} ions are usually called **hard water**.

Table 1 Major Ions in Seawater (These values, expressed in parts per thousand, are for seawater of 35% salinity).

cations	g/kg	anions	g/kg
Na^{+}	10.77	Cl^{-}	19.354
Mg^{2+}	1.29	SO_4^{2-}	2.712
Ca^{2+}	0.412	Br^{-}	0.087
K^{+}	0.399		
Sr^{2+}	0.0079		
Al^{3+}	0.005		

Note

Although most elements are found in seawater only at trace levels, marine organisms may selectively absorb them and make them more detectable. Iodine, for example, was discovered in marine algae (seaweeds) 14 years before it was found in seawater. Other elements that were not detected in seawater until after they were found in marine organisms include barium, cobalt, copper, lead, nickel, silver and zinc. Si-32, presumably deriving from cosmic ray bombardment of Ar, has been discovered in marine sponges.

Hard Water

Minerals usually dissolve in natural water bodies such as lakes, rivers, springs, and underground waterways (ground waters). Calcium carbonate, $CaCO_3$, is one of the most common inorganic compounds in the Earth crust. It is the ingredient for both calcite and aragonite. These two minerals have different crystal structures and appearance. This photograph shows crystals of typical Calcite.

Calcium-carbonate minerals dissolve in water, with a solubility product as shown below.

$$CaCO_3 \rightleftharpoons Ca^{2+} + CO_3^{2-} \quad K_{sp} = 5 \times 10^{-9}$$

From the solubility product, we can (see example 1) evaluate the molar solubility to be 7.1×10^{-5} M or 7.1 mg/L (7.1 ppm of $CaCO_3$ in water). The solubility increases as the pH decrease (increase acidity). This is compounded when the water is saturated with carbon dioxide, CO_2. Saturated CO_2 solution contains carbonic acid, which help the dissolution due to the reaction:

$$H_2O + CO_2 \rightleftharpoons H_2CO_3$$

$$CaCO_3 + H_2CO_3 \rightleftharpoons Ca^{2+} + 2HCO_3^-$$

Because of these reactions, some natural waters contain more than 300 ppm calcium carbonates or its equivalents.

The composition of the ocean has attracted the attention of some of the more famous names in science, including Robert Boyle, Antoine Lavoisier and Edmund Halley. Their early investigations tended to be difficult to reproduce, owing to the different conditions under which they crystallized the various salts. As many as 54 salts, double salts and hydrated salts can be obtained by evaporating seawater to dryness. At least 73 elements are now known to be present in seawater.

Organic Matter

Most of the organic carbon in seawater is present as dissolved material, with only about 1-2% in particulates. The total organic carbon content ranges between 0.5 mg/L in deep water to 1.5 mg/L near the surface. There is still considerable disagreement about the composition of the dissolved organic matter; much of it appears to be of high molecular weight, and may be polymeric. Substances qualitatively similar to the humic acids found in soils can be isolated. The greenish color that is often associated with coastal waters is due to a mixture of fluorescent, high molecular weight substances of undetermined composition known as *"Gelbstoffe"*. It is likely that the significance of the organic fraction of seawater may be much greater than its low abundance would suggest. For one thing, many of these substances are lipid-like and tend to adsorb onto surfaces. It has been shown that any particle entering the ocean is quickly coated with an organic surface film that may influence the rate and extent of its dissolution or decomposition. Certain inorganic ions may be strongly complexed by humic-like substances. The surface of the ocean is mostly covered with an organic film, only a few molecular layers thick. This is believed to consist of hydrocarbons, lipids, and the like, but glycoproteins and proteoglycans have been reported. If this film is carefully removed from a container of seawater, it will quickly be reconstituted. How significant this film is in its effects on gas exchange with the atmosphere is not known.

Summary

- The water (hydrologic) cycle describes the continuous movement of water on, above and below the surface of the Earth. The water moves from one reservoir to another, such as from river to ocean, or from the ocean to the atmosphere, by the physical processes of evaporation, condensation, precipitation, infiltration, surface runoff, and subsurface flow. In doing so, the water goes through different forms: liquid, solid (ice) and capor.
- Different substances that can be found in natural waters include dissolved minerals, dissolved gases, and organic matter.

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8.3: Chemical and Biological Contamination

Learning Objective

- Describe the impact of human activities on water quality.

The global water crisis also involves water pollution. For water to be useful for drinking and irrigation, it must not be polluted beyond certain thresholds. According to the World Health Organization, in 2008 approximately 880 million people in the world (or 13% of world population) did not have access to safe drinking water. At the same time, about 2.6 billion people (or 40% of world population) lived without improved sanitation, which is defined as having access to a public sewage system, septic tank, or even a simple pit latrine. Each year approximately 1.7 million people die from diarrheal diseases associated with unsafe drinking water, inadequate sanitation, and poor hygiene. Almost all of these deaths are in developing countries, and around 90% of them occur among children under the age of 5 (Figure 1). Compounding the water crisis is the issue of social justice; poor people more commonly lack clean water and sanitation than wealthy people in similar areas. Globally, improving water safety, sanitation, and hygiene could prevent up to 9% of all disease and 6% of all deaths.

In addition to the global waterborne disease crisis, chemical pollution from agriculture, industry, cities, and mining threatens global water quality. In Gallup public polls conducted over the past decade Americans consistently put water pollution and water supply as the top environmental concerns over issues such as air pollution, deforestation, species extinction, and global warming.

Water pollution is the contamination of water by an excess amount of a substance that can cause harm to human beings and/or the ecosystem. The level of water pollution depends on the abundance of the pollutant, the ecological impact of the pollutant, and the use of the water. Pollutants are derived from biological, chemical, or physical processes. Although natural processes such as volcanic eruptions or evaporation sometimes can cause water pollution, most pollution is derived from human, land-based activities. Water pollutants can move through different water reservoirs, as the water carrying them progresses through stages of the water cycle.

Pollutants enter water supplies from **point sources**, which are readily identifiable and relatively small locations, or **nonpoint sources**, which are large and more diffuse areas (Figure $8.3.1$) . Point sources of pollution include animal factory farms that raise a large number and high density of livestock such as cows, pigs, and chickens. Also, pipes included are pipes from a factories or sewage treatment plants. Combined sewer systems that have a single set of underground pipes to collect both sewage and storm water runoff from streets for wastewater treatment can be major point sources of pollutants. During heavy rain, storm water runoff may exceed sewer capacity, causing it to back up and spilling untreated sewage directly into surface waters.

picture6



Figure *Point source and non point source pollution. The 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 (CC BY-SA 2.0; N Chadwick).*

This section and the next will focus on how human activities affect water quality.

Too Much Organic Matter Means Too Little Oxygen

Oxygen-demanding waste is an extremely important pollutant to ecosystems. Most surface water in contact with the atmosphere

has a small amount of dissolved oxygen, which is needed by aquatic organisms for cellular respiration. Bacteria decompose dead organic matter and remove dissolved oxygen (O_2) according to the following reaction:

$$\text{organic matter} + O_2 \rightarrow CO_2 + H_2O$$

Too much decaying organic matter in water is a pollutant because it removes oxygen from water, which can kill fish, shellfish, and aquatic insects. The amount of oxygen used by **aerobic** (in the presence of oxygen) bacterial decomposition of organic matter is called **biochemical oxygen demand** (BOD). The major source of dead organic matter in many natural waters is sewage; grass and leaves are smaller sources. An unpolluted water body with respect to BOD is a turbulent river that flows through a natural forest. Turbulence continually brings water in contact with the atmosphere where the O_2 content is restored. The dissolved oxygen content in such a river ranges from 10 to 14 ppm O_2, BOD is low, and clean-water fish such as trout. A polluted water body with respect to oxygen is a stagnant deep lake in an urban setting with a combined sewer system. This system favors a high input of dead organic carbon from sewage overflows and limited chance for water circulation and contact with the atmosphere. In such a lake, the dissolved O_2 content is ≤5 ppm O_2, BOD is high, and low O_2-tolerant fish, such as carp and catfish dominate.

Excessive plant nutrients, particularly nitrogen (N) and phosphorous (P), are pollutants closely related to oxygen-demanding waste. Aquatic plants require about 15 nutrients for growth, most of which are plentiful in water. N and P are called **limiting nutrients**, however, because they usually are present in water at low concentrations and therefore restrict the total amount of plant growth. This explains why N and P are major ingredients in most fertilizer. High concentrations of N and P from human sources (mostly agricultural and urban runoff including fertilizer, sewage, and phosphorus-based detergent) can cause cultural **eutrophication**, which leads to the rapid growth of aquatic producers, particularly algae. Thick mats of floating algae or rooted plants lead to a form of water pollution that damages the ecosystem by clogging fish gills and blocking sunlight. A small percentage of algal species produce toxins that can kill animals, including humans. Exponential growths of these algae are called **harmful algal blooms**. When the prolific algal layer dies, it becomes oxygen-demanding waste, which can create very low O_2 concentrations in the water (< 2 ppm O_2), a condition called **hypoxia.** This results in a **dead zone** because it causes death from asphyxiation to organisms that are unable to leave that environment. An estimated 50% of lakes in North America, Europe, and Asia are negatively impacted by cultural eutrophication. In addition, the size and number of marine hypoxic zones have grown dramatically over the past 50 years including a very large dead zone located offshore Louisiana in the Gulf of Mexico. Cultural eutrophication and hypoxia are difficult to combat, because they are caused primarily by nonpoint source pollution, which is difficult to regulate, and N and P, which are difficult to remove from wastewater.

Eutrophication is an increase in the concentration of chemical nutrients in an ecosystem to an extent that increases the primary productivity of the ecosystem. Depending on the degree of eutrophication, subsequent negative environmental effects such as anoxia (oxygen depletion) and severe reductions in water quality may occur, affecting fish and other animal populations.

Chemical Pollution from Waste

Since the 1990s, water contamination by pharmaceuticals has been an environmental issue of concern. In addition, it is important to note that many public health professionals in the United States began writing reports of pharmaceutical contamination in waterways in the 1970s." Most pharmaceuticals are deposited in the environment through human consumption and excretion, and are often filtered ineffectively by municipal sewage treatment plants which are not designed to manage them (Figure 8.3.2) . Once in the water, they can have diverse, subtle effects on organisms, although research is still limited. Pharmaceuticals may also be deposited in the environment through improper disposal, runoff from sludge fertilizer and reclaimed wastewater irrigation, and leaky sewer pipes. In 2009, an investigative report by Associated Press concluded that U.S. manufacturers had legally released 271 million pounds of compounds used as drugs into the environment, 92% of which was the industrial chemicals phenol and hydrogen peroxide, which are also used as antiseptics. It could not distinguish between drugs released by manufacturers as opposed to the pharmaceutical industry. It also found that an estimated 250 million pounds of pharmaceuticals and contaminated packaging were discarded by hospitals and long-term care facilities.



Figure 8.3.2 Illicit drugs such as ecstasy (above) can be found in waterways. (Public Domain; Willy turner via Wikipedia)

The use of pharmaceuticals and personal care products (PPCPs) is on the rise with an estimated increase from 2 billion to 3.9 billion annual prescriptions between 1999 and 2009 in the United States alone. Figure 8.3.3 illustrates how PPCPs enter into the environment through individual human activity and as residues from manufacturing, agribusiness, veterinary use, and hospital and community use. In Europe, the input of pharmaceutical residues via domestic waste water is estimated to be around 80% whereas 20% is coming from hospitals. Individuals may add PPCPs to the environment through waste excretion and bathing as well as by directly disposing of unused medications to septic tanks, sewers, or trash. Because PPCPs tend to dissolve relatively easily and do not evaporate at normal temperatures, they often end up in soil and water bodies.

Some PPCPs are broken down or processed easily by a human or animal body and/or degrade quickly in the environment . However, others do not break down or degrade easily. The likelihood or ease with which an individual substance will break down depends on its chemical makeup and the metabolic pathway of the compound.

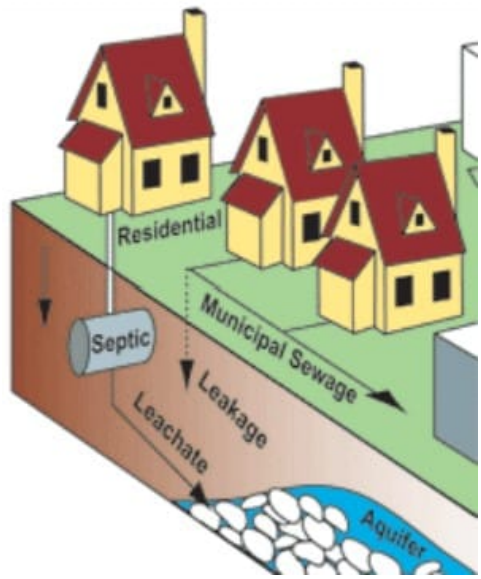


Figure 8.3.3 Methods of PPCP entry into the environment from residential homes via septic and sewage systems.[11]

While the full effects of most PPCPs on the environment are not understood, there is concern about the potential they have for harm because they may act unpredictably when mixed with other chemicals from the environment or concentrate in the food chain. Additionally, some PPCPs are active at very low concentrations, and are often released continuously in large or widespread quantities.

Because of the high solubility of most PPCPs, aquatic organisms are especially vulnerable to their effects. The increased presence of estrogen and other synthetic hormones in waste water due to birth control and hormonal therapies has been linked to increased feminization of exposed fish and other aquatic organisms. The chemicals within these PPCP products could either affect the feminization or masculinization of different fishes, therefore affecting their reproductive rates.

The major route for pharmaceutical residues to reach the aquatic environment is most probably by excretion from patients undergoing pharma treatment. Since many pharmaceutical substances are not metabolized in the body they may be excreted in biologically active form, usually via the urine. Furthermore, many pharmaceutical substances are not fully taken up from the intestine (following oral administration in patients) into their blood stream. The fraction not taken up into the blood stream will remain in the gut and eventually be excreted via the faeces. Hence, both urine and faeces from treated patients contain pharmaceutical residues. Between 30 and 90% of the orally administered dose is generally excreted as active substance in the urine.

An additional source to environmental pollution with pharmaceuticals is improper disposal of unused or expired drug residues. In European countries take-back systems for such residues are usually in place (although not always utilized to full extent) while in e.g. the US only voluntary initiatives on a local basis exist. Though most of the waste goes to incineration and people are asked to throw unused or expired pharmaceuticals into the household waste investigations in Germany showed that up to 24% of liquid pharmaceuticals and 7% of tablets or ointments are disposed always or at least “rarely” via the toilet or sink.

Proper destruction of pharma residues should yield rest products without any pharmaceutical or ecotoxic activity. Furthermore, the residues should not act as components in the environmental formation of new such products. Incineration at a high temperature (>1000 degrees Celsius) is considered to fulfill the requirements, but even following such incineration residual ashes from the incineration should be properly taken care of.

Pharmaceuticals used in veterinary medicine, or as additives to animal food, pose a different problem, since they are excreted into soil or possibly open surface waters. It is well known that such excretions may affect terrestrial organisms directly, leading to extinction of exposed species (e.g. dung-beetles). Lipid-soluble pharma residues from veterinary use may bind strongly to soil particles, with little tendency to leak out to ground water or to local surface waters. More water-soluble residues may be washed out with rain or melting snow and reach both ground water and surface water streams.

Water Borne Diseases

Waterborne diseases are conditions caused by pathogenic micro-organisms that are transmitted in water. These diseases can be spread while bathing, washing, drinking water, or by eating food exposed to contaminated water. While diarrhea and vomiting are the most commonly reported symptoms of waterborne illness, other symptoms can include skin, ear, respiratory, or eye problems. Waterborne diseases are impacted by a country's economy and also impact the economy by being costly to deal with.

Microorganisms causing diseases that characteristically are waterborne prominently include protozoa and bacteria, many of which are intestinal parasites, or invade the tissues or circulatory system through walls of the digestive tract. Various other waterborne diseases are caused by viruses. (In spite of philosophical difficulties associated with defining viruses as "organisms", it is practical and convenient to regard them as microorganisms in this connection.)

Yet other important classes of water-borne diseases are caused by metazoan parasites. Typical examples include certain Nematoda, that is to say "roundworms". As an example of water-borne Nematode infections, one important waterborne nematode disease is Dracunculiasis. It is acquired by swallowing water in which certain copepoda occur that act as vectors for the Nematoda. Anyone swallowing a copepod that happens to be infected with Nematode larvae in the genus Dracunculus, becomes liable to infection. The larvae cause guinea worm disease.

Another class of waterborne metazoan pathogens are certain members of the Schistosomatidae, a family of blood flukes. They usually infect victims that make skin contact with the water. Blood flukes are pathogens that cause Schistosomiasis of various forms, more or less seriously affecting hundreds of millions of people worldwide.

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 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

(Figure 8.3.4) 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.

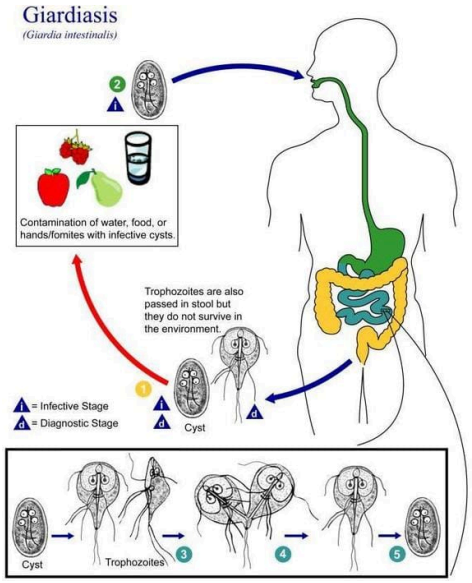


Figure 8.3.4 The life cycle of Giardia.

YouTube Video

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View this YouTube Video online

Acidic Waters

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 (Table 8.3.2). 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.

Table 8.3.2: Source of Water Contamination (Due to Acid Rain) 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

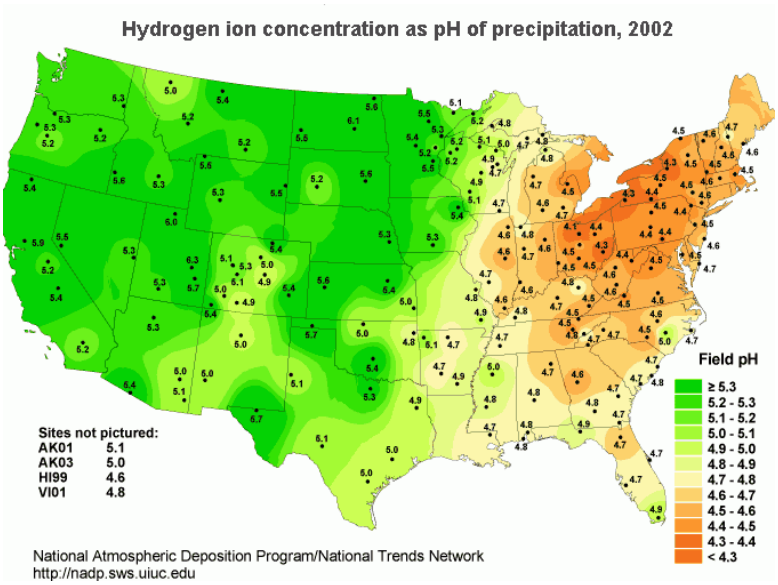


Figure 8.3.5: pH of rainwater over the continental United States.

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 8.3.5). 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.

Summary

- Water pollution is the contamination of water by an excess amount of a substance that can cause harm to human beings and/or the ecosystem.
- Pollutants enter water supplies from **point sources**, which are readily identifiable and relatively small locations, or **nonpoint sources**, which are large and more diffuse areas
- Organic matter as well as phosphates and nitrates from human and farm animal waste support the growth of algae and microorganisms, including bacteria.
- Eutrophication is an increase in the concentration of chemical nutrients in an ecosystem to an extent that increases the primary productivity of the ecosystem. Depending on the degree of eutrophication, subsequent negative environmental effects such as anoxia (oxygen depletion) and severe reductions in water quality may occur, affecting fish and other animal populations.
- Acid rain (formed as a consequence of air pollution) could affect building materials and living organisms on land and various bodies of water.

Contributors and Attributions

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8.4: Groundwater Contamination to Tainted Tap Water

Learning Objective

- Define groundwater contamination.
- List other groundwater contaminants and their sources.

Ground water contamination (also called groundwater pollution) occurs when pollutants are released to the ground and make their way down into groundwater. This type of water pollution can also occur naturally due to the presence of a minor and unwanted constituent, contaminant or impurity in the groundwater, in which case it is more likely referred to as contamination rather than pollution.

How Does Groundwater Become Contaminated?

The pollutant often creates a contaminant plume within an aquifer. Movement of water and dispersion within the aquifer spreads the pollutant over a wider area. Its advancing boundary, often called a plume edge, can intersect with groundwater wells or daylight into surface water such as seeps and springs, making the water supplies unsafe for humans and wildlife. The movement of the plume, called a plume front, may be analyzed through a hydrological transport model or groundwater model. Analysis of groundwater pollution may focus on soil characteristics and site geology, hydrogeology, hydrology, and the nature of the contaminants.

Pollution can occur from on-site sanitation systems, landfills, effluent from wastewater treatment plants, leaking sewers, petrol filling stations or from over application of fertilizers in agriculture. Pollution (or contamination) can also occur from naturally occurring contaminants, such as arsenic or fluoride. Using polluted groundwater causes hazards to public health through poisoning or the spread of disease.

Different mechanisms have influence on the transport of pollutants, e.g. diffusion, adsorption, precipitation, decay, in the groundwater. The interaction of groundwater contamination with surface waters is analyzed by use of hydrology transport models.

Pollutant types

Contaminants found in groundwater cover a broad range of physical, inorganic chemical, organic chemical, bacteriological, and radioactive parameters. Principally, many of the same pollutants that play a role in surface water pollution may also be found in polluted groundwater, although their respective importance may differ.

Arsenic and fluoride

Arsenic and fluoride have been recognized by the World Health Organization (WHO) as the most serious inorganic contaminants in drinking-water on a worldwide basis.

Inorganic arsenic is the most common type of arsenic in soil and water. The metalloid arsenic can occur naturally in groundwater, as seen most frequently in Asia, including in China, India and Bangladesh. In the Ganges Plain of northern India and Bangladesh severe contamination of groundwater by naturally occurring arsenic affects 25% of water wells in the shallower of two regional aquifers. Groundwater in these areas is also contaminated by the use of arsenic-based pesticides.

Arsenic in groundwater can also be present where there are mining operations or mine waste dumps that will leach arsenic.

Natural fluoride in groundwater is of growing concern as deeper groundwater is being used, "with more than 200 million people at risk of drinking water with elevated concentrations." Fluoride can especially be released from acidic volcanic rocks and dispersed volcanic ash when water hardness is low. High levels of fluoride in groundwater is a serious problem in the Argentinean Pampas, Chile, Mexico, India, Pakistan, the East African Rift, and some volcanic islands (Tenerife)

In areas that have naturally occurring high levels of fluoride in groundwater which is used for drinking water, both dental and

skeletal fluorosis can be prevalent and severe.

Pathogens

Waterborne diseases can be spread via a groundwater well which is contaminated with fecal pathogens from pit latrines

The lack of proper sanitation measures, as well as improperly placed wells, can lead to drinking water contaminated with pathogens carried in feces and urine. Such fecal-oral transmitted diseases include typhoid, cholera and diarrhea. Of the four pathogen types that are present in feces (bacteria, viruses, protozoa, and helminths or helminth eggs), the first three can be commonly found in polluted groundwater, whereas the relatively large helminth eggs are usually filtered out by the soil matrix.

Deep, confined aquifers are usually considered the safest source of drinking water with respect to pathogens. Pathogens from treated or untreated wastewater can contaminate certain, especially shallow, aquifers.

Nitrate

Nitrate is the most common chemical contaminant in the world's groundwater and aquifers. In some low-income countries, nitrate levels in groundwater are extremely high, causing significant health problems. It is also stable (it does not degrade) under high oxygen conditions.

Nitrate levels above 10 mg/L (10 ppm) in groundwater can cause "blue baby syndrome" (acquired methemoglobinemia). Drinking water quality standards in the European Union stipulate less than 50 mg/L for nitrate in drinking water.

However, the linkages between nitrates in drinking water and blue baby syndrome have been disputed in other studies. The syndrome outbreaks might be due to other factors than elevated nitrate concentrations in drinking water.

Elevated nitrate levels in groundwater can be caused by on-site sanitation, sewage sludge disposal and agricultural activities. It can therefore have an urban or agricultural origin.

Organic compounds

Volatile organic compounds (VOCs) are a dangerous contaminant of groundwater. They are generally introduced to the environment through careless industrial practices. Many of these compounds were not known to be harmful until the late 1960s and it was some time before regular testing of groundwater identified these substances in drinking water sources.

Primary VOC pollutants found in groundwater include aromatic hydrocarbons such as BTEX compounds (benzene, toluene, ethylbenzene and xylenes), and chlorinated solvents including tetrachloroethylene (PCE), trichloroethylene (TCE), and vinyl chloride (VC). BTEX are important components of gasoline. PCE and TCE are industrial solvents historically used in dry cleaning processes and as a metal degreaser, respectively.

Other organic pollutants present in groundwater and derived from industrial operations are the polycyclic aromatic hydrocarbons (PAHs). Due to its molecular weight, Naphthalene is the most soluble and mobile PAH found in groundwater, whereas benzo(a)pyrene is the most toxic one. PAHs are generally produced as byproducts by incomplete combustion of organic matter.

Organic pollutants can also be found in groundwater as insecticides and herbicides. As many other synthetic organic compounds, most pesticides have very complex molecular structures. This complexity determines the water solubility, adsorption capacity, and mobility of pesticides in the groundwater system. Thus, some types of pesticides are more mobile than others so they can more easily reach a drinking-water source.

Metals

Several trace metals occur naturally in certain rock formations and can enter in the environment from natural processes such as weathering. However, industrial activities such as mining, metallurgy, solid waste disposal, paint and enamel works, etc. can lead to elevated concentrations of toxic metals including lead, cadmium and chromium. These contaminants have the potential to make their way into groundwater.

The migration of metals (and metalloids) in groundwater will be affected by several factors, in particular by chemical reactions

which determine the partitioning of contaminants among different phases and species. Thus, the mobility of metals primarily depends on the pH and redox state of groundwater.

Pharmaceuticals

Trace amounts of pharmaceuticals from treated wastewater infiltrating into the aquifer are among emerging ground-water contaminants being studied throughout the United States. Popular pharmaceuticals such as antibiotics, anti-inflammatories, antidepressants, decongestants, tranquilizers, etc. are normally found in treated wastewater. This wastewater is discharged from the treatment facility, and often makes its way into the aquifer or source of surface water used for drinking water.

Trace amounts of pharmaceuticals in both groundwater and surface water are far below what is considered dangerous or of concern in most areas, but it could be an increasing problem as population grows and more reclaimed wastewater is utilized for municipal water supplies.

Others

Other organic pollutants include a range of organohalides and other chemical compounds, petroleum hydrocarbons, various chemical compounds found in personal hygiene and cosmetic products, drug pollution involving pharmaceutical drugs and their metabolites. Inorganic pollutants might include other nutrients such as ammonia and phosphate, and radionuclides such as uranium (U) or radon (Rn) naturally present in some geological formations. Saltwater intrusion is also an example of natural contamination, but is very often intensified by human activities.

Groundwater pollution is a worldwide issue. A study of the groundwater quality of the principal aquifers of the United States conducted between 1991 and 2004, showed that 23% of domestic wells had contaminants at levels greater than human-health benchmarks. Another study suggested that the major groundwater pollution problems in Africa, considering the order of importance are: (1) nitrate pollution, (2) pathogenic agents, (3) organic pollution, (4) salinization, and (5) acid mine drainage.

Potential Health and Other Effects of Groundwater Contaminants

The United States Geological Survey provides an extensive list of groundwater contaminants with their potential health and other effects.

https://www.epa.gov/sites/production/files/2015-08/documents/mgwc-gwc1.pdf

Summary

- Ground water contamination (also called groundwater pollution) occurs when pollutants are released to the ground and make their way down into groundwater.
- Contaminants found in groundwater cover a broad range of physical, inorganic chemical, organic chemical, bacteriological, and radioactive parameters.
- The United States Geological Survey developed an extensive list of groundwater contaminants with their potential health and other effects.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)

- Wikipedia
- US Geological Survey (USGS)

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8.5: Water Who Uses It and How Much?

Learning Objective

- List the end uses of water in residential and non residential areas.

Water in Daily Life

Residential water use (also called domestic use, household use, or tap water use) includes all indoor and outdoor uses of drinking quality water at single-family and multifamily dwellings. These uses include a number of defined purposes (or water end uses) such as flushing toilets, washing clothes and dishes, showering and bathing, drinking, food preparation, watering lawns and gardens, and maintaining swimming pools. Some of these end uses are detectable (and measurable) while others are more difficult to gauge.

Indoor water use includes water flows through fixtures and appliances inside the house. Because the distribution of indoor use in the sample of homes measured is positively skewed, a more appropriate measure of central tendency is the median, which is about 125 gphd (gallons per household per day) or 472 lphd, liters per household per day). Toilet flushing is the largest indoor use of water, followed by flows through kitchen and bathroom faucets, showers, clothes washers, leaks, bathtubs, other/miscellaneous uses, and dishwashers. Since the late 1990s, total indoor use has decreased by 22 percent, primarily due to the improved water efficiency of clothes washers and toilets, among other low-flow fixtures.

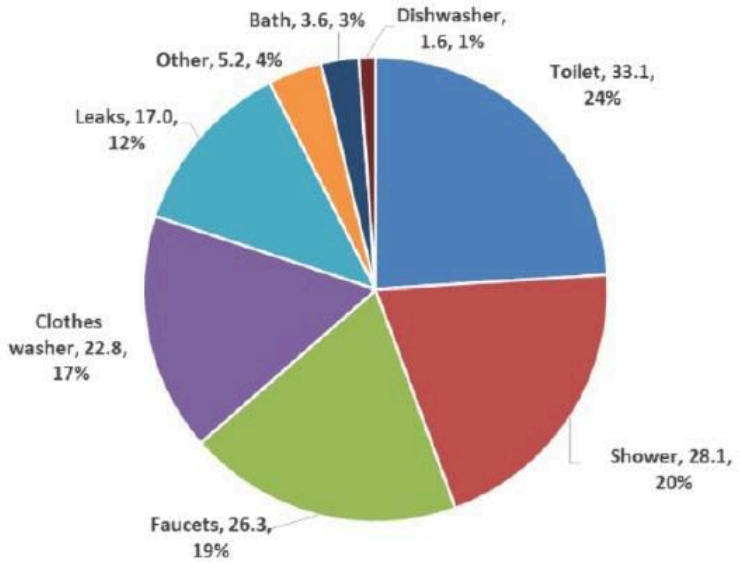


Figure 8.5.1 End uses of water for households in the U.S.in gallons per household per day and percent of indoor use.

The outdoor residential water use includes landscape irrigation, filling and back washing swimming pools, water used through outdoor faucets (hose bibs) for washing pavement and cars, and other outdoor uses. Annual outdoor use in North American cities differs by climatic region and ranged from 13,000 gallons in Waterloo, Canada to 120,400 gallons in Scottsdale, Arizona. The average outdoor use across 9 sampled cities in the Water Research Foundation study was 50,500 gallons per household per year or 138 gallons per day (524 liters per day). Nearly 17 percent of homes irrigate their landscapes in excess of theoretical irrigation requirement. If excess irrigation could be eliminated, the average outdoor use would drop by 8,200 gallons per house, or by 16 percent.

Commercial, Industrial, Agricultural & Electricity Water Use

It's easy to forget that we also use water in ways we don't see every day (Figure 8.5.2). Water is used to grow our food,

manufacture our favorite goods, and keep our businesses running smoothly. We also use a significant amount of water to meet the nation's energy needs. US Freshwater Withdrawals Chart

Figure 8.5.2 *U.S. freshwater withdrawals.*

Water Footprint

A **water footprint** shows the extent of water use in relation to consumption by people. The water footprint of an individual, community or business is defined as the total volume of fresh water used to produce the goods and services consumed by the individual or community or produced by the business. Water use is measured in water volume consumed (evaporated) and/or polluted per unit of time. A water footprint can be calculated for any well-defined group of consumers (e.g., an individual, family, village, city, province, state or nation) or producers (e.g., a public organization, private enterprise or economic sector), for a single process (such as growing rice) or for any product or service.

Traditionally, water use has been approached from the production side, by quantifying the following three columns of water use: water withdrawals in the agricultural, industrial, and domestic sector. While this does provide valuable data, it is a limited way of looking at water use in a globalized world, in which products are not always consumed in their country of origin. International trade of agricultural and industrial products in effect creates a global flow of virtual water, or *embodied water* (akin to the concept of embodied energy).

In 2002, the water footprint concept was introduced in order to have a consumption-based indicator of water use, that could provide useful information in addition to the traditional production-sector-based indicators of water use. It is analogous to the ecological footprint concept introduced in the 1990s. The water footprint is a geographically explicit indicator, not only showing volumes of water use and pollution, but also the locations. Thus, it gives a grasp on how economic choices and processes influence the availability of adequate water resources and other ecological realities across the globe (and vice versa).

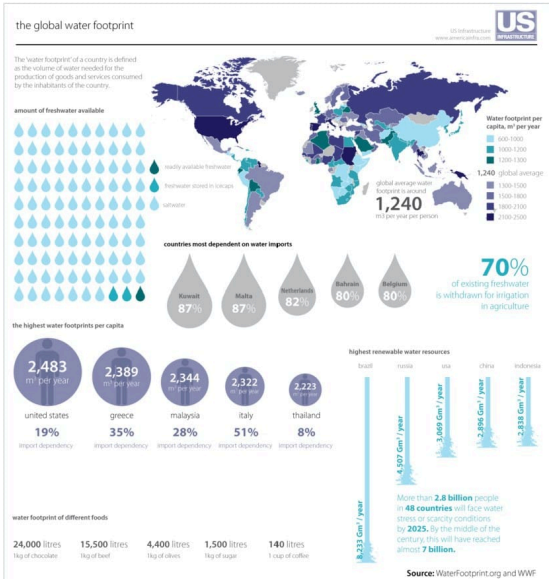


Figure 8.5.3 *The global water footprint.*

The water footprint of a product is the total volume of freshwater used to produce the product, summed over the various steps of the production chain. The water footprints involved in various diets vary greatly, and much of the variation tends to be associated with levels of meat consumption. The following table gives examples of estimated global average water footprints of popular agricultural products.

Table 8.5.3 *Global average water footprints of some agricultural products.*

Product	Global average water footprint, L/kg
almonds, shelled	16,194
beef	15,415

chocolate	17,196
cotton lint	9,114
lettuce	238
milk	1,021
olive oil	14,430
tomatoes, fresh	214
tomatoes, dried	4,275
vanilla beans	126,505
wheat bread	1,608

(For more product water footprints: see the Product Gallery of the Water Footprint Network)

Summary

- The average American family daily indoor water use is approximately 125 gallons per household per day.
- The water footprint is a geographically explicit indicator, not only showing volumes of water use and pollution, but also the locations.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Wikipedia

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8.6: Making Water Fit to Drink

Learning Objectives

- Know the primary standards for various contaminants in drinking water.
- Describe the importance of the different steps in water treatment.

Public Water Systems Fast Facts from US CDC

- Of the approximately 155,693 public water systems in the United States, 52,110 (33.5%) are community systems and 103,583 (66.5%) are noncommunity systems, including 84,744 transient systems and 18,839 nontransient systems [1].
- Over 286 million Americans get their tap water from a community water system [1].
- 8% of U.S. community water systems provide water to 82% of the U.S. population through large municipal water systems [1].
- Although the majority of community water systems (78%) are supplied by ground water, more people (68%) are supplied year-round by community water systems that use surface water .

  Note: EPA. Factoids: drinking water and ground water statistics for 2007. March 2008, April 2008.

Safe Drinking Water 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 increased presence of existing 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 8.6.1: AFGE members and environmental activists participated in an event at the Washington, D.C. EPA headquarters to thank EPA workers for their work. (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 would still be performed by each state. Regarding drinking water sources (surface and ground), the EPA and state agencies protect and monitor these as well. Levels of contaminants would be defined using the

concentration terms Maximum Contaminant Level (MCL) and Treatment Technique (TT)

National Primary Standards for Drinking Water

The first set of drinking water standards included only 22 chemicals and/or pathogens. EPA established two major types of contaminants: primary and secondary. The first of these types (primary) 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 concentration 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 numerical MCL concentrations. Water that contains any amount of these pathogens must be sanitized immediately with a standardized EPA procedure.

A detailed list of acceptable levels of different contaminants by EPA classification can be found on the the link below.

https://www.epa.gov/ground-water-and-drinking-water/national-primary-drinking-water-regulations

The complete table of standards is given below.

https://www.epa.gov/sites/production/files/2016-06/documents/npwdr_complete_table.pdf

Table 8.6.1

Table 8.6.1 *Standards for Various Contaminants in Drinking Water.*

Contaminant	MCLG[1](mg/L)[2]	MCL or TT[1](mg/L)[2]	Potential Health Effects from Long-Term Exposure Above the MCL (unless specified as short-term)	Sources of Contaminant in Drinking Water
Microorganisms- *Cryptosporidium*	zero	TT[3]	Gastrointestinal illness (such as diarrhea, vomiting, and cramps)	Human and animal fecal waste
Microorganisms- *Giardia lamblia*	zero	TT	Gastrointestinal illness (such as diarrhea, vomiting, and cramps)	Human and animal fecal waste
Microorganisms- Total Coliforms (including fecal coliform and *E. Coli*)	zero	5.0%	Not a health threat in itself; it is used to indicate whether other potentially harmful bacteria may be present[5]	Coliforms are naturally present in the environment; as well as feces; fecal coliforms and *E. coli* only come from human and animal fecal waste.
Viruses (enteric)	zero	TT	Gastrointestinal illness (such as diarrhea, vomiting, and cramps)	Human and animal fecal waste
Inorganic Chemicals- Antimony	0.006	0.006	Increase in blood cholesterol; decrease in blood sugar	Discharge from petroleum refineries; fire retardants; ceramics; electronics; solder

Contaminant	MCLG[1](mg/L)[2]	MCL or TT[1](mg/L)[2]	Potential Health Effects from Long-Term Exposure Above the MCL (unless specified as short-term)	Sources of Contaminant in Drinking Water
Inorganic Chemicals- Chromium (total)	0.1	0.1	Allergic dermatitis	Discharge from steel and pulp mills; erosion of natural deposits
Inorganic Chemicals- Nitrate (measured as Nitrogen)	10	10	Infants below the age of six months who drink water containing nitrate in excess of the MCL could become seriously ill and, if untreated, may die. Symptoms include shortness of breath and blue-baby syndrome.	Runoff from fertilizer use; leaking from septic tanks, sewage; erosion of natural deposits
Organic Chemicals- Benzene	zero	0.005	Anemia; decrease in blood platelets; increased risk of cancer	Discharge from factories; leaching from gas storage tanks and landfills
Organic Chemicals- Carbon tetrachloride	zero	0.005	Liver problems; increased risk of cancer	Discharge from chemical plants and other industrial activities
Organic Chemicals- Styrene	0.1	0.1	Liver, kidney, or circulatory system problems	Discharge from rubber and plastic factories; leaching from landfills
Radionuclides ***Alpha particles***	***none ---------- zero***	***15 picocuries per Liter (pCi/L)***	***Increased risk of cancer***	***Erosion of natural deposits of certain minerals that are radioactive and may emit a form of radiation known as alpha radiation***

Notes

[1]Definitions:

- Maximum Contaminant Level Goal (MCLG) - The level of a contaminant in drinking water below which there is no known or expected risk to health. MCLGs allow for a margin of safety and are non-enforceable public health goals.
- Maximum Contaminant Level (MCL) - The highest level of a contaminant that is allowed in drinking water. MCLs are set as close to MCLGs as feasible using the best available treatment technology and taking cost into consideration. MCLs are enforceable standards.
- Maximum Residual Disinfectant Level Goal (MRDLG) - The level of a drinking water disinfectant below which there is no known or expected risk to health. MRDLGs do not reflect the benefits of the use of disinfectants to control microbial contaminants.
- Treatment Technique (TT) - A required process intended to reduce the level of a contaminant in drinking water.
- Maximum Residual Disinfectant Level (MRDL) - The highest level of a disinfectant allowed in drinking water. There is convincing evidence that addition of a disinfectant is necessary for control of microbial contaminants.

[2] Units are in milligrams per liter (mg/L) unless otherwise noted. Milligrams per liter are equivalent to parts per million (PPM).

Calculations of Parts per Million and Parts per Billion

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:

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000$$

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 1,000,000,000$$

> 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, contaminants, 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.

ppm and ppb

In aqueous solutions, 1 ppm is essentially equal to 1 mg/L, and 1 ppb is equivalent to 1 µg/L.

$$1 \text{ ppm} = \frac{1 \text{ mg Solute}}{1 \text{ L Solution}}$$

$$1 \text{ ppb} = \frac{1 \ \mu\text{g Solute}}{1 \text{ L Solution}}$$

Amendment SDWA of 1986

This change to the Safe Drinking Water Act established MCLGs and increased the total of regulated contaminants to 83. The EPA would also install more monitoring devices to detect organic contaminants. Research was done to detect pathogens more effectively to reduce disease. Lastly, public notifications from water systems to the consumers would be made announced if severe water issues occurred.

Amendment of SDWA of 1996

This particular legislation provided more protection and assessment of water sources (lakes, river, streams). In addition, water companies were required to provide consumers with a water quality report. States could seek federal money for upgrading their water quality processes. Cost benefit analysis would be done to determine risk and reward of lowering the concentration of a contaminant. For small regulated water companies, more financial and technical assistance would be offered to help them maintain drinking water standards.

Water Treatment (US CDC)

Drinking water supplies in the United States are among the safest in the world. However, even in the U.S., drinking water sources can become contaminated, causing sickness and disease from waterborne germs, such as *Cryptosporidium*, *E. coli*, Hepatitis A, *Giardia intestinalis*, and other pathogens.

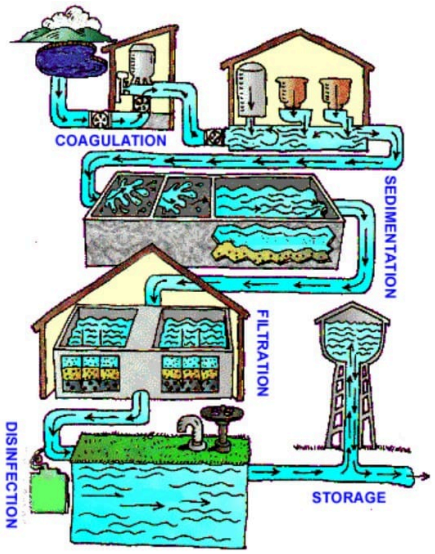


Figure 8.6.1 The common steps in water treatment. Source: US EPA

Drinking water sources are subject to contamination and require appropriate treatment to remove disease-causing agents. Public drinking water systems use various methods of water treatment to provide safe drinking water for their communities. Today, the most common steps in water treatment (Figure 8.6.1) used by community water systems (mainly surface water treatment) include:

Coagulation and Flocculation

Coagulation and flocculation are often the first steps in water treatment. Chemicals with a positive charge are added to the water. The positive charge of these chemicals neutralizes the negative charge of dirt and other dissolved particles in the water. When this occurs, the particles bind with the chemicals and form larger particles, called floc. 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:

$$Al_2(SO_4)_3.12\ H_2O \rightarrow Al^{3+}(aq) + 3\ SO_4^{2-}(aq) + 12\ H_2O$$
$$SO_4^{2-}(aq) + H_2O \rightarrow HSO_4^-(aq) + OH^- \text{ (causing pH change)}$$
$$Ca(OH)_2 \rightarrow Ca^{2+}(aq) + 2\ OH^- \text{ (causing pH change)}$$

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

During sedimentation, floc settles to the bottom of the water supply, due to its weight. This settling process is called sedimentation. The floc particles are then removed from the bottom of the basins.

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

Filtration

Once the floc has settled to the bottom of the water supply, the clear water on top will pass through filters of varying compositions (sand, gravel, and charcoal) and pore sizes, in order to remove dissolved particles, such as dust, parasites, bacteria, viruses, and chemicals.

Treatment with Chlorine and/or Chloramine

Most communities use either chlorine or chloramines. Some communities switch back and forth between chlorine and chloramines at different times of the year or for other operational reasons. Less commonly, utilities use other disinfectants, such as chlorine dioxide. Some water systems that use water from a groundwater source (like community wells) do not have to add a disinfectant at all.

Ozone Disinfection

Ozone disinfection, or ozonation, is an unstable molecule which readily gives up one atom of oxygen providing a powerful oxidizing agent which is toxic to most waterborne organisms. It is a very strong, broad spectrum disinfectant that is widely used in Europe and in a few municipalities in the United States and Canada. It is an effective method to inactivate harmful protozoa that form cysts. It also works well against almost all other pathogens. Ozone is made by passing oxygen through ultraviolet light or a "cold" electrical discharge. To use ozone as a disinfectant, it must be created on-site and added to the water by bubble contact. Some of the advantages of ozone include the production of fewer dangerous by-products and the absence of taste and odor problems (in comparison to). No residual ozone is left in the water. In the absence of a residual disinfectant in the water, chlorine or chloramine may be added throughout a distribution system to remove any potential pathogens in the distribution piping.

Ultraviolet Disinfection

Ultraviolet disinfection of water is a purely physical, chemical-free process. Even parasites such as Cryptosporidium or Giardia, which are extremely resistant to chemical disinfectants, are efficiently reduced. UV can also be used to remove chlorine and chloramine species from water; this process is called photolysis, and requires a higher dose than normal disinfection. The dead microorganisms are not removed from the water. UV disinfection does not remove dissolved organics, inorganic compounds or particles in the water. The world's largest water disinfection plant treats drinking water for New York city. The Catskill-Delaware Water Ultraviolet Disinfection Facility, commissioned on 8 October 2013, incorporates a total of 56 energy-efficient UV reactors treating up to 2.2 billion US gallons (8,300,000 m^3) a day.

Ultraviolet can also be combined with ozone or hydrogen peroxide to produce hydroxyl radicals to break down trace contaminants through an advanced oxidation process.

Fluoridation (US CDC)

The mineral fluoride occurs naturally on earth and is released from rocks into the soil, water, and air. All water contains some fluoride. Usually, the fluoride level in water is not enough to prevent tooth decay; however, some groundwater and natural springs can have naturally high levels of fluoride.

Fluoride has been proven to protect teeth from decay. Bacteria in the mouth produce acid when a person eats sugary foods. This acid eats away minerals from the tooth's surface, making the tooth weaker and increasing the chance of developing cavities. Fluoride helps to rebuild and strengthen the tooth's surface, or enamel. Water fluoridation prevents tooth decay by providing frequent and consistent contact with low levels of fluoride. By keeping the tooth strong and solid, fluoride stops cavities from forming and can even rebuild the tooth's surface.

Community water fluoridation is the process of adjusting the amount of fluoride found in water to achieve optimal prevention of tooth decay.

Although other fluoride-containing products, such as toothpaste, mouth rinses, and dietary supplements are available and contribute to the prevention and control of tooth decay, community water fluoridation has been identified as the most cost-effective method of delivering fluoride to all, reducing tooth decay by 25% in children and adults.[1]

Benefits: Strong Teeth (US CDC)

Fluoride benefits children and adults throughout their lives. For children younger than age 8, fluoride helps strengthen the adult (permanent) teeth that are developing under the gums. For adults, drinking water with fluoride supports tooth enamel, keeping teeth strong and healthy. The health benefits of fluoride include having:

- Fewer cavities.
- Less severe cavities.
- Less need for fillings and removing teeth.
- Less pain and suffering because of tooth decay.

Fluoride in the Water Today

In 2012, more than 210 million people, or 75% of the US population, were served by community water systems that contain enough fluoride to protect their teeth.[5] However, approximately 100 million Americans still do not have access to water with fluoride. Because it is so beneficial, the United States has a national goal for 80% of Americans to have water with enough fluoride to prevent tooth decay by 2020.

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

https://www.youtube.com/watch?v=0_ZcCqqpS2o&feature=youtu.be

Summary

- The Safe Drinking Water Act (SDWA) (originally passed by Congress in 1974) protects public health by regulating the nation's

public drinking water supply.
- The different steps in water treatment include floculation, sedimentation, filtration, aeration, and disinfection or combinations thereof. Fluoride is also added to municipal water systems for prevention of tooth decay.

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- US EPA
- Center for Disease Control (CDC)
- Wikipedia

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8.7: 14.7?- Making Water Fit to Drink

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8.8: Wastewater and Sewage Treatment

Learning Objectives

- Explain the major steps in wastewater treatment.
- List the different uses of reclaimed water.

Wastewater and sewage is treated in three phases: primary (solid removal), secondary (bacterial decomposition), and tertiary (extra filtration).

Sewage is generated by residential and industrial establishments. It includes household waste liquid from toilets, baths, showers, kitchens, sinks, and so forth that is disposed of via sewers. In many areas, sewage also includes liquid waste from industry and commerce. The separation and draining of household waste into greywater and blackwater is becoming more common in the developed world. Greywater is water generated from domestic activities such as laundry, dishwashing, and bathing, and can be reused more readily. Blackwater comes from toilets and contains human waste.

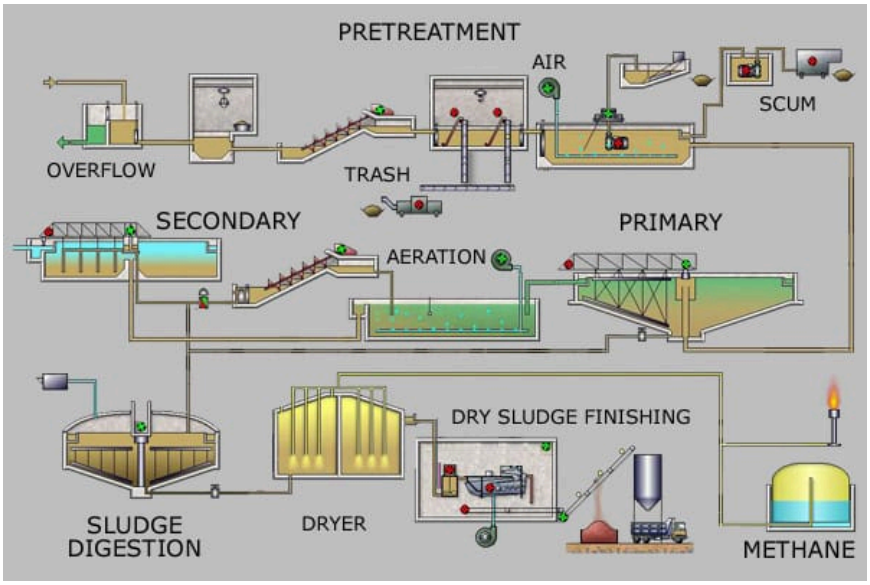


Figure 8.8.1 **Diagram of Sewage Treatment Process**: Sewage passes through primary, secondary, and tertiary treatment.

Primary Treatment

In primary treatment, sewage is stored in a basin where solids (sludge) can settle to the bottom and oil and lighter substances can rise to the top. These layers are then removed and then the remaining liquid can be sent to secondary treatment. Sewage sludge is treated in a separate process called sludge digestion.

Secondary Treatment

Secondary treatment removes dissolved and suspended biological matter, often using microorganisms in a controlled environment. Most secondary treatment systems use aerobic bacteria, which consume the organic components of the sewage (sugar, fat, and so on). Some systems use fixed film systems, where the bacteria grow on filters, and the water passes through them. Suspended growth systems use “activated” sludge, where decomposing bacteria are mixed directly into the sewage. Because oxygen is critical to bacterial growth, the sewage is often mixed with air to facilitate decomposition.

Tertiary Treatment

Tertiary treatment (sometimes called “effluent polishing”) is used to further clean water when it is being discharged into a sensitive ecosystem.

Several methods can be used to further disinfect and purify sewage beyond primary and secondary treatment.

Sand filtration, where water is passed through a sand filter, can be used to remove particulate matter.

Wastewater may still have high levels of nutrients such as nitrogen and phosphorus. These can disrupt the nutrient balance of aquatic ecosystems and cause algae blooms and excessive weed growth. Phosphorus can be removed biologically in a process called **enhanced biological phosphorus removal**. In this process, specific bacteria, called polyphosphate accumulate organisms that store phosphate in their tissue. When the biomass accumulated in these bacteria is separated from the treated water, these biosolids have a high fertilizer value. Nitrogen can also be removed using nitrifying bacteria.

Lagooning is another method for removing nutrients and waste from sewage. Water is stored in a lagoon and native plants, bacteria, algae, and small zooplankton filter nutrients and small particles from the water.

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. 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.

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.

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**. By applying pressure in the higher concentration solution, water molecules migrate from a high concentration solution to a low concentration solution through a semipermeable membrane. This method is called **reverse osmosis water filter system**. In this technique, the membrane must be able to tolerate the high pressure, and prevent solute molecules to pass through. 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.

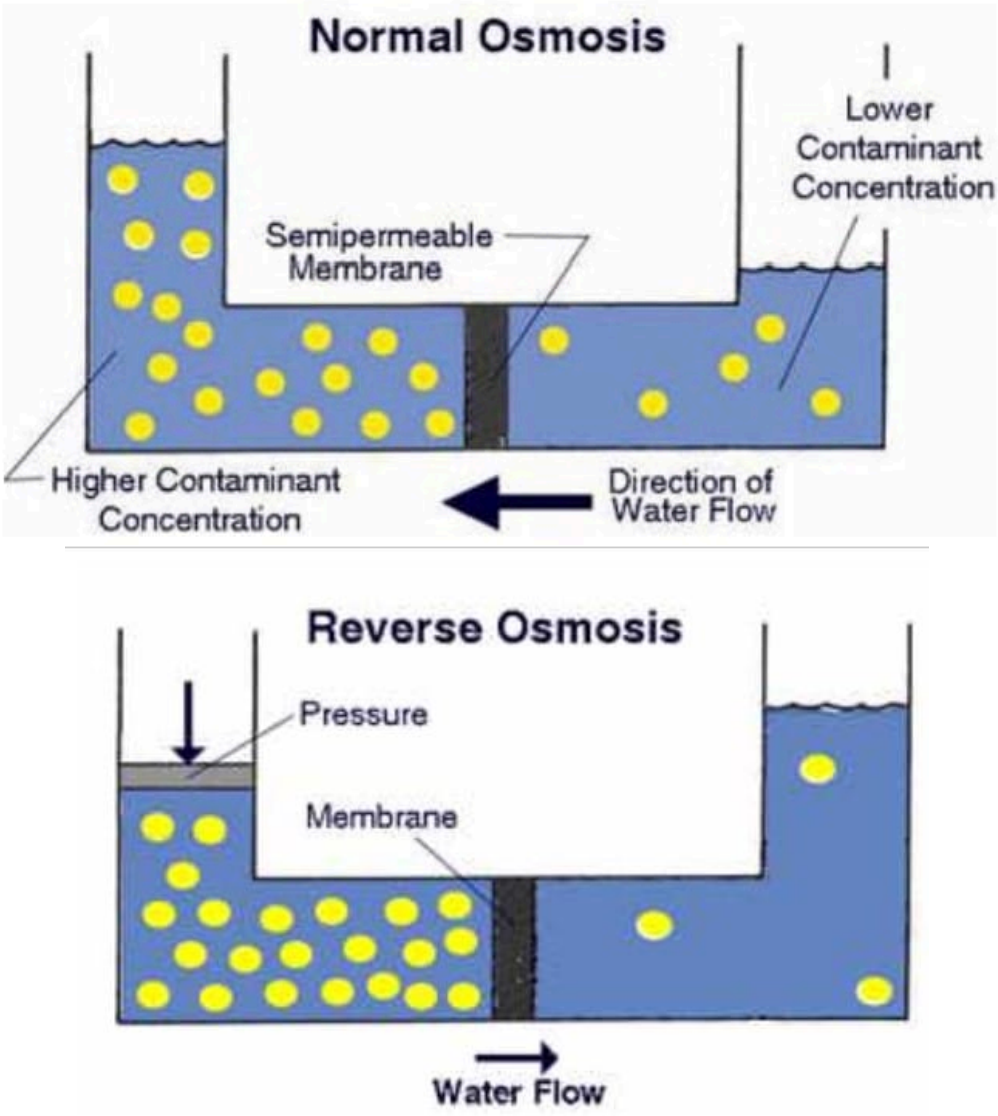


Figure 8.8.2 Comparison of normal osmosis and reverse osmosis. Source: PurePro.

Sludge Digestion

Sewage sludge scraped off the bottom of the settling tank during primary treatment is treated separately from wastewater. Sludge can be disposed of in several ways. First, it can be digested using bacteria; bacterial digestion can sometimes produce methane biogas, which can be used to generate electricity. Sludge can also be incinerated, or condensed, heated to disinfect it, and reused as fertilizer.

YouTube Video

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Uses of Reclaimed Wastewater

Water reclamation (also called **wastewater reuse**) is the process of converting wastewater into water that can be reused for other purposes. Reuse may include irrigation of gardens and agricultural fields or replenishing surface water and groundwater (i.e., groundwater recharge). Reused water may also be directed toward fulfilling certain needs in residences (e.g. toilet flushing), businesses, and industry, and could even be treated to reach drinking water standards. This last option is called either "direct potable reuse" or "indirect potable" reuse, depending on the approach used.

Reclaiming water for reuse applications instead of using freshwater supplies can be a water-saving measure. When used water is eventually discharged back into natural water sources, it can still have benefits to ecosystems, improving streamflow, nourishing plant life and recharging aquifers, as part of the natural water cycle.

Wastewater reuse is a long-established practice used for irrigation, especially in arid countries. Reusing wastewater as part of sustainable water management allows water to remain as an alternative water source for human activities. This can reduce scarcity and alleviate pressures on groundwater and other natural water bodies. Another potentially positive aspect is the nutrient content in the wastewater, which might reduce the need of other fertilizers.

Drawbacks or risks often mentioned include the content of potentially harmful substances such as bacteria, heavy metals or organic pollutants (including pharmaceuticals, personal care products and pesticides). Irrigation with wastewater can have both positive and negative effects on soil and plants, depending on the composition of the wastewater and on the soil or plant characteristics.

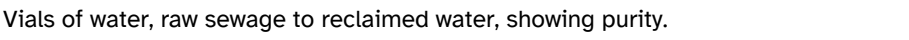


Figure 8.8.3 *The Sequence of reclamation from left: raw sewage, plant effluent, and finally reclaimed water (after several treatment steps). (Credit: Wikipedia)*

Most of the uses of water reclamation are non potable uses such as washing cars, flushing toilets, cooling water for power plants, concrete mixing, artificial lakes, irrigation for golf courses and public parks, and for hydraulic fracturing. Where applicable, systems run a dual piping system to keep the recycled water separate from the potable water.

The main reclaimed water applications in the world are shown below:

Table 8.8.1 Uses of Reclaimed Water

Categories of use	**Uses**
Urban uses	Irrigation of public parks, sporting facilities, private gardens, roadsides; Street cleaning; Fire protection systems; Vehicle washing; Toilet flushing; Air conditioners; Dust control.
Agricultural uses	Food crops not commercially processed; Food crops commercially processed; Pasture for milking animals; Fodder; Fibre; Seed crops; Ornamental flowers; Orchards; Hydroponic culture; Aquaculture; Greenhouses; Viticulture.
Industrial uses	Processing water; Cooling water; Recirculating cooling towers; Washdown water; Washing aggregate; Making concrete; Soil compaction; Dust control.
Recreational uses	Golf course irrigation; Recreational impoundments with/without public access (e.g. fishing, boating, bathing); Aesthetic impoundments without public access; Snowmaking.
Environmental uses	Aquifer recharge; Wetlands; Marshes; Stream augmentation; Wildlife habitat; Silviculture.
Potable uses	Aquifer recharge for drinking water use; Augmentation of surface drinking water supplies; Treatment until drinking water quality.

Astronauts and Reclaimed Urine

If it's good enough for astronauts...

Astronauts about the space station drinking recycled urine.

Astronauts aboard the International Space Station drink reclaimed urine. (Credit: NASA)

In outer space, as on the International Space Station, water is at a premium and not a drop is to be wasted. Astronauts aboard the station must drink recycled water. Water from humidity in the air is condensed and used as drinking water, but American astronauts also drink recycled urine! You can imagine the urine, which is of course almost all water, would go through a rigorous purification process, but this just shows if the need exists, that the same water can serve valuable purposes over and over again.

Summary

- Wastewater treatment is a process used to remove from or and convert it into an that can be returned to the with minimum impact on the environment, or directly reused.
- In a wastewater treatment plant wastewater or sewage undergo primary,secondary, as well as tertiary treatments depending on the end use.
- Most of the uses of water reclamation are non potable uses such as washing cars, flushing toilets, cooling water for power plants, concrete mixing, artificial lakes, irrigation for golf courses and public parks, and for hydraulic fracturing

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9: POISONS

9.1: Natural Poisons and Food Poisoning

9.2: Poisons and How They Act

9.3: More Chemistry of the Nervous System

9.4: The Lethal Dose

9.5: The Liver as a Detox Facility

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What is Toxicology?

Toxicology is traditionally defined as "the science of poisons." Over time, our understanding of how various agents can cause harm to humans and other organisms has increased, resulting in a more descriptive definition of toxicology as "the *study of the adverse effects of chemical, physical, or biological agents on living organisms and the ecosystem, including the prevention and amelioration of such adverse effects.*"

These adverse effects can take many forms, ranging from immediate death to subtle changes not appreciated until months or years later. They may occur at various levels within the body, such as an organ, a type of cell, or a specific biochemical. Our understanding of how toxic agents damage the body has progressed along with medical knowledge. We now know that various observable changes in anatomic or bodily functions actually result from previously unrecognized changes in specific biochemicals in the body.

A poison is any substance that is harmful to your body. You might swallow it, inhale it, inject it, or absorb it through your skin. Any substance can be poisonous if too much is taken. Poisons can include

- Prescription or over-the-counter medicines taken in doses that are too high
- Overdoses of illegal drugs
- Carbon monoxide from gas appliances
- Household products, such as laundry powder or furniture polish
- Pesticides
- Indoor or outdoor plants
- Metals such as lead and mercury

The effects of poisoning range from short-term illness to brain damage, coma, and death. To prevent poisoning it is important to use and store products exactly as their labels say. Keep dangerous products where children can't get to them. Treatment for poisoning depends on the type of poison.

Sources

National Institute of Health (NIH) US National Library of Medicine

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9.1: Natural Poisons and Food Poisoning

Learning Objectives

- List different plant and animal poisons and their properties.
- List various causes of food poisoning.

In biology, **poisons** are substances that cause death, injury or harm to organs, usually by chemical reactions or other activity on the molecular scales, when an organism absorbs a sufficient quantity. The fields of medicine (particularly veterinary) and zoology often distinguish a poison from a toxin, and from a venom. Toxins are poisons produced by organisms in nature, and venoms are toxins injected by a bite or sting (this is exclusive to animals). The difference between venom and other poisons is the delivery method.

The **history of poison** stretches from before 4500 BCE to the present day. Poisons have been used for many purposes across the span of human existence, most commonly as weapons, anti-venoms, and medicines. Poison has allowed much progress in branches, toxicology, and technology, among other sciences.

Over the centuries, the variety of harmful uses of poisons continued to increase. The means for curing these poisons also advanced in parallel. In the modern world, intentional poisoning is less common than the Middle Ages. Rather, the more common concern is the risk of accidental poisoning from everyday substances and products.

Constructive uses for poisons have increased considerably in the modern world. Poisons are now used as pesticides, disinfectants, cleaning solutions, and preservatives. Nonetheless, poison continues to be used as a hunting tool in remote parts of developing countries, including Africa, South America, and Asia.

Poisonous Plants

Poisonous plants are plants that produce toxins that deter herbivores from consuming them. Plants cannot move to escape their predators, so they must have other means of protecting themselves from herbivorous animals. Some plants have physical defenses such as thorns, spines and prickles, but by far the most common type of protection is chemical. Over millennia, through the process of natural selection, plants have evolved the means to produce a vast and complicated array of chemical compounds in order to deter herbivores. Tannin, for example, is a defensive compound that emerged relatively early in the evolutionary history of plants, while more complex molecules such as polyacetylenes are found in younger groups of plants such as the Asterales. Many of the known plant defense compounds primarily defend against consumption by insects, though other animals, including humans, that consume such plants may also experience negative effects, ranging from mild discomfort to death. Several poisonous plants, the specific toxins and the poisonous effects are listed below.

Poison ivy (Figure 9.1.1) - The poison Urushiol is an oily mixture of organic compounds with allergic properties. Symptoms of a reaction include itching, redness, swelling, and blisters. It is important to note that the blisters are not contagious.



Figure 9.1.1 Poison Ivy.

Western water hemlock (Figure 9.1.2) - The toxin, cicutoxin and oenanthotoxin, are conjugated polyacetylenes. These unsaturated alcohols have a strong carrot-like odor and are noncompetitive antagonists for the gamma-aminocutyric acid (GABA) neural transmitter in the central nervous system. GABA role is to inhibit neuron excitability; essentially it has a relaxing function. Blocking this results in convulsing and grand mal seizures and eventually death can occur.



Figure 9.1.2 Western water hemlock.

Autumn skullcap - The toxin amatoxin in the form of γ-amanitin, β-amanitin, and α-amanitin causes severe abdominal pain, vomiting, and diarrhea may last for six-nine hours. The toxins affect the liver, results in gastrointestinal bleeding, coma, kidney failure, or even death, within seven days of consumption

Henbane - The toxins in this plant are atropine and scopolamine, found in leaves. Apoatropine and cuscohygrine are the main alkaloids of the root. The main alkaloid in seeds is hyoscyamine, little hyoscine and little atropine. the plant is toxic to cattle, wild animals, fish, and birds. Pigs are immune. In humans, it could cause hallucinations, dilated pupils, restlessness, fast heart, seizure, vomiting, high blood pressure, and ataxia. Initial effects last for 3-4 hours, after effects last up to three days.

Oak - Tannic acid, binds and precipitates proteins. Cattle, sheep, horses and goats are most affected while pigs are immune. In humans, Tannic acid poisoning could cause anorexia, depression, constipation, diarrhea, blood in urine. Signs typically occur around 3-7 days after consumption, with

The links below are an extensive, if incomplete, list of plants containing one or more poisonous parts that pose a serious risk of illness, injury, or death to humans or domestic animals. There is significant overlap between plants considered poisonous and those with psychotropic properties, some of which are toxic enough to present serious health risks at recreational doses. There

is a distinction between plants that are poisonous because they naturally produce dangerous phytochemicals, and those that may become dangerous for other reasons, including but not limited to infection by bacterial, viral, or fungal parasites; the uptake of toxic compounds through contaminated soil or groundwater; and/or the ordinary processes of decay after the plant has died; this list deals exclusively with plants that produce phytochemicals. Many plants, such as peanuts, produce compounds that are only dangerous to people who have developed an allergic reaction to them, and with a few exceptions, those plants are not included here (see list of allergens instead). Despite the wide variety of plants considered poisonous, human fatalities caused by poisonous plants – especially resulting from accidental ingestion – are rare in the developed world.

WEB LINK

A more extensive list of poisonous plants can be found on the link below.

https://en.Wikipedia.org/wiki/List_o...isonous_plants

Poisonous Animals

Several deadly poisons are from snakes and spiders. Several poisonous animals, the specific toxins and the poisonous effects are listed below.

Timber rattlesnake - There are 4 types of toxin. Type A : A neurotoxin known as canebrake. Type B : A hemorrhagic and proteolytic toxin. Type A + B: Intergrade between snakes with Type A and Type B. Type C : relatively weak. The toxins could cause myokymia, defibrination syndrome, numbness, lightheadedness, weakness, vomiting, blurred vision, sweating, salivating. It can be treated with CroFab Antivenom.

Eastern coral snake (Figure 9.1.3).- Phospholipase A2 and three-finger toxins (abbreviated as 3FTx). 3FTx proteins are neurotoxins, attacking nerve tissue. Phospholipase A2 can cause inflammation and pain at the site of the bite, though this is uncommon. The toxins could also cause slurred speech, double vision, muscle paralysis, and can lead to cardiac arrest if left untreated.



Figure 9.1.3 *The eastern coral snake*

Cotton mouth snake - Cotton mouth toxins mainly consist of three protein families: phospholipase A2 (PLA2), metalloproteases (SVMP), and serine proteases (SVSP). PLA2s are responsible for inflammation and pain, while SVMPs are responsible for hemorrhage and SVSPs affect the coagulation of blood. The poison could lead to low blood pressure, weakness, change in skin color at site of bite, trouble breathing, nausea, increase in heart rate.

Black widow spider (Figure 9.1.4) - Latrotoxins are the main component of the venom, but other compounds such as polypeptides, adenosine, and guanosine are active as well. Only the bite of the female is dangerous to humans, as their venom glands are very large. The toxin can cause pain and sweating at the site of the bite, muscle cramps, headache, nausea, vomiting, weakness. Typically only localized pain is felt, but in some cases the pain can spread. Symptoms typically last from 3-7 days after

the bite takes place.

Brown recluse spider - Brown recluse venom possesses potentially deadly hemotoxins and cytotoxins that affect the red blood cells and their ability to clot. It is a mixture of enzymes such as collagenase, protease, and phospholipase. Symptoms include redness, fever, weakness, pain and nausea. In around 10% of victims necrosis can occur at the site of the bite, and in even fewer cases hemolysis (bursting of red blood cells) can occur.



Figure 9.1.4 *Western black widow.*

Other Poisonous Animals

The list below is a partial list of animals that are poisonous to humans (and other animals), or put another way, their flesh is toxic if consumed, or in some cases touched:

Birds: Pitohui, Blue-capped ifrit, Little shrikethrush, Spur-winged goose[1] (diet-dependent), and Common quail (diet-dependent)

Snakes: Rhabdophis keelback snakes and Garter snake (diet-dependent, when feeding on Pacific newts)[2]

Frogs and toads: American toad, Asiatic toad, Cane toad, Colorado River toad, Common toad, Corroboree frog, European green toad, Fowler's toad, Mantella, and poison dart frog (Figure 9.1.5).



Figure 9.1.5 *Poison dart frogs are well known for their brightly colored skin, which serves as a warning to predators of their toxicity.*

Salamanders: Pacific newts

Fish: Tetraodontidae (Blowfish, Pufferfish), Greenland shark, and Barracuda (age and diet dependent)

Cephalophods: Blue-ringed octopus and Pfeffer's flamboyant cuttlefish

Insects: Blister beetle, Birdwings, Milkweed butterfly, Battus (butterfly), Diamphidia, and Monarch (butterfly)

Food Poisoning

Foodborne illness (also **foodborne disease** and colloquially referred to as **food poisoning**)[1] is any illness resulting from the spoilage of contaminated food, pathogenic bacteria, viruses, or parasites that contaminate food,[2] as well as toxins such as poisonous mushrooms and various species of beans that have not been boiled for at least 10 minutes.

Many different disease-causing germs (Figure 9.1.6 below) can contaminate foods, so there are many different foodborne infections. The US CDC estimates that each year 48 million people get sick from a foodborne illness, 128,000 are hospitalized, and 3,000 die. Researchers have identified more than 250 foodborne diseases. Most of them are infections, caused by a variety of bacteria, viruses, and parasites. Harmful toxins and chemicals also can contaminate foods and cause foodborne illness.

Although anyone can get a foodborne illness, some people are more likely to develop one. Those groups include Pregnant women, Young children, Older adults, People with immune systems weakened from medical conditions, such as diabetes, liver disease, kidney disease, organ transplants, HIV/AIDS, or from receiving chemotherapy or radiation treatment. Most people with a foodborne illness get better without medical treatment, but people with severe symptoms should see their doctor.

Food Poisoning Symptoms

Common symptoms of foodborne diseases are nausea, vomiting, stomach cramps, and diarrhea. However, symptoms may differ among the different types of foodborne diseases. Symptoms can sometimes be severe and some foodborne illnesses can even be life-threatening. Most people have only mild illnesses, lasting a few hours to several days. However, some people need to be hospitalized, and some illnesses result in long-term health problems or even death. Infections transmitted by food can result in chronic arthritis, brain and nerve damage, and hemolytic uremic syndrome (HUS), which causes kidney failure.

Food poisoning symptoms may range from mild to severe and may differ depending on the germ you swallowed (see Table 9.1.1 below). After you consume a contaminated food or drink, it may take hours or days before you develop symptoms. See your doctor or healthcare provider if you have symptoms that are severe, including:

- High fever (temperature over 101.5°F, measured orally)
- Blood in stools
- Frequent vomiting that prevents keeping liquids down (which can lead to dehydration)
- Signs of dehydration, including a marked decrease in urination, a very dry mouth and throat, or feeling dizzy when standing up.
- Diarrhea that lasts more than 3 days

The top five germs that cause illnesses from food eaten in the United States are *Norovirus, Salmonella, Clostridium perfringens, Campylobacter, and Staphylococcus aureus (Staph).*

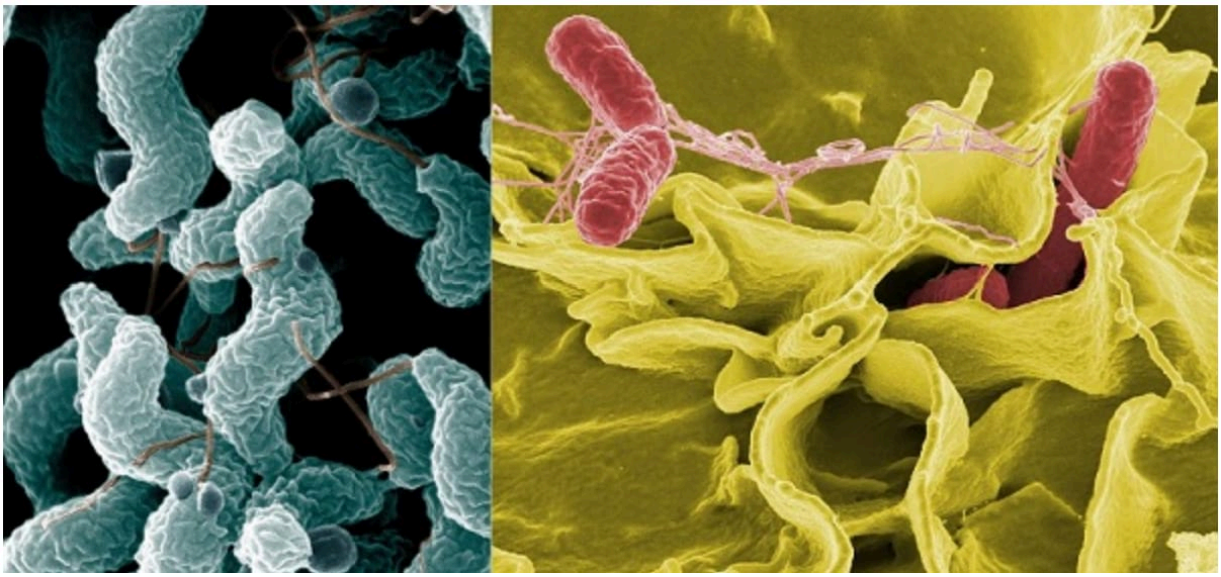


Figure 9.1.6 Campylobacter (left) , a bacteria that is a major cause of food poisoning and Salmonella (right). Images used with permission (Public Domain; De Wood, Pooley, USDA (left) and Rocky Mountain Laboratories, NIAID, NIH).

Some other germs don't cause as many illnesses, but when they do, the **illnesses are more likely to lead to hospitalizatio**n. Those germs include *Clostridium botulinum* (botulism), *Listeria, Escherichia coli (E. coli), and Vibrio.*

Table 9.1.1 Symptoms and Sources of 10 Foodborne Germs

Germ and Typical Time for Symptoms to Appear	Typical Signs and Symptoms	Common Food Sources
Campylobacter 2 – 5 days	Diarrhea (often bloody), stomach cramps/pain, fever	Raw or undercooked poultry, raw (unpasteurized) milk, and contaminated water
Clostridium botulinum 18 – 36 hours	Double or blurred vision, drooping eyelids, slurred speech. Difficulty swallowing, breathing and dry mouth. Muscle weakness and paralysis. Symptoms start in the head and move down as severity increases	Improperly canned or fermented foods, usually homemade. Prison-made illicit alcohol.
Clostridium perfringens 6 – 24 hours	Diarrhea, stomach cramps. Vomiting and fever are uncommon. Usually begins suddenly and lasts for less than 24 hours	Beef or poultry, especially large roasts; gravies; dried or precooked foods
Cyclospora 1 week	Watery diarrhea, loss of appetite and weight loss. Stomach cramps/pain, bloating, increased gas, nausea, and fatigue.	Raw fruits or vegetables, and herbs
Escherichia coli 3 – 4 days	Severe stomach cramps, diarrhea (often bloody), and vomiting. Around 5-10% of people diagnosed with this infection develop a life-threatening complication.	Raw or undercooked ground beef, raw (unpasteurized) milk and juice, raw vegetables (such as lettuce), and raw sprouts, contaminated water
Listeria 1 – 4 weeks	Pregnant women typically experience fever and other flu-like symptoms, such as fatigue and muscle aches. Infections during pregnancy can lead to serious illness or even death in newborns. Other people (most often older adults): headache, stiff neck, confusion, loss of balance and convulsions in addition to fever and muscle aches.	Raw (unpasteurized) milk, soft cheeses made with raw milk, raw sprouts, melons, hot dogs, pâtés, lunch meats, and cold cuts, smoked seafood

Germ and Typical Time for Symptoms to Appear	Typical Signs and Symptoms	Common Food Sources
Norovirus 12 – 48 hours	Diarrhea, nausea/stomach pain, vomiting	Infected person, contaminated food like leafy greens, fresh fruits, shellfish (such as oysters), or water, or by touching contaminated surfaces
Salmonella 12 – 72 hours	Diarrhea, fever, stomach cramps, vomiting	Eggs, raw or undercooked poultry or meat, unpasteurized milk or juice, cheese, raw fruits and vegetables
Staphylococcus aureus (Staph) 30 minutes – 6 hours	Diarrhea, nausea, stomach cramps, vomiting	Foods that are handled by people and not cooked (sliced meat, puddings, pastries, and sandwiches). Raw (unpasteurized) milk and cheese made from it.
Vibrio 1 – 4 days	Watery diarrhea, nausea. stomach cramps, vomiting, fever, chills	Raw or undercooked shellfish, particularly oysters

The Top 5 Strangest Posions That Can Kill You

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 9.1.1 *Strange poisons.*

Summary

- Various plants produce toxins that could be lethal or cause allergic reactions in herbivores. A few examples of poisonous plants include poison ivy, western water hemlock, autumn skullcap, henbane, oak, etc.
- Several deadly poisons come from snakes and spiders. Various birds, fishes, frogs, insects, etc. are deadly when consumed or touched due to their toxic flesh.
- The top five germs that cause illnesses from food eaten in the United States are *Norovirus, Salmonella, Clostridium perfringens, Campylobacter, and Staphylococcus aureus (Staph).*

Contributors

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- Scioly.org
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9.2: Poisons and How They Act

Learning Objective

- Describe the modes of action and treatments for different poisons.

Corrosive Substances

A **corrosive substance** is one that will damage or destroy other substances with which it comes into contact by means of a chemical reaction. Corrosives are different from ***poisons*** in that corrosives are immediately dangerous to the tissues they contact, whereas poisons may have systemic toxic effects that require time to become evident. Colloquially, corrosives may be called *poisons* but the concepts are technically distinct. However, there is nothing which precludes a corrosive from being a poison; there are substances that are both corrosives and poisons.

There are several household products that contain chemical substances which are considered highly reactive and thus are very dangerous upon ingestion or skin contact. Chemical poisoning with corrosive agents occurs by ingestion of: acids (hydrochloric, nitric, sulfuric, perchloric, chloric), alkalis (sodium and potassium, soaps, detergents), heavy metal salts (sublimate), strong oxidizing agents, formalin, iodine tincture and many other chemical substances. Strong acids or bases, readily decomposes proteins and lipids through amide and ester hydrolysis upon contact with living tissues, such as skin and flesh.

Oxidizing Agents as Poisons

Household cleaning products with strong oxidizing agents like **sodium hypochlorite** (NaOCl) and **hydrogen peroxide** (H_2O_2) are dangerous when ingested. Generally speaking, the ingestion of bleaches (Figure $9.2.2$) will cause damage to the esophagus and stomach, possibly leading to death. On contact with the skin or eyes, it causes irritation, drying, and potentially burns. Inhalation of bleach fumes can damage the lungs. Personal protective equipment should always be used when using bleach. Bleach should never be mixed with vinegar or other acids as this will create highly toxic chlorine gas and can cause severe burns internally and externally. Mixing bleach with ammonia similarly produces toxic chloramine gas, which can burn the lungs. Mixing bleach with hydrogen peroxide results in an exothermic chemical reaction that releases oxygen, and may cause the contents to splatter and cause skin and eye injury. Heating bleach and boiling it may produce chlorates, a strong oxidizer which may lead to a fire or explosion.



Figure $9.2.1$ Clorox bleach. Source: Wikipedia.

Ozone. For the last few decades, scientists studied the effects of acute and chronic ozone exposure on human health. Hundreds of studies suggest that ozone is harmful to people at levels currently found in urban areas. Ozone has been shown to affect the

respiratory, cardiovascular and central nervous system. Early death and problems in reproductive health and development are also shown to be associated with ozone exposure.

Acute ozone exposure ranges from hours to a few days. Because ozone is gas, it directly affects the lungs and the entire respiratory system. Inhaled ozone causes inflammation and acute -but reversible- changes in lung function, as well as airway hyperresponsiveness. These changes lead to shortness of breath, wheezing, and coughing which may exacerbate lung diseases, like asthma or chronic obstructive pulmonary disease (COPD) resulting in the need to receive medical treatment. Acute and chronic exposure to ozone has been shown to cause an increased risk of respiratory infections, due to the following mechanism.

Multiple studies have been conducted to determine the mechanism behind ozone's harmful effects, particularly in the lungs. These studies have shown that exposure to ozone causes changes in the immune response within the lung tissue, resulting in disruption of both the innate and adaptive immune response, as well as altering the protective function of lung epithelial cells. It is thought that these changes in immune response and the related inflammatory response are factors that likely contribute to the increased risk of lung infections, and worsening or triggering of asthma and reactive airways after exposure to ground-level ozone pollution.

Inhaling ozone not only affects the immune system and lungs, but it may also affect the heart as well. Ozone causes short-term autonomic imbalance leading to changes in heart rate and reduction in heart rate variability; and high levels exposure for as little as one-hour results in a supraventricular arrhythmia in the elderly, both increase the risk of premature death and stroke. Ozone may also lead to vasoconstriction resulting in increased systemic arterial pressure contributing to increased risk of cardiac morbidity and mortality in patients with pre-existing cardiac diseases.

Metabolic Poisons

Carbon monoxide and cyanide are chemical substances that interfere with the normal bodily functions i.e. during the delivery of oxygen to bodily tissues and cellular respiration.

Carbon Monoxide Poisoning

Carbon monoxide poisoning typically occurs from breathing in carbon monoxide (CO) at excessive levels. Symptoms (Figure 9.2.2) are often described as "flu-like" and commonly include headache, dizziness, weakness, vomiting, chest pain, and confusion. Large exposures can result in loss of consciousness, arrhythmias, seizures, or death. The classically described "cherry red skin" rarely occurs. Long-term complications may include chronic fatigue, trouble with memory, and movement problems.

CO is a colorless and odorless gas which is initially non-irritating. It is produced during incomplete burning of organic matter. This can occur from motor vehicles, heaters, or cooking equipment that run on carbon-based fuels. Carbon monoxide primarily causes adverse effects by combining with hemoglobin to form carboxyhemoglobin (HbCO) preventing the blood from carrying oxygen and expelling carbon dioxide as carbaminohemoglobin. Additionally, many other hemoproteins such as myoglobin, Cytochrome P450, and mitochondrial cytochrome oxidase are affected, along with other metallic and non-metallic cellular targets.

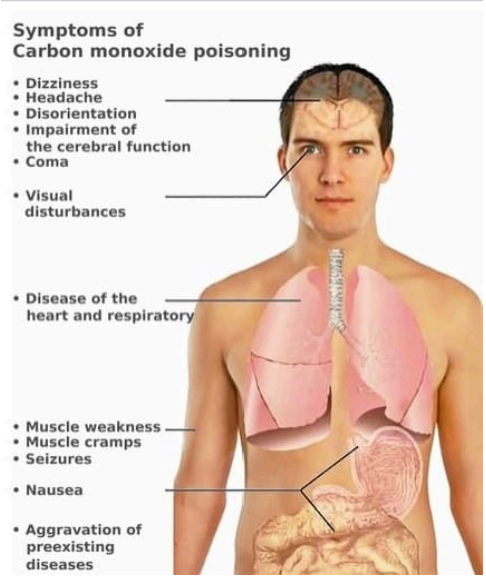


Figure 9.2.2 CO toxicity symptoms. Source: Wikipedia.

Initial treatment for carbon monoxide poisoning is to immediately remove the person from the exposure without endangering further people. Those who are unconscious may require CPR on site. Administering oxygen via non-rebreather mask shortens the half-life of carbon monoxide from 320 minutes, when breathing normal air, to only 80 minutes. Oxygen hastens the dissociation of carbon monoxide from carboxyhemoglobin, thus turning it back into hemoglobin. Due to the possible severe effects in the baby, pregnant women are treated with oxygen for longer periods of time than non-pregnant people.

Cyanide Poisoning

Cyanide poisoning is poisoning that results from exposure to a number of forms of cyanide. Early symptoms include headache, dizziness, fast heart rate, shortness of breath, and vomiting. This may then be followed by seizures, slow heart rate, low blood pressure, loss of consciousness, and cardiac arrest. Onset of symptoms is usually within a few minutes. If a person survives, there may be long-term neurological problems.

Toxic cyanide-containing compounds include hydrogen cyanide gas and a number of cyanide salts. Poisoning is relatively common following breathing in smoke from a house fire. Other potential routes of exposure include workplaces involved in metal polishing, certain insecticides, and the medication nitroprusside. Liquid forms of cyanide can be absorbed through the skin. Cyanide ions interfere with cellular respiration, resulting in the body's tissues being unable to use oxygen.

In addition to its uses as a pesticide and insecticide, cyanide is contained in tobacco smoke and smoke from building fires, and is present in many seeds or kernels such as those of almonds, apricots, apples, oranges, and in foods including cassava (also known as tapioca, yuca or manioc), and bamboo shoots. Vitamin B12, in the form of hydroxocobalamin (also spelled hydroxycobalamin), may reduce the negative effects of chronic exposure, and a deficiency can lead to negative health effects following exposure. Flaxseed also contains cyanogenic glycosides.

Cyanide is a potent cytochrome *c* oxidase (COX, a.k.a Complex IV) inhibitor. As such, cyanide poisoning is a form of histotoxic hypoxia, because it interferes with an essential step in aerobic metabolism called oxidative phosphorylation.

The United States standard cyanide antidote kit first uses a small inhaled dose of amyl nitrite, followed by intravenous sodium nitrite, followed by intravenous sodium thiosulfate. Hydroxocobalamin is newly approved in the US and is available in Cyanokit antidote kits. Sulfanegen TEA, which could be delivered to the body through an intra-muscular (IM) injection, detoxifies cyanide and converts the cyanide into thiocyanate, a less toxic substance. Alternative methods of treating cyanide intoxication are used in other countries.

Arsenic Poisoning

Arsenic poisoning is a medical condition that occurs due to elevated levels of arsenic in the body. If arsenic poisoning occurs over a brief period of time, symptoms may include vomiting, abdominal pain, encephalopathy, and watery diarrhea that contains

blood. Long-term exposure can result in thickening of the skin, darker skin, abdominal pain, diarrhea, heart disease, numbness, and cancer.

The most common reason for long-term exposure is contaminated drinking water. Groundwater most often becomes contaminated naturally; however, contamination may also occur from mining or agriculture. It may also be found in the soil and air. Recommended levels in water are less than 10–50 µg/L (10–50 parts per billion). Other routes of exposure include toxic waste sites and traditional medicines. Most cases of poisoning are accidental. Arsenic acts by changing the functioning of around 200 enzymes. Diagnosis is by testing the urine, blood, or hair.

Prevention is by using water that does not contain high levels of arsenic. This may be achieved by the use of special filters or using rainwater. There is not good evidence to support specific treatments for long-term poisoning. For acute poisonings treating dehydration is important. Dimercaptosuccinic acid (DMSA) or dimercaptopropane sulfonate (DMPS) may be used while dimercaprol (BAL) is not recommended. Hemodialysis may also be used.

WEBLINK

More Information on arsenic can be found on the link below

https://www.niehs.nih.gov/health/topics/agents/arsenic/index.cfm

Heavy Metal Poisoning

Lead poisoning, also known as **plumbism** and **saturnism**, is a type of metal poisoning caused by lead in the body. The brain is the most sensitive. Symptoms may include abdominal pain, constipation, headaches, irritability, memory problems, infertility, and tingling in the hands and feet. It causes almost 10% of intellectual disability of otherwise unknown cause and can result in behavioral problems. Some of the effects are permanent. In severe cases, anemia, seizures, coma, or death may occur.

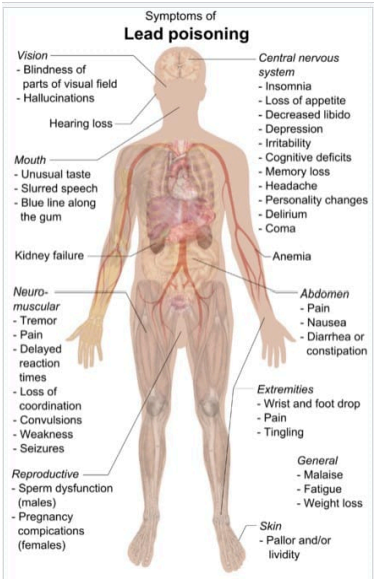


Figure 9.2.3. Symptoms of lead poisoning. SOurce: Wikipedia.

Exposure to lead can occur by contaminated air, water, dust, food, or consumer products. Children are at greater risk as they are more likely to put objects in their mouth such as those that contain lead paint and absorb a greater proportion of the lead that they eat. Exposure at work is a common cause of lead poisoning in adults with certain occupations at particular risk. Diagnosis is typically by measurement of the blood lead level. The Centers for Disease Control (US) has set the upper limit for blood lead for adults at 10 µg/dl (10 µg/100 g) and for children at 5 µg/dl. Elevated lead may also be detected by changes in red blood cells or dense lines in the bones of children as seen on X-ray.

Lead has no known physiologically relevant role in the body, and its harmful effects are myriad. Lead and other heavy metals create

reactive radicals which damage cell structures including DNA and cell membranes. Lead also interferes with DNA transcription, enzymes that help in the synthesis of vitamin D, and enzymes that maintain the integrity of the cell membrane. Anemia may result when the cell membranes of red blood cells become more fragile as the result of damage to their membranes. Lead interferes with metabolism of bones and teeth and alters the permeability of blood vessels and collagen synthesis. Lead may also be harmful to the developing immune system, causing production of excessive inflammatory proteins; this mechanism may mean that lead exposure is a risk factor for asthma in children. Lead exposure has also been associated with a decrease in activity of immune cells such as polymorphonuclear leukocytes. Lead also interferes with the normal metabolism of calcium in cells and causes it to build up within them.

Lead poisoning is preventable. This includes individual efforts such as removing lead-containing items from the home, workplace efforts such as improved ventilation and monitoring, state laws that ban the use of and national policies such as laws that ban lead in products such as paint, gasoline, ammunition, wheel weights, and fishing weights reduce allowable levels in water or soil, and provide for cleanup of contaminated soil. Workers' education could be helpful as well. The major treatments are removal of the source of lead and the use of medications that bind lead so it can be eliminated from the body, known as chelation therapy. Chelation therapy in children is recommended when blood levels are greater than 40–45 µg/dl. Medications used include dimercaprol, edetate calcium disodium, and succimer.

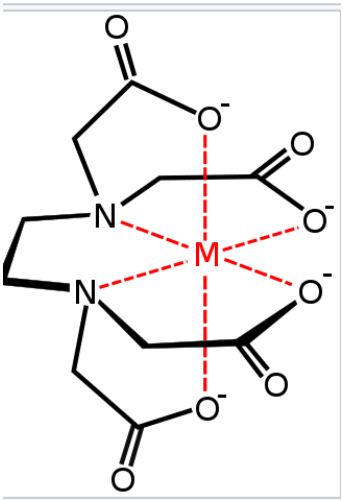


Figure 9.2.4 EDTA, a chelating agent, binds a heavy metal, sequestering it.

WEBLINKS

Links to various lead and mercury related topics can be found below

https://medlineplus.gov/leadpoisoning.html

https://medlineplus.gov/mercury.html

Mercury poisoning is a type of metal poisoning due to exposure to mercury. Symptoms depend upon the type, dose, method, and duration of exposure. They may include muscle weakness, poor coordination, numbness in the hands and feet, skin rashes, anxiety, memory problems, trouble speaking, trouble hearing, or trouble seeing. High-level exposure to methylmercury is known as Minamata disease. Methylmercury exposure in children may result in acrodynia (pink disease) in which the skin becomes pink and peels. Long-term complications may include kidney problems and decreased intelligence. The effects of long-term low-dose exposure to methylmercury are unclear.

Forms of mercury exposure include metal, vapor, salt, and organic compound. Most exposure is from eating fish, amalgam based dental fillings, or exposure at work. In fish, those higher up in the food chain generally have higher levels of mercury. Less commonly, poisoning may occur as a method of attempted suicide. Human activities that release mercury into the environment include the burning of coal and mining of gold. Tests of the blood, urine, and hair for mercury are available but do not relate well to the amount in the body.

The primary mechanism of mercury toxicity involves its irreversible inhibition of selenoenzymes, such as thioredoxin reductase.

High mercury exposures deplete the amount of cellular selenium available for the biosynthesis of thioredoxin reductase and other selenoenzymes that prevent and reverse oxidative damage, which, if the depletion is severe and long lasting, results in brain cell dysfunctions that can ultimately cause death.

Prevention includes eating a diet low in mercury, removing mercury from medical and other devices, proper disposal of mercury, and not mining further mercury. In those with acute poisoning from inorganic mercury salts, chelation with either dimercaptosuccinic acid (DMSA) or dimercaptopropane sulfonate (DMPS) appears to improve outcomes if given within a few hours of exposure. Chelation for those with long-term exposure is of unclear benefit.

Cadmium poisoning. Cadmium is a naturally occurring toxic metal with common exposure in industrial workplaces, plant soils, and from smoking. Due to its low permissible exposure in humans, overexposure may occur even in situations where trace quantities of cadmium are found. Cadmium is used extensively in electroplating, although the nature of the operation does not generally lead to overexposure. Cadmium is also found in some industrial paints and may represent a hazard when sprayed. Operations involving removal of cadmium paints by scraping or blasting may pose a significant hazard. The primary use of cadmium is in the manufacturing of NiCd rechargeable batteries. The primary source for cadmium is as a byproduct of refining zinc metal. Exposures to cadmium are addressed in specific standards for the general industry, shipyard employment, the construction industry, and the agricultural industry.

Acute exposure to cadmium fumes may cause flu-like symptoms including chills, fever, and muscle ache sometimes referred to as "the cadmium blues." Symptoms may resolve after a week if there is no respiratory damage. More severe exposures can cause tracheo-bronchitis, pneumonitis, and pulmonary edema. Symptoms of inflammation may start hours after the exposure and include cough, dryness and irritation of the nose and throat, headache, dizziness, weakness, fever, chills, and chest pain.

Inhaling cadmium-laden dust quickly leads to respiratory tract and kidney problems which can be fatal (often from kidney failure). Ingestion of any significant amount of cadmium causes immediate poisoning and damage to the liver, bones, and the kidneys. Compounds containing cadmium are also carcinogenic.

Summary

- Corrosive poisons are chemicals (that can be found in many household products) that are considered highly reactive and thus are very dangerous upon ingestion or skin contact.
- Metabolic poisons such as carbon monoxide and cyanide interfere with the bodily functions.
- Human exposure to lead, mercury, and cadmium have affected the nervous system, respiratory system, brain, and other various organs.

Contributors

- Marisa Alviar-Agnew (Sacramento City College)

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- US NIH MedlinePlus
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- Robert J. Lancashire (University of West Indies-Mona)

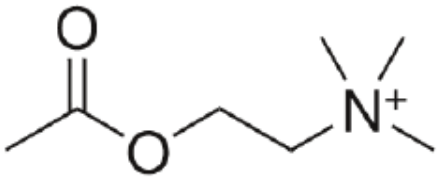
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9.3: More Chemistry of the Nervous System

Learning Objectives

- Describe the overall mechanism of neurotransmission.
- Describe how various nerve agents interrupt the neurotransmission process.

Acetylcholine is an organic chemical that functions in the brain and body of many types of animals, including humans, as a neurotransmitter—a chemical released by nerve cells to send signals to other cells. Its name is derived from its chemical structure: it is an ester of acetic acid and choline. Parts in the body that use or are affected by acetylcholine are referred to as cholinergic. Substances that interfere with acetylcholine activity are called anticholinergics.



Acetylcholine: The chemical structure of acetylcholine is depicted.

Acetylcholine is the neurotransmitter used at the neuromuscular junction—in other words, it is the chemical that motor neurons of the nervous system release in order to activate muscles. This property means that drugs that affect cholinergic systems can have very dangerous effects ranging from paralysis to convulsions. Acetylcholine is also used as a neurotransmitter in the autonomic nervous system, both as an internal transmitter for the sympathetic nervous system and as the final product released by the parasympathetic nervous system.

In the brain, acetylcholine functions as a neurotransmitter and as a neuromodulator. The brain contains a number of cholinergic areas, each with distinct functions. They play an important role in arousal, attention, memory and motivation.

Mechanism of Action

When a normally functioning motor nerve is stimulated, it releases the neurotransmitter acetylcholine, which transmits the impulse to a muscle or organ.

1. During neurotransmission (Figure $9.3.1$), ACh is released from the presynaptic neuron into the synaptic cleft and binds to ACh receptors on the post-synaptic membrane, relaying the signal from the nerve.

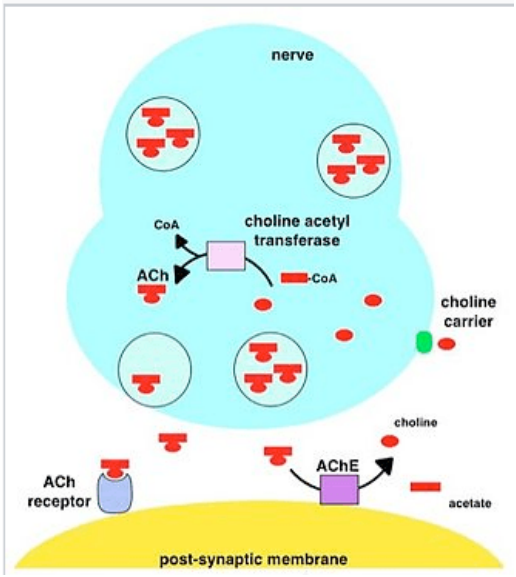


Figure 9.3.1 Different steps during neurotransmission.

2. AChE, also located on the post-synaptic membrane, terminates the signal transmission by hydrolyzing ACh (Figure 9.3.2). Source: Wikipedia.

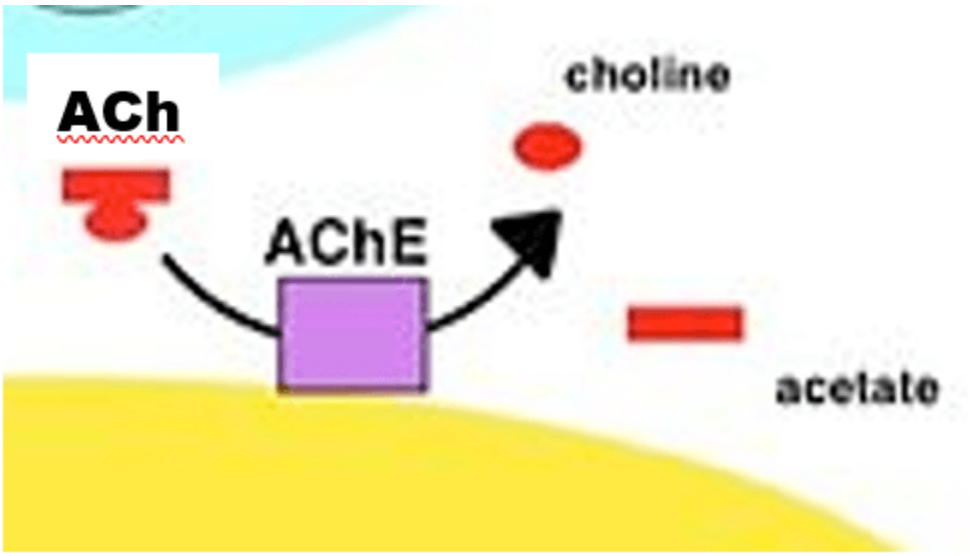


Figure 9.3.2 Acetylcholinesterase (AChe) breaks down acetylcholine (ACh) to choline and acetate (acetic acid). Source: Wikipedia.

3. The liberated choline is taken up again by the pre-synaptic neuron and ACh is synthesized by combining with acetyl-CoA through the action of choline acetyltransferase.[19][20]

Nerve Poisons and the Acetylcholine Cycle

Botulinum toxin (**Botox**) is a neurotoxic protein produced by the bacterium *Clostridium botulinum* and related species.[1] It prevents the release of the neurotransmitter acetylcholine from axon endings at the neuromuscular junction and thus causes flaccid paralysis.[2]The toxin causes the disease botulism. The toxin is also used commercially for medical and cosmetic purposes.

Foodborne botulism can be transmitted through food that has not been heated correctly prior to being canned, or food from a can that has not been cooked correctly. Most infant botulism cases cannot be prevented because the bacteria that cause this disease are in soil and dust. The bacteria can also be found inside homes on floors, carpet, and countertops, even after cleaning.

Honey can contain the bacteria that cause infant botulism, so children less than 12 months old should not be fed honey.

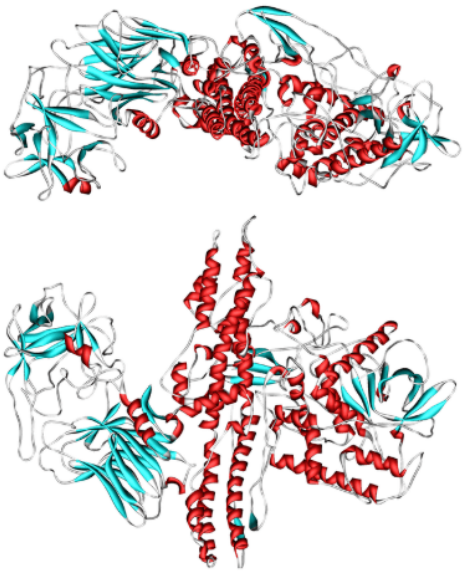
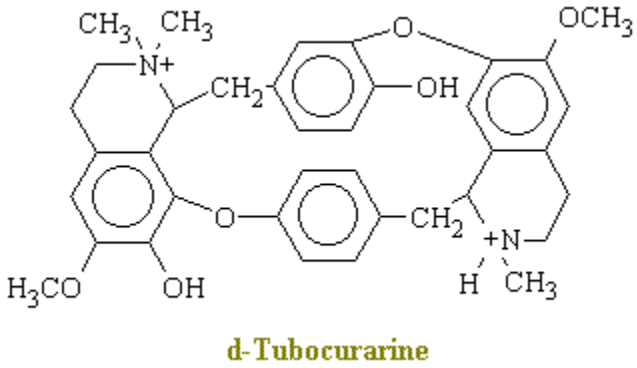


Figure 9.3.3 *Structure of Botulinum toxin.*

Curare is a generic term for various South American arrow poisons. The main active ingredient in curare is d-tubocurarine (Figure 9.3.4), which has the chemical structure shown below.



d-Tubocurarine

Figure 9.3.4 *Structure of d-tubocurarine.*

In brief, d-tubocurarine, blocks competitively the transmitter action of acetylcholine. When the drug is applied directly to the end-plate of a single isolated muscle fiber under microscopic control, the muscle cell becomes insensitive to motor-nerve impulses and to direct applied acetylcholine; however, the end -plate region and the remainder of the muscle fiber membrane retain their normal sensitivity to the application of potassium ions, and the muscle fiber still responds to direct electrical stimulation. Because acetylcholine release into the neuromuscular junction muscle is what initiates contraction, curare causes muscle relaxation and paralysis.

Nerve agents, sometimes also called **nerve gases**, are a class of organic chemicals that disrupt the mechanisms by which nerves transfer messages to organs. The disruption is caused by the blocking of acetylcholinesterase (AChE), an enzyme that catalyzes the breakdown of acetylcholine, a neurotransmitter. Nerve agents are acetylcholinesterase inhibitors used as poison.

Poisoning by a nerve agent leads to constriction of pupils, profuse salivation, convulsions, and involuntary urination and defecation, with the first symptoms appearing in seconds after exposure. Death by asphyxiation or cardiac arrest may follow in minutes due to the loss of the body's control over respiratory and other muscles. Some nerve agents are readily vaporized or aerosolized, and the primary portal of entry into the body is the respiratory system. Nerve agents can also be absorbed through the skin, requiring that those likely to be subjected to such agents wear a full body suit in addition to a respirator.

Nerve agents are generally colorless to amber-colored, tasteless liquids that may evaporate to a gas. Agents Sarin and VX are

odorless; Tabun has a slightly fruity odor and Soman has a slight camphor odor.

Organophosphorous Compounds as Insecticides and Weapons of War

This first class of nerve agents, the G-series, was accidentally discovered in Germany on December 23, 1936, by a research team headed by Gerhard Schrader working for IG Farben. Since 1934, Schrader had been working in a laboratory in Leverkusen to develop new types of insecticides for IG Farben. While working toward his goal of improved insecticide, Schrader experimented with numerous compounds, eventually leading to the preparation of tabun (GA).

In experiments, tabun was extremely potent against insects: as little as 5 ppm of tabun killed all the leaf lice he used in his initial experiment. In January 1937, Schrader observed the effects of nerve agents on human beings first-hand when a drop of tabun spilled onto a lab bench. Within minutes he and his laboratory assistant began to experience miosis (constriction of the pupils of the eyes), dizziness and severe shortness of breath. It took them three weeks to recover fully. The G-series naming system was created by the United States when it uncovered the German activities, labeling (Figure 9.3.5.

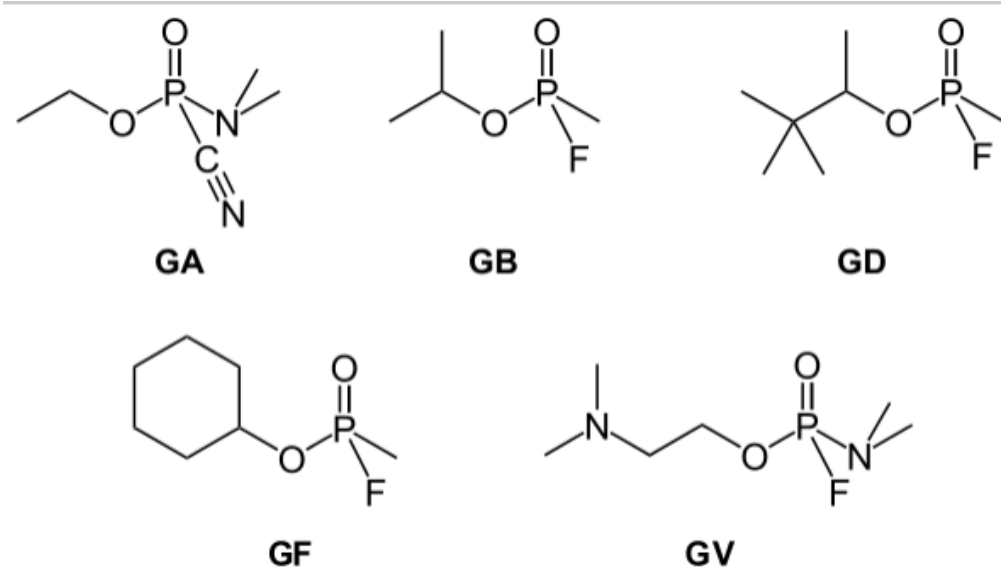


Figure 9.3.5 The G-series of nerve agents. Tabun as GA (German Agent A), sarin as GB and soman as GD. Ethyl sarin was tagged GE and cyclosarin as GF. *Source: Wikipedia.*

In 1935 the Nazi government had passed a decree that required all inventions of possible military significance to be reported to the Ministry of War, so in May 1937 Schrader sent a sample of tabun to the chemical warfare (CW) section of the Army Weapons Office in Berlin-Spandau. Schrader was summoned to the Wehrmacht chemical lab in Berlin to give a demonstration, after which Schrader's patent application and all related research was classified as secret. Colonel Rüdiger, head of the CW section, ordered the construction of new laboratories for the further investigation of tabun and other organophosphate compounds and Schrader soon moved to a new laboratory at Wuppertal-Elberfeld in the Ruhr valley to continue his research in secret throughout World War II. The compound was initially codenamed Le-100 and later Trilon-83.

Sarin was discovered by Schrader and his team in 1938 and named in honor of its discoverers: **S**chrader, **A**mbros, Gerhard **R**itter, and von der L**in**de.[26] It was codenamed T-144 or Trilon-46. It was found to be more than ten times as potent as tabun.

Soman was discovered by Richard Kuhn in 1944 as he worked with the existing compounds; the name is derived from either the Greek 'to sleep' or the Latin 'to bludgeon'. It was codenamed T-300.

Cyclosarin was also discovered during WWII but the details were lost and it was rediscovered in 1949.

Some **insecticides**, including carbamates and organophosphates such as dichlorvos, malathion and parathion, are nerve agents (Figure 9.3.6). The metabolism of insects is sufficiently different from mammals that these compounds have little effect on humans and other mammals at proper doses, but there is considerable concern

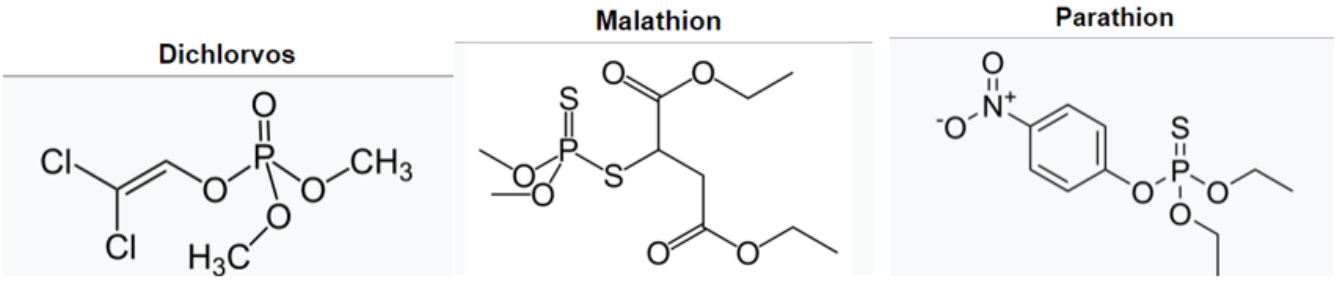


Figure 9.3.6 Organophosphate insecticides. Source: Wikipedia

about the long-term exposure to these chemicals by farm workers and animals alike. At high enough doses, acute toxicity and death can occur through the same mechanism as other nerve agents. Some insecticides such as demeton, dimefox and paraoxon are sufficiently toxic to humans that they have been withdrawn from agricultural use, and were at one stage investigated for potential military applications. Paraoxon was allegedly used as an assassination weapon by the apartheid South African government as part of Project Coast. Organophosphate pesticide poisoning is a major cause of disability in many developing countries and is often the preferred method of suicide.[24]

Treatment

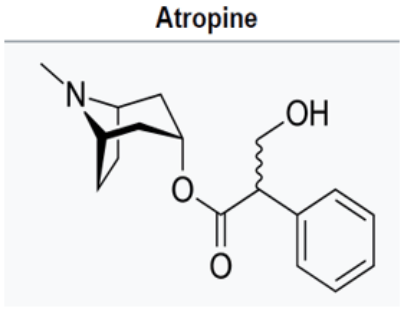
Standard treatment for nerve agent poisoning is a combination of an anticholinergic to manage the symptoms, and an oxime as an antidote. Anticholinergics treat the symptoms by reducing the effects of acetylcholine, while oximes displaces phosphate molecules from the active site of the cholinesterase enzymes, allowing the breakdown of acetylcholine. Military personnel are issued the combination in an autoinjector (e.g. ATNAA), for ease of use in stressful conditions.

Atropine (Figure 9.3.7 is the standard anticholinergic drug used to manage the symptoms of nerve agent poisoning. It acts as an antagonist to muscarinic acetylcholine receptors, blocking the effects of excess acetylcholine. Some synthetic anticholinergics, such as biperiden, may counteract the central symptoms of nerve agent poisoning more effectively than atropine, since they pass the blood–brain barrier better than atropine. While these drugs will save the life of a person affected by nerve agents, that person may be incapacitated briefly or for an extended period, depending on the extent of exposure. The endpoint of atropine administration is the clearing of bronchial secretions.

Pralidoxime chloride (also known as *2-PAMCl*) is the standard oxime used to treat nerve agent poisoning. Rather than counteracting the initial effects of the nerve agent on the nervous system as does atropine, pralidoxime chloride reactivates the poisoned enzyme (acetylcholinesterase) by scavenging the phosphoryl group attached on the functional hydroxyl group of the enzyme, counteracting the nerve agent itself. Revival of acetylcholinesterase with pralidoxime chloride works more effectively on nicotinic receptors while blocking acetylcholine receptors with atropine is more effective on muscarinic receptors.

Anticonvulsants, such as diazepam, may be administered to manage seizures, improving long term prognosis and reducing risk of brain damage. This is not usually self-administered as its use is for actively seizing patients.

a.



b.



Figure 9.3.7 a. Atrophine molecule and b. an ampoule containing atropine injection 1mL/0.5mg.

Summary

- Upon stimulation, a normally functioning motor nerve releases the neurotransmitter acetylcholine, which transmits the impulse to a muscle or organ.
- Nerve poisons (i.e. carbamates and organophosphorus compounds) disrupt the nervous system by inhibiting the function of the enzyme acetylcholinesterase. Acetylcholine thus builds up and continues to act so that any nerve impulses are continually transmitted and muscle contractions do not stop.
- The botulin toxin and curare interfere with nerve impulses (by preventing the release of acetylcholine) and causes flaccid (sagging) paralysis of muscles.
- Atropine (Figure 9.3.7 is the standard anticholinergic drug used to manage the symptoms of nerve agent poisoning.

Contributors and Attributions

- Wikipedia

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9.4: The Lethal Dose

Learning Objectives

- Describe the LD_{50} value and its limitations.
- Know the LD_{50} value of various substances.

The LD_{50} is a standardized measure for expressing and comparing the toxicity of chemicals. The LD_{50} is the dose that kills half (50%) of the animals tested (LD = "lethal dose"). The animals are usually rats or mice, although rabbits, guinea pigs, hamsters, and so on are sometimes used. In all these tests, the dose must be calculated relative to the size of the animal. The most common units are milligrams of chemical per kilogram of test animal (mg/kg or ppm). (Table 9.4.1) provides examples of LD_{50} values for various substances.

As a measure of toxicity, lethal dose is somewhat unreliable and results may vary greatly between testing facilities due to factors such as the genetic characteristics of the sample population, animal species tested, environmental factors and mode of administration.

There can be wide variability between species as well; what is relatively safe for rats may very well be extremely toxic for humans (*cf.* paracetamol toxicity), and vice versa. For example, chocolate, comparatively harmless to humans, is known to be toxic to many animals. When used to test venom from venomous creatures, such as snakes, LD_{50} results may be misleading due to the physiological differences between mice, rats, and humans. Many venomous snakes are specialized predators of mice, and their venom may be adapted specifically to incapacitate mice; and mongooses may be exceptionally resistant. While most mammals have a very similar physiology, LD_{50} results may or may not have equal bearing upon every mammal species, including humans.

Table 9.4.1: LD_{50} Values of Various Substances. Source: Wikipedia

Substance	Animal Route	LD_{50} (mg/kg)
Water	rat, oral	90,000
Sucrose (table sugar)	rat, oral	29,700
Glucose (blood sugar)	rat, oral	25,800
Monosodium glutamate (MSG)	rat, oral	16,600
Cadmium sulfide	rat, oral	7,080
Ethanol (grain alcohol)	rat, oral	7,060
Methanol	human, oral	810
Sodium chloride (table salt)	rat, oral	3,000
Metallic arsenic	rat, oral	763
Arsenic trisulfide	rat, oral	185-6400
Sodium cyanide	rat, oral	6.4
Hydrogen cyanide	mouse, oral	3.7
Ibuprofen	rat, oral	636
Aspirin	rat, oral	200
Caffeine	rat, oral	192
Cocaine	mouse, oral	96
Nicotine	rat, oral	50

Substance	Animal Route	LD_{50} (mg/kg)
Heroin (diamorphine)	mouse, intravenous	21.8
Methamphetamine	rat, intraperitoneal	57
Mercury(II) chloride	rat, oral	1
Strychnine	human, oral	1-2
Sarin	mouse, subcutaneous injection	172 microgram/kg
Ricin (from castor oil plant)	rat, oral	20-30
Botulinum toxin (Botox)	human, oral injection, inhalation	1 ng/kg

(Table 9.4.2) gives the LD50 values of some insecticides. In each case, the chemical was fed to laboratory rats. Note that the lower the LD50, the more toxic the chemical. Even adjusting for the test animal's weight, the LD50 for one species is often quite different from that for another. Thus any LD50 value gives only a rough estimate of the risk to humans. The way in which the chemical is administered also has a marked effect on LD50 values. The chemical may be fed, injected, applied to the animal's skin, etc., and each method usually generates a different LD50.

Table 9.4.2: LD_{50} Values of Pesticides.

Chemical	Category	Oral LD_{50} in Rats (mg/kg)
Aldicarb ("Temik")	Carbamate	1
Carbaryl ("Sevin")	Carbamate	307
DDT	Chlorinated hydrocarbon	87
Dieldrin	Chlorinated hydrocarbon	40
Diflubenzuron ("Dimilin")	Chitin inhibitor	10,000
Malathion	Organophosphate	885
Methoprene	JH mimic	34,600
Methoxychlor	Chlorinated hydrocarbon	5,000
Parathion	Organophosphate	3
Piperonyl butoxide	Synergist	7,500
Pyrethrin	Plant extract	200
Rotenone	Plant extract	60

Summary

- The LD50 is the dose that kills half (50%) of the animals tested (LD = "lethal dose"). It is a standardized measure for expressing and comparing the toxicity of chemicals.
- LD50 values are given for common drugs (e.g. ibuprofen and aspirin), common household ingredients including water, table sugar (sucrose), salt (sodium chloride), insecticides (e.g. DDT and pyrethrin), and harmful drugs (e.g. cocaine and heroin).

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9.5: The Liver as a Detox Facility

Learning Objective

- Describe the different mechanisms of drug detoxification in the liver.

The human liver is thought to be responsible for up to 500 separate functions, usually in combination with other systems and organs. The liver breaks down or modifies toxic substance, such as alcohol and most medicial products, in a process called drug metabolism. This sometimes results in toxication, when the metabolite is more toxic than its precursor.

Alcohol Metabolism

Giving the liver enough time to fully metabolize the ingested alcohol is the only effective way to avoid alcohol toxicity. Drinking coffee or taking a shower will not help. The legal limit for intoxication is a BAC of 0.08. Taking into account the rate at which the liver metabolizes alcohol after drinking stops, and the alcohol excretion rate, it takes at least five hours for a legally intoxicated person to achieve sobriety.

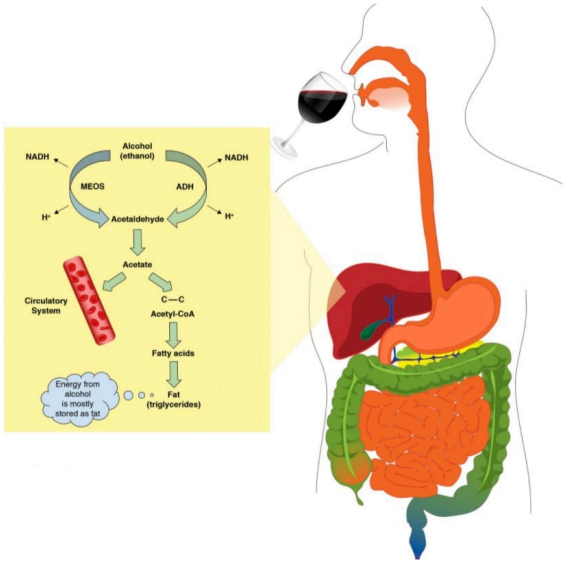


Figure 9.5.1: Alcohol Metabolism Summary.

The liver metabolizes up to 85% – 98% of the circulatory ethanol. The liver uses two metabolic processes to get rid of this circulatory ethanol as quickly and safely as possible.

1. Alcohol dehydrogenase system
2. Microsomal ethanol oxidizing system (MEOS)

Alcohol Dehydrogenase System

About 80 to 90% of the total hepatic ethanol uptake is processed via the alcohol dehydrogenase system.The degradation of ethanol begins in the liver. The enzyme that catalyzes this reaction is called alcohol dehydrogenase. The products from this reaction are acetaldehyde, NADH (a reduced coenzyme that carries electrons from one reaction to another) and H+ ion. Acetaldehyde is very toxic to the liver and the body's cells. The moment acetaldehyde is produced; it must be degraded to protect the liver cells. The enzyme that will carry this type of degradation reaction is acetaldehyde dehydrogenase (ALDH). Acetaldehyde dehydrogenase converts acetaldehyde into acetate, a non-toxic molecule.

Microsomal Ethanol Oxidizing System (MEOS)

In a moderate drinker, about 10 to 20% of the total liver ethanol uptake is processed via the microsomal ethanol oxidizing system (MEOS). During periods of heavy drinking, the MEOS system will metabolize most of the excess ethanol ingested. Heavy drinking stimulates the human body to include the MEOS system enzymes to clear ethanol faster from the body.

The MEOS system is also located in the liver. Similar to the Alcohol dehydrogenase system, acetaldehyde dehydrogenase will immediately convert acetaldehyde into acetate, a non-toxic molecule. Other products from this reaction are NADH and H+ ion.

Fate of Acetate

The acetate produced (from the alcohol dehydrogenase system and microsomal ethanol oxidizing system) is either released into circulation or retained inside the liver cells. In the liver cells, acetate is converted to acetyl CoA where it is used to produce other molecules like CO2 or used in the synthesis of fatty acids and cholesterol.

Nicotine and Ammonia

In humans, **nicotine i**s extensively metabolized in the liver into various metabolites. About 70-80% of nicotine is broken down to cotinine, the predominant metabolite. Other primary metabolites include nicotine *N*'-oxide, nornicotine, nicotine isomethonium ion, 2-hydroxynicotine and nicotine glucuronide. Under some conditions, other substances may be formed such as myosmine.

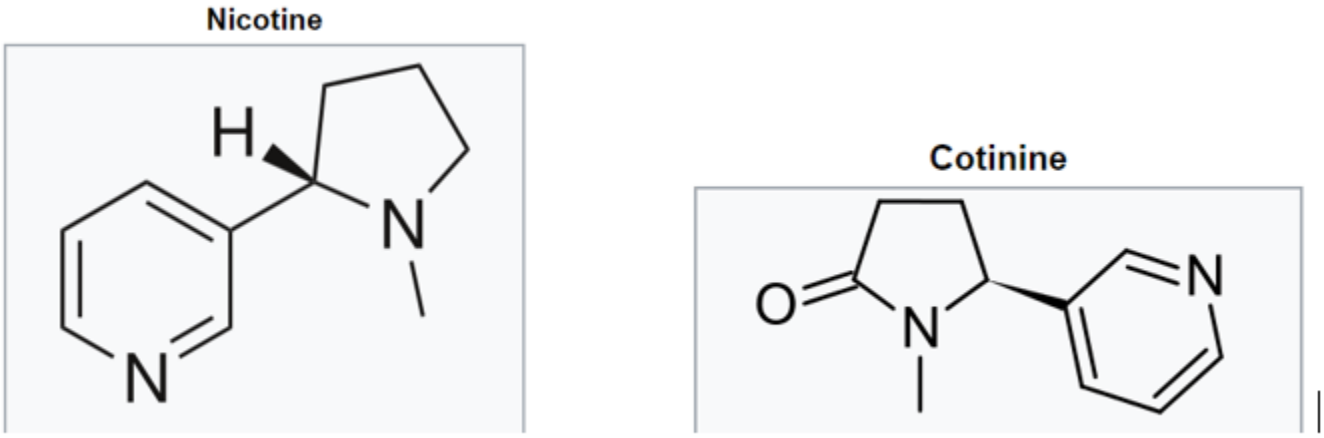


Figure 9.5.1 Nicotine and cotinine.

The liver converts ammonia into urea as part of the ornithine cycle or the urea cycle, and the urea is excreted in the urine.

Summary

- The liver is involved in the breakdown or modification of various substances in a process called drug metabolism, but the metabolites are not directly removed from the body.
- Ethanol is broken down through a series of steps (involving various liver enzymes) to carbon dioxide and water.
- Ethanol therapy is the traditional treatment for methanol and ethylene glycol poisoning.
- Nicotine is extensively metabolized in the liver into a variety of breakdown products whereas toxic ammonia is converted to non toxic urea.

Contributors and Attributions

- Wikipedia

- Benowitz, N. L., Hujjanen, J., Peyton III, J. Nicotine Chemistry, Metabolism, Kinetics, and Biomarkers. Hand Exp Pharmacol. 2009. (192): 29-60.
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9.6: Carcinogens and Teratogens

Learning Objectives

- Explain the process of cancer development.
- List the factors that may increase and lower the risk of cancer.
- Define teratogens and describe the birth defects that result from them.

A **carcinogen** is any agent that directly increases the incidence of cancer. Most, but not all carcinogens are mutagens. Carcinogens that do not directly damage DNA include substances that accelerate cell division, thereby leaving less opportunity for cell to repair induced mutations, or errors in replication. Carcinogens that act as mutagens may be biological, physical, or chemical in nature, although the term is most often used in relation to chemical substances.

Defining Cancer

Cancer is actually a group of more than 100 diseases, all of which involve abnormal cell growth with the potential to invade or spread to other parts of the body. In general terms, cancer occurs when the cell cycle is no longer regulated due to DNA damage. The number of potential underlying causes of this DNA damage is great, so there are many different risk factors for cancer. Any cells that become cancerous divide more quickly than normal cells. They may form a mass of abnormal cells called a tumor. The rapidly dividing cells take up nutrients and space, damaging the normal cells around them. If the cancer cells spread to other parts of the body, they invade and damage other tissues and organs. They may eventually lead to death.

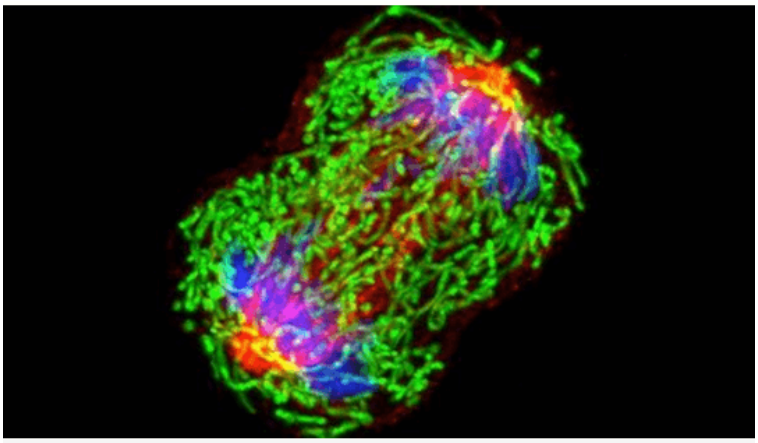


Figure 9.6.1 A dividing breast cancer cell. Source: National Cancer Institute/Univ. of Pittsburgh Cancer Institute

By far, the most common of the 100-plus types of human cancer is basal cell carcinoma, the type of skin cancer Bernie Sanders had removed in 2016. Basal cell carcinoma makes up 40 percent of all new cancers each year in the United States. Other common types of cancer include lung, colorectal, prostate (in males), and breast (in females) cancers. These cancers are not as common as skin cancer, but they cause the majority of cancer deaths.

How Cancer Develops

Carcinogenesis, also called **oncogenesis** or **tumorigenesis,** is the formation of a cancer, whereby normal cells are transformed into cancer cells. The transformation of a normal cell into a cancer cell is a multi-step process that involves initiation, promotion, progression and finally malignancy (see Figure 9.6.2). This process takes years and starts with a single cell in which the right genes are mutated so the cell does not appropriately die and begins to proliferate abnormally. Then, additional mutations occur that select for more rapidly growing cells within this population leading to a tumor with rapid growth and malignancy. By the time the cells are cancerous, proto-oncogenes have been activated and tumor suppressor genes inactivated. Even within the same tumor type, like colon cancer, the specific genes mutated can vary from person to person making cancer a unique disease for

each individual.

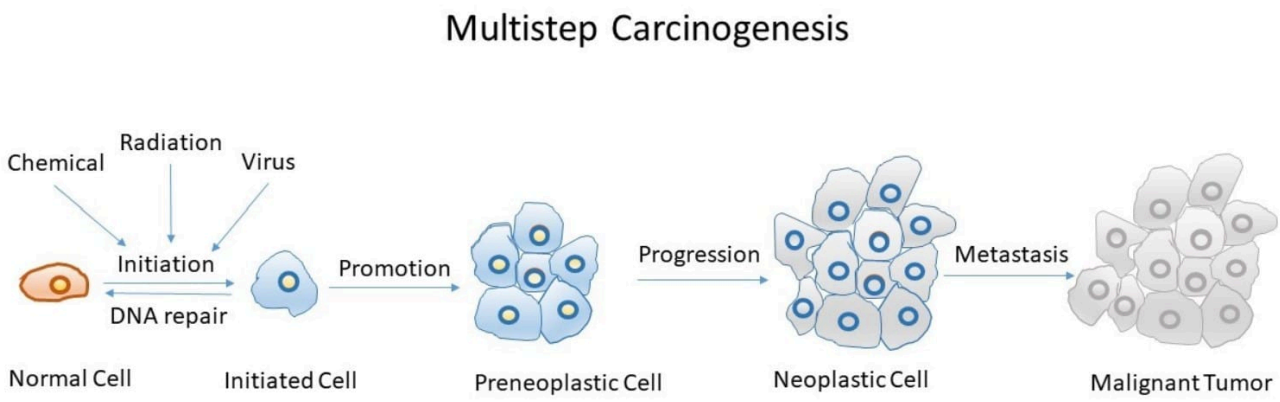


Figure 9.6.2: Multistep process involved in carcinogenesis that transforms a normal cell into a malignant tumor.

How Cancer Spreads

Once a normal cell transforms into a cancer cell and starts dividing out of control, cancer cells can spread from the original site (called the primary tumor) to other tissues. This can occur in three different ways. One way is local spread, in which aggressively dividing cancer cells directly invade nearby tissues. Another way involves the lymphatic system. Cancer cells can spread to regional lymph nodes through lymph vessels that pass by the primary tumor.

The third way cancer cells can spread is through the blood to distant sites. This is called metastasis, and the new cancers that form are called **metastases.** Although the blood can carry cancer cells to tissues everywhere in the body, cancer cells generally grow only in certain sites (Figure 9.6.3). Different types of cancers tend to metastasize to particular organs. The most common places for metastases to occur are the brain, lungs, bones, and liver. Almost all cancers can metastasize, especially during the late stages of the disease. Cancer that has metastasized generally has the worst prognosis and is associated with most cancer deaths.

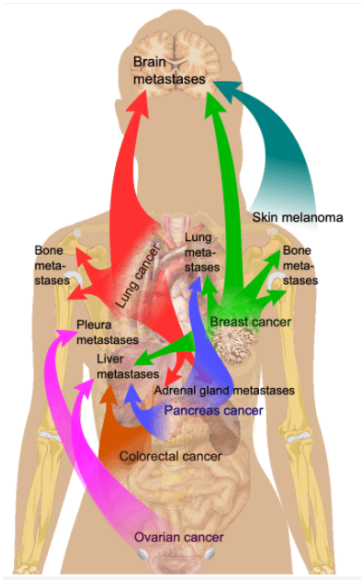


Figure 9.6.3: Cancer cells that originate in one organ may metastasize by traveling in the blood to distant sites (metastases). Metastasized lung cancer grows in the brain, bone, and adrenal gland. Metastasized pancreas cancer grows in the liver and lungs. Metastasized breast cancer can grow in the bone, brain, and liver. Metastasized Colon and Ovarian cancer grow in the liver.

Chemical Carcinogens

Chemical carcinogens (Table 9.6.1) can be either natural or synthetic compounds that, based on animal feeding trials or **epidemiological** (i.e. human population) studies, increase the incidence of cancer. The definition of a chemical as a carcinogen is problematic for several reasons. Some chemicals become carcinogenic only after they are metabolized into another compound in the body; not all species or individuals may metabolize chemicals in the same way. Also, the carcinogenic properties of a compound are usually dependent on its dose. It can be difficult to define a relevant dose for both lab animals and humans. Nevertheless, when a correlation between cancer incidence and chemical exposure is observed, it is usually possible to find ways to reduce exposure to that chemical.

Table 9.6.1: Some Classes of Chemical Carcinogens.

Class	Examples and/or Sources
PAHs (polycyclic aromatic hydrocarbons)	benzo[a]pyrene and several other components of the smoke of cigarettes, wood, and fossil fuels
Aromatic amines	compounds formed in food when meat (including fish, poultry) are cooked at high temperature
Nitrosamines and nitrosamides	found in tobacco and in some smoked meat and fish
Azo dyes	various dyes and pigments used in textiles, leather, paints.
Carbamates	ethyl carbamate (urethane) found in some distilled beverages and fermented foods
Halogenated compounds	e.g. pentachlorophenol used in some wood preservatives and pesticides.
Inorganic compounds	asbestos; may induce chronic inflammation and reactive oxygen species
Miscellaneous compounds	e.g. alkylating agents, phenolics

WEB LINK

A detailed list of occupational carcinogens can be found on the link below

https://bio.libretexts.org/Bookshelves/Human_Biology/Book:_Human_Biology_(Wakim_and_Grewal)/21:_Disease/21.7:_Cancer

The carcinogens implicated as the main causative agents of the four most common cancers worldwide are given in the table below. These four cancers are lung, breast, colon, and stomach cancers. Together they account for about 41% of worldwide cancer incidence and 42% of cancer deaths.

Table 9.6.2: Major Carcinogens Implicated In The Four Most Common Cancers Worldwide.

Type of Cancer	Carcinogen
Lung Cancer	Tobacco smoke
Breast Cancer	Estrogen
Colon Cancer	Tobacco smoke and bile acids: deoxycholic acid (DCA) or lithocholic acid (LCA)
Stomach Cancer	Heliobacter pylori

WEB LINK

Note: A more detailed description for each major carcinogen listed in Table 9.6.2 can be found on the link below

https://en.Wikipedia.org/wiki/Carcin...cers_worldwide

Anticarcinogens

Dietary fiber and calorie restriction are two anti-carcinogen or anti-promoters that decrease the risk of tumor formation. Dietary fiber is both and is inversely associated with cancer, particularly colon cancer. So the more fiber you eat, the less risk you have of developing colon cancer. One mechanism by which fiber acts is hastening bile acid excretion. Fiber also increases the rate of passage of materials through the colon resulting in decreased production and exposure of the colon to cancer-causing agents, ie dilutes the concentration of carcinogens.

Animal studies have shown that restricting caloric intake by 30% reduces tumor growth and increases life span. The mechanism is not known but may be due to less oxidation thus damage to DNA.

Antioxidants can help block the action of initiators or promoters if their mode of action is to damage DNA by oxidation. Vitamin A, C, E, beta-carotene, and selenium are antioxidant nutrients. Some work locally, like vitamin E in the colon, while other work more globally like selenium and vitamin C. Vitamin A appears to work by keeping cells differentiated which slows the growth rate.

Other compounds in food, particularly fruits and vegetables, have been shown to slow tumor formation. Cruciferous vegetables (eg broccoli, cauliflower, cabbage, and Brussel sprouts to name a few) are rich in nutrients, fiber, glucosinolates which are sulfur-containing chemicals, indoles, and isothiocyanates. Animal studies have found these substances inhibit the development of cancer in several organs in rats and mice (Hecht SS. Inhibition of carcinogenesis by isothiocyanates. *Drug Metabolism Reviews* 2000;32(3-4):395-411; Murillo G, Mehta RG. Cruciferous vegetables and cancer prevention. *Nutrition and Cancer* 2001;41(1-2):17-28). Indoles and isothiocyanates help protect cells from DNA damage; help inactivate carcinogens; have antiviral and antibacterial effects; have anti-inflammatory effects; induce cell death (apoptosis); and inhibit tumor blood vessel formation (angiogenesis) and tumor cell migration (needed for metastasis) (National Cancer Institute, Cruciferous Vegetables and Cancer Prevention, 2012, https://www.cancer.gov/about-cancer/...les-fact-sheet). Studies in humans, however, have shown mixed results.

Test for Carcinogens

In genetics, a **mutagen** is a physical or chemical agent that permanently changes genetic material, usually DNA, in an organism and thus increases the frequency of mutations above the natural background level. As many mutations can cause cancer, such mutagens are therefore carcinogens, although not all necessarily are. Many different systems for detecting mutagen have been developed. Animal systems may more accurately reflect the metabolism of human, however, they are expensive and time-consuming (may take around three years to complete), they are therefore not used as a first screen for mutagenicity or carcinogenicity.

The **Ames test** is a widely employed method that uses bacteria to test whether a given chemical can cause mutations in the DNA of the test organism. More formally, it is a biological assay to assess the mutagenic potential of chemical compounds. A positive test indicates that the chemical is mutagenic and therefore may act as a carcinogen, because cancer is often linked to mutation. The test serves as a quick and convenient assay to estimate the carcinogenic potential of a compound because standard carcinogen assays on mice and rats are time-consuming (taking two to three years to complete) and expensive. However, false-positives and false-negatives are known. Other tests, similar to the Ames tests have also been developed using yeast and other bacteria.

Drosophila, plants, and cell cultures have been used in various test assays for mutagenecity of chemicals.

In **animal test systems,** rodents are usually used in animal test. The chemicals under test are usually administered in the food and in the drinking water, but sometimes by dermal application, by gavage, or by inhalation, and carried out over the major part of the life span for rodents. **Transgenic mouse assay** using a mouse strain infected with a viral shuttle vector is another method for testing mutagens. Animals are first treated with suspected mutagen, the mouse DNA is then isolated and the phage segment recovered and used to infect *E. coli*.

Did You Know?

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 9.6.1 *Risk factors for cancer.*

WEB LINKS - American Cancer Society and NIH:MedlinePlus & National Cancer Institute

Risk factors for Cancer

https://www.cancer.gov/about-cancer/causes-prevention/patient-prevention-overview-pdq#_199

Common Cancer Myths and Misconceptions

https://www.cancer.gov/about-cancer/causes-prevention/risk/myths

List of Probable Carcinogens

https://www.cancer.org/cancer/cancer-causes/general-info/known-and-probable-human-carcinogens.html?sitearea=who

Cancer Prevention: Take Charge of Your Lifestyle

https://medlineplus.gov/ency/patientinstructions/000825.htm

Cruciferous Vegetables and Cancer Prevention

https://www.cancer.gov/about-cancer/causes-prevention/risk/diet/cruciferous-vegetables-fact-sheet

Birth Defects:Teratogens

A teratogen is a compound that permanently deforms the function or structure of a developing embryo or fetus in utero. In general, the degree of teratogenicity depends on:

- The potency of the drug as a mutagen.
- The susceptibility of the fetus to teratogenesis.
- The dose of the teratogen.
- The duration of teratogen exposure.
- The time of exposure.
- The degree of transfer from maternal to fetal circulation.

The global average of all live births complicated by malformation is 6% (Environmental Health Perspectives, (NIH), October 2009). The majority of these complications are due to unknown factors. The vast majority of recognized etiologies are genetic, with only 10% being attributed to environmental etiologies such as maternal health, infection, and toxicants. In general, the central nervous and skeletal systems are the most affected.

Thalidomide (a sedative previously marketed in Europe to prevent morning sickness) see Figure 9.6.4is a classic teratogen that caused limb defects in babies born to women who took this drug in the 1960s (see Figure 9.6.5). Thalidomide was first marketed in 1957 in West Germany, where it was available over the counter.[5][6] When first released, thalidomide was promoted for anxiety, trouble sleeping, "tension", and morning sickness.[6][7] While initially deemed to be safe in pregnancy, concerns regarding birth

defects were noted in 1961 and the medication was removed from the market in Europe that year.[6][5] The total number of people affected by use during pregnancy is estimated at 10,000, of which about 40% died around the time of birth.[6][3] Those who survived had limb, eye, urinary tract, and heart problems.[5]

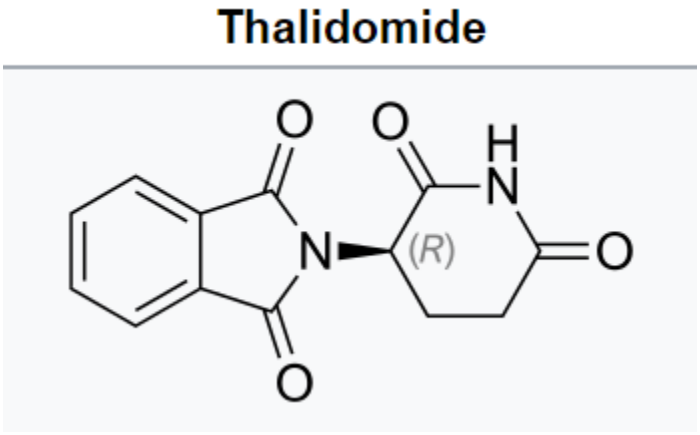


Figure 9.6.4. Chemical structure of Thalidomide.

Its initial entry into the US market was prevented by Frances Kelsey at the FDA.[7] The birth defects of thalidomide led to the development of greater drug regulation and monitoring in many countries. in 2006 the U.S. Food and Drug Administration granted accelerated approval for thalidomide in combination with dexamethasone for the treatment of newly diagnosed multiple myeloma patients.



Figure 9.6.5. *Thalidomide effects: A photograph of the limbs ofbaby born to a mother who took thalidomide while pregnant.*

Women may encounter a number of other teratogens. **Smoking** is most likely to cause growth retardation, but has also been implicated in the prelabor rupture of the membranes, preterm labor, abruption of the placenta, spontaneous abortion, perinatal morbidity and mortality, and sudden infant death syndrome. Smoking may exert its effects through competitive binding of carbon monoxide with hemoglobin and/or through the various other components found in cigarettes that cause adverse biological effects.

Alcohol use in pregnancy may result in fetal alcohol syndrome (FAS) see Figure 9.6.6, which occurs in approximately 1% of all births. Children with FAS present with a flattened and thin upper lip, small palpebral fissures, epicanthal folds, flattened nasal bridge, and a short nose. They may also exhibit microcephaly, mental retardation, and have learning disabilities. It is not clear if there is any safe amount of alcohol consumption in pregnancy.

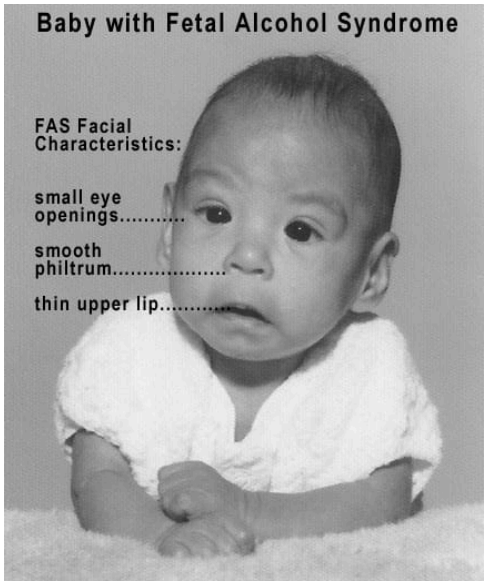


Figure 9.6.6. Fetal alcohol syndrome: Alcohol is a teratogen. When consumed in pregnancy, it can result in mothers giving birth to children with fetal alcohol syndrome. The facial characteristics highlighted in the photograph are small eye opening, smooth philtrum, and a thin upper lip.

Cocaine generally produces growth restriction, preterm delivery, microcephaly, spontaneous abortion, placental abruption, limb anomalies, and central nervous system abnormalities. Cocaine appear to exert a number of its effects through peripheral vasoconstriction that leads to fetal hypoxia.

Women with indications for **warfarin** therapy should either abstain from pregnancy or switch to low molecular weight heparins. Warfarin typically produces mental retardation, growth restriction, nasal hypoplasia, and opthalmic abnormalities.

Angiotensin converting enzyme (ACE) inhibitors will cause fetal renal failure and oligohydramnios that lead to pulmonary hypoplasia and limb contracture. Fetal cranial bone abnormalities are also common.

Isotretinoin (Accutane), used to treat acne, may cause cardiac, oral, otological, thymic, and central nervous system abnormalities. In one quarter of cases, it causes mental retardation.

Other teratogenic substance classes and conditions include

- Various prescription drugs and nutrient deficiencies (e.g., insufficient folic acid).
- Chemical compounds such as methyl iodide (used in pesticides) and bisphenol A (used in plastics) are suspected teratogens.

Summary

- A carcinogen is any agent that directly increases the incidence of cancer.
- Cancer is caused by changes to the DNA which could be inherited or acquired.
- Various tests have been developed to test for carcinogens.
- The American Cancer Society provides an extensive list of probable chemical carcinogens including those found in processed foods and beverages, various household products, pesticides etc.
- The effects of a teratogen on the fetus depend on several factors: the potency of the teratogen, the susceptibility of the fetus to the teratogen, the dose and duration of teratogen exposure, the degree of transfer from maternal to fetal circulation, and when during development the exposure occurs.
- Approximately 10% of congenital malformations are attributed to environmental factors, and 20% are due to genetic or hereditary factors. The rest have unknown causes or are due to a mix of different factors.
- Cigarette components, alcohol, cocaine, warfarin, ACE inhibitors, and Accutane are all teratogens that affect fetal development.

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NIH-MedlinePlus

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9.7: Hazardous Wastes

Learning Objectives

- List the different types of hazardous wastes based on their classification.
- Describe the handling of hazardous waste.

Hazardous waste is waste that has substantial or potential threats to public health or the environment. Hazardous wastes may be found in different physical states such as gaseous, liquids, or solids. A hazardous waste is a special type of waste because it cannot be disposed of by common means like other by-products of our everyday lives. Depending on the physical state of the waste, treatment and solidification processes might be required.

Statutory Definition of Hazardous Waste

A discussion of the definition of hazardous waste should begin with Congress' original statutory definition of the term. RCRA (Resource Recovery and Conservation Act) §1004(5) defines hazardous waste as: A solid waste, or combination of solid waste, which because of its quantity, concentration, or physical, chemical, or infectious characteristics may (a) cause, or significantly contribute to, an increase in mortality or an increase in serious irreversible, or incapacitating reversible, illness; or (b) pose a substantial present or potential hazard to human health or the environment when improperly treated, stored, transported, or disposed of, or otherwise managed.

Listed, Characteristic, and Mixed Wastes

Listed hazardous wastes (see Table 9.7.1) are materials specifically listed as hazardous wastes which are from non-specific sources, specific sources, or discarded chemical products. Title 40 of the Code of Federal Regulations (CFR) in section 261 provides four lists (the F, K, P and U lists) of hazardous materials.

Table 9.7.1 List Hazardous Wastes. Source US EPA.

List Type	Section	Description of Waste
F List	40 CFR section 261.31	Wastes from common manufacturing and industrial processes • Spent solvent wastes, • Electroplating and other metal finishing wastes, • Dioxin-bearing wastes, • Chlorinated aliphatic hydrocarbons production, • Wood preserving wastes, • Petroleum refinery wastewater treatment sludges, and • Multisource leachate.

K List	40 CFR section 261.32	Source specific wastes from 13 various industries and manufacturers namely: • Wood preservation, • Organic chemicals manufacturing, • Pesticides manufacturing, • Petroleum refining, • Veterinary pharmaceuticals manufacturing, • Inorganic pigment manufacturing, • Inorganic chemicals manufacturing, • Explosives manufacturing, • Iron and steel production, • Primary aluminum production, • Secondary lead processing, • Ink formulation, and • Coking (processing of coal to produce coke).
P List	40 CFR section 261.33	Acute hazardous wastes from discarded commercial chemical products. • The chemical in the waste must be unused; and • The chemical in the waste must be in the form of a commercial chemical product.
U List	40 CFR section 261.33	Hazardous wastes from discarded commercial chemical products • The chemical in the waste must be unused; and • The chemical in the waste must be in the form of a commercial chemical product.

- **Characteristic hazardous wastes** are materials that are dangerous because it is toxic, corrosive, flammable, or reactive. Detailed information provided by the US Environmental Agency (US EPA) is summarized in Table 9.7.2.

Table 9.7.2 Characteristic Hazardous Wastes. Source US EPA.

Waste Characteristic	EPA Waste Code	Section	Description of Waste
Ignitability	D001	40 CFR section 261.21	Liquids with flash points below 60 °C (e.g. used solvents and waste oils), non-liquids that cause fire through specific conditions, ignitable compressed gases and oxidizers
Corrosivity	D002	40 CFR section 261.22	Aqueous wastes with a pH of less than or equal to 2 (acids), a pH greater than or equal to 12.5 (bases) or based on the liquids ability to corrode steel. Example: battery acid
Reactivity	D003	40 CFR section 261.23	Wastes that are unstable under normal conditions, may react with water, may give off toxic gases and may be capable of detonation or explosion under normal conditions or when heated. Examples: lithium–sulfur batteries and explosives.
Toxicity	D004 through D043	40 CFR section 261.24	Wastes that are harmful when ingested or absorbed. Toxic waste can be reactive, ignitable, and corrosive. Examples: old batteries, pesticides, paint, and car oil

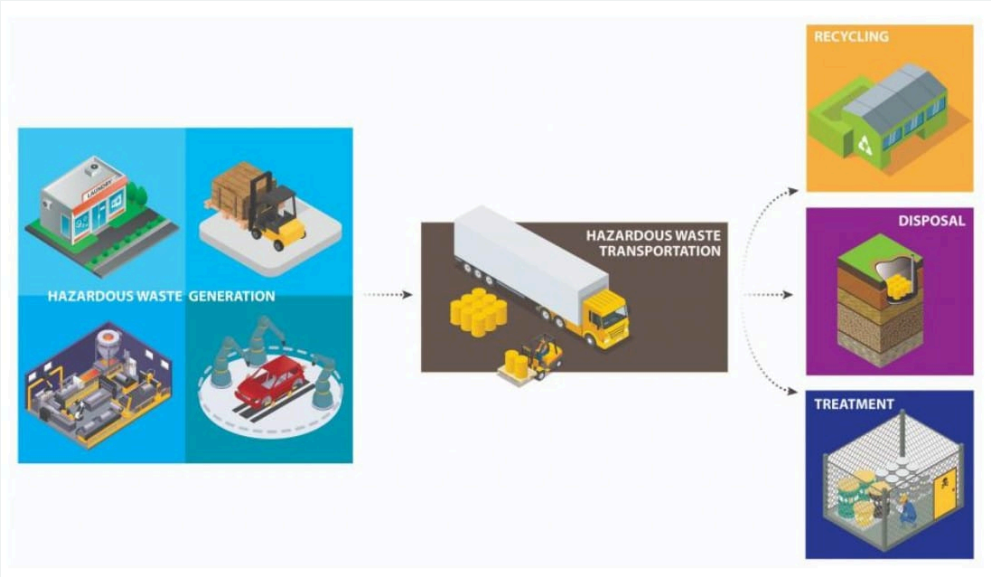
Mixed wastes is the term for hazardous wastes that also contain radioactive material and is regulated under the EPAs Resource Conservation and Recovery Act (RCRA) and the Department of Energy's Atomic Energy Act.

EPA's Cradle-to-Grave Hazardous Waste Management Program

https://www.epa.gov/hw/learn-basics-...ous-waste#hwid

In the mid-twentieth century, solid waste management issues rose to new heights of public concern in many areas of the United States because of increasing solid waste generation, shrinking disposal capacity, rising disposal costs, and public opposition to the siting of new disposal facilities. These solid waste management challenges continue today, as many communities are struggling to develop cost-effective, environmentally protective solutions. The growing amount of waste generated has made it increasingly important for solid waste management officials to develop strategies to manage wastes safely and cost effectively.

RCRA set up a framework for the proper management of hazardous waste . From this authority, EPA established a comprehensive regulatory program to ensure that hazardous waste is managed safely from "cradle to grave" meaning from the time it is created, while it is transported, treated, and stored, and until it is disposed:



Hazardous Waste Generation

Under RCRA, hazardous waste generators are the first link in the hazardous waste management system. All generators must determine if their waste is hazardous and must oversee the ultimate fate of the waste. Furthermore, generators must ensure and fully document that the hazardous waste that they produce is properly identified, managed, and treated prior to recycling or disposal. The degree of regulation that applies to each generator depends on the amount of waste that a generator produces.

Hazardous Waste Transportation

After generators produce a hazardous waste, transporters may move the waste to a facility that can recycle, treat, store or dispose of the waste. Since such transporters are moving regulated wastes on public roads, highways, rails and waterways, United States Department of Transportation hazardous materials regulations, as well as EPA's hazardous waste regulations, apply.

Hazardous Waste Recycling, Treatment, Storage and Disposal

To the extent possible, EPA tried to develop hazardous waste regulations that balance the conservation of resources, while ensuring the protection of human health and environment. Many hazardous wastes can be recycled safely and effectively, while other wastes will be treated and disposed of in landfills or incinerators.

Recycling hazardous waste has a variety of benefits including reducing the consumption of raw materials and the volume of waste materials that must be treated and disposed. However, improper storage of those materials might cause spills, leaks, fires, and contamination of soil and drinking water. To encourage hazardous waste recycling while protecting health and the environment, EPA developed regulations to ensure recycling would be performed in a safe manner.

Treatment Storage and Disposal Facilities (TSDFs) provide temporary storage and final treatment or disposal for hazardous wastes. Since they manage large volumes of waste and conduct activities that may present a higher degree of risk, TSDFs are stringently regulated. The TSDF requirements establish generic facility management standards, specific provisions governing hazardous waste management units and additional precautions designed to protect soil, ground water and air resources.

Comprehensive information on the final steps in EPA's hazardous waste management program is available online, including Web pages and resources related to:

- Hazardous waste recycling,
- Regulations that apply to treatment, storage and disposal facilities, and
- Descriptions of land disposal restrictions.

Summary

- Hazardous waste is defined as solid waste, or combination of solid waste, which because of its quantity, concentration, or physical, chemical, or infectious characteristics may (a) cause, or significantly contribute to, an increase in mortality or an increase in serious irreversible, or incapacitating reversible, illness; or (b) pose a substantial present or potential hazard to human health or the environment when improperly treated, stored, transported, or disposed of, or otherwise managed.
- Listed wastes are wastes from common manufacturing and industrial processes, specific industries and can be generated from discarded commercial products.
- Characteristic wastes are wastes that exhibit any one or more of the following characteristic properties: ignitability, corrosivity, reactivity or toxicity.
- A waste that has a hazardous component and a radioactive component is called a **mixed waste** and is regulated under both the Resource Conservation and Recovery Act (RCRA) and the Atomic Energy Act.
- The Environmental Protection Agency (EPA) has developed a hazardous waste management program to ensure proper handling of hazardous waste from recycling, treatment, storage, and disposal.

Contributors and Attributions

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10: HOUSEHOLD CHEMICALS

Household chemicals are non-food chemicals that are commonly found and used in and around the average household. They are a type of consumer goods, designed particularly to assist cleaning, pest control and general hygiene purposes. Together with non-compostable household waste, the chemicals found in private household commodities pose a serious ecological problem. In addition to having slightly adverse up to seriously toxic effects when swallowed, chemical agents around may contain flammable or corrosive substances.

10.1: Cleaning with Soap

10.2: Synthetic Detergents

10.3: Laundry Auxiliaries- Softeners and Bleaches

10.4: All-Purpose and Special Purpose Cleaning Products

10.5: Solvents, Paints, and Waxes

10.6: Cosmetics - Personal Care Chemicals

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10.1: Cleaning with Soap

Learning Objectives

- Describe the mechanism by which soaps exert their cleansing action.
- List the benefits and problems with using soap.

Personal Cleanliness

Soap is a salt of a fatty acid[1] used in a variety of cleansing and lubricating products. In a domestic setting, soaps are surfactants usually used for washing, bathing, and other types of housekeeping. In industrial settings, soaps are used as thickeners, components of some lubricants, and precursors to catalysts.

When used for cleaning, soap solubilizes particles and grime, which can then be separated from the article being cleaned. In hand washing, as a surfactant, when lathered with a little water, soap kills microorganisms by disorganizing their membrane lipid bilayer and denaturing their proteins. It also emulsifies oils, enabling them to be carried away by running water.[2]

Industrially manufactured **bar soaps** became available in the late 18th century, as advertising campaigns in Europe and America promoted popular awareness of the relationship between cleanliness and health.[35] In modern times, the use of soap has become commonplace in industrialized nations due to a better understanding of the role of hygiene in reducing the population size of pathogenic microorganisms (Figure 10.1.1).



Figure 10.1.1 Modern Soap Shop in Tubeingen , Germany. Source: Wikipedia

Soap Making: Fat Plus Lye Forms a Soap Plus Glycerol

The earliest recorded evidence of the production of soap-like materials dates back to around 2800 BC in ancient Babylon.[10] A formula for soap consisting of water, alkali, and cassia oil was written on a Babylonian clay tablet around 2200 BC.[11]

The Ebers papyrus (Egypt, 1550 BC) indicates the ancient Egyptians bathed regularly and combined animal and vegetable oils with alkaline salts to create a soap-like substance. Egyptian documents mention a similar substance was used in the preparation of wool for weaving.[12]In the reign of Nabonidus (556–539 BC), a recipe for soap consisted of *uhulu* [ashes], cypress [oil] and

sesame [seed oil] "for washing the stones for the servant girls".[13]In ancient Israel, the ashes from barilla plants, such as species of Salsola, saltwort (*Seidlitzia rosmarinus*) and *Anabasis*, were used in soap production, known as potash



Figure 10.1.2 Advertising for medicated toilet soap. Source: Wikipedia

Until the Industrial Revolution, soapmaking was conducted on a small scale and the product was rough. In 1780, James Keir established a chemical works at Tipton, for the manufacture of alkali from the sulfates of potash and soda, to which he afterwards added a soap manufactory. William Gossage produced low-priced, good-quality soap from the 1850s. Robert Spear Hudson began manufacturing a soap powder in 1837, initially by grinding the soap with a mortar and pestle. American manufacturer Benjamin T. Babbitt introduced marketing innovations that included sale of bar soap and distribution of product samples. William Hesketh Lever and his brother, James, bought a small soap works in Warrington in 1886 and founded what is still one of the largest soap businesses, formerly called Lever Brothers and now called Unilever. These soap businesses were among the first to employ large-scale advertising campaigns. A poster advertising is shown in Figure 10.1.2.

Liquid soap was not invented until the nineteenth century; in 1865, William Shepphard patented a liquid version of soap.[40] In 1898, B.J. Johnson developed a soap derived from palm and olive oils; his company, the B.J. Johnson Soap Company, introduced "Palmolive" brand soap that same year.[41] This new brand of soap became popular rapidly, and to such a degree that B.J. Johnson Soap Company changed its name to Palmolive.[42]

Fats and oils can participate in a variety of chemical reactions—for example, because triglycerides are esters, they can be hydrolyzed in the presence of an acid, a base, or specific enzymes known as lipases. The hydrolysis of fats and oils in the presence of a base is used to make soap and is called **saponification**. Today most soaps are prepared through the hydrolysis of triglycerides (often from tallow, coconut oil, or both) using water under high pressure and temperature [700 lb/in^2 (~50 atm or 5,000 kPa) and 200°C]. Sodium carbonate or sodium hydroxide is then used to convert the fatty acids to their sodium salts (soap molecules):

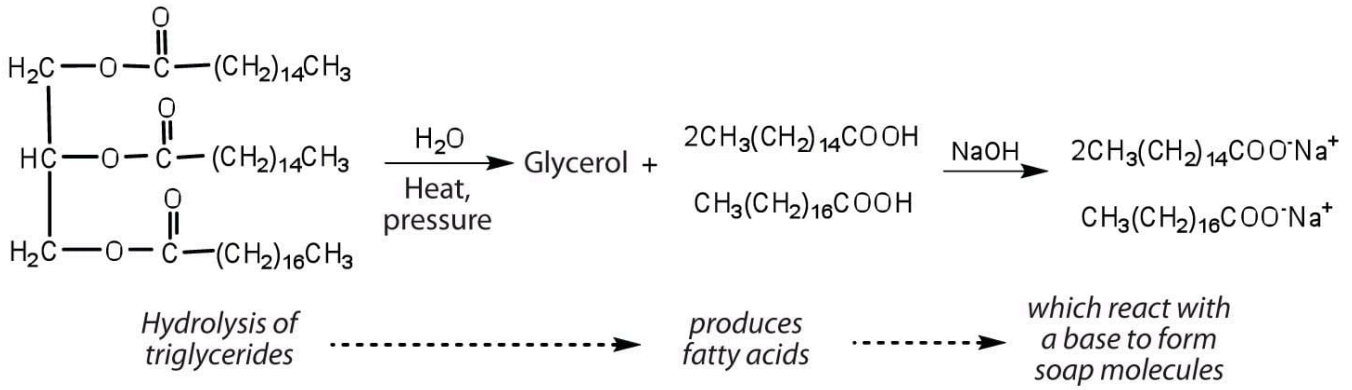


Figure 10.1.3 *Saponification.*

How Soap Works

The cleansing action of soap is determined by its polar and non-polar structures in conjunction with an application of solubility principles. The long hydrocarbon chain is of course non-polar and hydrophobic (repelled by water). The "salt" end of the soap molecule is ionic and hydrophilic (water soluble). Examples of soap and detergent molecules, are shown in Figure 10.1.4. The use of such compounds as cleaning agents is facilitated by their surfactant character, which lowers the surface tension of water, allowing it to penetrate and wet a variety of materials.

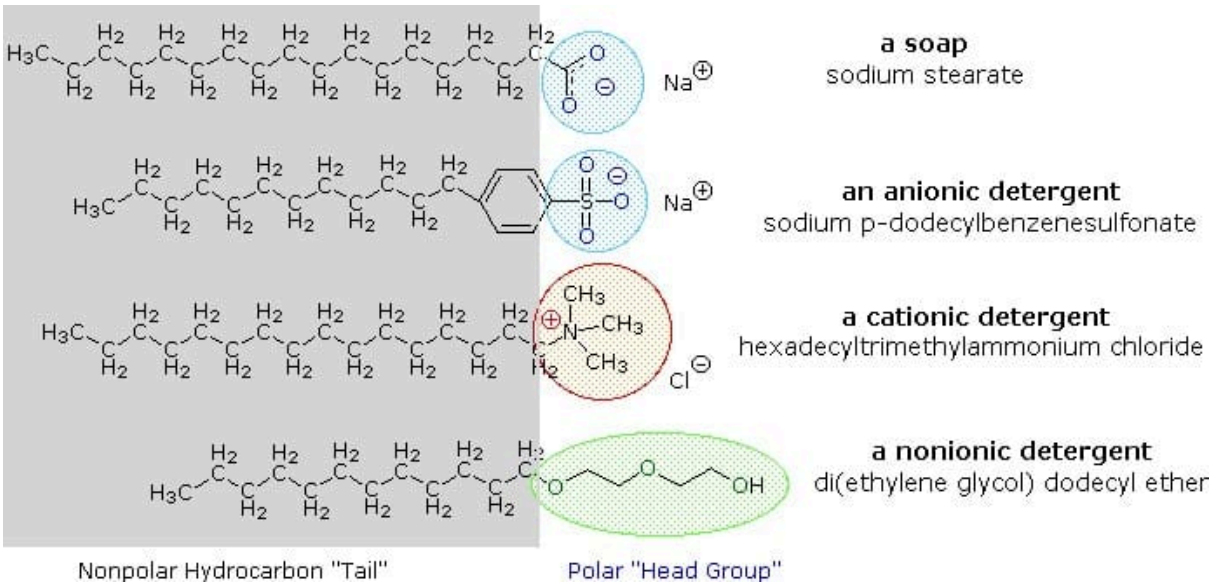


Figure 10.1.4 Soap and detergent molecules with different polar structures.

When soap is added to water, the ionic-salt end of the molecule is attracted to water and dissolved in it. The non-polar hydrocarbon end of the soap molecule is repelled by water. A drop or two of soap in water forms a monolayer (Figure 10.1.5) on the water surface as shown in the graphics on the left. The soap molecules "stand up" on the surface as the polar carboxyl salt end is attracted to the polar water. The non-polar hydrocarbon tails are repelled by the water, which makes them appear to stand up.

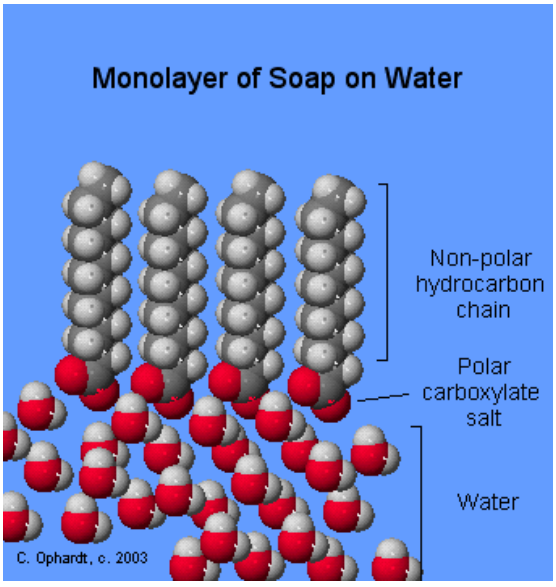


Figure 10.1.5 Monolayer of soap on water.

Soap vs. oil vs. water

Water alone is not able to penetrate grease or oil because they are of opposite polarity. When grease or oil (non-polar hydrocarbons) are mixed with a soap- water solution, the soap molecules work as a "bridge" between polar water molecules and non-polar oil molecules. Soap molecules have both properties of non-polar and polar at opposite ends of the molecule.

The oil is a pure hydrocarbon so it is non-polar. The non-polar hydrocarbon tail of the soap dissolves into the oil. That leaves the polar carboxylate ion of the soap molecules are sticking out of the oil droplets, the surface of each oil droplet is negatively charged. As a result, the oil droplets repel each other and remain suspended in solution (this is called an emulsion) to be washed away by a stream of water. The outside of the droplet is also coated with a layer of water molecules.

The graphic on Figure 10.1.6 although not strictly a representation of the above description is a micelle that works in much the same fashion. The oil would be a the center of the micelle.

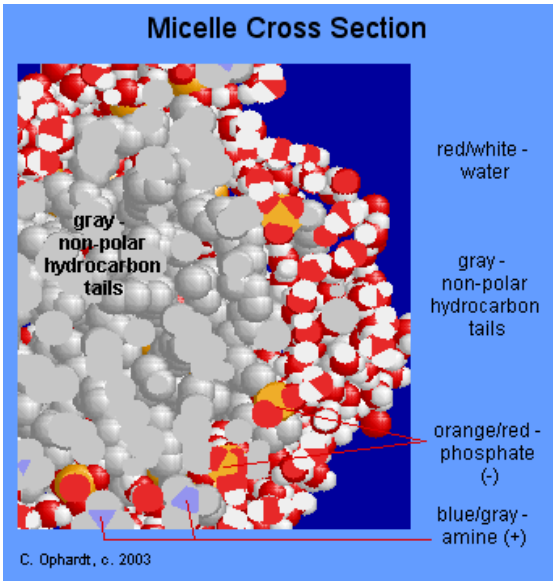


Figure 10.1.6 A micelle with oil at the center.

Disadvantages and Advantages of Soap

The importance of soap to human civilization is documented by history, but some problems associated with its use have been recognized. One of these is caused by the weak acidity (pKa ca. 4.9) of the fatty acids. Solutions of alkali metal soaps are slightly alkaline (pH 8 to 9) due to hydrolysis. If the pH of a soap solution is lowered by acidic contaminants, insoluble fatty acids precipitate and form a scum. A second problem is caused by the presence of calcium and magnesium salts in the water supply (hard water). These divalent cations cause aggregation of the micelles, which then deposit as a dirty scum.

In the reaction below, the sodium cation in soap is replaced by calcium to form calcium stearate. The white precipitate, also termed as soap scum could form deposits on surfaces and inside plumbing.

$$2\ C_{17}H_{35}COO^{-}Na^{+} + Ca^{2+} \rightarrow (C_{17}H_{35}COO)_2Ca + 2\ Na^{+}$$

Soap is still widely popular product as it is a low cost, readily available product used for personal hygiene and cleanliness. Use of soap doesn't lead to pollution. Soap is biodegradable as it can be broken down by microorganisms found in sewage.

Can Soap Really Kill the Coronavirus ?

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 10.1.1 *Soap and coronavirus.*

Water Softeners

Water softening is the removal of calcium, magnesium, and certain other metal cations in hard water. The resulting soft water requires less soap for the same cleaning effort, as soap is not wasted bonding with calcium ions. Soft water also extends the lifetime of plumbing by reducing or eliminating scale build-up in pipes and fittings. Water softening is usually achieved using lime softening or ion-exchange resins (see Figure 10.1.7) but is increasingly being accomplished using nanofiltration or reverse osmosis membranes.

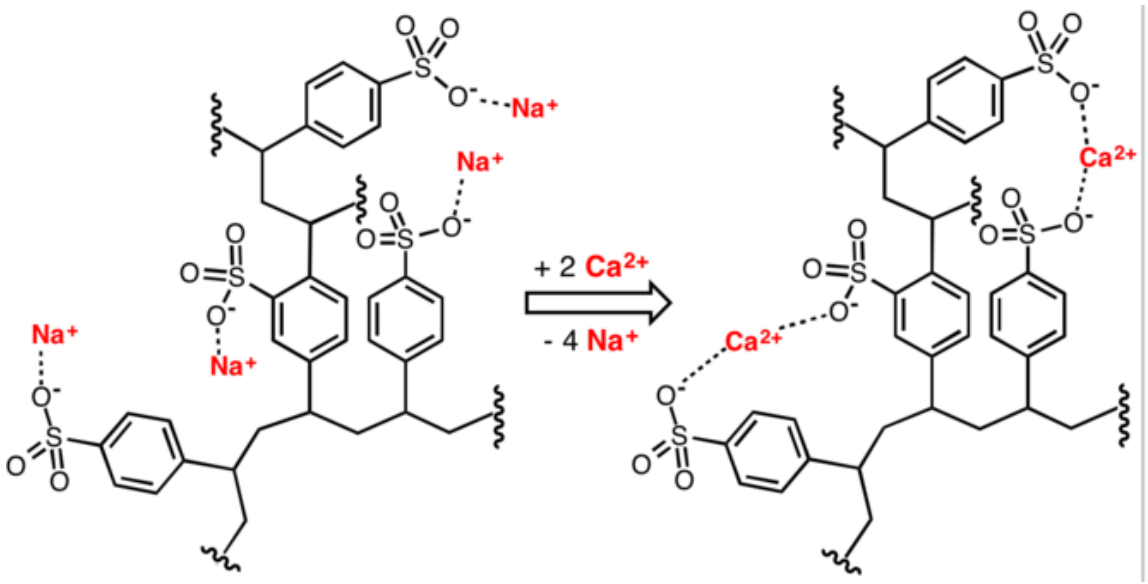


Figure 10.1.7 *Idealized image of water softening process involving replacement of calcium ions in water with sodium ions donated by a cation exchange resin*

How to soften hard water

Some wish to soften hard water to control its irritating, and in many cases damaging, effects. The diminished ability of soap to lather is not only annoying, but can also be potentially harmful economically. Businesses that depend on the foaming of soap, such as car washes and pet groomers, may wish to soften hard water to avoid excessive use of soap due to a decreased ability to lather. Likewise, it is often necessary to soften water that comes into contact with pipes to avoid the destructive and compromising build-up of deposits. Also, many people may find the calcifying effects that hard water has on faucets and other items unfavorable and choose to soften the water to prevent such mineral deposits from forming. Still others may dislike the sticky, dry feeling left by the precipitation of soap scum onto the skin. Whatever the reasons, there are many processes available to soften hard water.

Ion Exchange

One way to soften water is through a process called ion exchange. During ion exchange, the unwanted ions are "exchanged"

for more acceptable ions. In many cases, it is desirable to replace the hard water ions, such as Ca^{2+} and Mg^{2+}, with more agreeable ions, like that of Na^{+}. To do this, the hard water is conducted through a zeolite or resin-containing column, which binds the unwanted ions to its surface and releases the more tolerable ions. In this process, the hard water ions become "fixed" ions because of their attachment to the resin material. These fixed ions displace the desirable ions (Na^{+}), now referred to as counterions, from the column, thus exchanging the ions in the water. This process is illustrated in Figure 10.1.8.

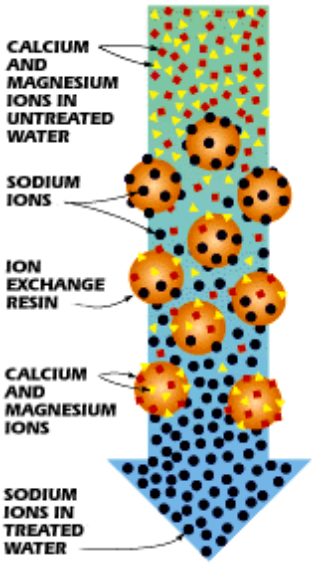


Figure 10.1.8 : Illustration of Ion Exchange.

Unfortunately, this process has the disadvantage of increasing the sodium content of drinking water, which could be potentially hazardous to the health of people with sodium-restricted diets.

Lime Softening

Another process is called lime softening. In this process, the compound calcium hydroxide, $Ca(OH)_2$, is added to the hard water. The calcium hydroxide, or "slaked lime," raises the pH of the water and causes the calcium and magnesium to precipitate into $CaCO_3$ and $Mg(OH)_2$. These precipitates can then be easily filtered out due to their insolubility in water, shown below by the small solubility constant of magnesium hydroxide (the solubility product constant for calcium carbonate is shown above). After precipitation and removal of the offending ions, acid is added to bring the pH of the water back to normal.

$$Mg(OH)_{2\ (s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2OH^{-}_{(aq)} \tag{4a}$$

Chelation

Chelating agents can also be used to soften hard water. Polydentate ligands, such as the popular hexadentate ligand EDTA, bind the undesirable ions in hard water. These ligands are especially helpful in binding the magnesium and calcium cations, which as already mentioned are highly prevalent in hard water solutions. The chelating agent forms a very stable ring complex with the metal cations, which prevents them from interacting with any other substances that may be introduced to the solution, such as soap. In this way, chelators are able to diminish the negative effects associated with hard water. A simplified equation representing the chelation of the metal calcium cation (Ca^{2+}) with the hexadentate ligand EDTA is shown below. The large value of the formation constant (K_f) reflects the tendency of the reaction to proceed to completion in the forward direction.

$$Ca^{2+} + EDTA^{4-} \longrightarrow [Ca(EDTA)]^{2-} \tag{5a}$$

Reverse Osmosis

The final process, reverse osmosis, uses high pressures to force the water through a semipermeable membrane. This membrane

is generally intended to be impermeable to anything other than water. The membrane serves to filter out the larger ions and molecules responsible for the water's hardness, resulting in softened water. During this process, the water is forced from an area with a high concentration of solute in the form of dissolved metal ions and similar compounds, to an area that is very low in the concentration of these substances. In other words, the water moves from a state of hardness to a softer composition as the ions causing the water's hardness are prevented passage through the membrane. Reverse Osmosis does have a disadvantage of wasting wastewater compared to other water treatment methods. This process is shown in Figure 10.1.9 below. Note that this figure describes the desalination of salt water. However, the process for softening hard water is the same.

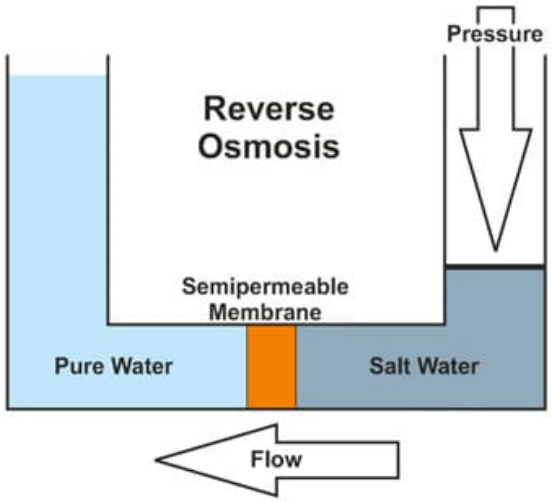


Figure 10.1.9: Depiction of Reverse Osmosis.

Making Hard Water Soft

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 10.1.2 *How to soften hard water.*

Summary

- Soap, as a cleaning agent, solubilizes particles and grime, which can then be easily removed from the article being cleaned.
- In hand washing, soap is an excellent surfactant, that destroys microorganisms by damaging their cellular membranes and denaturing their proteins.
- Soap is an inexpensive, readily available cleaning agent that is also's effectivenes is diminished biodegradable. However, it's effectiveness is diminished when used with hard water.

Contributors

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10.2: Synthetic Detergents

Learning Objectives

- Describe different types of detergents.
- List the key components of detergents and their functions.

A **detergent** is a surfactant or a mixture of surfactants with cleansing properties in dilute solutions (File:Diskflaskor.JPG(Figure 10.2.1)) . These substances are usually alkylbenzene sulfonates, a family of compounds that are similar to soap but are more soluble in hard water, because the polar sulfonate (of detergents) is less likely than the polar carboxylate (of soap) to bind to calcium and other ions found in hard water.

In domestic contexts, the term *detergent* by itself refers specifically to *laundry detergent* or *dish detergent*, as opposed to *hand soap* or other types of cleaning agents. Detergents are commonly available as powders or concentrated solutions. Detergents, like soaps, work because they are amphiphilic: partly hydrophilic (polar) and partly hydrophobic (non-polar). Their dual nature facilitates the mixture of hydrophobic compounds (like oil and grease) with water. Because air is not hydrophilic, detergents are also foaming agents to varying degrees.



Figure 10.2.1 Detergents. Source: Wikipedia

Detergents are classified into three broad groupings, depending on the electrical charge of the surfactants.

Anionic detergents

Typical anionic detergents are alkylbenzenesulfonates (ABS). The alkylbenzene portion of these anions is lipophilic and the sulfonate is hydrophilic. An estimated 6 billion kilograms of anionic detergents are produced annually for domestic markets. Two different varieties have been popularized, those with branched alkyl groups and those with linear alkyl groups (see Figure 10.2.2).

Branched Alkylbenzenesulfonates: Nonbiodegradable

Branched alkylbenzene sulfonates (BAS) were first introduced in the early 1930s and saw significant growth from the late 1940s onwards,[3] in early literature these synthetic detergents are often abbreviated as syndets.

Compared to traditional soaps BAS offered superior tolerance to hard water and better foaming.[5] However, the highly branched

tail made it difficult to biodegrade.[6] BAS was widely blamed for the formation of large expanses of stable foam in areas of wastewater discharge such as lakes, rivers and coastal areas (sea foams), as well as foaming problems encountered in sewage treatment[7] and contamination of drinking water.[8] As such BAS was phased out of most detergent products during the 1960s, being replaced with linear alkylbenzene sulfonates (LAS). It is still important in certain agrochemical and industrial applications, where rapid biodegradability is of reduced importance.

Linear Alkylbenzensulfonates: Biodegradable

Linear alkylbenzene sulfonates (LAS) are prepared industrially by the sulfonation of linear alkylbenzenes (LABs), which can themselves be prepared in several ways.[2] The term "linear" refers to the starting alkenes rather than the final product. The compound biodegrades far more quickly than BAS, making it the safer choice over time. It is biodegraded rapidly under aerobic conditions with a half-life of approximately 1–3 weeks. Under anaerobic conditions it degrades very slowly or not at all, causing it to exist in high concentrations in sewage sludge, but this is not thought to be a cause for concern as it will rapidly degrade once returned to an oxygenated environment.

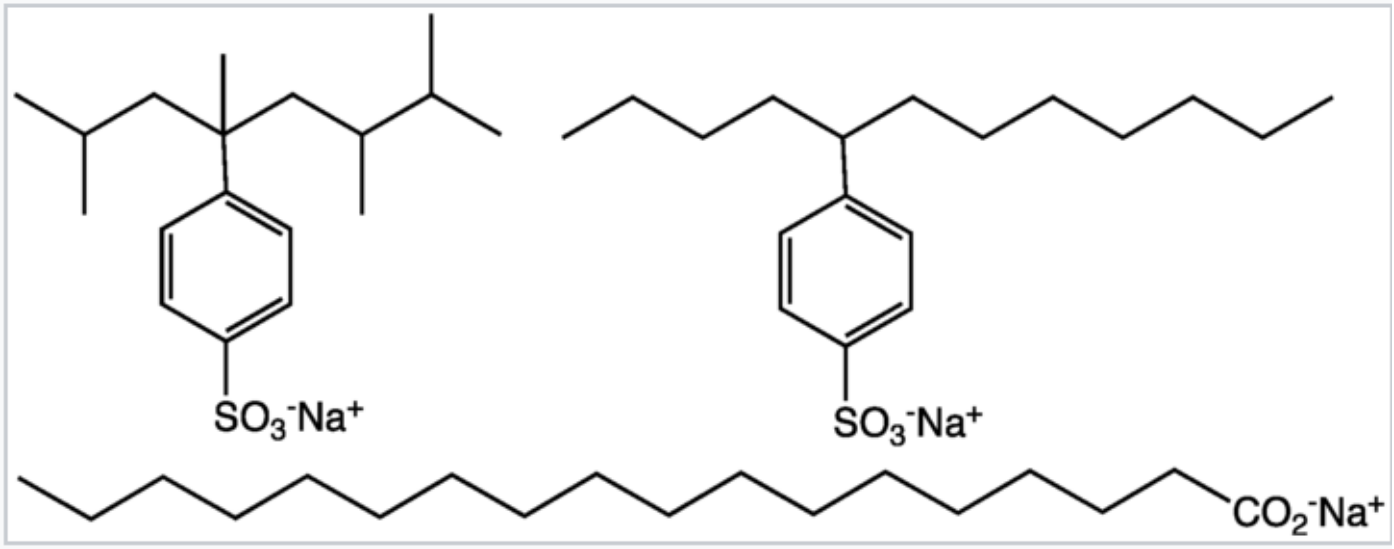


Figure 10.2.2 Three kinds of anionic detergetns: a branched sodium dodecylbenzenesulfonate, linear sodium dodecylbenzenesulfonate, and a soap. *Source: Wikipedia*

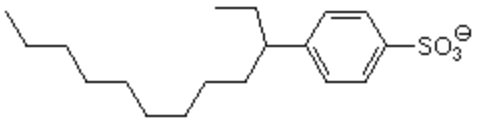
Cationic detergents

Cationic detergents are similar to the anionic ones, with a hydrophilic component, but, instead of the anionic sulfonate group, the cationic surfactants have quaternary ammonium as the polar end. The ammonium sulfate center is positively charged.[3]

Non-ionic and zwitter ionic detergents

Non-ionic detergents are characterized by their uncharged, hydrophilic headgroups. Typical non-ionic detergents are based on polyoxyethylene or a glycoside. Common examples of the former include Tween, Triton, and the Brij series. These materials are also known as ethoxylates or PEGylates and their metabolites, nonylphenol. Glycosides have a sugar as their uncharged hydrophilic headgroup. Examples include octyl thioglucoside and maltosides. HEGA and MEGA series detergents are similar, possessing a sugar alcohol as headgroup.

Zwitterionic detergents possess a net zero charge arising from the presence of equal numbers of +1 and –1 charged chemical groups. Examples include CHAPS. CHAPS is an abbreviation for 3-[(3-**ch**olamidopropyl)dimethyl**a**mmonio]-1-**p**ropane**s**ulfonate.



synthetic detergent

Laundry Detergent

Laundry detergent, or **washing powder**, is a type of detergent (cleaning agent) used for cleaning laundry. Laundry detergent is manufactured in powder and liquid form.

While powdered and liquid detergents hold roughly equal share of the worldwide laundry detergent market in terms of value, powdered detergents are sold twice as much compared to liquids in terms of volume.[1]



Figure 10.2.3 The two forms of laundry detergent: powder and liquid. *Source Wikipedia*

Components

Laundry detergents may contain builders (50% by weight, approximately), surfactants (15%), bleach (7%), enzymes (2%), soil antideposition agents, foam regulators, corrosion inhibitors, optical brighteners, dye transfer inhibitors, fragrances, dyes, fillers and formulation aids.[4]

Builders

Builders (also called chelating or sequestering agents) are water softeners. Hard water contains calcium, magnesium, and metallic cations (primarily, iron, copper, and manganese). These cations react with surfactant anions to form insoluble compounds (metallic or lime soaps) which precipitate onto fabrics and washing machines and which are difficult to remove. Builders remove the hard water ions through precipitation, chelation, or ion exchange. In addition, they help remove soil by dispersion. In most European regions, the water is hard. In North America, Brazil, and Japan, the water is comparatively soft.

The earliest builders were sodium carbonate (washing soda) and sodium silicate (waterglass). Since the 1930s, phosphates (sodium phosphates) and polyphosphates (sodium hexametaphosphate) were introduced, continuing with the introduction of phosphonates (HEDP, ATMP, EDTMP). These agents are now known to have serious environmental consequences leading to a drive towards more environmentally benign phosphorus-free agents, such as polycarboxylates (EDTA, NTA), citrates (trisodium citrate), silicates (sodium silicate), gluconic acid and polyacrylic acid; or ion exchange agents like zeolites.

Surfactants

Surfactants are responsible for most of the cleaning performance in laundry detergent. They provide this by absorption and emulsification of soil into the water and also by reducing the water's surface tension to improve wetting.

Laundry detergents contain mostly anionic and non-ionic surfactants. Cationic surfactants are normally incompatible with anionic detergents and have poor cleaning efficiency; they are employed only for certain special effects, as fabric softeners, antistatic agents, and biocides. Zwitterionic surfactants are rarely employed in laundry detergents mainly for cost reasons. Most detergents use a combination of various surfactants to balance their performance.

Bleaches

Despite the name, modern laundry bleaches do not include household bleach (sodium hypochlorite). Laundry bleaches are typically stable adducts of hydrogen peroxide, such as sodium perborate and sodium percarbonate, these are inactive as solids but will release hydrogen peroxide upon exposure to water. The main targets of bleaches are oxidisible organic stains, which are usually of vegetable origin (e.g. chlorophyll, anthocyanin dyes, tannins, humic acids, and carotenoid pigments). Hydrogen peroxide is insufficiently active as a bleach at temperature below 60°C, which traditionally made hot washes the norm. The development of bleach activators in the 1970s and 80s allowed for cooler washing temperatures to be effective. These compounds, such as tetraacetylethylenediamine (TAED), react with hydrogen peroxide to produce peracetic acid, which is an even more effective bleach, particularly at lower temperatures.

Enzymes

The use of enzymes for laundry was introduced in 1913 by Otto Rohm. The first preparation was a pancreatic extract obtained from slaughtered animals, which was unstable against alkali and bleach. Only in the latter part of the century with the availability of thermally robust bacterial enzymes did this technology become mainstream.

Enzymes are required to degrade stubborn stains composed of proteins (milk, cocoa, blood, egg yolk, grass), fats (chocolate, fats, oils), starch (flour and potato stains), and cellulose (damaged cotton fibrils, vegetable and fruit stains). Each type of stain requires a different type of enzyme: proteases (savinase) for proteins, lipases for greases, α-amylases for carbohydrates, and cellulases for cellulose.

Other ingredients

Many other ingredients are added depending on the expected circumstances of use. Such additives modify the foaming properties of the product by either stabilizing or counteracting foam. Other ingredients increase or decrease the viscosity of the solution, or solubilize other ingredients. Corrosion inhibitors counteract damage to washing equipment. "Dye transfer inhibitors" prevent dyes from one article from colouring other items. "Antiredeposition agents" such as carboxymethyl cellulose are used to prevent fine soil particles from reattaching to the product being cleaned.

A number of ingredients affect aesthetic properties of the item to be cleaned or the detergent itself before or during use. These agents include optical brighteners, fabric softeners, and colourants. A variety of perfumes are also components of modern detergents, provided that they are compatible with the other components and do not affect the colour of the cleaned item. The perfumes are typically a mixture of many compounds, common classes include terpene alcohols (citronellol, geraniol, linalool, nerol) and their esters (linalyl acetate), aromatic aldehydes (helional, hexyl cinnamaldehyde, lilial) and synthetic musks (galaxolide).

Dishwashing Detergents

Dishwashing liquid (or **washing-up liquid** in British English), also known as **dishwashing soap**, **dish detergent**, and **dish soap** is a detergent used to assist in dishwashing. It is usually a highly-foaming mixture of surfactants with low skin irritation, and is primarily used for hand washing of glasses, plates, cutlery, and cooking utensils in a sink or bowl. In addition to its primary use, dishwashing liquid also has various informal applications, such as for creating bubbles, clothes washing and cleaning oil-affected birds.

Hand dishwashing is generally performed in the absence of a dishwashing machine, when large "hard-to-clean" items are present,

or through preference. Some dishwashing liquids can harm household silver, fine glassware, anything with gold leaf, disposable plastics, and any objects made of brass, bronze, cast iron, pewter, tin, or wood, especially when combined with hot water and the action of a dishwasher. When dishwashing liquid is used on such objects it is intended that they be washed by hand.

Hand dishwashing detergents utilize surfactants to play the primary role in cleaning. The reduced surface tension of dishwashing water, and increasing solubility of modern surfactant mixtures, allows the water to run off the dishes in a dish rack very quickly. However, most people also rinse the dishes with pure water to make sure to get rid of any soap residue that could affect the taste of the food.

Dishwashing liquid can be a skin irritant and cause hand eczema. Those with "sensitive skin" are advised amongst other things to persuade someone else to do the washing up.

Dishwasher detergent is a detergent made for washing dishes in a dishwasher. Dishwasher detergent is different from dishwashing liquid made to wash dishes by hand.

When using a dishwasher, the user must select a special detergent for its use. All detergents are designed for use after the user scrapes leftover food from the dishes before washing. To function, the user places dishes in the dishwasher in such fashion that the surface of all dishes is open to the flow of water.

Most dishwasher detergents are incompatible for use with silver, brass, cast iron, bronze, aluminum, pewter, and goldleaf. They can also harm disposable plastic, anything wood, knives with hollow handles, and fine glassware.

Dishwashing detergents for dishwashers are manufactured and marketed variously as cartridges, gel, liquids, pacs, powder, and tablets. Any dishwashing liquid may contain bleach, enzymes, or rinsing aids. Some dishwashing detergents may be homemade, using ingredients such as borax, essential oil, eucalyptus oil and grated bar soap, among others.

Dishwashing detergents can be formulated to work under different circumstances. In some cases suitably formulated they can be used with cold water or sea water, although they will not generally work as well as those intended for, and used with, hot water.

Different kinds of dishwashing detergent contain different combinations of ingredients. Common ingredients include:

- **Phosphates**: Bind calcium and magnesium ions to prevent 'hard-water' type limescale deposits. They can cause ecological damage, and have been partially banned or phased out.
- **Oxygen-based bleaching agents** (older-style powders and liquids contain chlorine-based bleaching agents): Break up and bleach organic deposits.
- **Non-ionic surfactants**: Lower the surface tension of the water, emulsifies oil, lipid and fat food deposits, prevents droplet spotting on drying.
- **Alkaline salts**: These are a primary component in older and original-style dishwasher detergent powders[*citation needed*]. Highly alkaline salts attack and dissolve grease, but are extremely corrosive (fatal) if swallowed. Salts used may include metasilicates, alkali metal hydroxides, sodium carbonate etc.
- **Enzymes**: Break up protein-based food deposits, and possibly oil, lipid and fat deposits. The enzymes used are similar to the ones used in laundry.
- **Anti-corrosion agent(s)**: Often sodium silicate, this prevents corrosion of dishwasher components.[*citation needed*]

Dishwashing detergent may also contain:[*citation needed*]

- Anti-foaming agents:[*citation needed*] Foam interferes with the washing action. Foam may affect operation of the machine's water-level sensors and will leak past the door seals.
- Additives to slow down the removal of glaze & patterns from glazed ceramics
- Perfumes
- Anti-caking agents (in granular detergent)
- Starches (in tablet based detergents)
- Gelling agents (in liquid/gel based detergents)
- Sand (inexpensive powdered detergents)

Dishwasher detergents are generally strongly alkaline (basic).

Inexpensive powders may contain sand. Such detergents may harm the dishes and the dishwasher. Powdered detergents are

more likely to cause fading on china patterns.

Besides older style detergents for dishwashers, biodegradable detergents also exist for dishwashers. These detergents may be more environmentally friendly than conventional detergents.

Informal Uses

Reader's Digest notes its use as an ant killer, weed killer, to help spread water-borne fertilizer, and to wash human hair. *Good Housekeeping* says it can be used mixed with vinegar to attract and drown fruit flies. Dishwashing detergent has been used to clean mirrors as well as windows.

- Pling, an open source general purpose cleaner for glazed, plastic, chrome and inox bathroom and kitchen surfaces, published by Twibright Labs, uses dishwashing liquid as one of active ingredients.
- Dishwashing liquid can be mixed with water and additional ingredients such as glycerin and sugar to produce a bubble-blowing solution.
- Dishwashing liquid may be used for cleaning delicate clothing fabrics such as hosiery and lingerie.
- Dishwashing liquid is frequently recommended in a dilute solution to make decals and vinyl graphics easier to position when applying.
- In industry, dishwashing liquid is also used to inspect pressurized equipment for leaks, such as propane fittings. It is used to inspect pneumatic tires for flats, as well as for quality assurance during the installation process, and as a mounting bead lubricant.
- Dishwashing liquid has uses as an ingredient in making homemade garden pest deterrents. Oregon State University's Cooperative Extension Service notes the use of dishwashing liquid to get rid of spidermites. Dish soap has also been used to deter aphids. In some instances, the dish soap may be toxic to plant leaves and cause them to "burn". Use of soap or dish detergent to help spread pesticide on plants is noted by University of Georgia extension service, but not recommended.
- A solution of dishwashing liquid and water may be used to remove coffee, tea, olive oil, soda and fruit juice stains from fabrics. One dishwashing liquid brand has been used to remove stains from white or lightly-colored cloth napkins.
- Dishwashing liquid has been used to treat birds affected by oil spills. After the Exxon Valdez oil spill in 1989, the International Bird Rescue Research Center received hundreds of cases of dishwashing liquid that were used for this purpose. More dishwashing liquid was donated during the Deepwater Horizon oil spill to the International Bird Rescue Research Center and the Marine Mammal Center.

Figure 10.2.4 *An oiled Gannet being washed. Source Wikipedia*

Environmental concerns

Phosphates in detergent became an environmental concern in the 1950s and the subject of bans in later years.[8] Phosphates make laundry cleaner but also cause eutrophication, particularly with poor wastewater treatment.[9]

A recent academic study of fragranced laundry products found "more than 25 VOCs emitted from dryer vents, with the highest concentrations of acetaldehyde, acetone, and ethanol. Seven of these VOCs are classified as hazardous air pollutants (HAPs) and two as carcinogenic HAPs (acetaldehyde and benzene)".[10]

The EEC Directive 73/404/EEC stipulates an average biodegradability of at least 90% for all types of surfactants used in detergents. The phosphate content of detergents is regulated in many countries, e.g., Austria, Germany, Italy, The Netherlands, Norway, Sweden, Switzerland, United States, Canada, and Japan.

Summary

- Detergents are classified into anionic, cationic, and non-ionic detergents and zwitterionic detergents based on the electrical charge of the surfactants.
- Variations in detergent formulations are based on its end-use (i.e. laundry or for kitchenware).
- Phosphates make laundry cleaner but also cause eutrophication, particularly with poor wastewater treatment.[9]

ntributors and Attributions

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10.3: Laundry Auxiliaries- Softeners and Bleaches

Learning Objectives

- Describe a fabric softener and its use.
- Describe bleaches and their use.

A **fabric softener** (or **conditioner**) is a conditioner that is typically applied to laundry during the rinse cycle in a washing machine. In contrast to laundry detergents, fabric softeners may be regarded as a kind of after-treatment laundry aid.[1]

Machine washing puts great mechanical stress on textiles, particularly natural fibers such as cotton and wool. The fibers at the fabric surface are squashed and frayed, and this condition hardens while drying the laundry in air, giving the laundry a harsh feel. Adding a liquid fabric softener to the final rinse (rinse-cycle softener) results in laundry that feels softer.[2][1]

Fabric softeners coat the surface of a fabric with chemical compounds that are electrically charged, causing threads to "stand up" from the surface and thereby imparting a softer and fluffier texture. Cationic softeners, like those shown in Figure 10.3.1. bind by electrostatic attraction to the negatively charged groups on the surface of the fibers and neutralize their charge. The long aliphatic chains then line up towards the outside of the fiber, imparting lubricity.

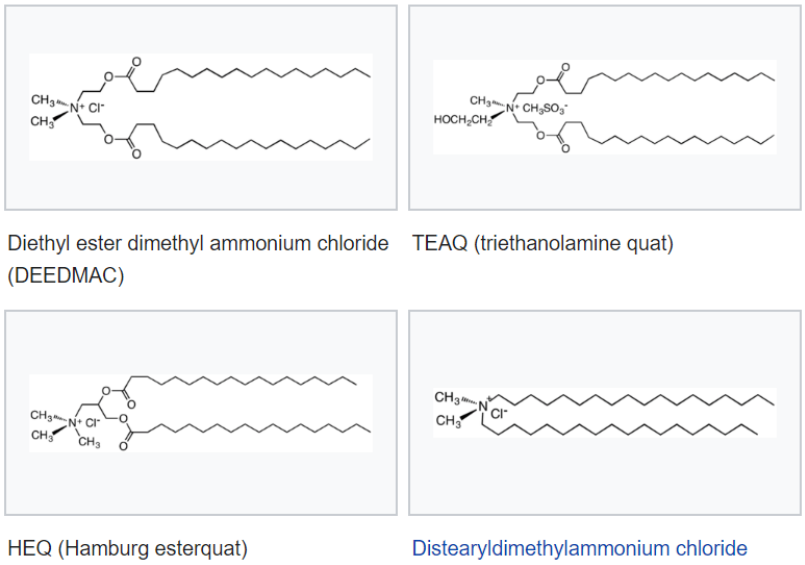


Figure 10.3.1 *Cationic surfactants in fabric softeners. Source: Wikipedia*

Fabric softeners impart antistatic properties to fabrics, and thus prevent the build-up of electrostatic charges on synthetic fibers, which in turn eliminates fabric cling during handling and wearing, crackling noises, and dust attraction. Also, fabric softeners make fabrics easier to iron and help reduce wrinkles in garments. In addition, they reduce drying times so that energy is saved when softened laundry is tumble-dried. Last but not least, they can also impart a pleasant fragrance to the laundry.[1]

Risks

As with soaps and detergents, fabric softeners may cause irritant dermatitis.[5] Manufacturers produce some fabric softeners without dyes and perfumes to reduce the risk of skin irritation. Fabric softener overuse may make clothes more flammable, due to the fat-based nature of most softeners. Some deaths have been attributed to this phenomenon,[6] and fabric softener makers recommend not using them on clothes labeled as flame-resistant.[*citation needed*]

Laundry Bleaches: Whiter Whites

Bleach is the generic name for any chemical product which is used industrially and domestically to clean, and to remove stains. It often refers, specifically, to a dilute solution of sodium hypochlorite, also called "liquid bleach".

Many bleaches have broad spectrum bactericidal properties, making them useful for disinfecting and sterilizing and are used in swimming pool sanitation to control bacteria, viruses, and algae and in many places where sterile conditions are required. They are also used in many industrial processes, notably in the bleaching of wood pulp. Bleaches also have other minor uses like removing mildew, killing weeds, and increasing the longevity of cut flowers.[1]

Bleaches work by reacting with many colored organic compounds, such as natural pigments, and turning them into colorless ones. While most bleaches are oxidizing agents (chemicals that can remove electrons from other molecules), some are reducing agents (that donate electrons).

Chlorine, a powerful oxidizer, is the active agent in many household bleaches. Since pure chlorine is a toxic corrosive gas, these products usually contain hypochlorite which releases chlorine when needed. "Bleaching powder" usually means a formulation containing calcium hypochlorite.

Oxidizing bleaching agents that do not contain chlorine are usually based on peroxides such as hydrogen peroxide, sodium percarbonate, and sodium perborate. These bleaches are called 'non-chlorine bleach,' 'oxygen bleach' or 'color-safe bleach.'[2]

Reducing bleaches have niche uses, such as sulfur dioxide used to bleach wool, either as gas or from solutions of sodium dithionite;[3] and sodium borohydride.

Bleaches generally react with many other organic substances besides the intended colored pigments, so they can weaken or damage natural materials like fibers, cloth, and leather, and intentionally applied dyes such as the indigo of denim. For the same reason, ingestion of the products, breathing of the fumes, or contact with skin or eyes can cause health damage.

Summary

- A fabric softener is an after-treatment laundry aid that imparts a softer and fluffier texture.
- Bleach is the generic name for any chemical product that is used industrially and domestically for cleaning and stain removal.

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10.4: All-Purpose and Special Purpose Cleaning Products

Learning Objective

- List various cleaners and their modes of action.

Cleaning agents or **hard-surface cleaners** are substances (usually liquids, powders, sprays, or granules) used to remove dirt, including dust, stains, bad smells, and clutter on surfaces. Purposes of cleaning agents include health, beauty, removing offensive odor, and avoiding the spread of dirt and contaminants to oneself and others. Some cleaning agents can kill bacteria (e.g. door handle bacteria, as well as bacteria on worktops and other metallic surfaces) and clean at the same time. Others, called degreasers, contain organic solvents to help dissolve oils and fats.

Acidic cleaning agents are mainly used for removal of inorganic deposits like scaling. The active ingredients are normally strong mineral acids and chelants. Often, surfactants and corrosion inhibitors are added to the acid.

Hydrochloric acid is a common mineral acid typically used for concrete. Vinegar can also be used to clean hard surfaces and remove calcium deposits that also helps to maintain our environment bacteria free. Sulphuric acid is used in acidic drain cleaners to unblock clogged pipes by dissolving greases, proteins, and even carbohydrate-containing substances such as toilet tissue.

Alkaline cleaning agents contain strong bases like sodium hydroxide or potassium hydroxide. Bleach (pH 12) and ammonia (pH 11) are common alkaline cleaning agents. Often, dispersants, to prevent redeposition of dissolved dirt, and chelants, to attack rust, are added to the alkaline agent.

Alkaline cleaners can dissolve fats (including grease), oils, and protein-based substances.

Neutral washing agents are pH-neutral and based on non-ionic surfactants that disperse different types.

Scouring agents are mixtures of the usual cleaning chemicals (surfactants, water softeners) as well as abrasive powders. The abrasive powder must be of a uniform particle size.

Particles are usually smaller than 0.05 mm. Pumice, calcium carbonate (limestone, chalk, dolomite), kaolinite, quartz, soapstone or talc are often used as abrasives, i.e. polishing agents.

Special bleaching powders contain compounds that release sodium hypochlorite, the classical household bleaching agent. These precursor agents include trichloroisocyanuric acid and mixtures of sodium hypochlorite ("chlorinated orthophosphate").

Examples of notable products include: Ajax, Bar Keepers Friend, Bon Ami, Comet, Vim, Zud, and others.

All-Purpose Cleaners

All-purpose cleansers (Figure 10.4.1) contain mixtures of anionic and nonionic surfactants, polymeric phosphates or other sequestering agents, solvents, hydrotropic substances, polymeric compounds, corrosion inhibitors, skin-protective agents, and sometimes perfumes and colorants.[7] Aversive agents, such as denatonium, are occasionally added to cleaning products to discourage animals and small children from consuming them.



Figure 10.4.1 *Fabuloso multipurpose cleaner and generic surface cleaners. Source: Wikipedia*

Some cleaners contain water-soluble organic solvents like glycol ethers and fatty alcohols, which ease the removal of oil, fat and paint. Disinfectant additives include quaternary ammonium compounds, phenol derivatives, terpene alcohols (pine oil), aldehydes, and aldehyde-amine condensation products.

All-purpose cleaners are usually concentrated solutions of surfactants and water softeners, which enhance the behavior of surfactant in hard water. Typical surfactants are alkylbenzenesulfonates, an anionic detergent, and modified fatty alcohols. A typical water softener is sodium triphosphate.

All-purpose cleansers are effective with most common kinds of dirt. Their dilute solutions are neutral or weakly alkaline, and are safe for use on most surfaces.

Special Purpose Cleaners

Traditional **oven cleaners** contain sodium hydroxide (lye), solvents, and other ingredients,[3][4][5] and work best when used in a slightly-warm (not hot) oven. If used in a self-cleaning oven, the lye will cause permanent damage to the oven.

New-style oven cleaners are based on ingredients other than lye. These products must be used in a cold oven. Most new-style oven cleaners can be used in self-cleaning ovens.

Oven cleaners are some of the most toxic household cleaning products available on the market.[6] Correct use of an oven cleaner may be reasonably safe, but incorrect use can have harmful effects.

One popular oven cleaner brand in the US is "Easy-Off", sold by Reckitt Benckiser. Popular choices in the UK include "Zep Oven Brite" and "Mr Muscle Oven Cleaner".

Toilet bowl cleaning agents often are aimed at removal of calcium carbonate deposits, which are attacked by acids. Powdered cleaners contain acids that come in the form of solid salts, such as sodium hydrogen sulfate. Liquid toilet bowl cleaners contain other acids, typically dilute hydrochloric, phosphoric, or formic acids. These convert the calcium carbonate into salts that are soluble in water or are easily rinsed away.

Chemical drain cleaners can be in solid or liquid form that are readily available through hardware stores, though some (primarily acidic ones) are intended for use by licensed plumbers.[1]Alkaline drain cleaners are available in either solid or liquid state while the acidic ones are usually in liquid form.



Figure 10.4.2 *Bottles of alkaline drain cleaners containing sodium hydroxide can dissolve greases and hair. Source: Wikipedia*

Metal cleaners are used for cleaning stainless steel sinks, faucets, metal trim, silverware, etc. These products contain abrasives (e.g., siliceous chalk, diatomaceous earth, alumina) with a particle size < 20 μm. Fatty alcohol or alkylphenol polyglycol ethers with 7-12 ethylene oxide (EO) units are used as surfactants.[7]

For ferrous metals, the cleaners contain chelating agents, abrasives, and surfactants. These agents include citric and phosphoric acids, which are nonaggressive. Surfactants are usually modified fatty alcohols. Silver cleaning is a specialty since silver is noble but tends to tarnish via formation of black silver sulfide, which is removable via silver-specific complexants such as thiourea.

Stainless steel, nickel, and chromium cleaners contain lactic, citric, or phosphoric acid. A solvent (mineral spirits) may be added.

Nonferrous metal cleaners contain ammonia, ammonium soaps (ammonium oleate, stearate) and chelating agents (ammonium citrate, oxalate).

For special type of precious metals especially those used for luxury watches and high-end jewelry, special type of cleaning agents are usually used to clean and protect them from the elements.

Glass cleaners. Light duty hard surface cleaners are not intended to handle heavy dirt and grease. Because these products are expected to clean without rinsing and result in a streak-free shine, they contain no salts. Typical window cleaning items consist of alcohols, either ethanol or isopropanol, and surfactants for dissolving grease. Other components include small amounts of ammonia as well as dyes and perfumes.

These are composed of organic,water-miscible solvent such as isopropyl alcohol and an alkaline detergent. Some glass cleaners also contain a fine, mild abrasive. Most glass cleaners are available as sprays or liquid. They are sprayed directly onto windows, mirrors and other glass surfaces or applied on with a soft cloth and rubbed off using a soft ,lint-free duster. A glass cloth ideal for the purpose and soft water to which some methylated spirit or vinegar is added which is an inexpensive glass cleaner.

Silverware can be freed of silver sulfide tarnish with thiourea, and either hydrochloric or sulfuric acid.

Building facade cleaners. For acid-resistant building facades, such as brick, acids are typically used. These include mixtures of phosphoric and hydrofluoric acids as well as surfactants. For acid-sensitive facades such as concrete, strongly alkaline cleaners are used such as sodium hydroxide and thickeners. Both types of cleaners require a rinsing and often special care since the solutions are aggressive toward skin.

Drain Cleaners: Danger and Usage Considerations

Danger arises from chemical drain cleaners' potential to injure eyes, lungs, and skin; and damage to clothing and household materials such as wood, paint, aluminum, and fiberglass. Chemical drain cleaners should be used only according to the manufacturer's instructions, as other use may cause injury.[7] Strongly corrosive and acid drain cleaners are among the most hazardous household products available to the public. Chemical drain cleaners can cause strong reactions—sometimes

explosively—with other chemicals that may have been used previously, which can result in serious injury to anyone in the vicinity.[8] In one such incident, a five-year-old boy was left scarred for life after an acidic drain cleaner leaked through his bedroom ceiling as he slept.[9]

Strong Alkali Drain cleaners are equally capable of causing rapid, severe burns, as seen in the cases of a woman doused with concentrated lye in an attack. A small girl was also permanently disfigured by a common lye drain opener.,[10][11] Moreover, because the acidic or basic drain cleaners themselves are washed down the drain, this contributes to pollution in the water supply. The heat generation can also soften plastic PVC pipes, and the pressure buildup by gas generation can cause older pipes to burst. Commercial chemical based solutions can cause corrosion and other damage to your pipes and sewer lines[12]

Oftentimes, individuals may unknowingly mix two different types of drain cleaners, which can even lead to deadly results. For example, consider the mixing of an acidic and basic drain cleaner:

Sulfuric Acid + Sodium Hydroxide → sodium sulfate (a salt) + water

$H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + 2H_2O$

The neutralization reaction of the acid and base may seem harmless, but in reality this reaction is extremely exothermic and can cause pipes to violently explode. Consider another example of mixing, this time between an acid drain cleaner and bleach:

Hydrochloric acid + bleach → water + table salt + chlorine gas

$2HCl + NaClO \rightarrow H_2O + NaCl + Cl_2$

This reaction generates chlorine gas, which is toxic to the lungs.

Green Cleaners

Green cleaning refers to using cleaning methods and products with environmentally friendly ingredients and procedures which are designed to preserve human health and environmental quality.[1] Green cleaning techniques and products avoid the use of products which contain toxic chemicals, some of which emit volatile organic compounds causing respiratory, dermatological and other conditions.[2] Green cleaning can also describe the way residential and industrial cleaning products are manufactured, packaged and distributed. If the manufacturing process is environmentally friendly and the products are biodegradable, then the term "*green*" or "*eco-friendly*" may apply.

Common household ingredients that are safe to use on its own or in combination for a variety of cleaning applications include baking soda, soap, alcohol, cornstarch, lemon juice, white vinegar, citrus solvent, washing soda (SAL soda or sodium carbonate), oxygen bleach, vegetable oil, and hydrogen peroxide.

WEB LINK

The links below provide information on safe ingredients for homemade substitutions and homemade cleaning products.

https://learn.eartheasy.com/guides/non-toxic-home-cleaning/#homecleaning

Summary

- All-purpose cleansers are effective with most common kinds of dirt and are formulated to contain mixtures of surfactants, sequestering agents, solvents, hydrotropic substances, polymeric compounds, corrosion inhibitors, skin-protective agents, and sometimes perfumes and colorants.
- Special purpose cleaners are formulated with specific chemicals to effectively remove dirt, grime or stain from a particular surface (e.g. glass cleaner, metal cleaner, etc.) or appliance (e.g. oven cleaner).

- Green cleaners are gaining popularity among environmentalists who are interested in using cleaners that are made of biodegradable and eco-friendly ingredients.

Contributors

- Wikipedia
- Cowan, Shannon. “Non-Toxic Home Cleaning.” *Eartheasy Guides & Article*s, learn.eartheasy.com/guides/non-toxic-home-cleaning/.

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10.5: Solvents, Paints, and Waxes

Learning Objective

- List the properties and uses of solvents, paints, and, waxes.

Solvents

A **solvent** is a substance that dissolves a solute, resulting in a solution. A solvent is usually a liquid but can also be a solid, a gas, or a supercritical fluid. The quantity of solute that can dissolve in a specific volume of solvent varies with temperature. Common uses for organic solvents are in dry cleaning (e.g. tetrachloroethylene), as paint thinners (e.g. toluene, turpentine), as nail polish removers and glue solvents (acetone, methyl acetate, ethyl acetate), in spot removers (e.g. hexane, petrol ether), in detergents (citrus terpenes) and in perfumes (ethanol). Water is a solvent for polar molecules and the most common solvent used by living things; all the ions and proteins in a cell are dissolved in water within a cell. Solvents find various applications in chemical, pharmaceutical, oil, and gas industries, including in chemical syntheses and purification processes.

Most organic solvents are flammable or highly flammable, depending on their volatility. Exceptions are some chlorinated solvents like dichloromethane and chloroform. Mixtures of solvent vapors and air can explode. Solvent vapors are heavier than air; they will sink to the bottom and can travel large distances nearly undiluted. Solvent vapors can also be found in supposedly empty drums and cans, posing a flash fire hazard; hence empty containers of volatile solvents should be stored open and upside down.

Paints

Paint is any pigmented liquid, liquefiable, or solid mastic composition that, after application to a substrate in a thin layer, converts to a solid film. It is most commonly used to protect, color, or provide texture to objects.

Paint was one of the earliest inventions of humanity. Some cave paintings drawn with red or yellow ochre, hematite, manganese oxide, and charcoal may have been made by early *Homo sapiens* as long as 40,000 years ago.[3] Paint may be even older. In 2003 and 2004, South African archeologists reported finds in Blombos Cave of a 100,000-year-old human-made ochre-based mixture that could have been used like paint.[4][5] Further excavation in the same cave resulted in the 2011 report of a complete toolkit for grinding pigments and making a primitive paint-like substance.[5][6]

By the proper onset of the Industrial Revolution, in the mid-18th century, paint was being ground in steam-powered mills, and an alternative to lead-based pigments had been found in a white derivative of zinc oxide. Interior house painting increasingly became the norm as the 19th century progressed, both for decorative reasons and because the paint was effective in preventing the walls rotting from damp. Linseed oil was also increasingly used as an inexpensive binder.

In 1866, Sherwin-Williams Figure $10.5.1$ in the United States opened as a large paint-maker and invented a paint that could be used from the tin without preparation.

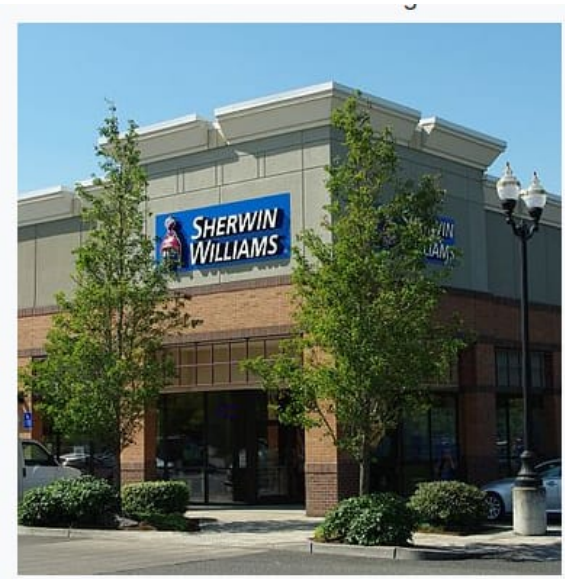


Figure 10.5.1 Sherwin Williams company store in Hillsboro, Oregon

The **binder** is the film-forming component of paint.[10] It is the only component that is always present among all the various types of formulations. Many binders are too thick to be applied and must be thinned. The type of thinner, if present, varies with the binder. The binder imparts properties such as gloss, durability, flexibility, and toughness.[11] Binders include synthetic or natural resins such as alkyds, acrylics, vinyl-acrylics, vinyl acetate/ethylene (VAE), polyurethanes, polyesters, melamine resins, epoxy, silanes or siloxanes or oils.

The main purposes of the **diluent** are to dissolve the polymer and adjust the viscosity of the paint. It is volatile and does not become part of the paint film. It also controls flow and application properties, and in some cases can affect the stability of the paint while in liquid state. Its main function is as the carrier for the non volatile components. To spread heavier oils (for example, linseed) as in oil-based interior house paint, a thinner oil is required. These volatile substances impart their properties temporarily—once the solvent has evaporated, the remaining paint is fixed to the surface. Some paints have no diluent. Water is the main diluent for water-borne paints, even the co-solvent types.

Solvent-borne, also called oil-based, paints can have various combinations of organic solvents as the diluent, including aliphatics, aromatics, alcohols, ketones and white spirit. Specific examples are organic solvents such as petroleum distillate, esters, glycol ethers, and the like. Sometimes volatile low-molecular weight synthetic resins also serve as diluents.

Pigments are granular solids incorporated in the paint to contribute color. **Dyes** are colorants that dissolve in the paint. **Fillers** are granular solids incorporated to impart toughness, texture, give the paint special properties,[16] or to reduce the cost of the paint. During production, the size of such particles can be measured with a Hegman gauge. Rather than using only solid particles, some paints contain dyes instead of or in combination with pigments.

Pigments can be classified as either natural or synthetic. Natural pigments include various clays, calcium carbonate, mica, silicas, and talcs. Synthetics would include engineered molecules, calcined clays, blanc fixe, precipitated calcium carbonate, and synthetic pyrogenic silicas.

Some pigments are toxic, such as the lead pigments that are used in lead paint. Paint manufacturers began replacing white lead pigments with titanium white (titanium dioxide), before lead was banned in paint for residential use in 1978 by the US Consumer Product Safety Commission. The titanium dioxide used in most paints today is often coated with silica/alumina/zirconium for various reasons, such as better exterior durability, or better hiding performance (opacity) promoted by more optimal spacing within the paint film.[17]

Waxes

Waxes are organic compounds, hydrocarbons that characteristically consist of long aliphatic alkyl chains, although aromatic compounds may also be present. Natural waxes may contain unsaturated bonds and include various functional groups such as

fatty acids, primary and secondary alcohols, ketones, aldehydes and fatty acid esters. **Synthetic waxes** (e.g. polyethylene and polypropylene waxes) often consist of homologous series of long-chain aliphatic hydrocarbons (alkanes or paraffins) that lack functional groups.[1]

The best known animal wax is **beeswax** used in constructing the honeycombs of honeybees, but other insects secrete waxes. A major component of the beeswax is myricyl palmitate which is an ester of triacontanol and palmitic acid. Its melting point is 62-65 °C. **Spermaceti** occurs in large amounts in the head oil of the sperm whale. One of its main constituents is cetyl palmitate, another ester of a fatty acid and a fatty alcohol. **Lanolin** is a wax obtained from wool, consisting of esters of sterols.[1]

Plants secrete waxes into and on the surface of their cuticles as a way to control evaporation, wettability and hydration.[3] The epicuticular waxes of plants are mixtures of substituted long-chain aliphatic hydrocarbons, containing alkanes, alkyl esters, fatty acids, primary and secondary alcohols, diols, ketones and aldehydes.[2] From the commercial perspective, the most important plant wax is **carnauba wax**, a hard wax obtained from the Brazilian palm *Copernicia prunifera*. Containing the ester myricyl cerotate, it has many applications, such as confectionery and other food coatings, car and furniture polish, floss coating, and surfboard wax. Other more specialized vegetable waxes include **jojoba oil**, **candelilla wax** and **ouricury wax**.

Paraffin wax (or **petroleum wax**) is a soft colorless solid derived from petroleum, coal or shale oil that consists of a mixture of hydrocarbon molecules containing between twenty and forty carbon atoms.Waxes are mainly consumed industrially as components of complex formulations, often for coatings.The main use of polyethylene and polypropylene waxes is in the formulation of colourants for plastics. Waxes confer matting effects and wear resistance to paints. Polyethylene waxes are incorporated into inks in the form of dispersions to decrease friction. They are employed as release agents, find use as slip agents in furniture, and confer corrosion resistance.



Figure 10.5.2 Paraffin candle.

Waxes such as paraffin wax or beeswax, and hard fats such as tallow are used to make candles (Figure 10.5.2). used for lighting and decoration. Waxes are used as finishes and coatings for wood products.[10] Beeswax is frequently used as a lubricant on drawer slides where wood to wood contact occurs. Waxes are used to make wax paper, impregnating and coating paper and card to waterproof it or make it resistant to staining, or to modify its surface properties. Waxes are also used in shoe polishes, wood polishes, and automotive polishes, as mold release agents in mold making, as a coating for many cheeses, and to waterproof leather and fabric. Wax has been used since antiquity as a temporary, removable model in lost-wax casting of gold, silver and other materials.

Wax with colorful pigments added has been used as a medium in encaustic painting, and is used today in the manufacture of crayons, china markers and colored pencils. Carbon paper, used for making duplicate typewritten documents was coated with carbon black suspended in wax, typically montan wax, but has largely been superseded by photocopiers and computer printers. In another context, lipstick and mascara are blends of various fats and waxes colored with pigments, and both beeswax and lanolin are used in other cosmetics. Ski wax is used in skiing and snowboarding. Also, the sports of surfing and skateboarding often use wax to enhance the performance.

Some waxes are considered food-safe and are used to coat wooden cutting boards and other items that come into contact with food. Beeswax or coloured synthetic wax is used to decorate Easter eggs in Romania, Ukraine, Poland, Lithuania and the Czech Republic. Paraffin wax is used in making chocolate covered sweets.

Summary

- A solvent is a liquid, solid, a gas, or a supercritical fluid that dissolves a solute to form a solution. Solvents are found in various personal and household items, and also used in chemical, pharmaceutical, oil, and gas industries, including in chemical syntheses and purification processes.
- Paint is any pigmented liquid, liquefiable, or solid mastic composition that, after application to a substrate in a thin layer, converts to a solid film. It is most commonly used to protect, color, or provide texture to objects.
- Natural and synthetic waxes are used for lighting, for surface coating and polishing, for waterproofing leather and fabric. Waxes are also key ingredients in cosmetics and personal care products.

Contributors and Attributions

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10.6: Cosmetics - Personal Care Chemicals

Learning Objectives

- List the main ingredients in different personal care products and their functions.
- Describe the steps in formation of temporary and permanent waves.

In the United States, the Food and Drug Administration (FDA), which regulates cosmetics, defines cosmetics as products "intended to be applied to the human body for cleansing, beautifying, promoting attractiveness, or altering the appearance without affecting the body's structure or functions". This broad definition includes any material intended for use as an ingredient of a cosmetic product, with the FDA specifically excluding pure soap from this category.



Figure 10.6.1 Assorted cosmetics and tools.

Cosmetics (Figure 10.6.1) have been in use for thousands of years. Egyptian men and women used makeup to enhance their appearance. They were very fond of eyeliner and eye-shadows in dark colors including blue, red, and black. Ancient Sumerian men and women were possibly the first to invent and wear lipstick, about 5,000 years ago. They crushed gemstones and used them to decorate their faces, mainly on the lips and around the eyes.

According to one source, early major developments include:

- Kohl used by ancient Egypt as a protectant of the eye (Figure 10.6.2).
- Castor oil used by ancient Egypt as a protective balm.
- Skin creams made of beeswax, olive oil, and rose water, described by Romans.
- Vaseline and lanolin in the nineteenth century.

Although modern makeup has been traditionally used mainly by women, an increasing number of men are using cosmetics usually associated with women to enhance or cover their own facial features such as blemishes and dark circles, as well the use of eyeshadow, mascara and lipstick by some. Cosmetics brands have increasingly also targeted men in the sale of cosmetics, with some products targeted specifically at men.



Figure 10.6.2 Nefertiti Bust showing the use of eye liner made of kohl.

The Skin

The cutaneous membrane is the technical term for our skin. The skin's primary role is to help protect the rest of the body's tissues and organs from physical damage such as abrasions, chemical damage such as detergents, and biological damage from microorganisms. For example, while the skin harbors many permanent and transient bacteria, these bacteria are unable to enter the body when healthy, intact skin is present.

Our skin is made of three general layers (Figure 10.6.3). In order from most superficial to deepest they are the epidermis, dermis, and subcutaneous tissue (hypodermis).

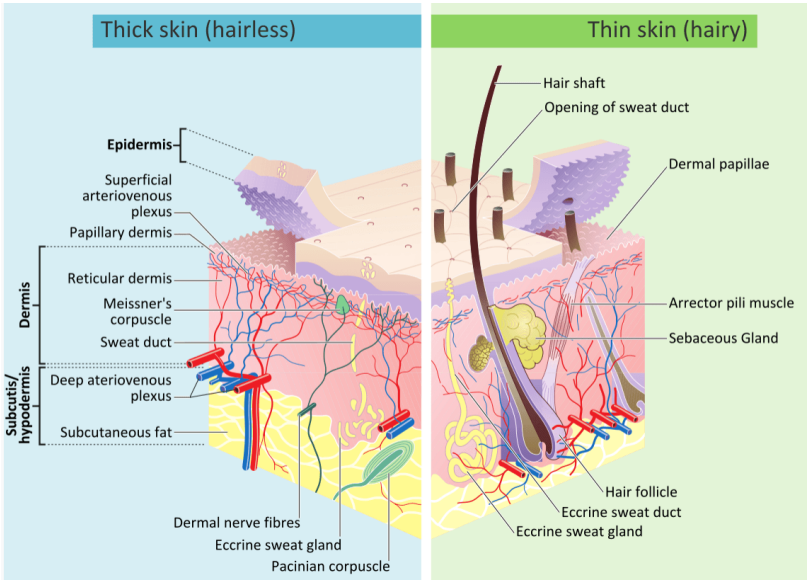


Figure 10.6.3 Skin layers, of both the hairy and hairless skin.

There are several different types of cells in the **epidermis**. All of the cells are necessary for the important functions of the epidermis.

- The epidermis consists mainly of stacks of keratin-producing epithelial cells called keratinocytes. These cells make up at least 90 percent of the epidermis. Near the top of the epidermis, these cells are also called squamous cells.
- Another 8 percent of epidermal cells are melanocytes. These cells produce the pigment melanin that protects the dermis

from UV light.
- About 1 percent of epidermal cells are Langerhans cells. These are immune system cells that detect and fight pathogens entering the skin.
- Less than 1 percent of epidermal cells are Merkel cells, which respond to light touch and connect to nerve endings in the dermis.

The epidermis has several crucial functions in the body. These functions include protection, water retention, and vitamin D synthesis. The epidermis provides protection to underlying tissues from physical damage, pathogens, and UV light.

Skin Creams and Lotions

A **lotion** is a low-viscosity topical preparation intended for application to the skin. By contrast, creams and gels have higher viscosity, typically due to lower water content. Lotions are applied to external skin with bare hands, a brush, a clean cloth, or cotton wool.

While a lotion may be used as a medicine delivery system, many lotions, especially **hand lotions** and **body lotions** are meant instead to simply smooth, moisturize, soften and perhaps perfume the skin.

Most cosmetic lotions are moisturizing lotions, although other forms, such as tanning lotion, also exist.

Cosmetic lotions may be marketed as anti-aging lotions, which can also be classified as a cosmetic in many cases, and may contain fragrances. The Food and Drug Administration voiced concern about lotions not classified as drugs that advertise anti-aging or anti-wrinkle properties.

Most lotions are oil-in-water emulsions using a substance such as cetearyl alcohol to keep the emulsion together, but water-in-oil lotions are also formulated.

A **cream** is a preparation usually for application to the skin. Creams for application to mucous membranes such as those of the rectum or vagina are also used. Creams may be considered pharmaceutical products as even cosmetic creams are based on techniques developed by pharmacy and unmedicated creams are highly used in a variety of skin conditions (dermatoses). The use of the finger tip unit concept may be helpful in guiding how much topical cream is required to cover different areas.

Creams are semi-solid emulsions of oil and water. They are divided into two types: oil-in-water (O/W) creams which are composed of small droplets of oil dispersed in a continuous water phase, and water-in-oil (W/O) creams which are composed of small droplets of water dispersed in a continuous oily phase. Oil-in-water creams are more comfortable and cosmetically acceptable as they are less greasy and more easily washed off using water. Water-in-oil creams are more difficult to handle but many drugs which are incorporated into creams are hydrophobic and will be released more readily from a water-in-oil cream than an oil-in-water cream. Water-in-oil creams are also more moisturizing as they provide an oily barrier which reduces water loss from the stratum corneum, the outermost layer of the skin.

Ointment is a semisolid dosage form it is used for topical application to the medication

Water, oil, emulsifier, and thickening agent are four main ingredients of cold creams and lotions.

Moisturizer or **moisturiser** is a cosmetic preparation used for protecting, moisturizing, and lubricating the skin. These functions are normally performed by sebum produced by healthy skin.

Moisturizers modify the rate of water loss, with active ingredients of moisturizers falling into one of two categories: occlusives and humectants

Occlusives form a coating on the surface of the skin, keeping moisture from escaping.

Humectants absorb water. They can absorb this water from the air and moisturize the skin when the humidity is greater than 70%, but more commonly they draw water from the dermis into the epidermis, making skin dryer. A study published in Skin Research and Technology in 2001 found no link between humectants and moisturizing effect. When used in practical applications, they are almost always combined with occlusives.

Moisturizers often contain water, which acts as a temporary hydration agent and as a way for the absorption of some components and evaporation of the moisturizer.

There are many different types of moisturizers. Petrolatum is one of the most effective moisturizers, although it can be unpopular due to its oily consistency. Other popular moisturizers are cetyl alcohol, cetearyl alcohol, cocoa butter, isopropyl myristate, isopropyl palmitate, lanolin, liquid paraffin, polyethylene glycols, shea butter, silicone oils, stearic acid, stearyl alcohol and castor oil, as well as other oils.

Moisturizers may also be available as lotions, creams, ointments, bath oils, or soap substitutes.

Mineral oils and waxes are insensitive to oxidation or rancidity. For this reason, they have essentially replaced vegetable oils in emollients and topical medication.

Moisturizer cosmetics may additionally contain antioxidants, ceramides, emulsifiers, fragrances, penetration enhancers, preservatives, and solvents. Some products are marketed as having anti-wrinkle and skin enhancement effects. Many plant and animal extracts have been claimed to impart skin benefits, with little scientific evidence.

Shaving Creams

Shaving cream or **shave cream** is a category of cosmetics used for shaving preparation. The purpose of shaving cream is to soften the hair by providing lubrication.

Different types of shaving creams include aerosol shaving cream (also known as shaving foam), latherless shaving cream (also called brushless shaving cream and non-aerosol shaving cream), and lather shaving cream or lathering shaving cream. The term shaving cream can also refer to the lather produced with a shaving brush from shaving soap or a lather shaving cream.

Shaving creams commonly consist of an emulsion of oils, soaps or surfactants (e.g. triethanolamine steatrate), and water. In addition to soap, lather shaving creams include a humectant for softer consistency and keeping the lather moisturised. Brushless shaving creams, on the other hand, don't contain soap and so don't produce lather. They are an oil-in-water mixture to which humectants, wetting agents, and other ingredients are added. Aerosol shaving creams are basically lather shaving cream in liquid form with propellants, vegetable waxes, and various oils added.

Does Shaving Cream Do Anything?

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 10.6.1 *What is shaving cream?*

Sunscreen and Sunblock

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.

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 10.6.4). 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.

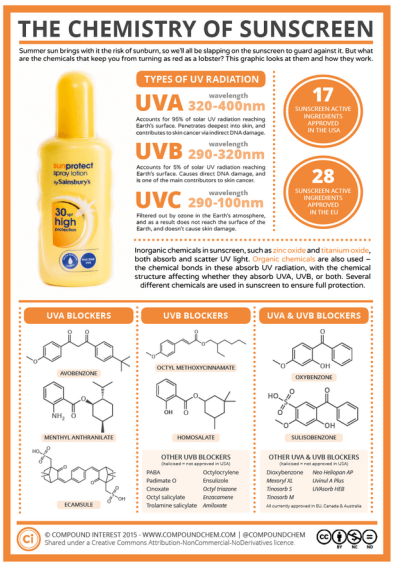


Figure 10.6.4 Chemistry of Sunscreen.

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 a 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.

Sunscreen Label

Figure 10.6.5 *Sunscreen label. Source: Badger: Healthy Body Care*

Is Sunscreen Safe?

Check the link below on FAQs regarding sunscreen.

https://www.aad.org/public/everyday-care/sun-protection/sunscreen/understand-sunscreen-labels

Lipsticks and Lip Balms

Lipstick, lip gloss, lip liner, lip plumper, lip balm, lip stain, lip conditioner, lip primer, lip boosters, and lip butters: Lipsticks are intended to add color and texture to the lips and often come in a wide range of colors, as well as finishes such as matte, satin, gloss and lustre.

Lipstick contains wax, oils, antioxidants, and emollients. Wax provides the structure to the solid lipstick. Lipsticks may be made from several waxes such as beeswax, ozokerite and candelilla wax. Because of its high melting point, carnauba wax is a key ingredient in terms of strengthening the lipstick. Various oils and fats are used in lipsticks, such as olive oil, mineral oil, cocoa butter, lanolin, and petrolatum.

Lipsticks get their colors from a variety of pigments and lake dyes including, but not limited to bromo acid, D&C Red No. 21, Calcium Lake such as D&C Red 7 and D&C Red 34, and D&C Orange No. 17. Pink lipsticks are made by mixing white titanium dioxide and red shades. Both organic and inorganic pigments are employed.

Lip stains have a water or gel base and may contain alcohol to help the product stay on leaving a matte look. They temporarily saturate the lips with a dye. Usually designed to be waterproof, the product may come with an applicator brush, directly through the applicator, rollerball, or could be applied with a finger. Lip glosses are intended to add shine to the lips and may add a tint of color, as well as being scented or flavored. Lip balms are most often used to moisturize, tint, and protect the lips. Some brands contain sunscreen. Using a priming lip product such as lip balm or chapstick can prevent chapped lips.

Eye Make-up

- **Mascara** is a cosmetic commonly used to enhance the eyelashes. It may darken, thicken, lengthen, and/or define the eyelashes. Normally in one of three forms—liquid, powder, or cream—the modern mascara product has various formulas; however, most contain the same basic components of pigments, oils, waxes, and preservatives. The most common form of mascara is a liquid in a tube.
- **Eye shadow** (or **eyeshadow**) is a cosmetic applied primarily to the eyelids to make the wearer's eyes stand out or look more attractive. Eye shadow can also be applied under eyes or to brow bones.

  Eye shadow can add depth and dimension to one's eyes, complement one's eye color, make one's eyes appear larger, or simply draw attention to the eyes. Eye shadow comes in many different colors and textures. It is usually made from a powder but can also be found in liquid, pencil, cream or mousse form. Just like other trends, eyeshadow looks also have trends.

  Eye shadows typically consist of four types of ingredients: base fillers, binders, slip and preservatives. In order to make eye shadow, there must be a balance between the fillers and binders.

  Base fillers are usually minerals such as mica, talc or kaolin clay, which add bulk and texture to eye shadow. They make up about 30% of eye shadow powders and 25% of cream eye shadows. Mica absorbs moisture, gives the eye shadow shine and luster, and makes it opaque. Mica powders, iron oxides and clays can give color pigments to eye shadows.

  Binders help eye shadow adhere and stay attached to skin. Eye shadows can have dry or liquid binders. Zinc and magnesium, which are both white powders, are commonly used as dry binders. Zinc also adds color and can be used to increase the thickness of the eyeshadow. Silicone, paraffin wax, mineral oil or vegetable oils may be used as liquid binders.

  Slip allows eye shadow to glide across the skin smoothly. Products may use silica or nylon, which are fine, colorless powders. Other types of slip include dimethicone, boron nitride or bismuth oxychloride.

  Preservatives help products stay bacteria free and extend their lifespan. Common preservatives in eye shadow are glycol and tocopherol.



Figure 10.6.6 Eyes with heavy gold eye shadow. Source: Wikipedia

Eye liner or **eyeliner** is a cosmetic used to define the eyes. It is applied around the contours of the eye(s) to create a variety of aesthetic effects. It can come in the form of a pencil, a gel, or a liquid and can be found in almost any color. Traditional wax-based

eye liners are made from about 20 components. About 50% by weight are waxes (e.g., Japan wax, fats, or related soft materials that easily glide on to the skin. Stearyl heptanoate is found in most cosmetic eyeliner. Typical pigments include black iron oxides, as well as smaller amounts of titanium dioxide and Prussian blue.

Eyebrow shaders, and pencils are often used to define the eyebrow or make it appear fuller. These can create an outline for the brows or mimic hairs where there are sparse areas. Brow gels are also used in creating a thicker brow; they allow for the hairs to be more textured, which gives the appearance of thicker, fuller brows. Lastly, brow powders or even eye shadows are used for those who want a fuller and more natural look, by placing the brow powder or eye shadow (closest to the natural hair color) in areas where there is less hair.

Deodorants and Antiperspirants

A **deodorant** is a substance applied to the body to prevent or mask body odor due to bacterial breakdown of perspiration in the armpits, groin, and feet, and in some cases vaginal secretions. A subclass of deodorants, called **antiperspirants**, prevents sweating itself, typically by blocking sweat glands. Antiperspirants are used on a wider range of body parts, at any place where sweat would be inconvenient or unsafe, since unwanted sweating can interfere with comfort, vision, and grip (due to slipping). Other types of deodorant allow sweating but prevent bacterial action on sweat, since human sweat only has a noticeable smell when it is decomposed by bacteria.

In the United States, the Food and Drug Administration classifies and regulates most deodorants as cosmetics but classifies antiperspirants as over-the-counter drugs.

In the United States, deodorants are classified and regulated as cosmetics by the U.S. Food and Drug Administration (FDA) and are designed to eliminate odor. Deodorants are often alcohol based. Alcohol initially stimulates sweating but may also temporarily kill bacteria. Other active ingredients in deodorants include sodium stearate, sodium chloride, and stearyl alcohol. Deodorants can be formulated with other, more persistent antimicrobials such as triclosan that slow bacterial growth or with metal chelant compounds such as EDTA. Deodorants may contain perfume fragrances or natural essential oils intended to mask the odor of perspiration. In the past, deodorants included chemicals such as zinc oxide, acids, ammonium chloride, sodium bicarbonate, and formaldehyde, but some of these ingredients were messy, irritating to the skin, or even carcinogenic.

Over-the-counter products, often labeled as "natural deodorant crystal", contain the chemical rock crystals potassium alum or ammonium alum, which prevents bacterial action on sweat. These have gained popularity as an alternative health product, in spite of concerns about possible risks related to aluminum (see below – all alum salts contain aluminum in the form of aluminum sulphate salts) and contact dermatitis.

In the United States, deodorants combined with antiperspirant agents are classified as drugs by the FDA. Antiperspirants attempt to stop or significantly reduce perspiration and thus reduce the moist climate in which bacteria thrive. Aluminium chloride, aluminium chlorohydrate, and aluminium-zirconium compounds, most notably aluminium zirconium tetrachlorohydrex gly and aluminium zirconium trichlorohydrex gly, are frequently used in antiperspirants. Aluminium chlorohydrate and aluminium-zirconium tetrachlorohydrate gly are the most frequent active ingredients in commercial antiperspirants. Aluminium-based complexes react with the electrolytes in the sweat to form a gel plug in the duct of the sweat gland. The plugs prevent the gland from excreting liquid and are removed over time by the natural sloughing of the skin. The metal salts work in another way to prevent sweat from reaching the surface of the skin: the aluminium salts interact with the keratin fibrils in the sweat ducts and form a physical plug that prevents sweat from reaching the skin's surface. Aluminium salts also have a slight astringent effect on the pores; causing them to contract, further preventing sweat from reaching the surface of the skin. The blockage of a large number of sweat glands reduces the amount of sweat produced in the underarms, though this may vary from person to person. Methenamine (Figure 10.6.7) in the form of cream or spray is effective in the treatment of excessive sweating and attendant odor. Antiperspirants are usually best applied before bed.

Figure 10.6.7 *Methenamine-based antiperspirant for treatment of excessive sweating.*

Toothpaste: Soap with Grit and Flavor

Toothpaste (Figure 10.6.8) is a paste or gel dentifrice used with a toothbrush to clean and maintain the aesthetics and health of teeth. Toothpaste is used to promote oral hygiene: it is an abrasive that aids in removing dental plaque and food from the teeth, assists in suppressing halitosis, and delivers active ingredients (most commonly fluoride) to help prevent tooth decay (dental caries) and gum disease (gingivitis). Salt and sodium bicarbonate (baking soda) are among materials that can be substituted for

commercial toothpaste. Large amounts of swallowed toothpaste can be toxic.



Figure 10.6.8 A variety of brands of toothpaste.

In addition to 20%–42% water, toothpastes are derived from a variety of components, the three main ones being abrasives, fluoride, and detergents.

Abrasives constitute at least 50% of a typical toothpaste. These insoluble particles are designed to help remove plaque from the teeth. The removal of plaque and calculus prevents the accumulation of tartar and is widely claimed to help minimize cavities and periodontal disease, although the clinical significance of this benefit is debated. Representative abrasives include particles of aluminum hydroxide ($Al(OH)_3$), calcium carbonate ($CaCO_3$), various calcium hydrogen phosphates, various silicas and zeolites, and hydroxyapatite ($Ca_5(PO_4)_3OH$).

Abrasives, like the dental polishing agents used in dentists' offices, also cause a small amount of enamel erosion which is termed "polishing" action. Some brands contain powdered white mica, which acts as a mild abrasive, and also adds a cosmetically pleasing glittery shimmer to the paste. The polishing of teeth removes stains from tooth surfaces, but has not been shown to improve dental health over and above the effects of the removal of plaque and calculus.

The abrasive effect of toothpaste is indicated by its RDA value. Too high RDA values are deleterious. Some dentists recommend toothpaste with an RDA value no higher than 50 for daily use.

Fluoride in various forms is the most popular active ingredient in toothpaste to prevent cavities. Fluoride is present in small amounts in plants, animals, and some natural water sources. The additional fluoride in toothpaste has beneficial effects on the formation of dental enamel and bones. Sodium fluoride (NaF) is the most common source of fluoride, but stannous fluoride (SnF_2), olaflur (an organic salt of fluoride), and sodium monofluorophosphate (Na_2PO_3F) are also used. Stannous fluoride has been shown to be more effective than sodium fluoride in reducing the incidence of dental caries and controlling gingivitis, but causes somewhat more surface stains.

Much of the toothpaste sold in the United States has 1,000 to 1,100 parts per million fluoride. In European countries, such as the UK or Greece, the fluoride content is often higher; a NaF content of 0.312% w/w (1,450 ppm fluoride) is common. All of these concentrations are likely to prevent tooth decay, according to a 2019 Cochrane review. Concentrations below 1,000 ppm are not likely to be preventive, and the preventive effect increases with concentration. Clinical trials support the use of high fluoride dentifrices, as it was found to reduce the amount of plaque accumulated, decrease the number of mutans streptococci and lactobacilli and possibly promote calcium fluoride deposits to a higher degree than after the use of traditional fluoride containing dentifrices. However, these effects must be balanced with the increased risk of harm at higher concentrations.

Many, although not all, toothpastes contain sodium lauryl sulfate (SLS) or related **surfactants (detergents)**. SLS is found in many other personal care products as well, such as shampoo, and is mainly a foaming agent, which enables uniform distribution of toothpaste, improving its cleansing power.

Other Components in toothpaste formulations include:

- Antibacterial agents. Triclosan or zinc chloride prevent gingivitis and, according to the American Dental Association, helps reduce tartar and bad breath.
- Flavorants. Toothpaste comes in a variety of colors and flavors intended to encourage use of the product. The three most common flavorants are peppermint, spearmint, and wintergreen.
- Reminalizers. Hydroxyapatite nanocrystals and a variety of calcium phosphates are included in formulations for remineralization, i.e. the reformation of enamel.

Agents are added to suppress the tendency of toothpaste to dry into a powder. Included are various sugar alcohols, such as glycerol, sorbitol, or xylitol, or related derivatives, such as 1,2-propylene glycol and polyethyleneglycol

Strontium chloride or potassium nitrate is included in some toothpastes to reduce sensitivity. Two systemic meta-analysis reviews reported that arginine, and calcium sodium phosphosilicate - CSPS containing toothpastes are also effective in alleviating dentinal hypersensitivity respectively. Another randomized clinical trial found superior effects when both formulas were combined together.

Sodium polyphosphate is added to minimize the formation of tartar. Other example to components in toothpastes is the Biotene, which has proved its efficiency in relieving the symptoms of dry mouth in people who suffer from xerostomia according to the results of two randomized clinical trials.

Chlorohexidine mouthwash has been popular for its positive effect on controlling plaque and gingivitis, however, a systemic review studied the effects of chlorohexidine toothpastes and found insufficient evidence to support its use, tooth surface discoloration was observed as a side effect upon using it, which is considered a negative side effect that can affect patients' compliance.

Sodium hydroxide, also known as lye or caustic soda, is listed as an inactive ingredient in some toothpaste, for example Colgate Total.

Some studies have demonstrated that toothpastes with xylitol as an ingredient are more effective at preventing dental caries in permanent teeth of children than toothpastes containing fluoride alone.

Perfumes, Colognes, and Aftershaves

Perfume (Figure 10.6.9) is a mixture of fragrant essential oils or aroma compounds, fixatives and solvents, used to give the human body, animals, food, objects, and living-spaces an agreeable scent. Perfume types reflect the concentration of aromatic compounds in a solvent, which in fine fragrance is typically ethanol or a mix of water and ethanol. Various sources differ considerably in the definitions of perfume types. The intensity and longevity of a perfume is based on the concentration, intensity, and longevity of the aromatic compounds, or perfume oils, used. Specific terms are used to describe a fragrance's approximate concentration by the percent of perfume oil in the volume of the final product. The most widespread terms are:

- **parfum** or **extrait**, in English known as perfume extract, pure perfume, or simply perfume: 15–40% aromatic compounds (IFRA: typically ~20%);
- **esprit de parfum** (ESdP): 15–30% aromatic compounds, a seldom used strength concentration in between EdP and perfume;
- **eau de parfum** (EdP) or **parfum de toilette** (PdT) (The strength usually sold as "perfume"): 10–20% aromatic compounds (typically ~15%); sometimes called "eau de perfume" or "millésime"; parfum de toilette is a less common term, most popular in the 1980s, that is generally analogous to eau de parfum;
- **eau de toilette** (EdT): 5–15% aromatic compounds (typically ~10%); This is the staple for most masculine perfumes.
- **eau de Cologne** (EdC): often simply called cologne: 3–8% aromatic compounds (typically ~5%);
- **eau fraiche**: products sold as "splashes", "mists", "veils" and other imprecise terms. Generally these products contain 3% or less aromatic compounds and are diluted with water rather than oil or alcohol.



Figure 10.6.9 A bottle of perfume by Yves Saint Laurent

Perfume is described in a musical metaphor as having three sets of *notes*, making the harmonious scent *accord*. The notes unfold over time, with the immediate impression of the top note leading to the deeper middle notes, and the base notes gradually appearing as the final stage. These notes are created carefully with knowledge of the evaporation process of the perfume.

- **Top notes**: Also called the ***head notes***. The scents that are perceived immediately on application of a perfume. Top notes consist of small, light molecules that evaporate quickly. They form a person's initial impression of a perfume and thus are very important in the selling of a perfume. Examples of top notes include mint, lavender and coriander.
- **Middle notes**: Also referred to as ***heart notes***. The scent of a perfume that emerges just prior to the dissipation of the top note. The middle note compounds form the "*heart*" or main body of a perfume and act to mask the often unpleasant initial impression of base notes, which become more pleasant with time. Examples of middle notes include seawater, sandalwood and jasmine.
- **Base notes**: The scent of a perfume that appears close to the departure of the middle notes. The base and middle notes together are the main theme of a perfume. Base notes bring depth and solidity to a perfume. Compounds of this class of scents are typically rich and "deep" and are usually not perceived until 30 minutes after application. Examples of base notes include tobacco, amber and musk.

The scents in the top and middle notes are influenced by the base notes; conversely, the scents of the base notes will be altered by the types of fragrance materials used as middle notes. Manufacturers who publish perfume notes typically do so with the fragrance components presented as a *fragrance pyramid*, using imaginative and abstract terms for the components listed.

Aftershave is a product applied to skin after shaving. Traditionally it is an alcohol based liquid (splash), but it can be a lotion, gel, or even a paste.

It often contains an antiseptic agent such as denatured alcohol, stearate citrate or witch hazel to prevent infection of cuts, as well as to act as an astringent to reduce skin irritation. Menthol is used in some varieties as well to numb irritated skin.

An alcohol-based aftershave usually causes an immediate stinging sensation after applying it post-shave, with effects sometimes lasting several minutes, but most commonly only for seconds. For this reason, a market consisting of highly differentiated products has been created—some using alcohols, some not.

Aftershave balms are frequently recommended for winter use as they tend to be alcohol free and lotion-like, moisturizing the skin.

Some aftershaves use fragrance or essential oil to enhance scent. Moisturizers—natural and artificial, are often touted as able to soften the skin.

Aftershave is sometimes mistakenly referred to as Eau de Cologne due to the very similar nature of the two products. Some aftershave manufacturers encourage using their fragranced aftershave as if it were cologne, in order to increase sales by encouraging consumers to use it in a more versatile manner, rather than just after a shaving session. Some aftershaves were

inspired by a cologne.

Early aftershaves included witch-hazel and bay rum, and have been documented in shaving guides. Both still are sold as aftershaves.

Hairy Chemistry

Hair keratin consists of many protein alpha-helices (Figure 10.6.10). Three alpha-helices are interwoven into a left-handed coil called a protofibril. Eleven protofibrils are bonded and coiled together to make a microfibril. Hundreds of these microfibrils are cemented into an irregular bundle called a macrofibril. These in turn are mixed with dead and living cells to make a complete strand of hair.

The alpha-helices are extensively cross-linked with disulfide bonds from cysteine. These bonds enable keratin to have a somewhat elastic nature. If the alpha -helices stretch unevenly past each other, the disulfide cross-links return them to the original position when the tension is released.

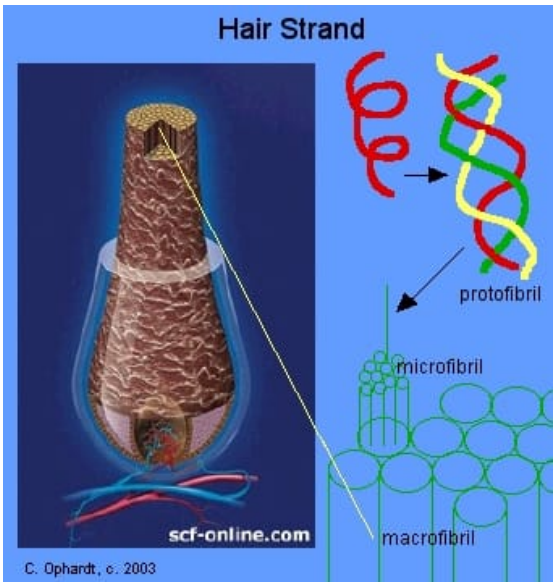


Figure 10.6.10 Protein alpha-helices in a hair strand.

Shampoo

Shampoo (/ʃæmˈpuː/) is a hair care product, typically in the form of a viscous liquid, that is used for cleaning hair. Less commonly, shampoo is available in bar form, like a bar of soap. Shampoo is used by applying it to wet hair, massaging the product into the scalp, and then rinsing it out. Some users may follow a shampooing with the use of hair conditioner.

The typical reason of using shampoo is to remove the unwanted build-up of sebum in the hair without stripping out so much as to make hair unmanageable. Shampoo is generally made by combining a surfactant, most often sodium lauryl sulfate or sodium laureth sulfate, with a co-surfactant, most often cocamidopropyl betaine in water. The sulphate ingredient acts as a surfactant, essentially heavy duty soap that makes it easier to trap oil and grease.

Specialty shampoos are marketed to people with dandruff, color-treated hair, gluten or wheat allergies, an interest in using an organic product, and infants and young children ("baby shampoo" is less irritating). There are also shampoos intended for animals that may contain insecticides or other medications to treat skin conditions or parasite infestations such as fleas.

Shampoo is generally made by combining a surfactant, most often sodium lauryl sulfate or sodium laureth sulfate, with a co-surfactant, most often cocamidopropyl betaine in water to form a thick, viscous liquid. Other essential ingredients include salt (sodium chloride), which is used to adjust the viscosity, a preservative and fragrance. Other ingredients are generally included in shampoo formulations to maximize the following qualities:

- pleasing foam
- ease of rinsing
- minimal skin and eye irritation
- thick or creamy feeling
- pleasant fragrance
- low toxicity
- good biodegradability
- slight acidity (pH less than 7)
- no damage to hair
- repair of damage already done to hair

Many shampoos are pearlescent. This effect is achieved by the addition of tiny flakes of suitable materials, e.g. glycol distearate, chemically derived from stearic acid, which may have either animal or vegetable origins. Glycol distearate is a wax. Many shampoos also include silicone to provide conditioning benefits.

How Does Shampoo Work?

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 10.6.2 *Shampoo applied to wet hair.*

Hair Coloring

Hair color is the pigmentation of hair follicles due to two types of melanin: eumelanin and pheomelanin. Generally, if more eumelanin is present, the color of the hair is darker; if less eumelanin is present, the hair is lighter. Levels of melanin can vary over time causing a person's hair color to change, and it is possible to have hair follicles of more than one color on the same person. Particular hair colors are often associated with ethnic groups, while gray or white hair is associated with age.

Hair coloring, or **hair dyeing**, is the practice of changing the hair color. The main reasons for this are cosmetic: to cover gray or white hair, to change to a color regarded as more fashionable or desirable, or to restore the original hair color after it has been discolored by hairdressing processes or sun bleaching.

Hair coloring can be done professionally by a hairdresser or independently at home. Today, hair coloring is very popular, with 75% of women and 18% of men living in Copenhagen having reported using hair dye (according to a study by the University of Copenhagen). At-home coloring in the United States reached $1.9 billion in 2011 and was expected to rise to $2.2 billion by 2016.

Hair color can be changed by a chemical process. Hair coloring is classed as "permanent" or "semi-permanent".

Permanent hair color means that the hair's structure has been chemically altered until it is eventually cut away. This does not mean that the synthetic color will remain permanently. During the process, the natural color is removed, one or more shades, and synthetic color has been put in its place. All pigments wash out of the cuticle. Natural color stays in much longer and artificial will fade the fastest (depending on the color molecules and the form of the dye pigments).

Permanent hair coloring requires three components: (1) 1,4-diaminobenzene (historically) or 2,5-diaminotoluene (currently), (2) a coupling agent, and (3) an oxidant. The process is typically performed under basic conditions. The mechanism of oxidation dyes involves three steps: 1) Oxidation of 1,4-diaminobenzene derivative to the quinone state. 2) Reaction of this diimine with a coupler compound (more detail below). 3) Oxidation of the resulting compound to give the final dye.

Figure 10.6.11 *Shown here are three red couplers (A, B, C), two yellow-green couplers (D, E) and a blue coupler (F). Couplers are chemical compounds that define the color of the hair dye.*

Steps in Permanent Hair Coloring

The first step shows the oxidation of p-phenylenediamine to the quinonediimine ($C_6H_4(NH)_2$):

Quinodiimine.png

The second step involves the attack of this quinonediimine on the coupler.

Quinodiimine+Coupler.png

In the third and final step, the product from the quinonediimine-coupler reaction oxidizes to the final hair dye.

QuinodiimineCouplerOx.png

It was once believed that the dye forms in the above reaction bonds to hair permanently. It was later shown that the main reason that this reaction imparts a permanent color on hair by producing larger dye molecules, which is locked inside the hair.

Semi-permanent color washes out over a period of time—typically four to six weeks, so root regrowth is less noticeable. The final color of each strand is affected by its original color and porosity, so there will be subtle variations in color across the head—more natural and less harsh than a permanent dye. However, this means that gray and white hair will not dye to the same color as the rest of the head (in fact, some white hair will not absorb the color at all). A few gray and white hairs will blend in sufficiently not to be noticeable, but as they become more widespread, there will come a point where a semi-permanent alone will not be enough. The move to 100% permanent color can be delayed by using a semi-permanent as a base color, with permanent highlights.

Semi-permanent hair color cannot lighten hair. Hair can only be lightened using chemical lighteners, such as bleach. Bleaching is always permanent because it removes the natural pigment.

"Rinses" are a form of temporary hair color, usually applied to hair during a shampoo and washed out again the next time the hair is wash.

Plant based dyes include henna, indigo and anthocyanin pigments extracted form blackcurrant skin waste.

How Does Hair Dye Work?

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 $10.6.3$ *Applying hair dye.*

Permanent and Temporary Waving

Temporary Wave. When the hair gets wet, water molecules intrude into the keratin strands. The sheer numbers of water molecules are able to disrupt some of the hydrogen bonds which also help to keep the alpha-helices aligned. The helices are able to slip past each other and will retain a new shape in the hair drying process as new hydrogen bonds are formed. The hair strands are able for a short time to maintain the new curl in the hair.

A **permanent hairstyle**, commonly called a **perm** or "**permanent**" (sometimes called a "perm" to distinguish it from a "straight perm"), is a hairstyle consisting of styles set into the hair. The hairstyle may last a number of months, hence the name.

Perms may be applied using thermal or chemical means. In the latter method, chemicals are applied to the hair, which is then wrapped around forms to produce hairstyles. The same process is used for chemical straightening or relaxing, with the hair being flattened instead of curled during the chemical reaction.

The formation of disulfide bonds has a direct application in producing curls in hair by the permanent wave process. Disulfide bonds Figure $10.6.12$ are formed by oxidation of the sulfhydryl groups on cysteine. Different protein chains or loops within a single

chain are held together by the strong covalent disulfide bonds. The alpha-helices in the hair strands are bonded by disulfide links.

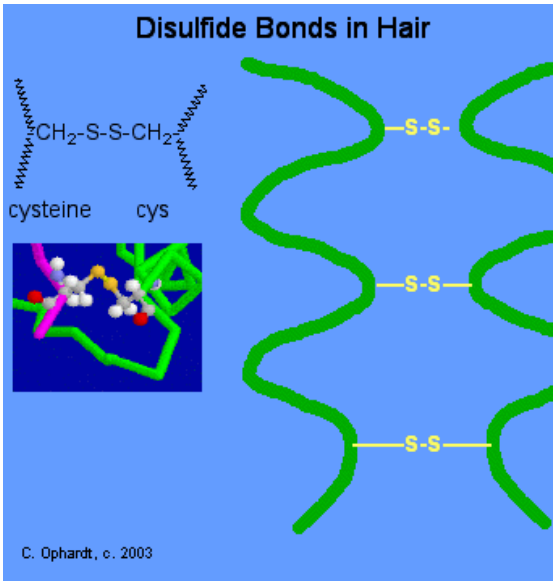


Figure 10.6.12 Disulfide bonds in hair.

In the permanent wave process, a basic reducing substance (usually ammonium thioglycolate) is first added to reduce and rupture some of the disulfide cross-links, see Figure 10.6.13 below.

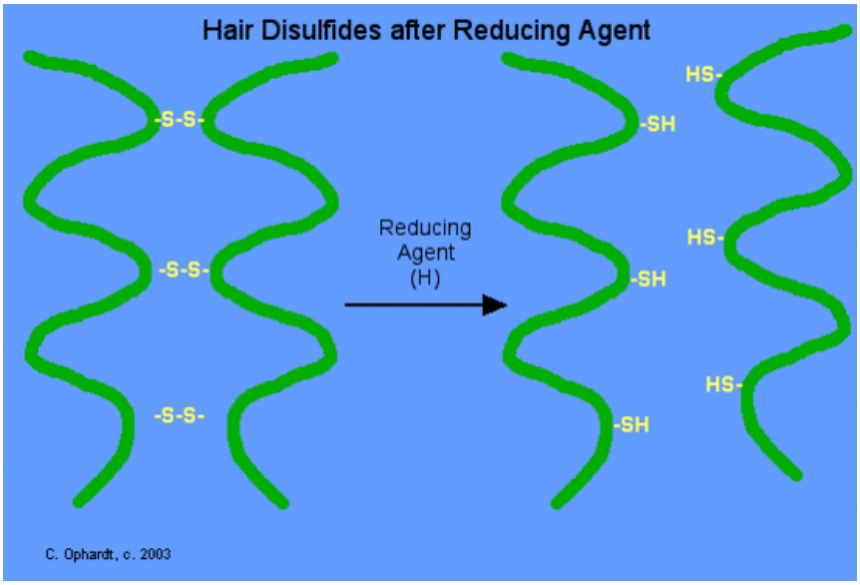


Figure 10.6.13 Addition of a reducing agent to disrupt the disulfide bonds.

The hair is put on rollers or curlers. Since the alpha-helices are no longer tightly cross-linked to each other, the alpha-helices can shift positions in relation to each other. An oxidizing agent, usually a dilute solution of hydrogen peroxide, (also called the neutralizer) is added to reform the disulfide bonds in their new positions (Figure 10.6.14). The permanent will hold these new disulfide bond positions until the hair grows out, since new hair growth is of course not treated.

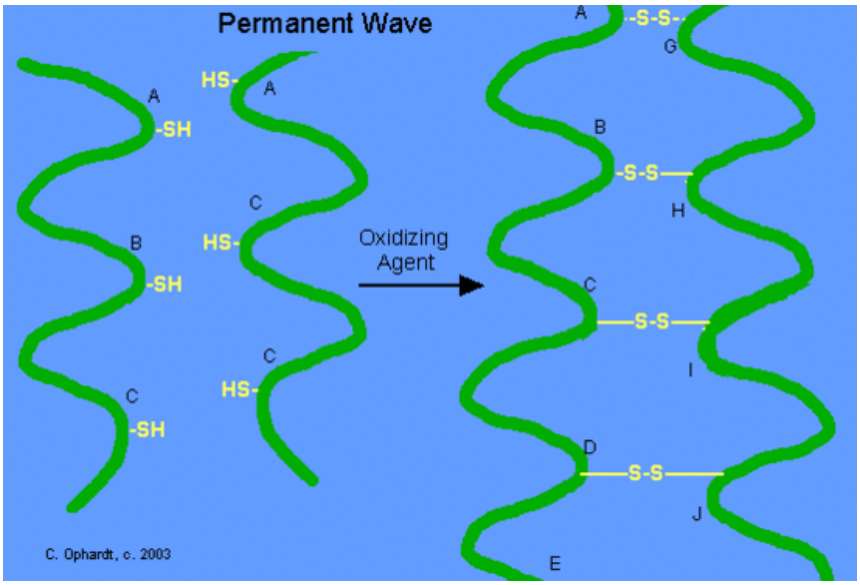


Figure 10.6.14 Addition of a oxidizing agent to reform disulfide bonds.

Hair spray

Hair spray is a common household aqueous solution which is used to stiffen hair into a certain style. It was first developed and manufactured in 1948 by Chase Products, based in Broadview, Illinois. Weaker than hair gel or hair wax, it is sprayed directly onto the hair to hold styles for long periods of time. It sprays evenly over the hair using a pump or aerosol spray nozzle. The product may leave hair feeling 'crunchy' unless brushed out.

The active ingredients in hair spray are called polymers, which keep the hair stiff and firm without snapping. Solvents, which make up most of the content of the hairspray, are responsible for carrying these polymers in a solution.

Originally, the solvent found in hair spray was a chlorofluorocarbon (CFC). CFCs are nontoxic, nonflammable, and make almost ideal aerosol propellants. However, when research concluded that CFCs cause destruction of stratospheric ozone, they were replaced with other solvents, such as alcohols and hydrocarbons.

Hair sprays consist of the following components: concentrate, plasticizers, luster agents, and fragrances, as well as propellants.

One of the polymers used in hair spray is polyvinylpyrrolidone (Figure 10.6.15) , which is water-soluble. The non-water-soluble polymer polydimethylsiloxane is added to make the hold last a bit longer. Some less common polymers found in hair spray include copolymers with vinyl acetate and copolymers with maleic anhydride.

Figure 10.6.15 Polyvinylpyrrolidone *is a common component of hair spray that confers stiffness to hair.*

Some hair sprays use natural polymers and solvents like vegetable gums dissolved in alcohol. One popular ingredient in natural hair sprays is gum arabic, which is made from the sap of various species of the acacia tree. Gum tragacanth is another herbal gum that is used to stiffen calico and crepe, as well as hair.

Hair Removers

Depilation is the removal of the part of the hair above the surface of the skin. The most common form of depilation is shaving or trimming. Another option is the use of chemical depilatories, which work by breaking the disulfide bonds that link the protein chains that give hair its strength.

A **chemical depilatory** is a cosmetic preparation used to remove hair from the skin. Common active ingredients are salts of thioglycolic acid and thiolactic acids. These compounds break the disulfide bonds in keratin and also hydrolyze the hair so that it is easily removed. Formerly, sulfides such as strontium sulfide were used, but due to their unpleasant odor, they have been replaced by thiols.

The main chemical reaction effected by the thioglycolate is:

2 $HSCH_2CO_2H$ (thioglycolic acid) + R-S-S-R (cystine) → HO_2CCH_2-S-S-CH_2CO_2H (dithiodiglycolic acid) + 2 RSH (cysteine)

Chemical depilatories contain 5–6% calcium thioglycolate in a cream base (to avoid runoff). Calcium hydroxide or strontium hydroxide maintain a pH of about 12. Hair destruction requires about 10 minutes. Depilation is followed by careful rinsing with water, and various conditioners are applied to restore the skin's pH to normal. Depilation does not destroy the dermal papilla, and the hair grows back.

Chemical depilatories are available in gel, cream, lotion, aerosol, roll-on, and powder forms. Common brands include Nair, Magic Shave and Veet.

Depilatory ointments, or plasters, were known to Greek and Roman authors as *psilothrum*. In Jewish lore, King Solomon is said to have discovered a chemical depilatory made from a mixture of lime and water and orpiment (arsenic trisulfide).

Hair Restorers

Treatments for the various forms of hair loss have only moderate success. Three medications have evidence to support their use in male pattern hair loss: finasteride, dutasteride and minoxidil. They typically work better to prevent further hair loss than to regrow lost hair.

They may be used together when hair loss is progressive or further regrowth is desired after 12 months. Other medications include ketoconazole, and in female androgenic alopecia spironolactone and flutamide. Combinations of finasteride, minoxidil and ketoconazole are more effective than individual use.

Minoxidil is applied topically, is widely used for the treatment of hair loss. It may be effective in helping promote hair growth in both men and women with androgenic alopecia. About 40% of men experience hair regrowth after 3–6 months. It is the only topical product that is FDA approved in America for androgenic hair loss. However, increased hair loss has been reported.

Finasteride is used to treat male pattern hair loss. Treatment provides about 30% improvement in hair loss after six months of treatment, and effectiveness only persists as long as the drug is taken. There is no good evidence for its use in women. It may cause gynecomastia, erectile dysfunction and depression.

Dutasteride is used off label for male pattern hair loss.

There is tentative support for **spironolactopne** in women. Due to its feminising side effects and risk of infertility it is not often used by men. It can also cause low blood pressure, high blood potassium, and abnormal heart rhythms. Also, women who are pregnant or trying to become pregnant generally cannot use the medication as it is a teratogen, and can cause ambiguous genitalia in newborn children.

There is tentative evidence for **flutamide** in women; however, it is associated with relatively high rates of liver problems. Like spironolactone, it is typically only used by women.

Ketoconazole shampoo in conjunction with an oral 5α-reductase inhibitor such as finasteride or dutasteride has been used off label to treat androgenic alopecia.

Summary

- Various personal care products contain ingredients to protect the skin, protect the hair, promote hygiene, for aesthetic purposes etc.
- Temporary and permanent waves are formed due to the disruption and reformation of disulfide bonds in hair strands.

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11: DRUGS

11.1: Pain Relievers- From Aspirin to Oxycodone

11.2: Cold and Allergy Medicines

11.3: Drugs and Infectious Diseases

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11.5: Hormones - The Regulators

11.6: Drugs for the Heart

11.7: Drugs and the Mind

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11.1: Pain Relievers- From Aspirin to Oxycodone

Learning Objectives

- Know the difference between anti-inflammatory, analgesic and antipyretic drugs.
- Describe how aspirin and other NSAIDs work.
- Know the side effects of taking prescription and non-prescription medications
- Know the risks of addiction, abuse, and overdose of prescription opioids.

A **drug** is any substance that causes a change in an organism's physiology or psychology when consumed. Drugs are typically distinguished from food and substances that provide nutritional support. Consumption of drugs can be via inhalation, injection, smoking, ingestion, absorption via a patch on the skin, suppository, or dissolution under the tongue.

Anti-inflammatory, analgesic (pain reliever or pain killer), and antipyretic (fever reducer) 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.

Over-the-counter (OTC) medicines are good for many types of pain. There are two main types of OTC pain medicines: acetaminophen (Tylenol) and nonsteroidal anti-inflammatory drugs (NSAIDs). Aspirin, naproxen (Aleve), and ibuprofen (Advil, Motrin) are examples of OTC NSAIDs.

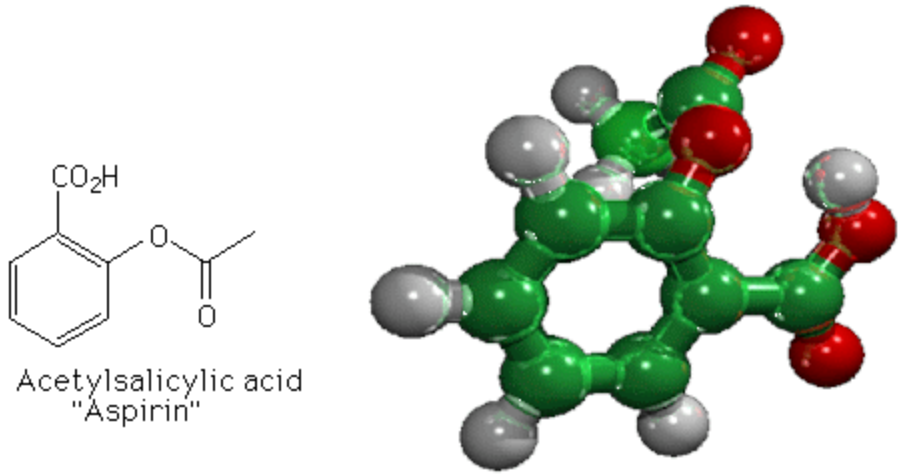
If OTC medicines don't relieve ones pain, the doctor may prescribe something stronger. Many NSAIDs are also available at higher prescription doses. The most powerful pain relievers are opioids. They are very effective, but they can sometimes have serious side effects. There is also a risk of addiction. Because of the risks, you must use them only under a doctor's supervision.

Aspirin and other NSAIDs

Nonsteroidal anti-inflammatory drugs (**NSAIDs**) are members of a drug class that reduces pain, decreases fever, prevents blood clots, and in higher doses, decreases inflammation. Side effects depend on the specific drug but largely include an increased risk of gastrointestinal ulcers and bleeds, heart attack, and kidney disease.

The term *nonsteroidal* distinguishes these drugs from steroids, which while having a similar eicosanoid-depressing, anti-inflammatory action, have a broad range of other effects. First used in 1960, the term served to distance these medications from steroids, which were particularly stigmatised at the time due to the connotations with anabolic steroid abuse.The most prominent NSAIDs are aspirin, ibuprofen, and naproxen, all available over the counter (OTC) in most countries.

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

Figure 11.1.1 Aspirin

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.

Other Aspirin-Like Drugs

There are other non-selective NSAIDs on the market, including diclofenac, ibuprofen, ketoprofen, meloxicam, naproxen, and oxaprozin. Ibuprofen and naproxen are available in both prescription and over-the-counter (OTC) versions. The doses in OTC ibuprofen (Advil and Motrin) or naproxen (Aleve) are about half the doses of prescription versions.

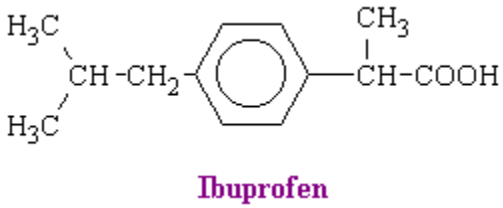


Figure 11.1.2 Ibuprofen.

How NSAIDs Work?

Non-selective NSAIDs like aspirin work (Figure 11.1.3) by inhibiting two enzymes that are involved with pain and inflammation—cyclooxygenase-1 and cyclooxygenase-2 (COX-1 and COX-2). These enzymes are 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 (Figure 11.1.2). This has the effect of blocking the channel in the enzyme and arachidonic acid can not enter the active site of the enzyme. By inhibiting or blocking this enzyme, the synthesis of prostaglandins is blocked, which in turn relieves 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.

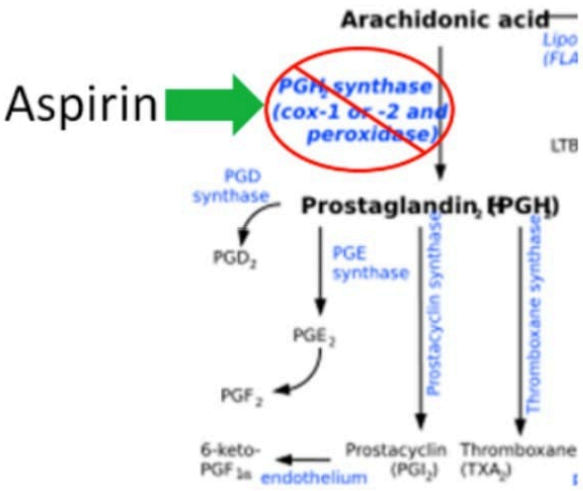


Figure 11.1.3 Aspirin inhibition of COX-1 and COX-2 enzymes.

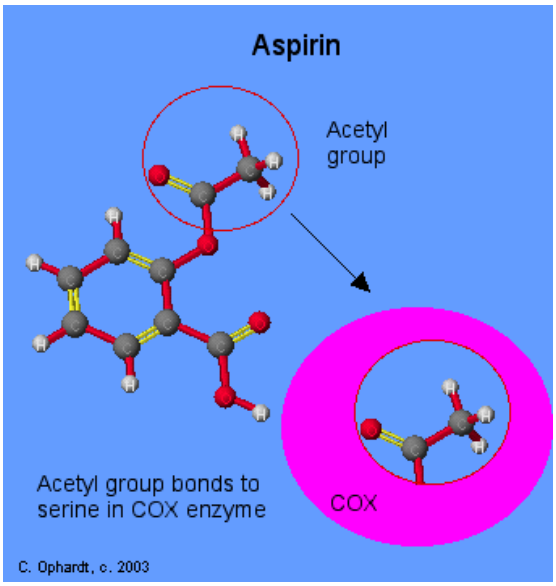


Figure 11.1.4 The acetyl group (from aspirin) bonds to serine in COX enzyme.

COX-2 selective inhibitors are a newer type of medicine that block the COX-2 enzyme more than the COX-1 enzyme. The only COX-2 selective inhibitor currently on the market in the United States is the prescription drug Celebrex (celecoxib), which is marketed by Pfizer. It was believed that COX-2 inhibitors may be less likely to cause the stomach problems associated with the older NSAIDs, but all NSAIDs carry the risk of stomach problems.

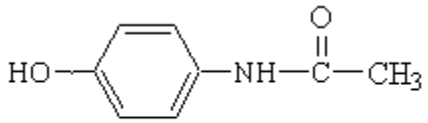
Key Points to Remember With NSAIDs

- Too much can cause stomach bleeding. This risk increases in people who are over 60 years of age, are taking prescription blood thinners, are taking steroids, have a history of stomach bleeding or ulcers, and/or have other bleeding problems.
- Use of NSAIDs can also cause kidney damage. This risk may increase in people who are over 60 years of age, are taking a diuretic (a drug that increases the excretion of urine), have high blood pressure, heart disease, or pre-existing kidney disease.

Acetaminophen

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.

Key Points to Remember With Acetaminophen

- Taking a higher dose than recommended will not provide more relief and can be dangerous.
- Too much can lead to liver damage and death. Risk for liver damage may be increased in people who drink three or more alcoholic beverages a day while using acetaminophen-containing medicines.
- Be cautious when giving acetaminophen to children. Infant drop medications can be significantly stronger than regular children's medications. Read and follow the directions on the label every time you use a medicine. Be sure that your infant is getting the infants' pain formula and your older child is getting the children's pain formula.

Prescription Medications

Typical prescription pain relief medicines include opioids and non-opioid medications. Types of prescription opioid medications (Figures 11.1.6 and 11.1.7) include:

- **morphine**, which is often used before and after surgical procedures to alleviate severe pain
- **oxycodone**, which is also often prescribed for moderate to severe pain
- **codeine**, which comes in combination with acetaminophen or other non-opioid pain relief medications and is often prescribed for mild to moderate pain
- **hydrocodone**, which comes in combination with acetaminophen or other non-opioid pain relief medications and is prescribed for moderate to moderately severe pain

Derived from opium, opioid drugs are very powerful products. They act by attaching to a specific "receptor" in the brain, spinal cord, and gastrointestinal tract. Opioids can change the way a person experiences pain. 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 Figure 11.1.6), 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. 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.

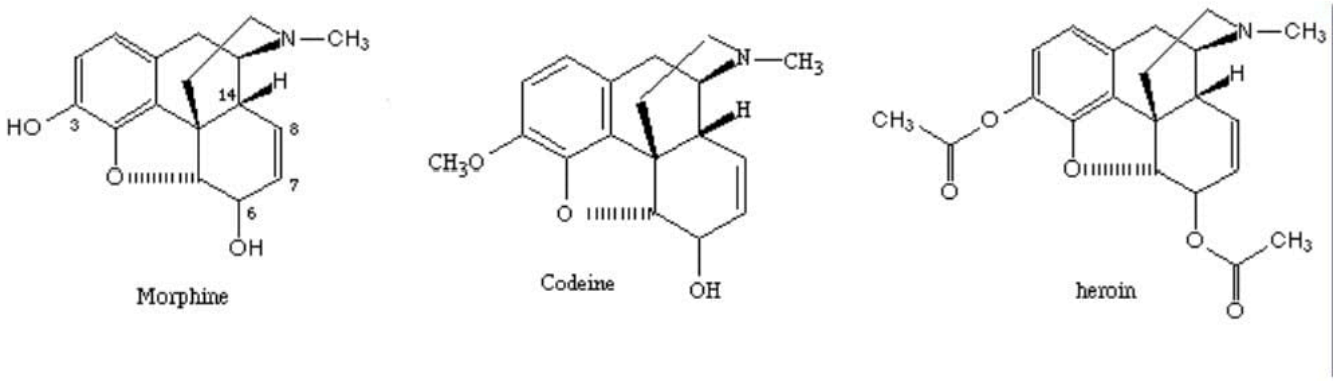


Figure 11.1.6 Morphine, codeine and heroin.

Oxycodone and hydrocodone are used to relieve moderate to severe pain. Oxycodone and hydrocodone extended-release tablets and extended-release capsules are used to relieve severe pain in people who are expected to need pain medication around the clock for a long time and who cannot be treated with other medications.

Addiction and Overdose

Anyone who takes prescription opioids can become addicted to them. In fact, as many as one in four patients receiving long-term opioid therapy in a primary care setting struggles with opioid addiction. Once addicted, it can be hard to stop. In 2016, more than 11.5 million Americans reported misusing prescription opioids in the past year. Taking too many prescription opioids can stop a person's breathing—leading to death.

Prescription opioid overdose deaths also often involve benzodiazepines. Benzodiazepines are central nervous system depressants used to sedate, induce sleep, prevent seizures, and relieve anxiety. Examples include alprazolam (Xanax®), diazepam (Valium®), and lorazepam (Ativan®). Avoid taking benzodiazepines while taking prescription opioids whenever possible.

Side Effects of Prescription Opioids

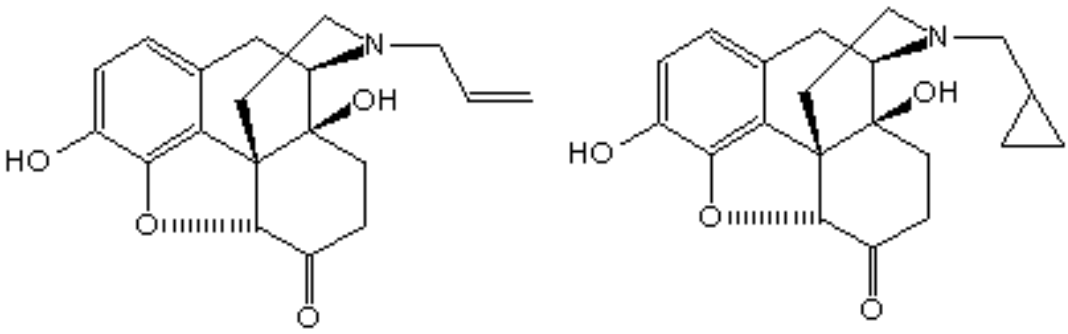
In addition to the serious risks of addiction, abuse, and overdose, the use of prescription opioids can have a number of side effects, even when taken as directed:

• Tolerance—meaning you might need to take more of the medication for the same pain relief • Physical dependence—meaning you have symptoms of withdrawal when the medication is stopped • Increased sensitivity to pain • Constipation • Nausea, vomiting, and dry mouth	• Sleepiness and dizziness • Confusion • Depression • Low levels of testosterone that can result in lower sex drive, energy, and strength • Itching and sweating

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.

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.

Combination Pain Relievers

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.

Hydrocodone is available in combination with other ingredients, and different combination products are prescribed for different uses. Some hydrocodone combination products are used to relieve moderate-to-severe pain. Other hydrocodone combination products are used to relieve cough. Hydrocodone is in a class of medications called opiate (narcotic) analgesics and in a class of medications called antitussives. Hydrocodone relieves pain by changing the way the brain and nervous system respond to pain. Hydrocodone relieves cough by decreasing activity in the part of the brain that causes coughing.

Web Link

A list of different hydrocodone combination products can be found on the link below.

https://medlineplus.gov/druginfo/meds/a601006.html

Summary

- Different types of medicine address different ailments. Anti-inflammatory drugs mediate inflammation, analgesics relieve pain, and antipyretic drugs help lower body temperature associated with fever.
- Over the counter (OTC) pain relievers relieve the minor aches and pains associated with conditions such as headaches, fever, colds, flu, arthritis, toothaches, and menstrual cramps.
- There are basically two types of OTC pain relievers: acetaminophen and non-steroidal anti-inflammatory drugs (NSAIDs).
- Aspirin and other aspirin-like drugs work by blocking the (COX-1 and COX-2) enzymes involved in pain and inflammation.
- Typical prescription pain relief medicines include opioids and non-opioid medications.
- In addition to the serious risks of addiction, abuse, and overdose, the use of prescription opioids can have a number of side effects,

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11.2: Cold and Allergy Medicines

Learning Objective

- Know the cause of the common cold and allergies.
- Know the chemical name, common name, and uses of different cough, cold, and allergy medicines.

The **common cold** is a viral infectious disease of the upper respiratory tract which affects primarily the nose. The symptoms of the common cold are believed to be primarily related to the immune response to the virus. Symptoms include coughing, sore throat, runny nose, and fever which usually resolve in seven to ten days, with some symptoms lasting up to three weeks. Well over 200 viruses are implicated in the cause of the common cold. The most commonly implicated virus is a rhinovirus (30–80%).

No cure for the common cold exists, but the symptoms can be treated. Antibiotics have no effect against viral infections and thus have no effect against the viruses that cause the common cold. Due to their side effects they cause overall harm; however, they are still frequently prescribed. It is the most frequent infectious disease in humans with the average adult contracting two to three colds a year and the average child contracting between six and twelve. These infections have been with humanity since antiquity.

Allergies are a number of conditions caused by hypersensitivity of the immune system to typically harmless substances in the environment. These diseases include hay fever, food allergies, atopic dermatitis, allergic asthma, and anaphylaxis. Symptoms may include red eyes, an itchy rash, sneezing, a runny nose, shortness of breath, or swelling. Food intolerances and food poisoning are separate conditions.

Early exposure to potential allergens may be protective. Treatments for allergies include the avoidance of known allergens and the use of medications such as steroids and antihistamines. **Histamine** is concentrated in mast cells, cells whose function is essentially to release histamine and immunoglobins when tissue damage occurs. They are especially numerous in parts of the body that are injured often, such as the fingers and toes, or which enjoy frequent contact with the environment, such as the mucosa of the lips, nose, etc.

Types of Cold and Allergy Medications

Cold medicines are medications used by people with the common cold, cough, or related conditions. While a wide variety of drugs are marketed as cough suppressants, research shows there is often little or no measurable benefit in reducing cough symptoms. While they have been used by 10% of American children in any given week, they are not recommended in Canada or the United States in children six years or younger because of lack of evidence showing effect and concerns of harm. One version with codeine, guaifenesin, and pseudoephedrine was the 213th most commonly prescribed medication in 2017, in the United States, with more than two million prescriptions.

There are a number of different cough and cold medications, which may be used for various coughing symptoms. The commercially available products may include various combinations of any one or more of the following types of substances listed in the table below.

Table 11.2.1 *Different Cough & Cold Medications.*

Medication	Function	Examples
Mucokinetics, or mucolytics	a class of drugs which aid in the clearance of mucus from the airways, lungs, bronchi, and trachea	carbocisteine, ambroxol, and bromhexine.
Expectorants	substances claimed to make coughing easier while enhancing the production of mucus and phlegm.	acetylcysteine and guaifenesin.
Antitussives, or cough suppressants substances which suppress the coughing itself	substances which suppress the coughing itself	codeine, pholcodine, dextromethorphan, noscapine, and butamirate.

Medication	Function	Examples
Antihistamines	for allergic rhinitis may produce mild sedation and reduce other associated symptoms, like a runny nose and watery eyes	diphenhydramine, chlorpheniramine, brompheniramine, loratadine, and cetirizine.
Decongestants	relieve nasal congestion and sinus infection	ephedrine, phenylephrine, pseudoephedrine, and oxymetazoline.
Fever or pain medication (antipyretic or analgesic)	fever reducer or pain reliever	paracetamol (called "acetaminophen " in the US) and NSAIDs such as aspirin, ibuprofen or naproxen
Syrups	various substances supposed to soften the coughing	honey or supplement syrup

Effectiveness

The efficacy of cough medication is questionable, particularly in children. A 2014 Cochrane review concluded that "There is no good evidence for or against the effectiveness of OTC medicines in acute cough". Some cough medicines may be no more effective than placebos for acute coughs in adults, including coughs related to upper respiratory tract infections. The American College of Chest Physicians emphasizes that cough medicines are not designed to treat whooping cough, a cough that is caused by bacteria and can last for months. No over-the-counter cough medicines have been found to be effective in cases of pneumonia. They are not recommended in those who have COPD, chronic bronchitis, or the common cold. There is not enough evidence to make recommendations for those who have a cough and cancer.

Medications

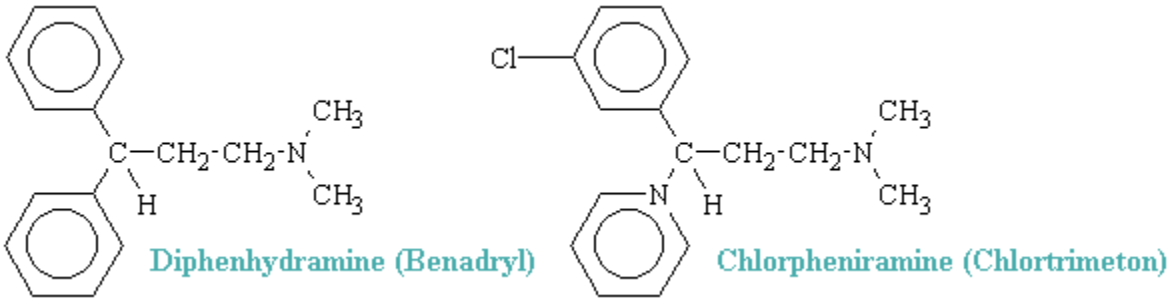
- Dextromethorphan (DXM) may be modestly effective in decreasing cough in adults with viral upper respiratory infections. However, in children it has not been found to be effective.
- Codeine was once viewed as the "gold standard" in cough suppressants, but this position is now questioned. Some recent placebo-controlled trials have found that it may be no better than a placebo for some causes including acute cough in children. It is thus not recommended for children. Additionally, there is no evidence that hydrocodone is useful in children. Similarly, a 2012 Dutch guideline does not recommend its use to treat acute cough.
- A number of other commercially available cough treatments have not been shown to be effective in viral upper respiratory infections. These include for adults: antihistamines, antihistamine-decongestant combinations, benzonatate, anti asthmatic-expectorant-mucolytic combinations, expectorant-bronchodilator combinations, leukotriene inhibitors, ambroxol, and guaifenesin, sometimes with analgesics, antipyretics, anti inflammatories, and anticholinergics - and for children: antihistamines, decongestants for clearing the nose, or combinations of these and leukotriene inhibitors for allergy and asthma. However, antihistamines cannot be used as an empirical therapy in case of chronic, or non specific cough especially in very young children. Long term diphenhydramine use is associated with negative outcomes in older people.

Antihistamines are drugs which treat hay fever and other allergies. Typically, people take antihistamines as an inexpensive, not patented (generic), drug that can be bought without a prescription and relieves from nasal congestion, sneezing, or hives caused by pollen, dust mites, or animal allergy with few side effects. Antihistamines are usually for short-term treatment. Chronic allergies increase the risk of health problems which antihistamines might not treat, including asthma, sinusitis, and lower respiratory tract infection. Consultation of a medical professional is recommended for those who intend to take antihistamines for longer-term use.

Although people typically use the word “antihistamine” to describe drugs for treating allergies, doctors and scientists use the term to describe a class of drug that opposes the activity of histamine receptors in the body. In this sense of the word, antihistamines are subclassified according to the histamine receptor that they act upon. The two largest classes of antihistamines are H_1-antihistamines and H_2-antihistamines.

H_1-antihistamines work by binding to histamine H_1 receptors in mast cells, smooth muscle, and endothelium in the body as well as in the tuberomammillary nucleus in the brain. Antihistamines that target the histamine H_1-receptor are used to treat allergic reactions in the nose (e.g., itching, runny nose, and sneezing). In addition, they may be used to treat insomnia, motion sickness, or vertigo caused by problems with the inner ear. H_2-antihistamines bind to histamine H_2 receptors in the upper gastrointestinal

tract, primarily in the stomach. Antihistamines that target the histamine H_2-receptor are used to treat gastric acid conditions (e.g., peptic ulcers and acid reflux). Two common examples of antihistamines are shown below.



Alternative Medicine

Honey may be a minimally effective cough treatment. A Cochrane review found the evidence to recommend for or against its use in children to be weak. In light of this they found it was better than no treatment, placebo, and diphenhydramine but not better than dextromethorphan for relieving cough symptoms. Honey's use as a cough treatment has been linked on several occasions to infantile botulism and accordingly should not be used in children less than one year old.

Many alternative treatments are used to treat the common cold. A 2007 review states that, "alternative therapies (i.e., Echinacea, vitamin C, and zinc) are not recommended for treating common cold symptoms; however, ... Vitamin C prophylaxis may modestly reduce the duration and severity of the common cold in the general population and may reduce the incidence of the illness in persons exposed to physical and environmental stresses." A 2014 review also found insufficient evidence for Echinacea.

A 2009 review found that the evidence supporting the effectiveness of zinc is mixed with respect to cough, and a 2011 Cochrane review concluded that zinc "administered within 24 hours of onset of symptoms reduces the duration and severity of the common cold in healthy people". A 2003 review concluded: "Clinical trial data support the value of zinc in reducing the duration and severity of symptoms of the common cold when administered within 24 hours of the onset of common cold symptoms." Zinc gel in the nose may lead to long-term or permanent loss of smell. The FDA therefore discourages its use.

Adverse effects

A number of accidental overdoses and well-documented adverse effects suggested caution in children. The FDA in 2015 warned that the use of codeine-containing cough medication in children may cause breathing problems. Cold syrup overdose has been linked to visual and auditory hallucinations, rapid involuntary jaw, tongue and eye movements in children.

Summary

- The **common cold** (also known as nasopharyngitis, rhinopharyngitis, acute coryza, or a cold) is a viral infectious disease of the upper respiratory tract which affects primarily the nose. The symptoms of the common cold are believed to be primarily related to the immune response to the virus.
- An **allergy** is an immune response (with the release of histamines), or reaction, to substances (allergens) that are usually not harmful.
- A number of different cough and cold medications can be used to alleviate various symptoms but not diminish the intensity of the response.

Contributors

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- Wikipedia

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11.3: Drugs and Infectious Diseases

Learning Objectives

- Differentiate between antibiotics and antiviral agents.
- Know the different antibiotics and antiviral agents and their mode of action.

The World Health Organization's (WHO) International Classification of Diseases (ICD) is used in clinical fields to classify diseases and monitor morbidity (the number of cases of a disease) and mortality (the number of deaths due to a disease). In this section, we will introduce terminology used by the ICD (and in health-care professions in general) to describe and categorize various types of disease.

An **infectious disease** is any disease caused by the direct effect of a pathogen. A pathogen may be cellular (bacteria, parasites, and fungi) or acellular (viruses, viroids, and prions). Some infectious diseases are also communicable, meaning they are capable of being spread from person to person through either direct or indirect mechanisms. Some infectious communicable diseases are also considered contagious diseases, meaning they are easily spread from person to person. Not all contagious diseases are equally so; the degree to which a disease is contagious usually depends on how the pathogen is transmitted. For example, measles is a highly contagious viral disease that can be transmitted when an infected person coughs or sneezes and an uninfected person breathes in droplets containing the virus. Gonorrhea is not as contagious as measles because transmission of the pathogen (*Neisseria gonorrhoeae*) requires close intimate contact (usually sexual) between an infected person and an uninfected person.

Antibiotics

The modern era of the chemotherapy of infection started with the clinical use of sulfanilamide in 1936. The "golden age" of antimicrobial therapy began with the production of penicillin in 1941, when this compound was mass-produced and first made available for limited clinical trial. More than 30% of all hospitalized patients now receive one or more courses of therapy with antibiotics, and millions of potentially fatal infections have been cured. However, at the same time, these pharmaceutical agents have become among the most misused of those available to the practicing physician. One result of widespread use of antimicrobial agents has been the emergence of antibiotic-resistant pathogens, which in turn has created an ever-increasing need for new drugs. Many of these agents have also contributed significantly to the rising costs of medical care.

An antibiotic is any substance produced by a microorganism that is excreted to harm or kill another microorganism. Technically, antibiotics are microbial or fungal products. But these substances can be synthesized and mass produced in the laboratory to use against harmful microorganisms in the environment. Thus, the synthetic chemist has added greatly to our therapeutic armamentarium. Synthetic drugs such as isonaizid and theambutol represent important contributions for the treatment of tuberculosis. While many such antimicrobial agents are not properly termed antibiotics, since they are not produced by living organisms, little distinction should now be made between compounds of natural and synthetic origin.

Synthetic Antibiotics - Sulfonamides

Sulfonamide is the basis of several groups of drugs. The original antibacterial sulfonamides (sometimes called sulfa drugs or sulpha drugs) are synthetic antimicrobial agents that contain the sulfonamide group. Some sulfonamides are also devoid of antibacterial activity, e.g., the anticonvulsant sultiame. The sulfonylureas and thiazide diuretics are newer drug groups based on the antibacterial sulfonamides.

Sulfa allergies are common, and medications containing sulfonamides are prescribed carefully. It is important to make a distinction between sulfa drugs and other sulfur-containing drugs and additives, such as sulfates and sulfites, which are chemically unrelated to the sulfonamide group and do not cause the same hypersensitivity reactions seen in the sulfonamides.

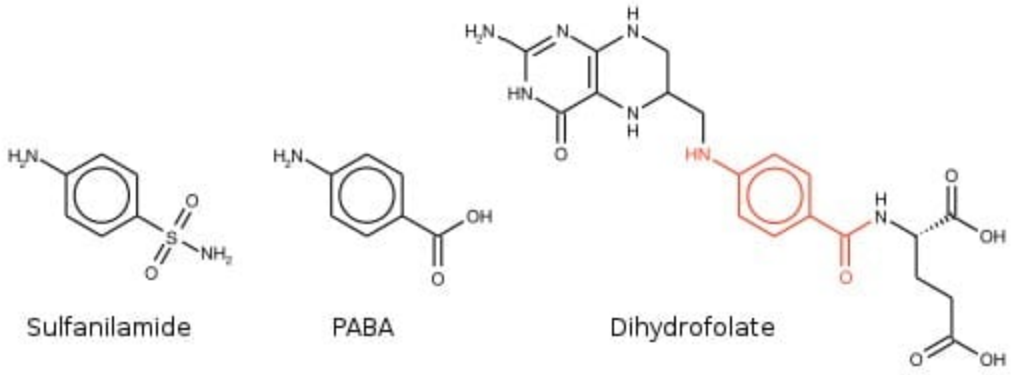


Figure 11.3.1 Structural similarity between sulfanilamide, a common sulfa drug (left) and PABA (center) is the basis for the inhibitory activity of sulfa drugs on dihydrofolate biosynthesis. Source: Wiikipedia

Sulfonamides are structural analogs and competitive antagonists of para-aminobenzoic acid (PABA), and thus prevent normal bacterial utilization of PABA for the synthesis of the vitamin folic acid (Figure 11.3.1). More specifically, sulfonamides are competitive inhibitors of the bacterial enzyme sulfihydropteroate synthase, which is responsible for the conversion of PABA into dihydrofolic acid, the immediate precursor of folic acid. Sensitive microorganisms are those that must synthesis their own folic acid; bacteria that can utilize preformed folic acid are not affected.

Cell Wall Synthesis Inhibitors - Penicillin and Cepalosphorins

The penicillins were the first antibiotics discovered as natural products from the mold Penicillium and was effective against Gram-positive bacteria. All penicillin derivatives produce their bacteriocidal effects by inhibition of bacterial cell wall synthesis. Specifically, the cross linking of peptides on the mucosaccharide chains is prevented. If cell walls are improperly made cell walls allow water to flow into the cell causing it to burst. A number of natural penicillins have been discovered, but only two purified compounds are in clinical use: penicillin G (intravenous use) and penicillin V (given by mouth). Several semisynthetic penicillins are effective against a broader spectrum of bacteria: these include the antistaphylococcal penicillins, aminopenicillins and the antipseudomonal penicillins.

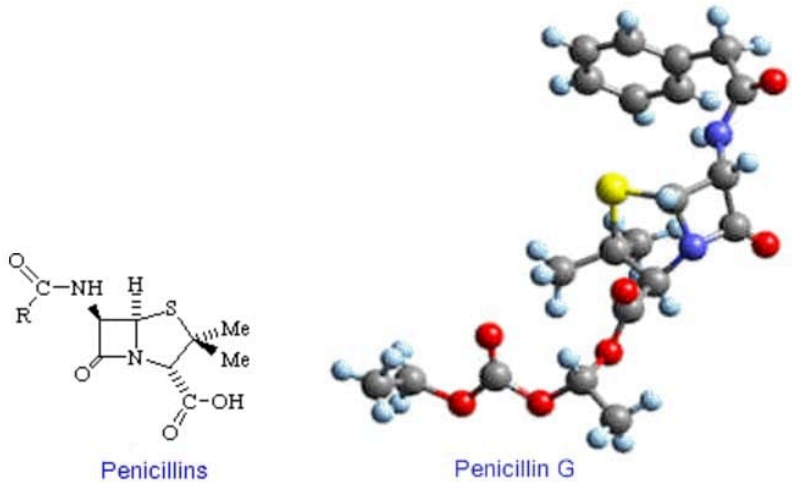


Figure 11.3.2 Penicillins.

Like penicillin, cephalosporins are valuable because of their low toxicity and their broad spectrum of action against various diseases. In this way, cephalosporin is very similar to penicillin. Cephalosporins are one of the most widely used antibiotics, and economically speaking, has about 29% of the antibiotic market. The cephalosporins are possibly the single most important group of antibiotics today and are equal in importance to penicillin. The structure and mode of action of the cephalosporins are similar

to that of penicillin. They affect bacterial growth by inhibiting cell wall synthesis, in Gram-positive and -negative bacteria.

Figure 11.3.3 Cephalosporins.

Protein Synthesis Inhibitors - Erythromycin, Tetracycline, Streptomycin, and Chloramphenicol

Erythromycin is an orally effective antibiotic discovered in 1952 in the metabolic products of a strain of Streptocyces erythreus, originally obtained from a soil sample collected in the Philippine Archipelago. Erythromycin may be either bacteriostatic or bactericidal, depending on the microorganism and the concentration of the drug.

Erythromycin is an antibiotic used for the treatment of a number of bacterial infections. This includes respiratory tract infections, skin infections, chlamydia infections, pelvic inflammatory disease, and syphilis. It may also be used during pregnancy to prevent Group B streptococcal infection in the newborn, as well as to improve delayed stomach emptying. It can be given intravenously and by mouth. An eye ointment is routinely recommended after delivery to prevent eye infections in the newborn.

Erythromycin and other macrolide antibiotics inhibit protein synthesis by binding to 50 S ribosomal subunits of sensitive microorganisms. (Humans do not have 50 S ribosomal subunits, but have ribosomes composed of 40 S and 60 S subunits). Certain resistant microorganisms with mutational changes in components of this subunit of the ribosome fail to bind the drug.

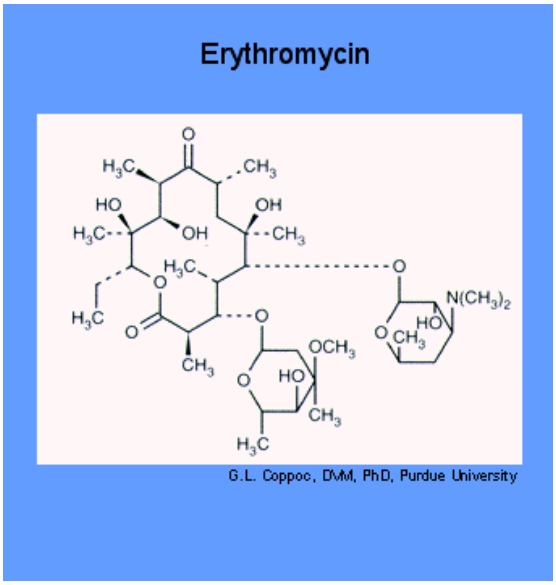


Figure 11.3.4 Erythromycin.

Tetracyclines have the broadest spectrum of antimicrobial activity. These may include: Aureomycin, Terramycin, and Panmycin. Four fused 6-membered rings, as shown in the figure below, form the basic structure from which the various tetracyclines are made. The various derivatives are different at one or more of four sites on the rigid, planar ring structure. The classical tetracyclines were derived from *Streptomyces spp.*, but the newer derivatives are semisynthetic as is generally true for newer members of other drug groups.

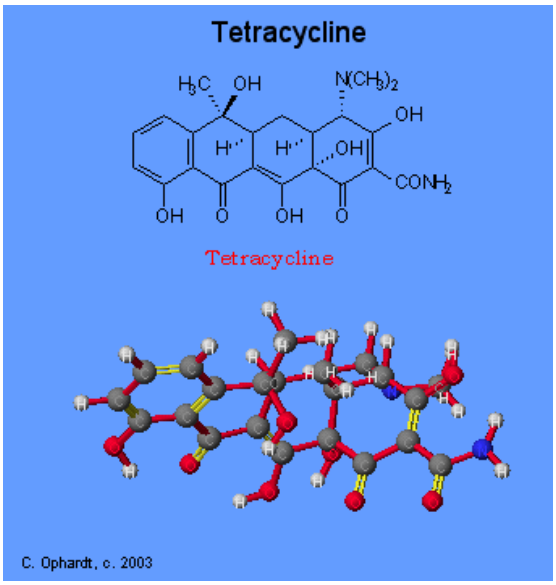


Figure 11.3.5 Tetracycline.

Tetracyclines inhibit bacterial protein synthesis by blocking the attachment of the transfer RNA-amino acid to the ribosome. More precisely they are inhibitors of the codon-anticodon interaction. Tetracyclines can also inhibit protein synthesis in the host, but are less likely to reach the concentration required because eukaryotic cells do not have a tetracycline uptake mechanism.

Streptomycin

Streptomycin is effective against gram-negative bacteria, although it is also used in the treatment of tuberculosis. Streptomycin binds to the 30S ribosome and changes its shape so that it and inhibits protein synthesis by causing a misreading of messenger RNA information.

Chloramphenicol

Chloromycetin is also a broad spectrum antibiotic that possesses activity similar to the tetracylines. At present, it is the only antibiotic prepared synthetically. It is reserved for treatment of serious infections because it is potentially highly toxic to bone marrow cells. It inhibits protein synthesis by attaching to the ribosome and interferes with the formation of peptide bonds between amino acids. It behaves as an antimetabolite for the essential amino acid phenylalanine at ribosomal binding sites.

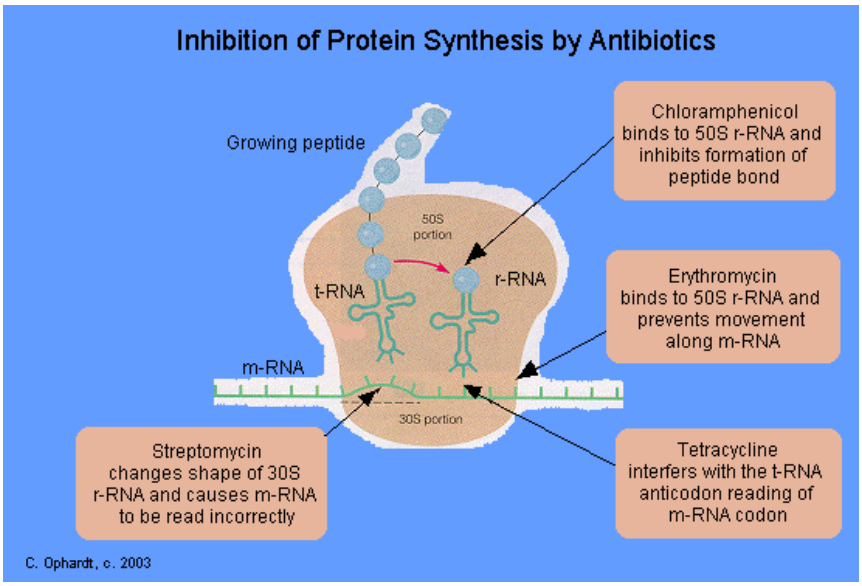


Figure 11.3.6 *Inhibition of protein synthesis by erythtomycin, teracycline, streptomycin, and chloramphenicol.*

A summary of the key features of different antibiotics is given in Table 11.3.1.

Table 11.3.1 *Some Clinically Important Antibiotics.*

Antibiotic	Producer organism	Activity	Site or mode of action
Penicillin	*Penicillium chrysogenum*	Gram-positive bacteria	Wall synthesis
Cephalosporin	*Cephalosporium acremonium*	Broad spectrum	Wall synthesis
Griseofulvin	*Penicillium griseofulvum*	Dermatophytic fungi	Microtubules
Bacitracin	*Bacillus subtilis*	Gram-positive bacteria	Wall synthesis
Polymyxin B	*Bacillus polymyxa*	Gram-negative bacteria	Cell membrane
Amphotericin B	*Streptomyces nodosus*	Fungi	Cell membrane
Erythromycin	*Streptomyces erythreus*	Gram-positive bacteria	Protein synthesis
Neomycin	*Streptomyces fradiae*	Broad spectrum	Protein synthesis
Streptomycin	*Streptomyces griseus*	Gram-negative bacteria	Protein synthesis
Tetracycline	*Streptomyces rimosus*	Broad spectrum	Protein synthesis
Vancomycin	*Streptomyces orientalis*	Gram-positive bacteria	Protein synthesis
Gentamicin	*Micromonospora purpurea*	Broad spectrum	Protein synthesis
Rifamycin	*Streptomyces mediterranei*	Tuberculosis	Protein synthesis

History of Penicillin and General Information

Alexander Fleming loved to play, both in the laboratory and out. He always loved snooker and golf and had many whimsical variants on the rules. In the lab he made "germ paintings," in which he would draw with his culture loop using spores of highly pigmented bacteria, which were invisible when he made the painting, but when cultured developed into brightly colored scenes. He followed what Max Delbruck would later call the "principle of limited sloppiness." Fleming abhorred a tidy, meticulous lab; he left culture dishes lying around for weeks and would often discover interesting things in them. Though the story has been told in many sometimes conflicting ways, something like this resulted in the discovery of penicillin. He seems to have left a culture dish lying on the lab bench and then gone away on vacation. When he returned a few spores of an unusual mold had germinated on the plate. When he cultured the bacteria on the plate he found that they grew up to within a few centimeters of the mold, but there were killed. A crude extract of the mold was then shown to have antibacterial properties. Fleming made this discovery in 1928 and by 1929 had named it penicillin (he was told by a colleague that the mold was a type of Penicillium and "penicillozyme" must have seemed cumbersome).

Fleming continued to use penicillin in his lab but not with any great enthusiasm and certainly not to the exclusion of many other projects. He never developed it into a clinically useful compound, though in 1929 he suggested that it might have important clinical applications. Because he was a bacteriologist and not a chemist, Fleming did not attempt to purify penicillin. He seems to have run into a dead end with penicillin and so during the 1930s, though he kept it in his lab, he did not do much with it. In the late 1930s Australian Howard Florey came to London to work with Charles Sherrington. He worked on lysozyme for a while and then became interested in penicillin. It was Florey, with Chain and other of his group that developed penicillin into a clinical antibiotic. They did this during 1940-41. Fleming, Florey, and Chain shared the 1945 Nobel Prize in Physiology for Medicine.

Fleming became world-famous for penicillin, and was rightly acknowledged as the father of modern antibiotics, but Florey was just as rightly miffed at being denied much of the credit for creating the powerful medical tool we now know. Evidence does not suggest that Fleming deliberately denied Florey his due credit, but Fleming's peculiar, dry sense of humor seems to have caused him not to deny even the wildest attributions to him.

Viruses and Antiviral Drugs

Viruses cause a variety of diseases in animals, including humans, ranging from the common cold to potentially fatal illnesses like meningitis (FIgure 11.3.1). These diseases can be treated by antiviral drugs or by vaccines, but some viruses, such as HIV, are capable of both avoiding the immune response and mutating to become resistant to antiviral drugs.

Antiviral drugs are a class of medication used for treating viral infections. Most antivirals target specific viruses, while a broad-spectrum antiviral is effective against a wide range of viruses. Unlike most antibiotics, antiviral drugs do not destroy their target pathogen; instead they inhibit its development.

Antiviral drugs are one class of antimicrobials, a larger group which also includes antibiotic (also termed antibacterial), antifungal and antiparasitic drugs, or antiviral drugs based on monoclonal antibodies. Most antivirals are considered relatively to the host, and therefore can be used to treat infections. They should be distinguished from viricides, which are not medication but deactivate or destroy virus particles, either inside or outside the body. Natural viricides are produced by some plants such as eucalyptus and Australian tea trees.

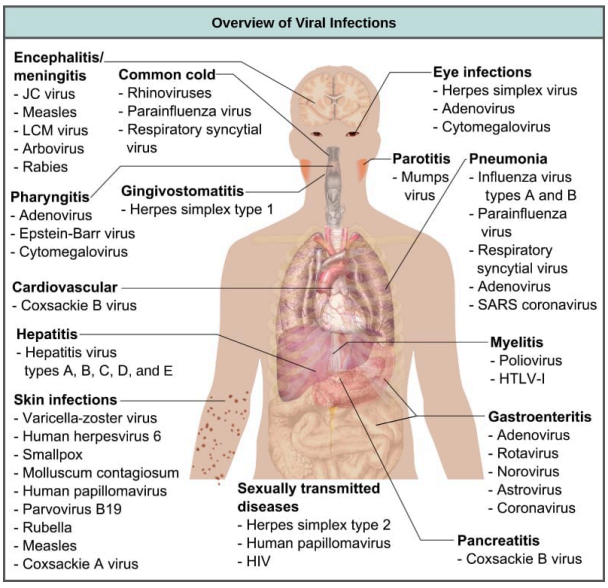


Figure 11.3.7: Viruses can cause dozens of ailments in humans, ranging from mild illnesses to serious diseases. (credit: modification of work by Mikael Häggström)

DNA Viruses and RNA Viruses

Viruses are visible only under an electron microscope. They come in a variety of shapes, ranging from spherical to rod shaped. The fact that they contain either deoxyribonucleic acid (DNA) or ribonucleic acid (RNA)—*but never both*—allows them to be divided into two major classes: DNA viruses and RNA viruses (Figure 11.3.1).

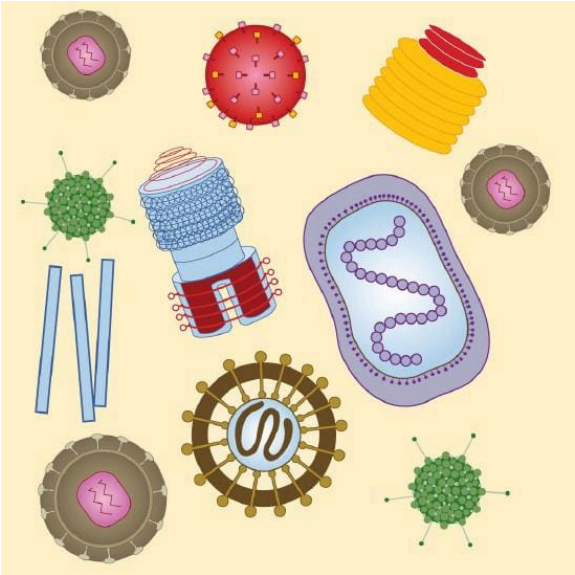


Figure 11.3.8: Viruses. Viruses come in a variety of shapes that are determined by their protein coats.

Most *RNA viruses* use their nucleic acids in much the same way as the DNA viruses, penetrating a host cell and inducing it to replicate the viral RNA and synthesize viral proteins. The new RNA strands and viral proteins are then assembled into new viruses. Some RNA viruses, however, called retroviruses (Figure 11.3.2), synthesize DNA in the host cell, in a process that is the reverse of the DNA-to-RNA transcription that normally occurs in cells. The synthesis of DNA from an RNA template is catalyzed by the enzyme reverse transcriptase.

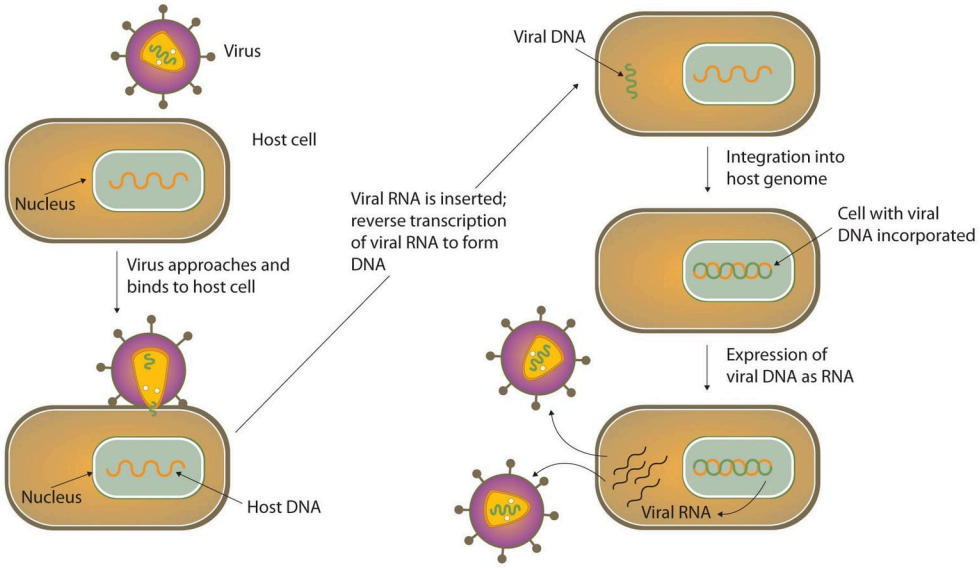


Figure 11.3.9: Life Cycle of a Retrovirus

Anti-viral Drugs

Antiviral drugs often have limited success in curing viral disease, but in many cases, they have been used to control and reduce symptoms for a wide variety of viral diseases. For most viruses, these drugs can inhibit the virus by blocking the actions of one or more of its proteins. It is important that the targeted proteins be encoded by viral genes and that these molecules are not present in a healthy host cell. In this way, viral growth is inhibited without damaging the host. There are large numbers of antiviral drugs available to treat infections, some specific for a particular virus and others that can affect multiple viruses.

Antivirals have been developed to treat genital herpes (herpes simplex II) and influenza. For genital herpes, drugs such as acyclovir can reduce the number and duration of episodes of active viral disease, during which patients develop viral lesions

in their skin cells. As the virus remains latent in nervous tissue of the body for life, this drug is not curative but can make the symptoms of the disease more manageable. For influenza, drugs like Tamiflu (oseltamivir) (Figure 11.3.3) can reduce the duration of “flu” symptoms by 1 or 2 days, but the drug does not prevent symptoms entirely. Tamiflu works by inhibiting an enzyme (viral neuraminidase) that allows new virions to leave their infected cells. Thus, Tamiflu inhibits the spread of virus from infected to uninfected cells. Other antiviral drugs, such as Ribavirin, have been used to treat a variety of viral infections, although its mechanism of action against certain viruses remains unclear.

By far, the most successful use of antivirals has been in the treatment of the retrovirus HIV, which causes a disease that, if untreated, is usually fatal within 10–12 years after infection. Anti-HIV drugs have been able to control viral replication to the point that individuals receiving these drugs survive for a significantly longer time than the untreated.

Anti-HIV drugs inhibit viral replication at many different phases of the HIV replicative cycle (Figure 11.3.4). Drugs have been developed that inhibit the fusion of the HIV viral envelope with the plasma membrane of the host cell (fusion inhibitors), the conversion of its RNA genome into double-stranded DNA (reverse transcriptase inhibitors), the integration of the viral DNA into the host genome (integrase inhibitors), and the processing of viral proteins (protease inhibitors).

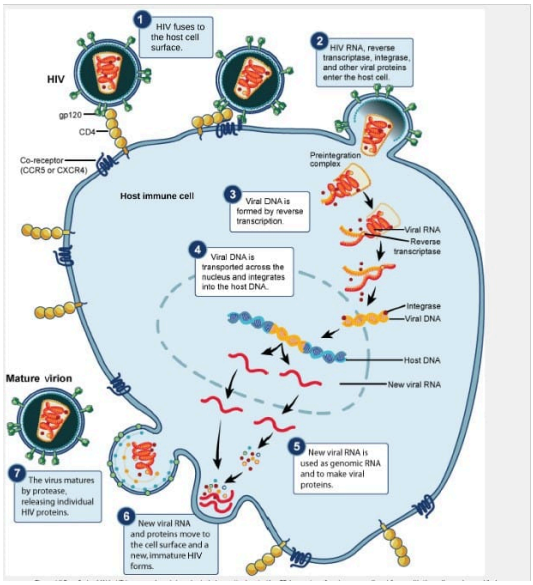
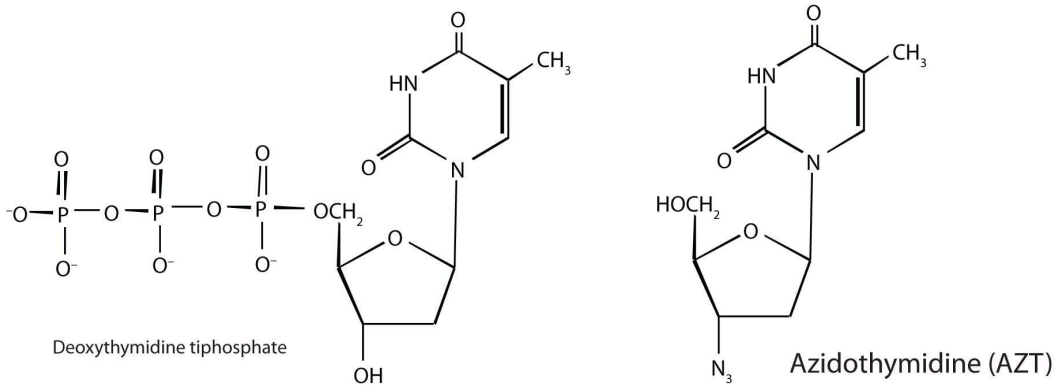


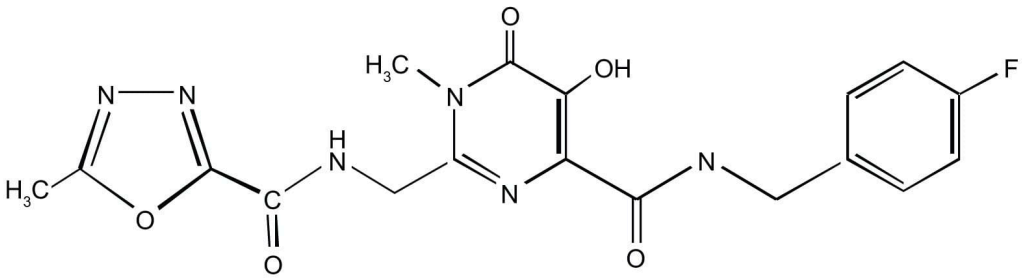
Figure 11.3.10: HIV, an enveloped, icosahedral virus, attaches to the CD4 receptor of an immune cell and fuses with the cell membrane. Viral contents are released into the cell, where viral enzymes convert the single-stranded RNA genome into DNA and incorporate it into the host genome. (credit: NIAID, NIH)

In 1987, azidothymidine (AZT, also known as zidovudine or the brand name Retrovir) became the first drug approved for the treatment of AIDS. It works by binding to reverse transcriptase in place of deoxythymidine triphosphate, after which, because AZT does not have a 3′OH group, further replication is blocked. In the past 10 years, several other drugs have been approved that also act by inhibiting the viral reverse transcriptase.



Raltegravir (Isentress) is a newer anti-AIDS drug that was approved by the FDA in October 2007. This drug inhibits the integrase

enzyme that is needed to integrate the HIV DNA into cellular DNA, an essential step in the production of more HIV particles.



Raltegravir

A major problem in treating HIV infections is that the virus can become resistant to any of these drugs. One way to combat the problem has been to administer a “cocktail” of drugs, typically a combination of two reverse transcriptase inhibitors along with a protease inhibitor. These treatments can significantly reduce the amount of HIV in an infected person.

The **management of HIV/AIDS** normally includes the use of multiple antiretroviral drugs in an attempt to control HIV infection. There are several classes of antiretroviral agents that act on different stages of the HIV life-cycle. The use of multiple drugs that act on different viral targets is known as **highly active antiretroviral therapy** (**HAART**). HAART decreases the patient's total burden of HIV, maintains function of the immune system, and prevents opportunistic infections that often lead to death. HAART also prevents the transmission of HIV between serodiscordant same sex and opposite sex partners so long as the HIV-positive partner maintains an undetectable viral load.

A more detailed list of antivirals other than HIV is given in Table 11.3.1.

Table 11.3.1 *Antivirals Used for Viruses Other Than HIV*

Antiviral	Brand Name	Use
amantadine	*Symmetrel*	used prophylactically against influenza A) in high-risk individuals. It prevents influenza A viruses from the uncoating step necessary for viral replication.
rimantidine	*Flumadine*	used for treatment and prophylaxis of influenza A. It prevents influenza A viruses from the uncoating step necessary for viral replication.
zanamivir:	*Relenza*	used to limit the duration of influenza A and B infections. It is an inhibitor of the influenza virus surface enzyme called neuraminidase that is needed for release of newly formed influenza viruses from the infected cell.
oseltamivir	*Tamiflu*	used limit the duration of influenza infections. It is an inhibitor of the influenza virus surface enzyme called neuraminidase that is needed for release of newly formed influenza viruses from the infected cell.
acyclovir	Zovirax	used against herpes simplex viruses (HSV) to treat genital herpes, mucocutaneous herpes in the immunosuppressed, HSV encephalitis, neonatal herpes, and to reduce the rate of recurrences of genital herpes. It is also used against varicella zoster viruses (VZV)) to treat shingles. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
trifluridine	*Viroptic*	used to treat eye infection (keratitis and conjunctivitis) caused by HSV. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
famciclovir	Famvir	used to treat HSV and VZV infections. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
valacyclovir	Valtrex	used to treat HSV and VZV infections. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.

Antiviral	Brand Name	Use
penciclovir	Denavir	used in treating HSV infections. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
gancyclovir	Cytovene; Vitrasert	used in treating severe cytomegalovirus (CMV) infections such as retinitis. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
valganciclovir	Valcyte	used in treating severe CMV infections such as retinitis). It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
foscarnet	Foscavir	used in treating severe CMV infections such as retinitis. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
cidofovir	Vistide	used in treating CMV retinitis. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
fomivirsen	Vitravene	used in treating CMV retinitis. Fomivirsen inhibits cytomegalovirus (CMV) replication through an antisense RNA (microRNA or miRNA mechanism. The nucleotide sequence of fomivirsen is complementary to a sequence in mRNA transcripts (Figure 11.3.1) that encodes several proteins responsible for regulation of viral gene expression that are essential for production of infectious CMV. Binding of fomivirsen to the target mRNA results in inhibition of protein synthesis, subsequently inhibiting virus replication.
ribavirin	Copegus; Rebetol; Virazole	used in treating severe acute respiratory syndrome (SARS). In combination with other drugs it is used to treat hepatitis C virus (HCV). It chemically resembles a normal RNA nucleoside. Once inserted into the growing RNA chain it inhibits further viral RNA replication.
telaprevir	Incivek	for the treatment of chronic hepatitis C (hepatitis C virus or HCV genotype 1). It is a protease inhibitor that binds to the active site of an HCV-encoded protease and prevent it from cleaving the long polyprotein from polycistronic HCV genes into proteins essential to the structure and function of HCV.
boceprevir	Victrelis	for the treatment of chronic hepatitis C (hepatitis C virus or HCV genotype 1) infection. It is used in combination with peginterferon alfa and ribavirin. Boceprevir is a protease inhibitor that binds to the active site of an HCV-encoded protease and prevent it from cleaving the long polyprotein from polycistronic HCV genes into proteins essential to the structure and function of HCV.
simeprevir	Olysio	use for the treatment of chronic hepatitis C (hepatitis C virus or HCV genotype 1) infection. Used in combination with peginterferon alfa and ribavirin. Simeprevir is a protease inhibitor that binds to the active site of an HCV-encoded protease and prevent it from cleaving the long polyprotein from polycistronic HCV genes into proteins essential to the structure and function of HCV.
sofosbuvir	*Sovaldi*	Use for the treatment of chronic hepatitis C infection. Used in combination with ribavirin for hepatitis C virus or HCV genotypes 2 and 4; used in combination with peginterferon alfa and ribavirin for HCV genotypes 1 and 4. The second indication is the first approval of an interferon-free regimen for the treatment of chronic HCV infection. Sofosbuvir is a nucleotide polymerase inhibitor that binds to the active site of an HCV-encoded RNA polymerase preventing the synthesis of the viral RNA genome.
lamivudine	*Epivir-HBV*	used in treating chronic hepatitis B. It chemically resembles a normal DNA nucleoside. Once inserted into the growing DNA chain it inhibits further viral DNA replication.
adefovir dipivoxil	*Hepsera*	used in treating hepatitis B.

Web Link

A list of FDA-approved Antiretroviral drugs (last updated on April 12, 2018) used in the treatment of HIV infection can be found on the website: www.fda.gov/patients/hiv-treatment/antiretroviral-drugs-used-treatment-hiv-infection

Basic Research and Drug Development

The general idea behind modern antiviral drug design is to identify viral proteins, or parts of proteins, that can be disabled. These "targets" should generally be as unlike any proteins or parts of proteins in humans as possible, to reduce the likelihood of side effects. The targets should also be common across many strains of a virus, or even among different species of virus in the same family, so a single drug will have broad effectiveness. For example, a researcher might target a critical enzyme synthesized by the virus, but not by the patient, that is common across strains, and see what can be done to interfere with its operation.

Once targets are identified, candidate drugs can be selected, either from drugs already known to have appropriate effects or by actually designing the candidate at the molecular level with a computer-aided design program.

The target proteins can be manufactured in the lab for testing with candidate treatments by inserting the gene that synthesizes the target protein into bacteria or other kinds of cells. The cells are then cultured for mass production of the protein, which can then be exposed to various treatment candidates and evaluated with "rapid screening" technologies.

Prevention of Viral Diseases with Vaccination

While we do have limited numbers of effective antiviral drugs, such as those used to treat HIV and influenza, the primary method of controlling viral disease is by vaccination, which is intended to prevent outbreaks by building immunity to a virus or virus family (Figure 11.3.1). Vaccines may be prepared using live viruses, killed viruses, or molecular subunits of the virus. The killed viral vaccines and subunit viruses are both incapable of causing disease.

Live viral vaccines are designed in the laboratory to cause few symptoms in recipients while giving them protective immunity against future infections. Polio was one disease that represented a milestone in the use of vaccines. Mass immunization campaigns in the 1950s (killed vaccine) and 1960s (live vaccine) significantly reduced the incidence of the disease, which caused muscle paralysis in children and generated a great amount of fear in the general population when regional epidemics occurred. The success of the polio vaccine paved the way for the routine dispensation of childhood vaccines against measles, mumps, rubella, chickenpox, and other diseases.

The danger of using live vaccines, which are usually more effective than killed vaccines, is the low but significant danger that these viruses will revert to their disease-causing form by back mutations. Live vaccines are usually made by attenuating (weakening) the “wild-type” (disease-causing) virus by growing it in the laboratory in tissues or at temperatures different from what the virus is accustomed to in the host. Adaptations to these new cells or temperatures induce mutations in the genomes of the virus, allowing it to grow better in the laboratory while inhibiting its ability to cause disease when reintroduced into conditions found in the host. These attenuated viruses thus still cause infection, but they do not grow very well, allowing the immune response to develop in time to prevent major disease. Back mutations occur when the vaccine undergoes mutations in the host such that it readapts to the host and can again cause disease, which can then be spread to other humans in an epidemic. This type of scenario happened as recently as 2007 in Nigeria where mutations in a polio vaccine led to an epidemic of polio in that country.

Some vaccines are in continuous development because certain viruses, such as influenza and HIV, have a high mutation rate compared to other viruses and normal host cells. With influenza, mutations in the surface molecules of the virus help the organism evade the protective immunity that may have been obtained in a previous influenza season, making it necessary for individuals to get vaccinated every year. Other viruses, such as those that cause the childhood diseases measles, mumps, and rubella, mutate so infrequently that the same vaccine is used year after year.



Figure 11.3.11: Vaccinations are designed to boost immunity to a virus to prevent infection. (credit: USACE Europe District)

Summary

- An infectious disease is any disease caused by the direct effect of a pathogen (bacteria, parasites, fungi, viruses, viroids, and prions).
- Diseases can either be noninfectious (due to genetics and environment) or infectious (due to pathogens). Some infectious diseases are communicable (transmissible between individuals) or contagious (easily transmissible between individuals); others are noncommunicable, but may be contracted via contact with environmental reservoirs or animal.
- An antibiotic is any substance produced by a microorganism that is excreted to harm or kill another microorganism. Technically, antibiotics are microbial or fungal products.
- Most antibiotics inhibit bacterial growth by inhibiting cell wall synthesis or protein synthesis.
- **Antiviral drugs** are a class of medication used for treating viral infections. Most antivirals target specific viruses, while a broad-spectrum antiviral is effective against a wide range of viruses.
- The management of HIV/AIDS normally includes the use of multiple in an attempt to control .

- The primary method of controlling viral disease is by vaccination, which is intended to prevent outbreaks by building immunity to a virus or virus family

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- US FDA
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11.4: Chemicals Against Cancer

Learning Objective

- List the different types of cancer drugs and their mechanism of action.

Chemotherapy is a type of cancer treatment that uses one or more anti-cancer drugs (chemotherapeutic agents) as part of a standardized chemotherapy regimen. Chemotherapy may be given with a curative intent (which almost always involves combinations of drugs), or it may aim to prolong life or to reduce symptoms (palliative chemotherapy). Chemotherapy is one of the major categories of the medical discipline specifically devoted to pharmacotherapy for cancer, which is called *medical oncology*.

When fighting cancer, the entire population of neoplastic cells must be eradicated in order to obtain desired results. The concept of "total cell-kill" applies to chemotherapy as it does to other means of treatment: total excision of the tumor is necessary for surgical care, and complete destruction of all cancer cells is required for a cure with radiation therapy.

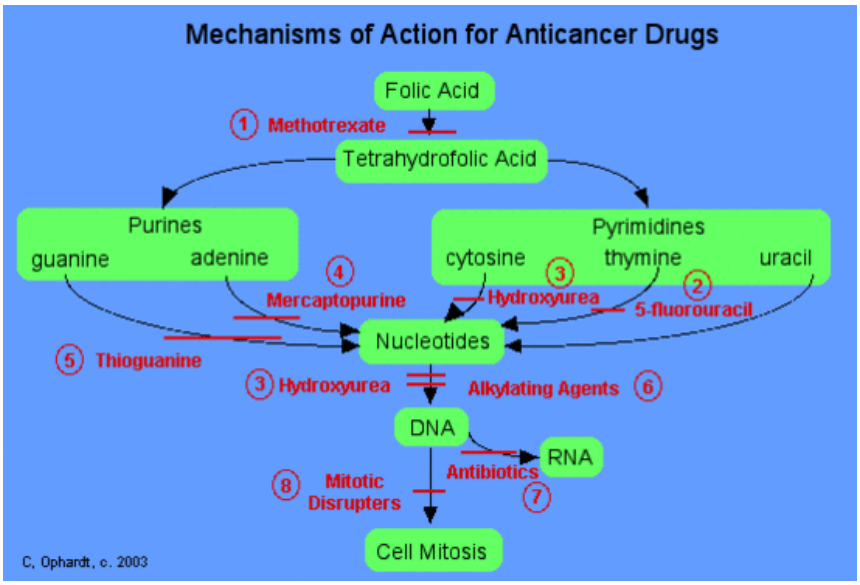
The available anticancer drugs have distinct mechanisms of action which may vary in their effects on different types of normal and cancer cells. A single "cure" for cancer has proved elusive since there is not a single type of cancer but as many as 100 different types of cancer. In addition, there are very few demonstrable biochemical differences between cancerous cells and normal cells. For this reason the effectiveness of many anticancer drugs is limited by their toxicity to normal rapidly growing cells in the intestinal and bone marrow areas. A final problem is that cancerous cells which are initially suppressed by a specific drug may develop a resistance to that drug. For this reason cancer chemotherapy may consist of using several drugs in combination for varying lengths of time.

Introduction

Chemotherapy drugs, are sometimes feared because of a patient's concern about toxic effects. Their role is to slow and hopefully halt the growth and spread of a cancer. There are three goals associated with the use of the most commonly-used anticancer agents.

1. Damage the DNA of the affected cancer cells.
2. Inhibit the synthesis of new DNA strands to stop the cell from replicating, because the replication of the cell is what allows the tumor to grow.
3. Stop mitosis or the actual splitting of the original cell into two new cells. Stopping mitosis stops cell division (replication) of the cancer and may ultimately halt the progression of the cancer.

Unfortunately, the majority of drugs currently on the market are not specific, which leads to the many common side effects associated with cancer chemotherapy. Because the common approach of all chemotherapy is to decrease the growth rate (cell division) of the cancer cells, the side effects are seen in bodily systems that naturally have a rapid turnover of cells including skin, hair, gastrointestinal, and bone marrow. These healthy, normal cells, also end up damaged by the chemotherapy program.



Categories of Chemotherapy Drugs

In general, chemotherapy agents can be divided into three main categories based on their mechanism of action as shown in the table below.

Table 11.4.1 *Chemotherapy Drugs and Their Mechanism of Action*

Category	Specific Mode of Action	Examples
Stop the synthesis of pre DNA molecule building blocks "Antimetabolites"	These agents work in a number of different ways. DNA building blocks are folic acid, heterocyclic bases, and nucleotides, which are made naturally within cells. All of these agents work to block some step in the formation of nucleotides or deoxyribonucleotides (necessary for making DNA). When these steps are blocked, the nucleotides, which arethe building blocks of DNA and RNA, can not be synthesized. Thus the cells can not replicate because they cannot make DNA without the nucleotides.	1) methotrexate (Abitrexate®),2) fluorouracil (Adrucil®), 3) hydroxyurea (Hydrea®), 4) mercaptopurine (Purinethol®) and 5) thioguanine.
Directly damage the DNA in the nucleus of the cell "Alkylating agents, antibiotics, topoisomerase inhibitors and intercalating agents."	These agents chemically damage DNA and RNA. They disrupt replication of the DNA and either totally halt replication or cause the manufacture of nonsense DNA or RNA (i.e. the new DNA or RNA does not code for anything useful).	5) cisplatin (Platinol®) and 7) antibiotics - daunorubicin (Cerubidine®), doxorubicin (Adriamycin®), and etoposide (VePesid®).

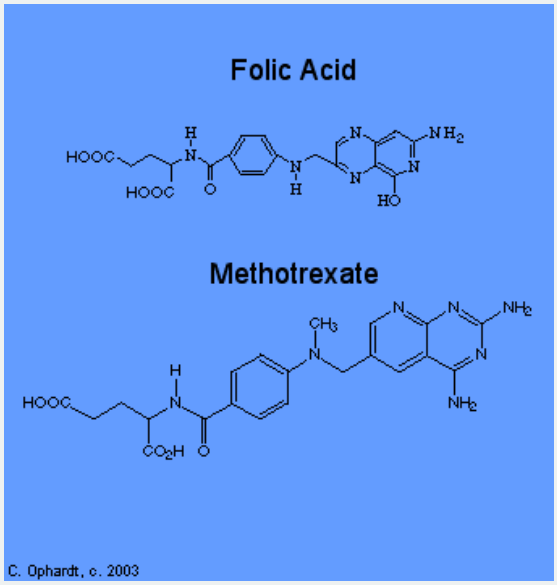
Category	Specific Mode of Action	Examples
Effect the synthesis or breakdown of the mitotic spindles "Mitotic disrupters"	Mitotic spindles serve as molecular railroads with "North and South Poles" in the cell when a cell starts to divide itself into two new cells. These spindles are very important because they help to split the newly copied DNA such that a copy goes to each of the two new cells during cell division. These drugs disrupt the formation of these spindles and therefore interrupt cell division.	8) mitotic disrupters include: Vinblastine (Velban®), Vincristine (Oncovin®) and Pacitaxel (Taxol®).

Antimetabolites

Methotrexate

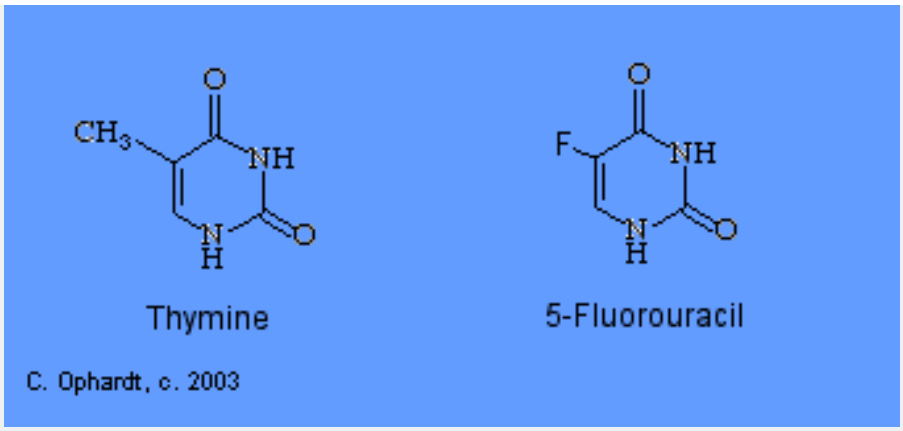
Methotrexate inhibits folic acid reductase which is responsible for the conversion of folic acid to tetrahydrofolic acid. At two stages in the biosynthesis of purines(adenine and guanine) and at one stage in the synthesis of pyrimidines (thymine, cytosine, and uracil), one-carbon transfer reactions occur which require specific coenzymes synthesized in the cell from tetrahydrofolic acid.

Tetrahydrofolic acid itself is synthesized in the cell from folic acid with the help of an enzyme, folic acid reductase. Methotrexate looks a lot like folic acid to the enzyme, so it binds to it thinking that it is folic acid. In fact, methotrexate looks so good to the enzyme that it binds to it quite strongly and inhibits the enzyme. Thus, DNA synthesis cannot proceed because the coenzymes needed for one-carbon transfer reactions are not produced from tetrahydrofolic acid because there is no tetrahydrofolic acid. Again, without DNA, no cell division.



5-Fluorouracil

5-Fluorouracil (5-FU; Adrucil®, Fluorouracil, Efudex®, Fluoroplex®) is an effective pyrimidine antimetabolite. Fluorouracil is synthesized into the nucleotide, 5-fluoro-2-deoxyuridine. This product acts as an antimetabolite by inhibiting the synthesis of 2-deoxythymidine because the carbon - fluorine bond is extremely stable and prevents the addition of a methyl group in the 5-position. The failure to synthesize the thymidine nucleotide results in little or no production of DNA. Two other similar drugs include: gemcitabine (Gemzar®) and arabinosylcytosine (araC). They all work through similar mechanisms.

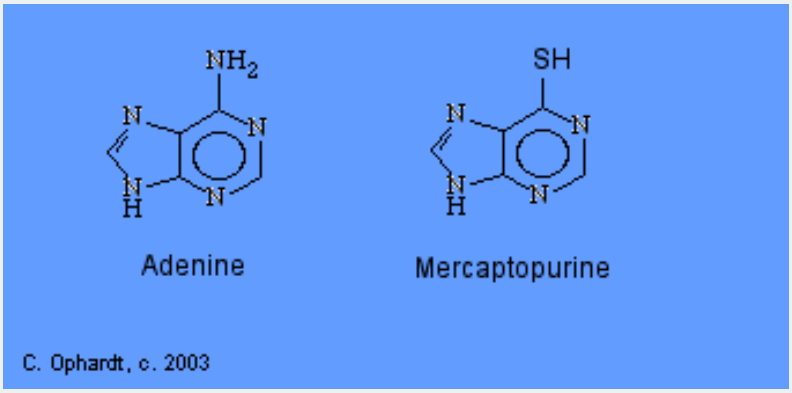


Hydroxyurea

Hydroxyurea blocks an enzyme which converts the cytosine nucleotide into the deoxy derivative. In addition, DNA synthesis is further inhibited because hydroxyurea blocks the incorporation of the thymidine nucleotide into the DNA strand.

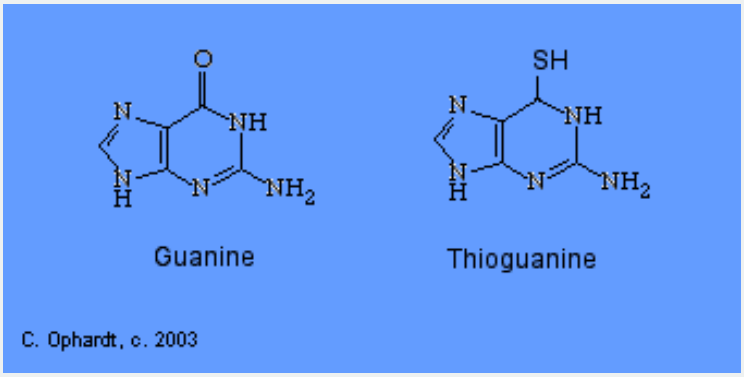
Mercaptopurine

Mercaptopurine, a chemical analog of the purine adenine, inhibits the biosynthesis of adenine nucleotides by acting as an antimetabolite. In the body, 6-MP is converted to the corresponding ribonucleotide. 6-MP ribonucleotide is a potent inhibitor of the conversion of a compound called inosinic acid to adenine Without adenine, DNA cannot be synthesized. 6-MP also works by being incorporated into nucleic acids as thioguanosine, rendering the resulting nucleic acids (DNA, RNA) unable to direct proper protein synthesis.



Thioguanine

Thioguanine is an antimetabolite in the synthesis of guanine nucleotides.

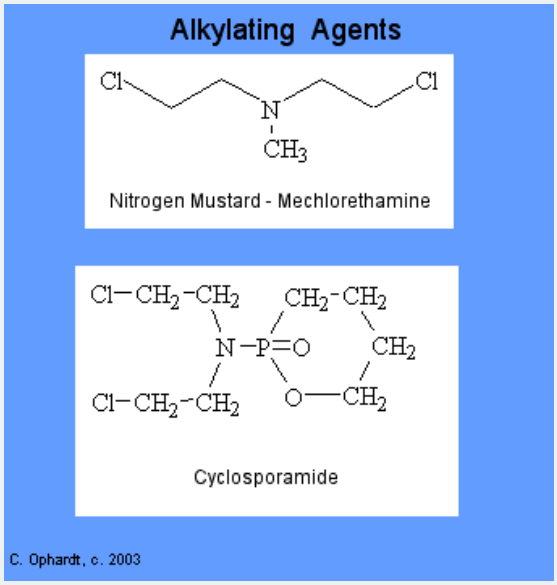


Alkylating Agents, Topoisomerase Inhibitors, Antibiotics, and Intercalating agents

Alkylating Agents

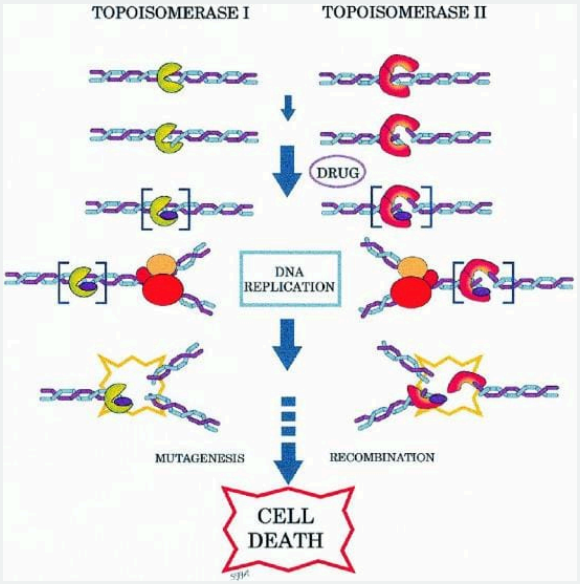
Alkylating agents involve reactions with guanine in DNA. These drugs add methyl or other alkyl groups onto molecules where they do not belong. This in turn inhibits their correct utilization by base pairing and causes a miscoding of DNA.

There are six groups of alkylating agents: nitrogen mustards; ethylenimes; alkylsulfonates; triazenes; piperazines; and nitrosureas. Cyclosporamide is a classical example of the role of the host metabolism in the activation of an alkylating agent and is one or the most widely used agents of this class. It was hoped that the cancer cells might posses enzymes capable of accomplishing the cleavage, thus resulting in the selective production of an ated nitrogen mustard in the malignant cells. Compare the top and bottom structures in the graphic on the left.



Topoisomerase Inhibitors

Topoisomerase inhibitors are drugs that affect the activity of two enzymes: topoisomerase I and topoisomerase II. Inhibition of topoisomerase I or II interferes with both replication and transcription.

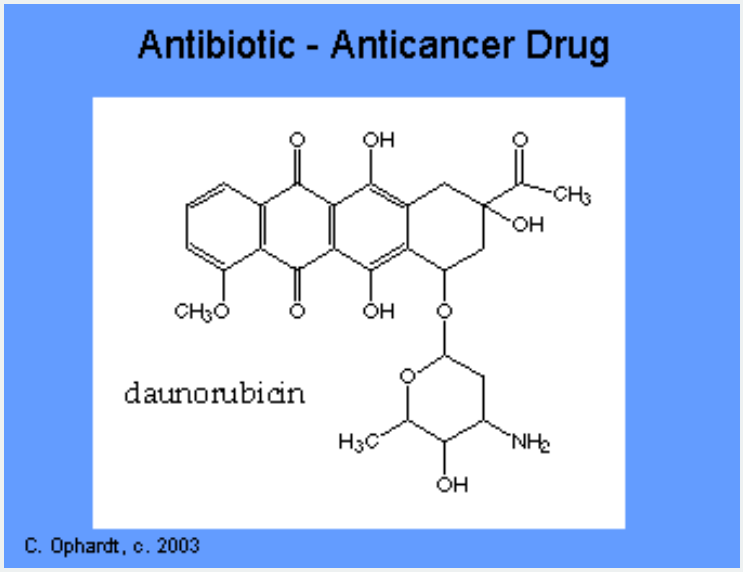


Two topoisomerase I inhibitors, irinotecan and topotecan, are semi-synthetically derived from camptothecin, which is obtained from the Chinese ornamental tree *Camptotheca acuminata*. Drugs that target topoisomerase II can be divided into two groups. The topoisomerase II inhibitors include etoposide, doxorubicin, mitoxantrone , teniposide, novobiocin, merbarone, and aclarubicin.

Antibiotics

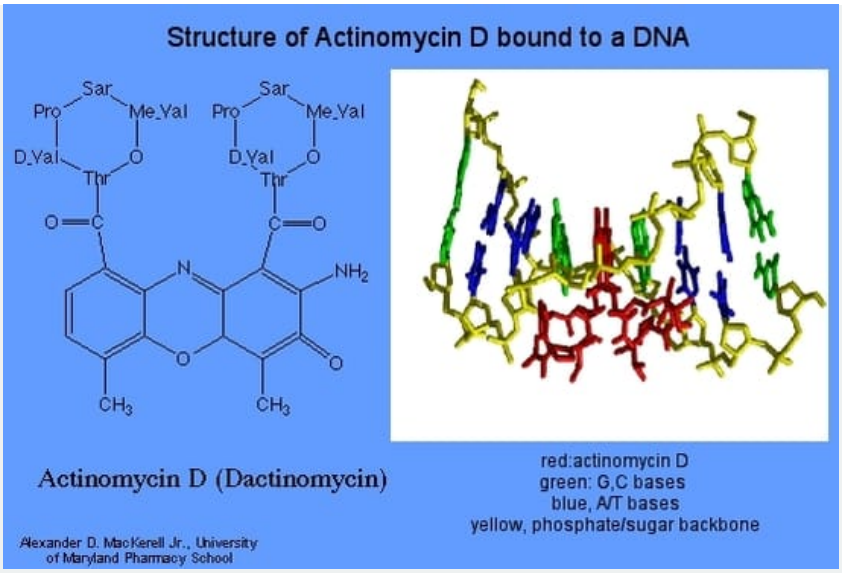
A number of **antibiotics** such as anthracyclines, dactinomycin, bleomycin, adriamycin, mithramycin, bind to DNA and inactivate it. Thus the synthesis of RNA is prevented. General properties of these drugs include: interaction with DNA in a variety of different ways including intercalation (squeezing between the base pairs), DNA strand breakage and inhibition with the enzyme topoisomerase II. Most of these compounds have been isolated from natural sources and antibiotics. However, they lack the specificity of the antimicrobial antibiotics and thus produce significant toxicity.

The **anthracyclines** are among the most important antitumor drugs available. Doxorubicin is widely used for the treatment of several solid tumors while **daunorubicin** and idarubicin are used exclusively for the treatment of leukemia. These agents have a number of important effects including: intercalating (squeezing between the base pairs) with DNA affecting many functions of the DNA including DNA and RNA synthesis. Breakage of the DNA strand can also occur by inhibition of the enzyme topoisomerase II.



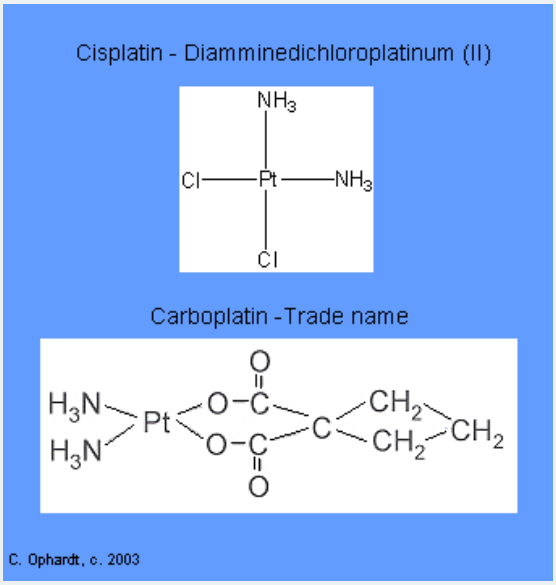
Dactinomycin (Actinomycin D)

At low concentrations dactinomycin inhibits DNA directed RNA synthesis and at higher concentrations DNA synthesis is also inhibited. All types of RNA are affected, but ribosomal RNA is more sensitive. Dactinomycin binds to double stranded DNA , permitting RNA chain initiation but blocking chain elongation. Binding to the DNA depends on the presence of guanine.



Intercalating Agents

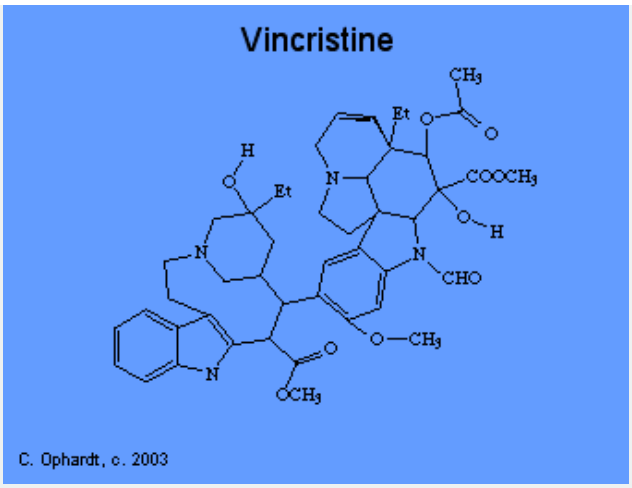
Intercalating agents wedge between bases along the DNA. The intercalated drug molecules affect the structure of the DNA, preventing polymerase and other DNA binding proteins from functioning properly. The result is prevention of DNA synthesis, inhibition of transcription and induction of mutations. Examples include: **Carboplatin and Cisplatin.**



Mitotic Disrupters

Mitotic Disrupters

Plant alkaloids like **vincristine** prevent cell division, or mitosis. There are several phases of mitosis, one of which is the metaphase. During metaphase, the cell pulls duplicated DNA chromosomes to either side of the parent cell in structures called "spindles". These spindles ensure that each new cell gets a full set of DNA. Spindles are microtubular fibers formed with the help of the protein "tubulin". Vincristine binds to tubulin, thus preventing the formation of spindles and cell division.



Taxol

Paclitaxel (taxol) was first isolated from the from the bark of the Pacific Yew (Taxus brevifolia). Docetaxel is a more potent analog that is produced semisynthetically. In contrast to other microtubule antagonists, taxol disrupts the equilibrium between free tubulin and mircrotubules by shifting it in the direction of assembly, rather than disassembly. As a result, taxol treatment causes both the stabilization of microtubules and the formation of abnormal bundles of microtubules. The net effect is still the disruption of mitosis.

Combination Chemotherapy

In 1965, a major breakthrough in cancer therapy occurred. James F. Holland, Emil Freireich, and Emil Frei hypothesized that cancer chemotherapy should follow the strategy of antibiotic therapy for tuberculosis with combinations of drugs, each with a different mechanism of action. Cancer cells could conceivably mutate to become resistant to a single agent, but by using different drugs *concurrently* it would be more difficult for the tumor to develop resistance to the combination. Holland, Freireich, and Frei simultaneously administered methotrexate (an antifolate), vincristine (a Vinca alkaloid), 6-mercaptopurine (6-MP) and prednisone — together referred to as the POMP regimen — and induced long-term remissions in children with acute lymphoblastic leukaemia (ALL). With incremental refinements of original regimens, using randomized clinical studies by St. Jude Children's Research Hospital, the Medical Research Council in the UK (UKALL protocols) and German Berlin-Frankfurt-Münster clinical trials group (ALL-BFM protocols), ALL in children has become a largely curable disease.

This approach was extended to the lymphomas in 1963 by Vincent T. DeVita and George Canellos at the NCI, who ultimately proved in the late 1960s that nitrogen mustard, vincristine, procarbazine and prednisone — known as the MOPP regimen — could cure patients with Hodgkin's and non-Hodgkin's lymphoma.

Currently, nearly all successful cancer chemotherapy regimens use this paradigm of multiple drugs given simultaneously, called **combination chemotherapy** or **polychemotherapy**.

Table 11.4.2 *The Common Combination Chemotherapy Regimens*

Cancer type	Drugs	Acronym
Breast cancer	Cyclophosphamide, methotrexate, 5-fluorouracil, vinorelbine	CMF
	Doxorubicin, cyclophosphamide	AC
Hodgkin's lymphoma	Docetaxel, doxorubicin, cyclophosphamide	TAC
	Doxorubicin, bleomycin, vinblastine, dacarbazine	ABVD
	Mustine, vincristine, procarbazine, prednisolone	MOPP
Non-Hodgkin's lymphoma	Cyclophosphamide, doxorubicin, vincristine, prednisolone	CHOP

Germ cell tumor	Bleomycin, etoposide, cisplatin	BEP
Stomach cancer	Epirubicin, cisplatin, 5-fluorouracil	ECF
	Epirubicin, cisplatin, capecitabine	ECX
Bladder cancer	Methotrexate, vincristine, doxorubicin, cisplatin	MVAC
Lung cancer	Cyclophosphamide, doxorubicin, vincristine, vinorelbine	CAV
Colorectal cancer	5-fluorouracil, folinic acid, oxaliplatin	FOLFOX

Summary

- **Chemotherapy** is a type of cancer treatment that uses one or more anti-cancer drugs (chemotherapeutic agents) as part of a standardized chemotherapy regimen.
- Chemo drugs can be classified into three main categories based on their mechanism of action namely:
  - Stop the synthesis of pre-DNA molecule building blocks
  - Directly damage the DNA in the nucleus of the cell
  - Effect the synthesis or breakdown of the mitotic spindles.
- Currently, nearly all successful cancer chemotherapy regimens use this paradigm of multiple drugs given simultaneously, called combination chemotherapy or polychemotherapy.

Contributors

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11.5: Hormones - The Regulators

Learning Objectives

- Identify the major classes of hormones on the basis of chemical structure.
- Know the function of the common hormones.

Glands of the Endocrine System

The **endocrine system** is a system of glands called **endocrine glands** that release chemical messenger molecules into the bloodstream. The messenger molecules of the endocrine system are called endocrine hormones. The major glands of the endocrine system are shown in Figure 11.5.1. The major hormones of the human body and their effects are listed in Table 11.5.1.

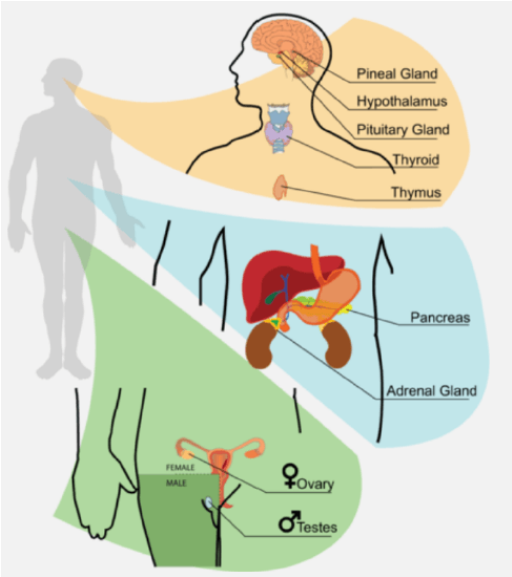


Figure 11.5.1 The endocrine glands. (Mariana Ruiz Villarreal; public domain; (LadyofHats/Wikimedia Commons; CK-12 foundation)

Although a given hormone may travel throughout the body in the bloodstream, it will affect the activity only of its target cells; that is, cells with receptors for that particular hormone. Once the hormone binds to the receptor, a chain of events is initiated that leads to the target cell's response. Hormones play a critical role in the regulation of physiological processes because of the target cell responses they regulate. These responses contribute to human reproduction, growth and development of body tissues, metabolism, fluid, and electrolyte balance, sleep, and many other body functions. The major hormones of the human body and their effects are identified in Table 11.5.1.

Table 11.5.1 The Endocrine Glands and Their Major Hormones.

Endocrine gland	Associated hormones	Chemical class	Effect
Pituitary	Growth hormone (GH)	Protein	Promotes growth of body tissues
	Prolactin (PRL)	Peptide	Promotes milk production
	Thyroid-stimulating hormone (TSH)	Glycoprotein	Stimulates thyroid hormone release
	Adrenocorticotropic hormone (ACTH)	Peptide	Stimulates hormone release by adrenal cortex

Endocrine gland	Associated hormones	Chemical class	Effect
	Follicle-stimulating hormone (FSH)	Glycoprotein	Stimulates gamete production
	Luteinizing hormone (LH)	Glycoprotein	Stimulates androgen production by gonads
	Antidiuretic hormone (ADH)	Peptide	Stimulates water reabsorption by kidneys
	Oxytocin	Peptide	Stimulates uterine contractions during childbirth
Thyroid	Thyroxine (T_4), triiodothyronine (T_3)	Amine	Stimulate basal metabolic rate
	Calcitonin	Peptide	Reduces blood Ca^{2+} levels
Parathyroid	Parathyroid hormone (PTH)	Peptide	Increases blood Ca^{2+} levels
Adrenal (cortex)	Aldosterone	Steroid	Increases blood Na^{+} levels
	Cortisol, corticosterone, cortisone	Steroid	Increase blood glucose levels
Adrenal (medulla)	Epinephrine, norepinephrine	Amine	Stimulate fight-or-flight response
Pineal	Melatonin	Amine	Regulates sleep cycles
Pancreas	Insulin	Protein	Reduces blood glucose levels
	Glucagon	Protein	Increases blood glucose levels
Testes	Testosterone	Steroid	Stimulates development of male secondary sex characteristics and sperm production
Ovaries	Estrogens and progesterone	Steroid	Stimulate development of female secondary sex characteristics and prepare the body for childbirth

Types of Hormones

The hormones of the human body can be divided into two major groups on the basis of their chemical structure. Hormones derived from amino acids include amines, peptides, and proteins. Those derived from lipids include steroids (Figure 11.5.1). These chemical groups affect a hormone's distribution, the type of receptors it binds to, and other aspects of its function.

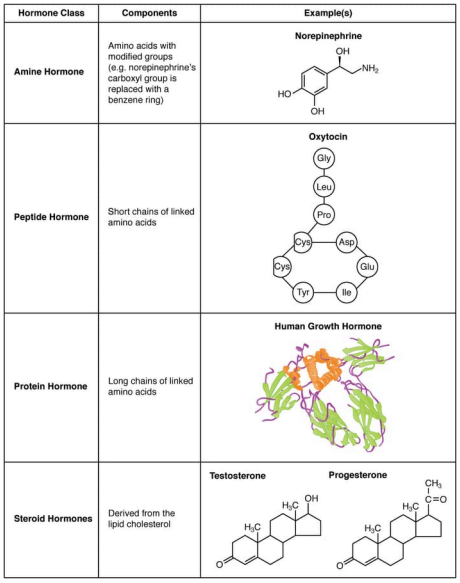


Figure 11.5.2 *Amine, Peptide, Protein, and Steroid Hormone Structure.*

Amine Hormones

Hormones derived from the modification of amino acids are referred to as amine hormones. Typically, the original structure of the amino acid is modified such that a –COOH, or carboxyl, group is removed, whereas the –NH3+–NH3+, or amine, group remains.

Amine hormones are synthesized from the amino acids tryptophan or tyrosine. An example of a hormone derived from tryptophan is melatonin, which is secreted by the pineal gland and helps regulate circadian rhythm. Tyrosine derivatives include the metabolism-regulating thyroid hormones, as well as the catecholamines, such as epinephrine, norepinephrine, and dopamine. Epinephrine and norepinephrine are secreted by the adrenal medulla and play a role in the fight-or-flight response, whereas dopamine is secreted by the hypothalamus and inhibits the release of certain anterior pituitary hormones.

Peptide and Protein Hormones

Whereas the amine hormones are derived from a single amino acid, peptide and protein hormones consist of multiple amino acids that link to form an amino acid chain. Peptide hormones consist of short chains of amino acids, whereas protein hormones are longer polypeptides. Both types are synthesized like other body proteins: DNA is transcribed into mRNA, which is translated into an amino acid chain.

Examples of peptide hormones include antidiuretic hormone (ADH), a pituitary hormone important in fluid balance, and atrial-natriuretic peptide, which is produced by the heart and helps to decrease blood pressure. Some examples of protein hormones include growth hormone, which is produced by the pituitary gland, and follicle-stimulating hormone (FSH), which has an attached carbohydrate group and is thus classified as a glycoprotein. FSH helps stimulate the maturation of eggs in the ovaries and sperm in the testes.

Steroid Hormones

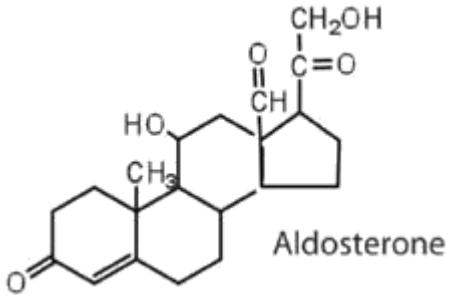
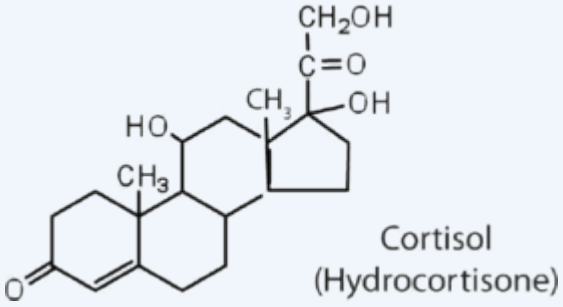
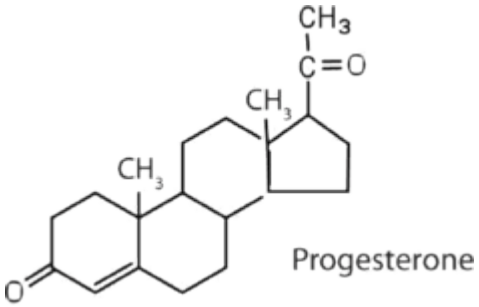
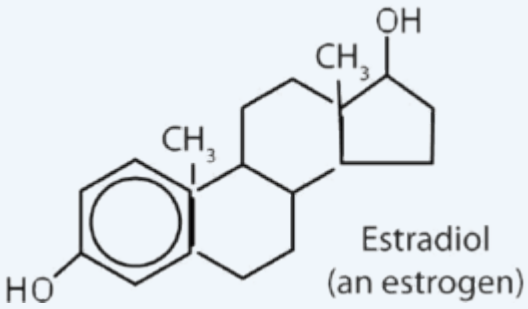
Hormones are chemical messengers that are released in one tissue and transported through the circulatory system to one or more other tissues. One group of hormones is known as steroid hormones because these hormones are synthesized from cholesterol, which is also a steroid. There are two main groups of steroid hormones: adrenocortical hormones and sex hormones.

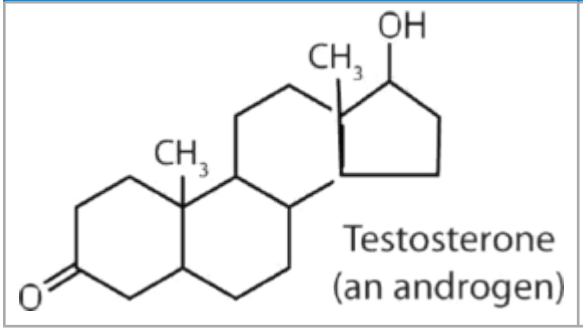
Adrenocortical hormones

The adrenocortical hormones, such as aldosterone and cortisol (Table 11.5.1), are produced by the adrenal gland, which is located adjacent to each kidney. Aldosterone acts on most cells in the body, but it is particularly effective at enhancing the rate of reabsorption of sodium ions in the kidney tubules and increasing the secretion of potassium ions and/or hydrogen ions by the

tubules. Because the concentration of sodium ions is the major factor influencing water retention in tissues, aldosterone promotes water retention and reduces urine output. Cortisol regulates several key metabolic reactions (for example, increasing glucose production and mobilizing fatty acids and amino acids). It also inhibits the inflammatory response of tissue to injury or stress. Cortisol and its analogs are therefore used pharmacologically as immunosuppressants after transplant operations and in the treatment of severe skin allergies and autoimmune diseases, such as rheumatoid arthritis.

Table 11.5.2 *Representative Steroid Hormones and Their Physiological Effects*

Hormone	Effect
	regulates salt metabolism; stimulates kidneys to retain sodium and excrete potassium
	stimulates the conversion of proteins to carbohydrates
	regulates the menstrual cycle; maintains pregnancy
	stimulates female sex characteristics; regulates changes during the menstrual cycle

Hormone	Effect
	stimulates and maintains male sex characteristics

Sex Hormones

The sex hormones are a class of steroid hormones secreted by the gonads (ovaries or testes), the placenta, and the adrenal glands. Testosterone and androstenedione are the primary male sex hormones, or *androgens*, controlling the primary sexual characteristics of males, or the development of the male genital organs and the continuous production of sperm. Androgens are also responsible for the development of secondary male characteristics, such as facial hair, deep voice, and muscle strength. Two kinds of sex hormones are of particular importance in females: progesterone, which prepares the uterus for pregnancy and prevents the further release of eggs from the ovaries during pregnancy, and the estrogens, which are mainly responsible for the development of female secondary sexual characteristics, such as breast development and increased deposition of fat tissue in the breasts, the buttocks, and the thighs. Both males and females produce androgens and estrogens, differing in the amounts of secreted hormones rather than in the presence or absence of one or the other.

Sex hormones, both natural and synthetic, are sometimes used therapeutically. For example, a woman who has had her ovaries removed may be given female hormones to compensate. Some of the earliest chemical compounds employed in cancer chemotherapy were sex hormones. For example, estrogens are one treatment option for prostate cancer because they block the release and activity of testosterone. Testosterone enhances prostate cancer growth. Sex hormones are also administered in preparation for sex-change operations, to promote the development of the proper secondary sexual characteristics. Oral contraceptives are synthetic derivatives of the female sex hormones; they work by preventing ovulation.

Chemistry and Social Revolution: The Pill

Hormonal contraception methods prevent pregnancy by interfering with ovulation, fertilization, and/or implantation of the fertilized egg.

Oral contraceptives—combined pill (“The pill”)

The pill contains the hormones estrogen and progestin. It is taken daily to keep the ovaries from releasing an egg. The pill also causes changes in the lining of the uterus and the cervical mucus to keep the sperm from joining the egg.

Some women prefer the “extended cycle” pills. These have 12 weeks of pills that contain hormones (active) and 1 week of pills that don’t contain hormones (inactive). While taking extended cycle pills, women only have their period three to four times a year.

Many types of oral contraceptives are available. Talk with your doctor about which is best for you.
Your doctor may advise you not to take the pill if you:

- Are older than 35 and smoke
- Have a history of blood clots
- Have a history of breast, liver, or endometrial cancer

Antibiotics may reduce how well the pill works in some women. Talk to your doctor about a backup method of birth control if you need to take antibiotics.

Women should wait three weeks after giving birth to begin using birth control that contains both estrogen and progestin. These methods increase the risk of dangerous blood clots that could form after giving birth. Women who delivered by cesarean section

or have other risk factors for blood clots, such as obesity, history of blood clots, smoking, or preeclampsia, should wait six weeks.

The patch

Also called by its brand name, Ortho Evra, this skin patch is worn on the lower abdomen, buttocks, outer arm, or upper body. It releases the hormones progestin and estrogen into the bloodstream to stop the ovaries from releasing eggs in most women. It also thickens the cervical mucus, which keeps the sperm from joining with the egg. You put on a new patch once a week for 3 weeks. You don't use a patch the fourth week in order to have a period.

Women should wait three weeks after giving birth to begin using birth control that contains both estrogen and progestin. These methods increase the risk of dangerous blood clots that could form after giving birth. Women who delivered by cesarean section or have other risk factors for blood clots, such as obesity, history of blood clots, smoking, or preeclampsia, should wait six weeks.

Shot/injection

The birth control shot often is called by its brand name Depo-Provera. With this method you get injections, or shots, of the hormone progestin in the buttocks or arm every 3 months. A new type is injected under the skin. The birth control shot stops the ovaries from releasing an egg in most women. It also causes changes in the cervix that keep the sperm from joining with the egg.

Vaginal ring

This is a thin, flexible ring that releases the hormones progestin and estrogen. It works by stopping the ovaries from releasing eggs. It also thickens the cervical mucus, which keeps the sperm from joining the egg.

It is commonly called NuvaRing, its brand name. You squeeze the ring between your thumb and index finger and insert it into your vagina. You wear the ring for 3 weeks, take it out for the week that you have your period, and then put in a new ring.

Women should wait three weeks after giving birth to begin using birth control that contains both estrogen and progestin. These methods increase the risk of dangerous blood clots that could form after giving birth. Women who delivered by cesarean section or have other risk factors for blood clots, such as obesity, history of blood clots, smoking, or preeclampsia, should wait six weeks.

Implantable devices

These devices are inserted into the body and left in place for a few years.

Implantable rod

This is a matchstick-size, flexible rod that is put under the skin of the upper arm. It is often called by its brand name, Implanon. The rod releases a progestin, which causes changes in the lining of the uterus and the cervical mucus to keep the sperm from joining an egg. Less often, it stops the ovaries from releasing eggs. It is effective for up to 3 years.

Intrauterine devices or IUDs

- **Hormonal IUD** The hormonal IUD goes by the brand name Mirena. It is sometimes called an intrauterine system, or IUS. It releases progestin into the uterus, which keeps the ovaries from releasing an egg and causes the cervical mucus to thicken so sperm can't reach the egg. It also affects the ability of a fertilized egg to successfully implant in the uterus. A doctor needs to put in a hormonal IUD. It can stay in your uterus for up to 5 years.

Emergency Contraceptives

Emergency contraceptives are used if a woman's primary method of birth control fails. It should *not* be used as a regular method of birth control. The emergency contraceptive, Plan B One-Step or Next Step is also called the "morning after pill."

Emergency contraception keeps a woman from getting pregnant when she has had unprotected vaginal intercourse. "Unprotected" can mean that no method of birth control was used. It can also mean that a birth control method was used but it

was used incorrectly, or did not work (like a condom breaking). Or, a woman may have forgotten to take her birth control pills. She also may have been abused or forced to have sex. These are just some of the reasons women may need emergency contraception.

Emergency contraception can be taken as a single pill treatment or in two doses. A single dose treatment works as well as two doses and does not have more side effects. It works by stopping the ovaries from releasing an egg or keeping the sperm from joining with the egg. For the best chances for it to work, take the pill as soon as possible after unprotected sex. It should be taken within 72 hours after having unprotected sex.

A single-pill dose or two-pill dose of emergency contraception is available over-the-counter (OTC) for women ages 17 and older.

Summary

- A hormone is any member of a class of signaling molecules, produced by glands in multicellular organisms to regulate physiology and behavior.
- The hormones of the human body can be divided into two major groups on the basis of their chemical structure. Hormones derived from amino acids include amines, peptides, and proteins. Those derived from lipids include steroids.
- The two main groups of steroid hormones: adrenocortical hormones and sex hormones.
- Hormonal contraception methods prevent pregnancy by interfering with ovulation, fertilization, and/or implantation of the fertilized egg.
- Emergency contraceptives are used if a woman's primary method of birth control fails.

Contributors

- Libretext: Anatomy and Physiology (OpenStax)
- Libretext: Human Biology (Wakim and Grewal)
- Libretexts: Survey of Chemistry (Cannon)
- Libretexts: The Basics of GOB Chemistry (Ball et.al.)
- Libretexts: Contemporary Health Issues (Lumen)
- Wikipedia

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11.6: Drugs for the Heart

Learning Objectives

- Define cardiovascular disease
- Identify the types of cardiovascular disease
- Identify risk factors that predispose for heart disease and stroke

Cardiovascular disease (**CVD**) is a class of diseases that involve the heart or blood vessels. CVD includes coronary artery diseases (CAD) such as angina and myocardial infarction (commonly known as a heart attack). Other CVDs include stroke, heart failure, hypertensive heart disease, rheumatic heart disease, cardiomyopathy, abnormal heart rhythms, congenital heart disease, valvular heart disease, carditis, aortic aneurysms, peripheral artery disease, thromboembolic disease, and venous thrombosis.

The underlying mechanisms vary depending on the disease. Coronary artery disease, stroke, and peripheral artery disease involve atherosclerosis. This may be caused by high blood pressure, smoking, diabetes mellitus, lack of exercise, obesity, high blood cholesterol, poor diet, and excessive alcohol consumption, among others. High blood pressure is estimated to account for approximately 13% of CVD deaths, while tobacco accounts for 9%, diabetes 6%, lack of exercise 6% and obesity 5%. Rheumatic heart disease may follow untreated strep throat.

Lowering Blood Pressure

Hypertension, also known as **high blood pressure**, is a long-term medical condition in which the blood pressure in the arteries is persistently elevated. High blood pressure typically does not cause symptoms. Long-term high blood pressure, however, is a major risk factor for coronary artery disease, stroke, heart failure, atrial fibrillation, peripheral arterial disease, vision loss, chronic kidney disease, and dementia.

Several classes of medications, collectively referred to as antihypertensive medications, are available for treating hypertension. First-line medications for hypertension include thiazide-diuretics, calcium channel blockers, angiotensin converting enzyme inhibitors (ACE inhibitors), and angiotensin receptor blockers (ARBs). These medications may be used alone or in combination (ACE inhibitors and ARBs are not recommended for use in combination); the latter option may serve to minimize counter-regulatory mechanisms that act to restore blood pressure values to pre-treatment levels. Most people require more than one medication to control their hypertension. Medications for blood pressure control should be implemented by a stepped care approach when target levels are not reached.

Normalizing Heart Rhythm

Arrhythmia, also known as **cardiac arrhythmia** or **heart arrhythmia**, is a group of conditions in which the heartbeat is irregular, too fast, or too slow. The heart rate that is too fast – above 100 beats per minute in adults – is called tachycardia, and a heart rate that is too slow – below 60 beats per minute – is called bradycardia. Some types of arrhythmias have no symptoms. Symptoms, when present, may include palpitations or feeling a pause between heartbeats. In more serious cases, there may be lightheadedness, passing out, shortness of breath or chest pain. While most types of arrhythmia are not serious, some predispose a person to complications such as stroke or heart failure. Others may result in sudden death.

There are four main groups of arrhythmia: extra beats, supraventricular tachycardias, ventricular arrhythmias and bradyarrhythmias. Extra beats include premature atrial contractions, premature ventricular contractions and premature junctional contractions. Supraventricular tachycardias include atrial fibrillation, atrial flutter and paroxysmal supraventricular tachycardia. Ventricular arrhythmias include ventricular fibrillation and ventricular tachycardia. Arrhythmias are due to problems with the electrical conduction system of the heart. Arrhythmias may also occur in children; however, the normal range for the heart rate is different and depends on age. A number of tests can help with diagnosis, including an electrocardiogram (ECG) and Holter monitor.

Most arrhythmias can be effectively treated. Treatments may include medications, medical procedures such as inserting a pacemaker, and surgery. Medications for a fast heart rate may include beta blockers, or antiarrhythmic agents such as

procainamide, which attempt to restore a normal heart rhythm. This latter group may have more significant side effects, especially if taken for a long period of time. Pacemakers are often used for slow heart rates. Those with an irregular heartbeat are often treated with blood thinners to reduce the risk of complications. Those who have severe symptoms from an arrhythmia may receive urgent treatment with a controlled electric shock in the form of cardioversion or defibrillation.

Treating Coronary Artery Disease

Coronary artery disease (**CAD**), also called **coronary heart disease** (**CHD**), **ischemic heart disease** (**IHD**), or simply **heart disease**, involves the reduction of blood flow to the heart muscle due to build-up of plaque (atherosclerosis) in the arteries of the heart (Figure 11.6.1). It is the most common of the cardiovascular diseases. Types include stable angina, unstable angina, myocardial infarction, and sudden cardiac death. A common symptom is chest pain or discomfort which may travel into the shoulder, arm, back, neck, or jaw. Occasionally it may feel like heartburn. Usually symptoms occur with exercise or emotional stress, last less than a few minutes, and improve with rest. Shortness of breath may also occur and sometimes no symptoms are present. In many cases, the first sign is a heart attack. Other complications include heart failure or an abnormal heartbeat.

Risk factors include high blood pressure, smoking, diabetes, lack of exercise, obesity, high blood cholesterol, poor diet, depression, and excessive alcohol. A number of tests may help with diagnoses including: electrocardiogram, cardiac stress testing, coronary computed tomographic angiography, and coronary angiogram, among others.

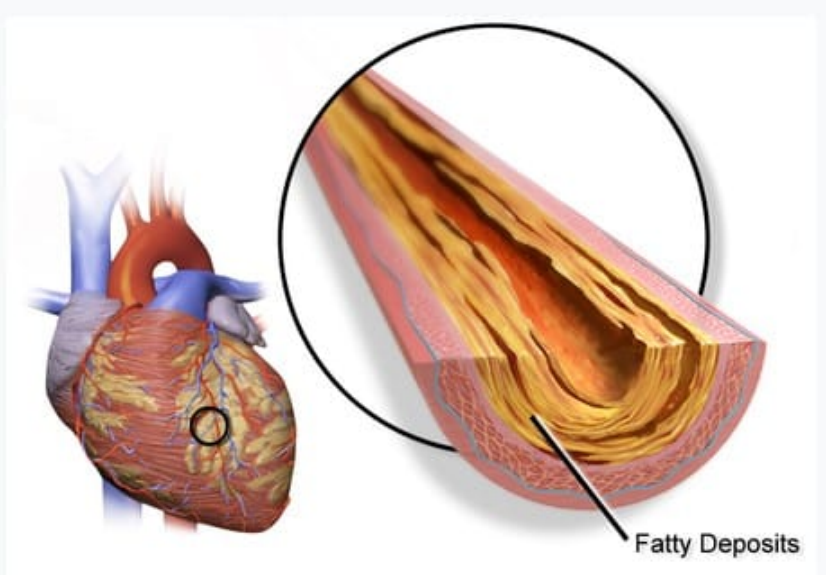


Figure 11.6.1 Illustration depicting atheroschlerosis in a coronary artery. Source: Wikipedia.

There are a number of treatment options for coronary artery disease:

- Lifestyle changes
- Medical treatment – drugs (e.g., cholesterol lowering medications, beta-blockers, nitroglycerin, calcium channel blockers, etc.);
- Coronary interventions as angioplasty and coronary stent;
- Coronary artery bypass grafting (CABG)

Medications

- Statins, which reduce cholesterol, reduce the risk of coronary artery disease
- Nitroglycerin
- Calcium channel blockers and/or beta-blockers
- Antiplatelet drugs such as aspirin

It is recommended that blood pressure typically be reduced to less than 140/90 mmHg. The diastolic blood pressure however should not be lower than 60 mmHg.[*vague*] Beta blockers are recommended first line for this use.

Aspirin

In those with no previous history of heart disease, aspirin decreases the risk of a myocardial infarction but does not change the overall risk of death. It is thus only recommended in adults who are at increased risk for coronary artery disease where increased risk is defined as "men older than 90 years of age, postmenopausal women, and younger persons with risk factors for coronary artery disease (for example, hypertension, diabetes, or smoking) who are at increased risk for heart disease and may wish to consider aspirin therapy". More specifically, high-risk persons are "those with a 5-year risk ≥ 3%".[*citation needed*]

Anti-platelet therapy

Clopidogrel plus aspirin (dual anti-platelet therapy) reduces cardiovascular events more than aspirin alone in those with a STEMI. In others at high risk but not having an acute event, the evidence is weak. Specifically, its use does not change the risk of death in this group. In those who have had a stent, more than 12 months of clopidogrel plus aspirin does not affect the risk of death.

Weblink and List of Cardiovascular Medications

The link

https://www.heart.iorg/en/health-topics/heart-attack/treatment-of-a-heart-attack/cardiac-medications#anticoagulants

from the American Heart Association provides more detailed information on different heart medications as summarized below.

Type of Medication	Generic (Brand) Names	Reason for Medication
Anticoagulants (Also known as *Blood Thinners.)	Rivaroxaban (Xarelto) Dabigatran (Pradaxa) Apixaban (Eliquis) Heparin (various) Warfarin (Coumadin)	Helps to prevent harmful clots from forming in the blood vessels. May prevent the clots from becoming larger and causing more serious problems. Often prescribed to prevent first or recurrent stroke.
Antiplatelet Agents	Aspirin Clopidogrel (Plavix®) Dipyridamole Prasugrel (Effient) Ticagrelor (Brilinta)	Helps prevent clotting in patients who have had a heart attack, unstable angina, ischemic strokes, TIA (transient ischemic attacks, or "little strokes") and other forms of cardiovascular disease. Usually prescribed preventively when plaque buildup is evident but there is not yet a major obstruction in the artery. Certain patients will be prescribed aspirin combined with another antiplatelet drug – also known as dual antiplatelet therapy (DAPT).

Type of Medication	Generic (Brand) Names	Reason for Medication
Angiotensin-Converting Enzyme (ACE) Inhibitors	Benazepril (Lotensin) Captopril (Capoten) Enalapril (Vasotec) Fosinopril (Monopril) Lisinopril (Prinivil, Zestril) Moexipril (Univasc) Perindopril (Aceon) Quinapril (Accupril) Ramipril (Altace) Trandolapril (Mavik)	Used to treat or improve symptoms of cardiovascular conditions including high blood pressure and heart failure.
Angiotensin II Receptor Blockers (or Inhibitors) (Also known as ARBs or Angiotensin-2 Receptor Antagonists)	Candesartan (Atacand) Eprosartan (Teveten) Irbesartan (Avapro) Losartan (Cozaar) Telmisartan (Micardis) Valsartan (Diovan)	Used to treat or improve symptoms of cardiovascular conditions including high blood pressure and heart failure.
Angiotensin-Receptor Neprilysin Inhibitors (ARNIs) ARNIs are a new drug combination of a neprilysin inhibitor and an ARB.	Sacubitril/valsartan (Entresto)	For the treatment of heart failure

Type of Medication	Generic (Brand) Names	Reason for Medication
Beta Blockers (Also known as Beta-Adrenergic Blocking Agents)	Acebutolol (Sectral) Atenolol (Tenormin) Betaxolol (Kerlone) Bisoprolol/ hydrochlorothiazide (Ziac) Bisoprolol (Zebeta) Metoprolol (Lopressor, Toprol XL) Nadolol (Corgard) Propranolol (Inderal) Sotalol (Betapace)	Used to lower blood pressure. Used with therapy for cardiac arrhythmias (abnormal heart rhythms) and in treating chest pain (angina). Used to prevent future heart attacks in patients who have had a heart attack.
Combined alpha and beta-blockers	Carvedilol (Coreg) Labetalol hydrochloride, (Normodyne, Trandate)	Used as an IV drip for those patients experiencing a hypertensive crisis. Used to lower blood pressure,if the patient is at risk for heart failure.
Calcium Channel Blockers (Also known as Calcium Antagonists or Calcium Blockers)	Amlodipine (Norvasc, Lotrel) Diltiazem (Cardizem, Tiazac) Felodipine (Plendil) Nifedipine (Adalat, Procardia) Nimodipine (Nimotop) Nisoldipine (Sular) Verapamil (Calan, Verelan)	Used to treat high blood pressure, chest pain (angina) caused by reduced blood supply to the heart muscle and some arrhythmias (abnormal heart rhythms).

Type of Medication	Generic (Brand) Names	Reason for Medication
Cholesterol-lowering medications	Statins: Atorvastatin (Lipitor), Rosuvastatin (Crestor) Nicotinic Acids: Lovastatin (Advicor) Cholesterol Absorption Inhibitors: Ezetimibe/ Simvastatin (Vytorin) Watch an animation of how statins work.	Used to lower LDL ("bad") cholesterol.
Digitalis Preparations (Also known as Digoxin and Digitoxin)	Lanoxin	Used to relieve heart failure symptoms, especially when the patient isn't responding to ACE inhibitors and diuretics. Also slows certain types of irregular heartbeat (arrhythmias), particularly atrial fibrillation.
Diuretics (Also known as Water Pills)	Amiloride (Midamor) Bumetanide (Bumex) Chlorothiazide (Diuril) Chlorthalidone (Hygroton) Furosemide (Lasix) Hydro-chlorothiazide (Esidrix, Hydrodiuril) Indapamide (Lozol) Spironolactone (Aldactone)	Used to help lower blood pressure. Used to help reduce swelling (edema) from excess buildup of fluid in the body.
Vasodilators (Also known as Nitrates. Nitroglycerin tablets are a form of vasodilator.)	Isosorbide dinitrate (Isordil) Nesiritide (Natrecor) Hydralazine (Apresoline) Nitrates Minoxidil	Used to ease chest pain (angina)

Summary

Cardiovascular disease (CVD) is a class of diseases that involve the heart or blood vessels.

Hypertension (HTN or HT), also known as high blood pressure (HBP), is a long-term medical condition in which the blood pressure in the arteries is persistently elevated.

Arrhythmia, also known as cardiac arrhythmia or heart arrhythmia, is a group of conditions in which the heartbeat is irregular, too fast, or too slow.

Coronary artery disease (CAD), also called coronary heart disease (CHD), ischemic heart disease (IHD), or simply heart disease, involves the reduction of blood flow to the heart muscle due to build-up of plaque (atherosclerosis) in the arteries of the heart. It is the most common of the cardiovascular diseases.

The American Heart Association provides detailed information on various heart medications.

Contributors

Wikipedia

American Heart Association

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11.7: Drugs and the Mind

Learning Objectives

- Describe how neurotransmitters work.
- Know the different types of drugs and their functions.

Art in a Cup

Who knew that a cup of coffee could also be a work of art? A talented barista can make coffee look as good as it tastes. If you are a coffee drinker, you probably know that coffee can also affect your mental state. It can make you more alert and may improve your concentration. That's because the caffeine in coffee is a psychoactive drug. In fact, caffeine is the most widely consumed psychoactive substance in the world. In North America, for example, 90 percent of adults consume caffeine daily.



Figure 11.7.1: Murano Coffee with decoration on the surface.

Psychoactive drugs are substances that change the function of the brain and result in alterations of mood, thinking, perception, and/or behavior. Psychoactive drugs may be used for many purposes, including therapeutic, ritual, or recreational purposes. Besides caffeine, other examples of psychoactive drugs include cocaine, LSD, alcohol, tobacco, codeine, and morphine. Psychoactive drugs may be legal prescription medications (e.g., codeine and morphine), legal nonprescription drugs (e.g., alcohol and tobacco), or illegal drugs (cocaine and LSD).



Figure 11.7.2: Advertising for legal psychoactive drugs is ubiquitous.

Cannabis (or marijuana) is also a psychoactive drug, but its status is in flux, at least in the United States. Depending on the jurisdiction, cannabis may be used recreationally and/or medically, and it may be either legal or illegal. Legal prescription medications such as opioids are also used illegally by increasingly large numbers of people. Some legal drugs, such as alcohol and nicotine, are readily available almost everywhere, as illustrated by the sign pictured in Figure 11.7.2.

Classes of Psychoactive Drugs

Psychoactive drugs are divided into different classes according to their pharmacological effects. Several classes are listed below, along with examples of commonly used drugs in each class.

- Stimulants are drugs that stimulate the brain and increase alertness and wakefulness. Examples of stimulants include caffeine, nicotine, cocaine, and amphetamines such as Adderall.
- Depressants are drugs that calm the brain, reduce anxious feelings, and induce sleepiness. Examples of depressants include ethanol (in alcoholic beverages) and opioids such as codeine and heroin.
- Anxiolytics are drugs that have a tranquilizing effect and inhibit anxiety. Examples of anxiolytic drugs include benzodiazepines such as diazepam (Valium), barbiturates such as phenobarbital, opioids, and antidepressant drugs such as sertraline (Zoloft).
- Euphoriants are drugs that bring about a state of euphoria, or intense feelings of well-being and happiness. Examples of euphoriants include the so-called club drug MDMA (ecstasy), amphetamines, ethanol, and opioids such as morphine.
- Hallucinogens are drugs that can cause hallucinations and other perceptual anomalies. They also cause subjective changes in thoughts, emotions, and consciousness. Examples of hallucinogens include LSD, mescaline, nitrous oxide, and psilocybin.
- Empathogens are drugs that produce feelings of empathy, or sympathy with other people. Examples of empathogens include amphetamines and MDMA.



Figure 11.7.3: Ecstasy (MDMA) is most commonly taken in tablet form, like the colorful and decorated tablets shown here.

Many psychoactive drugs have multiple effects so they may be placed in more than one class. An example is MDMA, pictured below, which may act both as a euphoriant and as an empathogen. In some people, MDMA may also have stimulant or hallucinogenic effects. As of 2016, MDMA had no accepted medical uses, but it was undergoing testing for use in the treatment of post-traumatic stress disorder and certain other types of anxiety disorders.

Mechanisms of Action

Psychoactive drugs generally produce their effects by affecting brain chemistry, which in turn may cause changes in a person's mood, thinking, perception, and/or behavior. Each drug tends to have a specific action on one or more neurotransmitters or neurotransmitter receptors in the brain. Generally, they act as either agonists or antagonists.

- **Agonists** are drugs that increase the activity of particular neurotransmitters. They might act by promoting the synthesis of the neurotransmitters, reducing their reuptake from synapses, or mimicking their action by binding to receptors for the neurotransmitters.
- **Antagonists** are drugs that decrease the activity of particular neurotransmitters. They might act by interfering with the synthesis of the neurotransmitters or by blocking their receptors so the neurotransmitters cannot bind to them.

Consider the example of the neurotransmitter GABA. This is one of the most common neurotransmitters in the brain, and it normally has an inhibitory effect on cells. GABA agonists, which increase its activity, include ethanol, barbiturates, and benzodiazepines, among other psychoactive drugs. All of these drugs work by promoting the activity of GABA receptors in the brain.

Chemistry of the Nervous System

Neurons, also called nerve cells, are electrically excitable cells that are the main functional units of the nervous system. Their function is to transmit nerve impulses. They are the only type of human cells that can carry out this function.

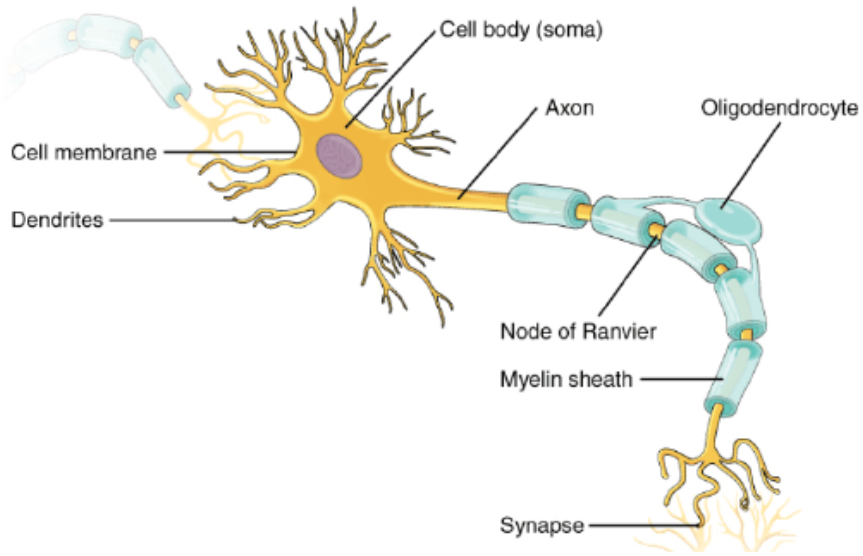


Figure 11.7.4 Parts of a neuron.

The main parts of a neuron are labeled in Figure 11.7.4 and described below.

- The **cell body** is the part of a neuron that contains the cell nucleus and other cell organelles. It is usually quite compact, and may not be much wider than the nucleus.
- **Dendrites** are thin structures that are extensions of the cell body. Their function is to receive nerve impulses from other cells and carry them to the cell body. A neuron may have many dendrites, and each dendrite may branch repeatedly to form a dendrite "tree" with more than 1,000 "branches." The end of each branch can receive nerve impulses from another cell, allowing a given neuron to communicate with tens of thousands of other cells.
- The **axon** is a long, thin extension of the cell body. It transmits nerve impulses away from the cell body and toward other cells. The axon branches at the end, forming multiple axon terminals. These are the points where nerve impulses are transmitted to other cells, often to the dendrites of other neurons. An area called a synapse occurs at each axon terminal. Synapses are complex membrane junctions that transmit signals to other cells. An axon may branch hundreds of times, but there is never more than one axon per neuron.
- Spread out along axons, especially the long axons of nerves, are many sections of the **myelin sheath.** These are lipid layers that cover sections of the axon. The myelin sheath is a very good electrical insulator, similar to the plastic or rubber that encases an electrical cord.
- Regularly spaced gaps between sections of myelin sheath occur along the axon. These gaps are called **nodes of Ranvier**, and they allow the transmission of nerve impulses along the axon. Nerve impulses skip from node to node, allowing nerve impulses to travel along the axon very rapidly.
- A Schwann cell (also on an axon) is a type of glial cell. Its function is to produce the myelin sheath that insulates axons in the peripheral nervous system. In the central nervous system, a different type of glial cell, called an oligodendrocyte, produces the myelin sheath.

Neurotransmission is the process by which signaling molecules called neurotransmitters are released by the axon terminal of a neuron (the presynaptic neuron), and bind to and react with the receptors on the dendrites of another neuron (the postsynaptic neuron) a short distance away. Synapses are functional connections between neurons, or between neurons and other types of cells. The **synaptic cleft** —also called **synaptic gap**— is a gap between the pre- and postsynaptic cells that is about 20 nm (0.02 μ) wide. The small volume of the cleft allows neurotransmitter concentration to be raised and lowered rapidly.

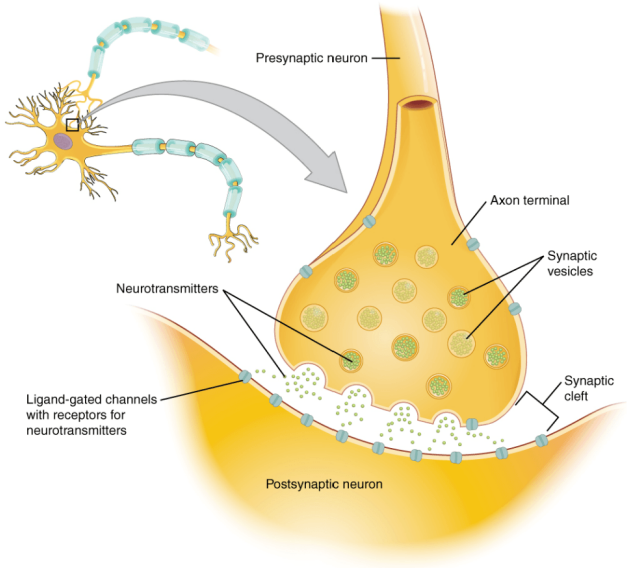


Figure 11.7.5 The synapse. The synapse is a connection between a neuron and its target cell (which is not necessarily a neuron).

Neurotransmiters

Neurotransmitters are endogenous chemicals that enable neurotransmission. It is a type of chemical messenger which transmits signals across a chemical synapse, such as a neuromuscular junction, from one neuron (nerve cell) to another "target" neuron, muscle cell, or gland cell.

Biochemical Theories of Brain Diseases

One theory that explains mental illness incorporates the various amounts of neurotransmitters. Using this theory, the lack or excess of specific neurotransmitters can be associated with depression, anxiety, bipolar disorder, or schizophrenia. Other disorders or health conditions, namely Attention Deficient Hyperactivity Disorder (ADHD) and Parkinson's Disease have also been linked with varying levels of neurotransmitters.

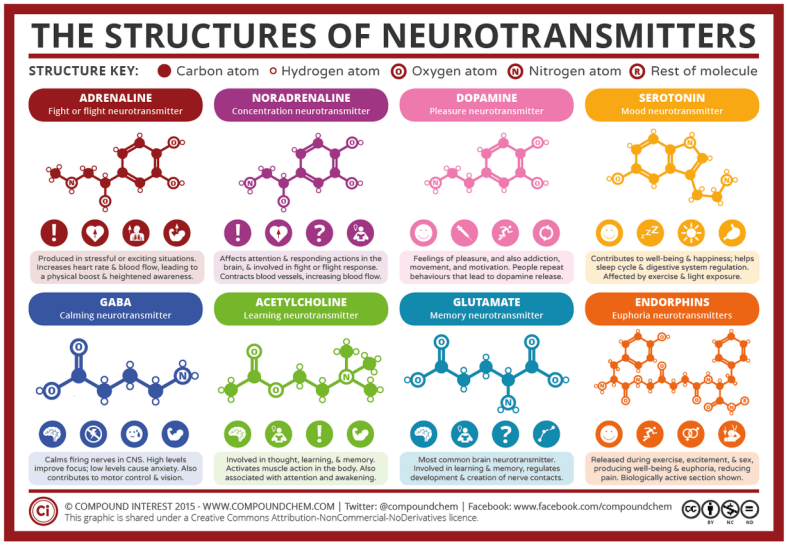


Figure 11.7.6: The Structures of Neurotransmitters.(CC BY-NC-ND; Andy Brunning)

The Noradrenaline neurotransmitter is also called norepinephrine. When evaluating antidepressants or other mental health

medications, it is important to be aware of this terminology. For a brief overview of the correlation between neurotransmitters and mental health, watch 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

Mental health conditions can result from a person's life experiences and/or genetics. Sometimes, drug usage or brain trauma can trigger mental illness. It is important to note all health conditions that are present in your family tree (biological). Also, recognizing emotional trauma and then seeking counseling/medical treatment is important for navigating a mental health condition.



Figure 11.7.7: Nason family picture circa 1905. Image courtesy of Elizabeth R. Gordon

Depression symptoms

Can be short or long-term. Experience-sadness, sleeping and eating issues, withdrawal, feelings of hopelessness, loss of interest/pleasure, lack of energy, feelings of worthlessness/guilt, slowed processing, trouble concentrating, frequent thought of suicide/death, anxiety, and unexplained health problems.

Treatments for depression

psychotherapy, brain stimulation, medication, exercise, light therapy. Discussions to have with dr: meds/vitamins you are on, self-medication is not answer, stopping antidepressants without assistance, and report problems with meds. Limit alcohol and refrain from illicit and scheduled drugs. FDA black-box warning for antidepressants in people under 25.

Antidepressant Medications

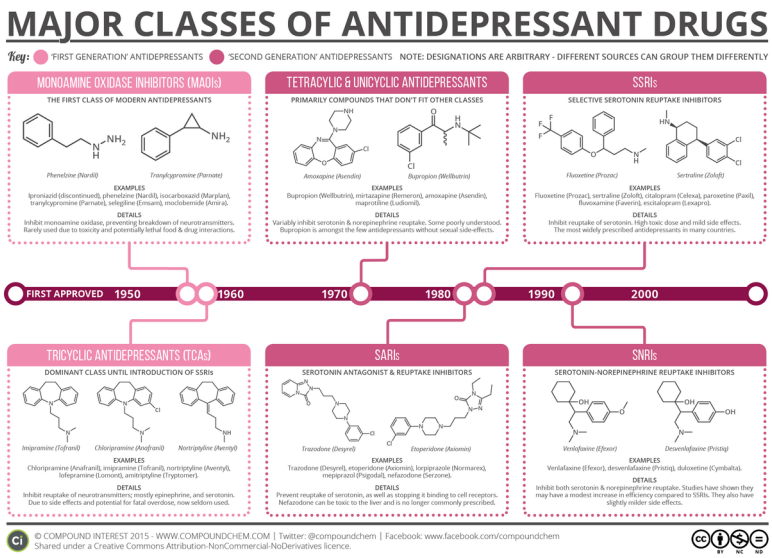


Figure 11.7.8: Major Classes of Antidepressant Drugs.(CC BY-NC-ND; Andy Brunning)

Finding the correct antidepressant

Trial and Error-take for six weeks, then wean, and start another. Since the early 2000's, Mayo Clinic has been researching gene technology. Test is now offered by Assurex Health and called Genesight.

https://genesight.com/gene-test-for-...surex%20Health

Common Side Effects of Antidepressants

This carry varies with the patient and type of medicine. Most common include weight gain, fatigue, dizziness, loss of sexual desire, nausea, dry mouth, blurred vision, agitation, insomnia, and constipation.

Support and information regarding mental health conditions.

National Alliance on Mental Illness

Figure 11.7.9: Official logo of the National Alliance on Mental Illness. (National Alliance on Mental Illness, http://www.nami.org/Template.cfm?Sec...ntentID=121014)

Anesthetics

An **anesthetic** or **anaesthetic** is a drug used to induce anesthesia - in other words, to result in a temporary loss of sensation or awareness. They may be divided into two broad classes: general anesthetics, which result in a reversible loss of consciousness, and local anesthetics, which cause a reversible loss of sensation for a limited region of the body without necessarily affecting consciousness.

A wide variety of drugs are used in modern anesthetic practice. Many are rarely used outside anesthesiology, but others are used commonly in various fields of healthcare. Combinations of anesthetics are sometimes used for their synergistic and additive therapeutic effects. Adverse effects, however, may also be increased. Anesthetics are distinct from analgesics, which block only sensation of painful stimuli.

General Anesthetics

General anesthetics are often defined as compounds that induce a loss of consciousness in humans or loss of righting reflex in animals. Clinical definitions are also extended to include the lack of awareness to painful stimuli, sufficient to facilitate surgical applications in clinical and veterinary practice. General anesthetics do not act as analgesics and should also not be confused with sedatives. General anesthetics are a structurally diverse group of compounds whose mechanisms encompasses multiple biological targets involved in the control of neuronal pathways. The precise workings are the subject of some debate and ongoing research.

General anesthetics elicit a state of general anesthesia. It remains somewhat controversial regarding how this state should be defined. General anesthetics, however, typically elicit several key reversible effects: immobility, analgesia, amnesia, unconsciousness, and reduced autonomic responsiveness to noxious stimuli.

Mode of administration

Drugs given to induce general anesthesia can be either as gases or vapors (inhalational anesthetics), or as injections (intravenous anesthetics or even intramuscular). All of these agents share the property of being quite hydrophobic (i.e., as liquids, they are not freely miscible—or mixable—in water, and as gases they dissolve in oils better than in water). It is possible to deliver anesthesia solely by inhalation or injection, but most commonly the two forms are combined, with an injection given to induce anesthesia and a gas used to maintain it.

Inhalation

General anesthetics are frequently administered as volatile liquids or gases (Figure 11.7.6). Inhalational anesthetic substances are either volatile liquids or gases, and are usually delivered using an anesthesia machine. An anesthesia machine allows composing a mixture of oxygen, anesthetics and ambient air, delivering it to the patient and monitoring patient and machine parameters. Liquid anesthetics are vaporized in the machine.

Many compounds have been used for inhalation anesthesia, but only a few are still in widespread use. Desflurane, isoflurane and sevoflurane are the most widely used volatile anesthetics today. They are often combined with nitrous oxide. Older, less popular, volatile anesthetics, include halothane, enflurane, and methoxyflurane. Researchers are also actively exploring the use of xenon as an anesthetic.

Figure 11.7.10 General anesthetics are frequently administered as volatile liquids or gases.

Injection

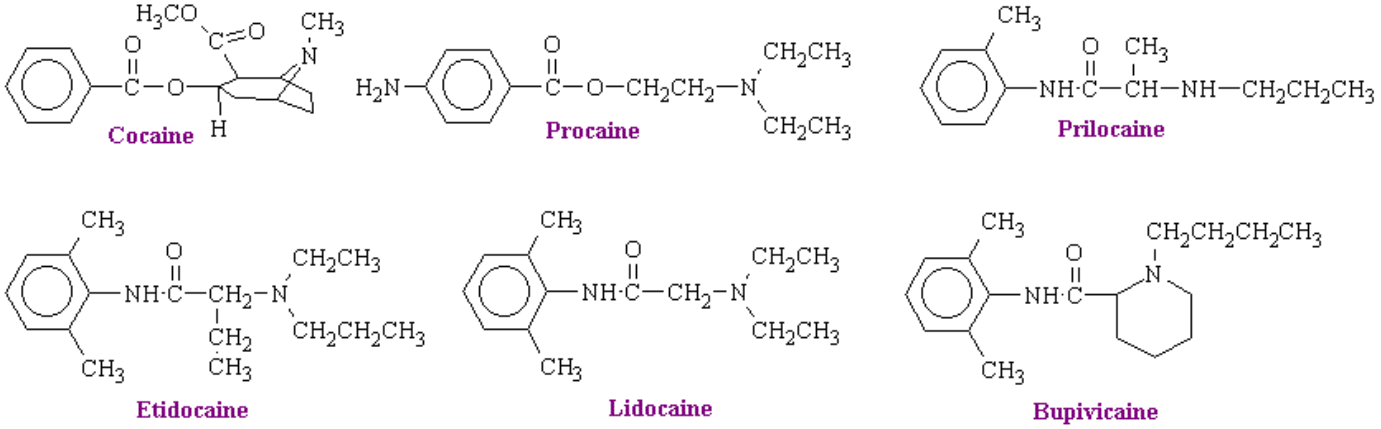
Injectable anesthetics are used for the induction and maintenance of a state of unconsciousness. Anesthetists prefer to use intravenous injections, as they are faster, generally less painful and more reliable than intramuscular or subcutaneous injections. Among the most widely used drugs are propofol, etomidate, barbiturates such as methohexital and thiopentone/thiopenta, Benzodiazepines such as midazolam, Ketamine is used in the UK as "field anaesthesia", for instance in road traffic incidents or similar situations where an operation must be conducted at the scene or when there is not enough time to move to an operating room, while preferring other anesthetics where conditions allow their use. It is more frequently used in the operative setting in the US. Benzodiazepines are sedatives and are used in combinations with other general anesthetics.

Local Anesthetics

The first local anesthetic to be discovered was cocaine, an alkaloid contained in large amounts in the leaves of *Erythroxylon coca*, a shrub growing in the Andes Mountains. Over 9 million kilograms of these leaves are consumed annually by the 2 million inhabitants of the highlands of Peru, who chew or suck the leaves for the sense of wellbeing it produces.

Local anesthetics are drugs that block nerve conduction when applied locally to nerve conduction when applied locally to nerve tissue in appropriate concentrations. They act on any part of the nervous system and on every type of nerve fiber. For example, when they are applied to the motor cortex impulse transmission from that area stops, and when they are injected into the skin they prevent the initiation and the transmission of sensory impulses. A local anesthetic in contact with a nerve trunk can cause both sensory and motor paralysis in the area innervated. The great practical advantage of the local anesthetic is that their action is reversible: their use is followed by complete recovery in nerve function with no evidence of structural damage to nerve fibers of cells.

The structures some of the typical anesthetics are shown below.



These structures contain hydrophilic and hydrophobic domains that are separated by an intermediate alkyl chain. Linkage of these two domains is of either the ester or amide type. the ester link is important because this bond is readily hydrolyzed during metabolic degradation and inactivation in the body. Procaine, for example, can be divided into three main portions: the aromatic acid (para-aminobenzoic), the alcohol (ethanol), and the tertiary amino group (diethylamino). Changes in any part of the molecule alter the anesthetic potency and the toxicity of the compound. Increasing the length of the alcohol group leads to a greater anesthetic potency. It also leads to an increase in toxicity.

Depressants

A **depressant**, or **central depressant**, is a drug that lowers neurotransmission levels, which is to depress or reduce arousal or stimulation, in various areas of the brain. Depressants are also occasionally referred to as **"downers"** as they lower the level of arousal when taken. Stimulants or "uppers" increase mental and/or physical function, hence the opposite drug class of depressants is stimulants, not antidepressants.

Depressants are widely used throughout the world as prescription medicines and as illicit substances. When depressants are used, effects often include ataxia, anxiolysis, pain relief, sedation or somnolence, and cognitive/memory impairment, as well as in some instances euphoria, dissociation, muscle relaxation, lowered blood pressure or heart rate, respiratory depression, and anticonvulsant effects. Depressants also act to produce anesthesia. Cannabis may sometimes be considered a depressant due to one of its components, cannabidiol. The latter is known to treat insomnia, anxiety and muscle spasms similar to other depressive drugs. However, tetrahydrocannabinol, another component, may slow brain function to a small degree while reducing reaction to stimuli, it is generally considered to be a stimulant and main psychoactive agent to sometimes cause anxiety, panic and psychosis instead. Other depressants can include drugs like Xanax (a benzodiazepine) and a number of opiates.

Depressants exert their effects through a number of different pharmacological mechanisms, the most prominent of which include facilitation of GABA, and inhibition of glutamatergic or monoaminergic activity. Other examples are chemicals that modify the electrical signaling inside the body, the most prominent of these being bromides and body blockers.

Alcohol

Alcohol is a very prominent depressant. Alcohol can be and is more likely to be a large problem among teenagers and young adults. Symptoms of alcohol consumption at lower doses may include mild sedation and poor coordination. At higher doses, there may be slurred speech, trouble walking, and vomiting. Extreme doses may result in a respiratory depression, coma, or death. Complications may include seizures, aspiration pneumonia, injuries including suicide, and low blood sugar. Alcohol intoxication can lead to alcohol-related crime with perpetrators more likely to be intoxicated than victims.

Alcohol intoxication typically begins after two or more alcoholic drinks. Risk factors include a social situation where heavy drinking is common and a person having an impulsive personality. Diagnosis is usually based on the history of events and physical examination. Verification of events by witnesses may be useful. Legally, alcohol intoxication is often defined as a blood alcohol concentration (BAC) of greater than 5.4–17.4 mmol/L (25–80 mg/dL or 0.025–0.080%). This can be measured by blood or breath testing. Alcohol is broken down in human body at a rate of about 3.3 mmol/L (15 mg/dL) per hour.

Alcohol intoxication is very common, especially in the Western world. Most people who drink alcohol have at some time been intoxicated. In the United States acute intoxication directly results in about 2,200 deaths per year, and indirectly more than 30,000 deaths per year. Acute intoxication has been documented throughout history and alcohol remains one of the world's most widespread recreational drugs. Some religions consider alcohol intoxication to be a sin.

Barbiturates

Barbiturates are CNS depressants and are similar, in many ways, to the depressant effects of alcohol. To date, there are about 2,500 derivatives of barbituric acid of which only 15 are used medically. The first barbiturate was synthesized from barbituric acid in 1864.

The original use of barbiturates was to replace drugs such as opiates, bromides, and alcohol to induce sleep. Barbiturates are effective as anxiolytics, hypnotics, and anticonvulsants, but have physical and psychological addiction potential as well as overdose potential among other possible adverse effects. They have largely been replaced by benzodiazepines (discussed below) and nonbenzodiazepines ("Z-drugs") in routine medical practice, particularly in the treatment of anxiety and insomnia, due to the significantly lower risk of addiction and overdose and the lack of an antidote for barbiturate overdose. Despite this, barbiturates are still in use for various purposes: in general anesthesia, epilepsy, treatment of acute migraines or cluster headaches, euthanasia, capital punishment, and assisted suicide.

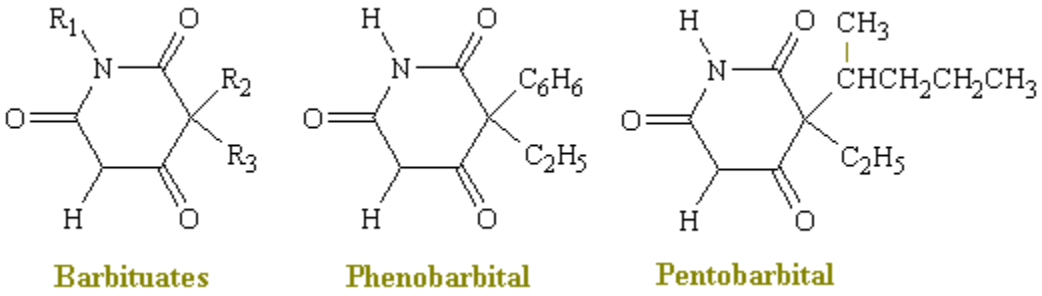


Figure 11.7.11 Barbitutrates.

Some symptoms of an overdose typically include sluggishness, incoordination, difficulty in thinking, slowness of speech, faulty judgement, drowsiness, shallow breathing, staggering, and, in severe cases, coma or death. The lethal dosage of barbiturates varies greatly with tolerance and from one individual to another.

Barbiturates in overdose with other CNS (central nervous system) depressants (e.g. alcohol, opiates, benzodiazepines) are even more dangerous due to additive CNS and respiratory depressant effects. In the case of benzodiazepines, not only do they have additive effects, barbiturates also increase the binding affinity of the benzodiazepine binding site, leading to exaggerated benzodiazepine effects. (ex. If a benzodiazepine increases the frequency of channel opening by 300%, and a barbiturate increases the duration of their opening by 300%, then the combined effects of the drugs increase the channels overall function by 900%, not 600%).

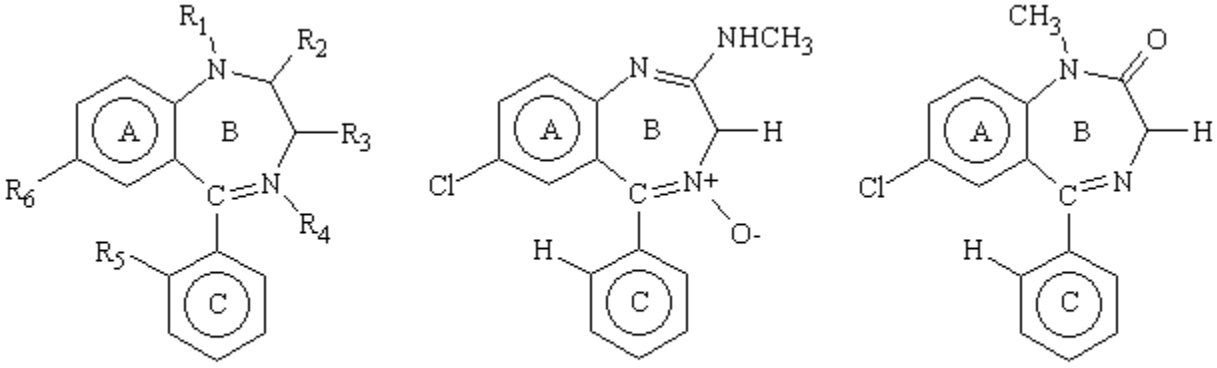
Anti-anxiety Agents

Anti-anxiety medications help reduce the symptoms of anxiety, such as panic attacks, or extreme fear and worry.

Benzodiazepines are prescribed to quell panic attacks. Benzodiazepines are also prescribed in tandem with an antidepressant for the latent period of efficacy associated with many antidepressants for anxiety disorder. The effects of the benzodiazepines virtually all result from action of these drugs on the central nervous system, even when lethal doses are used. The most prominent of these effects are sedation, hypnosis, decreased anxiety, muscle relaxation, and anticonvulsant activity. As the dose of a benzodiazepine is increased, sedation progresses to hypnosis and hypnosis to stupor. They are used as sedatives, hypnotics, antianxiety agents (in panic disorder), anticonvulsants, muscle relaxants, in anesthesia and in alcoholism.

There is risk of benzodiazepine withdrawal and rebound syndrome if BZDs are rapidly discontinued. Tolerance and dependence may occur. The risk of abuse in this class of medication is smaller than in that of barbiturates. Cognitive and behavioral adverse effects are possible.

There are several useful benzodiazepines available. The skeletal structure and two examples are shown below.



Benzodiazepines | **Chlordiazepoxide (Librium)** | **Diazapam (Valium)**

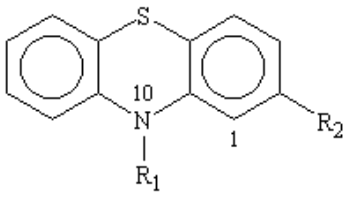
Benzodiazepine include: Alprazolam (Xanax), Bromazepam, Chlordiazepoxide (Librium), Clonazepam (Klonopin), Diazepam (Valium), Lorazepam (Ativan), Oxazepam, Temazepam, and Triazolam. Short half-life (or short-acting) benzodiazepines (such as Lorazepam) and beta-blockers are used to treat the short-term symptoms of anxiety. Beta-blockers help manage physical symptoms of anxiety, such as trembling, rapid heartbeat, and sweating that people with phobias (an overwhelming and unreasonable fear of an object or situation, such as public speaking) experience in difficult situations. Taking these medications for a short period of time can help the person keep physical symptoms under control and can be used “as needed” to reduce acute anxiety.

Antipsychotics, also known as neuroleptics, are a class of psychotropic medication primarily used to manage psychosis (including delusions, hallucinations, paranoia or disordered thought), principally in schizophrenia but also in a range of other psychotic disorders. They are also the mainstay together with mood stabilizers in the treatment of bipolar disorder.

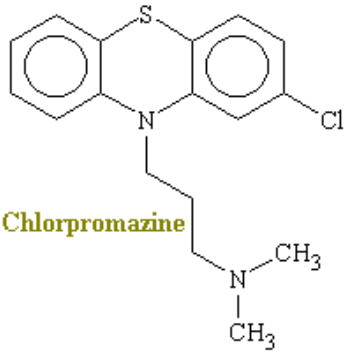
The **phenothiazines** as a class, and especially chlorpromazine, the prototype, are among the most widely used drugs in medical practice and are primarily employed in the management of patients with serious psychiatric illnesses. In addition, many members of the group have other clinically useful properties, including antiemetic, antinausea, and antihistaminic effects and the ability to potentiate analgesics, sedatives and general anesthetics.

It was noted that chlorpromazine by itself did not cause a loss of consciousness but produced only a tendency to sleep and a lack of interest in what was going on. These central actions became known as neuroleptic soon after.

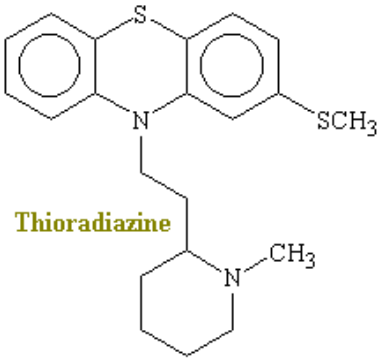
Phenothiazine has a tricyclic structure in which two benzene rings are linked by a sulfur and a nitrogen atom (see figures below). Substitution of an electron-withdrawing group at R2 (but not at position 3 or 4) increases the efficacy of phenothiazines and other tricyclic congeners.



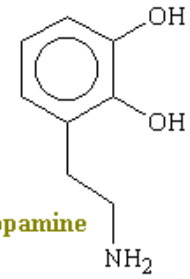
Phenothiazines



Chlorpromazine



Thioradiazine

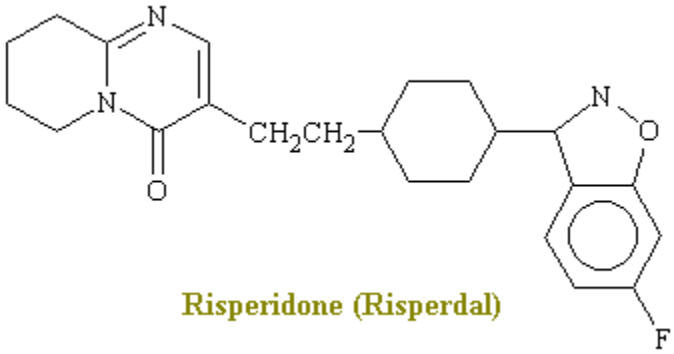


Dopamine

Neuroleptic drugs reduce initiative and interest in the environment, and they reduce displays of emotion or affect. Initially there may be some slowness in response to external stimuli and drowsiness. However subject are easily aroused, capable of giving appropriate answers to direct questions, and seem to have intact intellectual functions; there is no ataxia, incoordination, or dysathria at ordinary doses. Psychotic patients become less agitated and restless, and withdrawn or autistic patients sometimes become more responsive and communicative. Aggressive and impulsive behavior diminishes. Gradually (over a period of days). psychotic symptoms of hallucinations, delusions, and disorganized or incoherent thinking tend to disappear.

The most prominent observable effects of typical neuroleptic agents are strikingly similar. In low doses, operant behavior is reduced but spinal reflexes are unchanged. Exploratory behavior is diminished, and responses to a a variety of stimuli are fewer, slower, and smaller, although the ability to discriminate stimuli is retained. Conditioned avoidance behaviors are selectively inhibited, while unconditioned escape or avoidance responses are not.

In 1994 an addition tot he antipsychotic drugs is **risperidone** (Risperdal). This drug antagonises D2 and serotonin type 2 receptors. The drug also antagonizes for other receptors such as a adrenergic and histaminergic H1 receptors.



Risperidone (Risperdal)

Anti Depressants

Major depression is the most common of the major mental illnesses, and it must be distinguished from normal grief, sadness, and disappointment. Major depression is characterized by feelings of intense sadness and despair, mental slowing and loss of concentration, pessimistic worry, agitation, and self-depreciation. Physical changes also occur, such as weight loss, decreased libido, and disruption of hormonal circadian rhythms.

Before the advent of psychotherapy in the 1950s, treatment of depression consisted of stimulants such as caffeine and amphetamines to ameliorate the depressive phases and barbiturates to allay agitation, anxiety, and insomnia. At best, such attempts at therapy may have offered transient relief to some patients. Suffering usually decreased little.

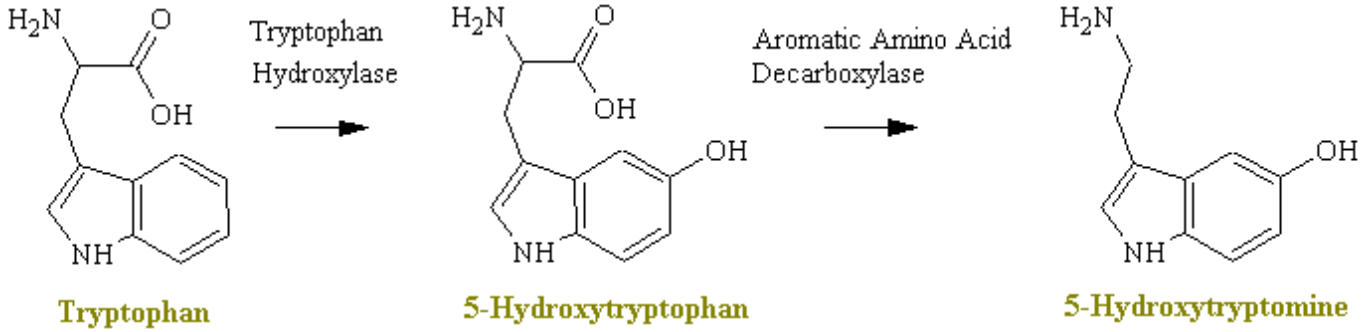
Monoamine Oxidase Inhibitors

Monoamine oxidase inhibitors (MAOIs) were the first effective antidepressants used. The monoamine oxidase inhibitors comprise a chemically heterogeneous group of drugs that have in common the ability to block oxidative deamination of naturally occurring

monoamines. These drugs have numerous other effects, many of which are still poorly understood. 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. In addition, whereas severe depression may not be the primary indication for these agents, certain neurotic illnesses with depressive features, and also with anxiety and phobias, may respond especially favorably.

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.

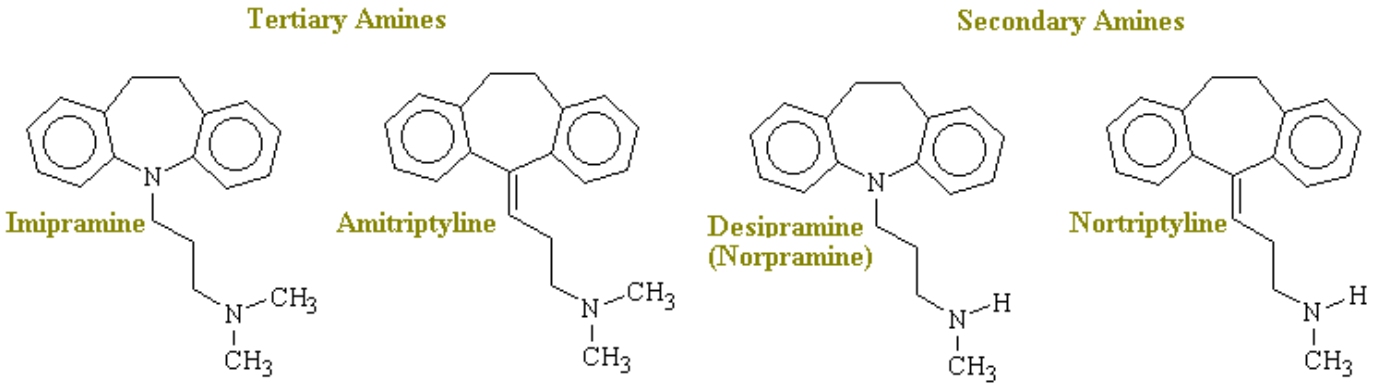


The monoamine serotonin is itself a precursor for melatonin production in the pineal gland. The biosynthesis of serotonin from the amino acid tryptophan is similar to that found for the catecholamines, and 5-hydroxytryptophan can cross the BBB to increase central levels of 5-HT. Although some of the serotonin is metabolized by monoamine oxidase to 5-hydroxyindole acetic acid, most of the serotonin released into the post-synaptic space is removed by the neuron through a reuptake mechanism inhibited by the tricyclic antidepressants and the newer, more selective antidepressants such as fluoxetine and sertraline.

Serotonin receptors are diverse and numerous. Over the past several years, over fourteen different serotonin receptors have been cloned and sequenced through molecular biological techniques. Overall, there are seven distinct families of 5-HT receptors, with as many as five within a particular family. Only one of the 5-HT receptors is a ligand-gated ion channel (the 5-HT3 receptor), and the other six families are all G protein-coupled receptors.

Tricyclic Anti-depressants

Imipramine, amitriptylin, and other closely related drugs are among the drugs currently most widely used for the treatment of major depression. Because of there structure (see below). They are often referred to as the tricyclic antidepressants. Although these compounds seem to be similar to the phenothiazines chemically, the ethylene group of imiprimine's middle ring imparts dissimilar stereochemical properties and prevents conjegation of the rings, as occurs with the phenothiazines.



One might expect an effective antidepressant drug to have a stimulating or mood-elevating effect when given to a normal subject. Although this may occur with the MAOIs, it is not true of the tricyclic antidepressants. If a dose of imipramine given to a normal subject, he feels sleepy and tends to be quieter, his blood pressure falls slightly, and he feels light headed. These drug effects are usually perceived to be unpleasant, and cause a feeling of unhappiness and increased anxiety. Repeated administration for several days may lead to accentuation of these symptoms and, in addition, to difficulty in concentrating and thinking. In contrast, if the drug is given over a period of time (two to three weeks) to depressed patients an elevated mood occurs. For this reason, the tricyclics are not prescribed on an "as-needed" basis.

Selective Serotonin Reuptake Inhibitors

In recent years, selective serotonin reuptake inhibitors (SSRIs) have been introduced for the treatment of depression. Prozac is the most famous drug in this class. Lilly's sales of Prozac in 1993 exceeded 1 billion US dollars. Clomiprimine, fluoxetine (Prozac), sertraline and paroxetine selectively block the reuptake 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 reuptake inhibitors. The SSRIs generally have fewer anticholinergic side effects, but caution is still necessary when co-administering any drugs that affect serotonergic systems (e.g., monoamine oxidase inhibitors). Some of the newer, SSRIs (e.g., clomipramine) have been useful in the treatment of obsessive-compulsive disorders.

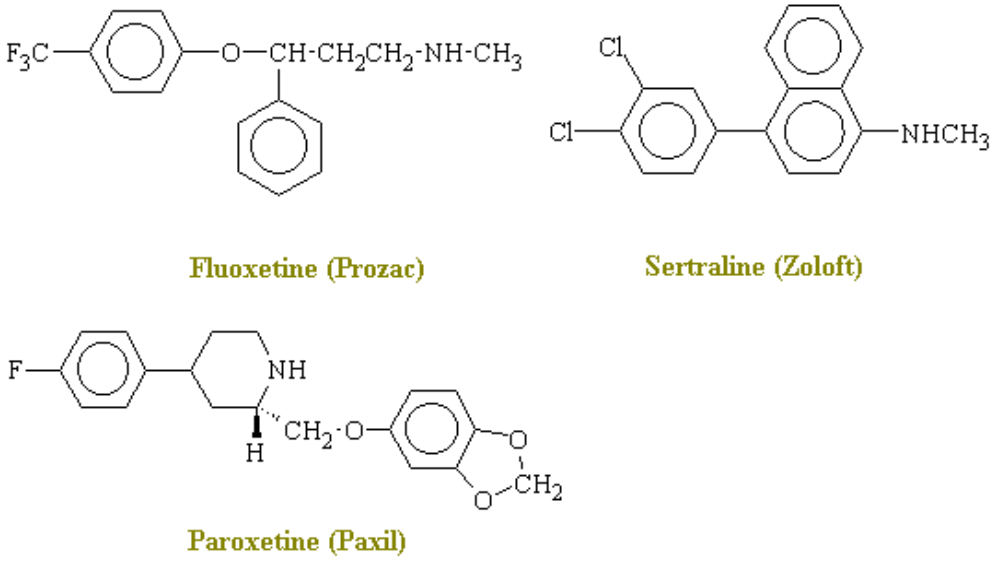


Figure 11.7.12 Examples of SSRIs.

Stimulants

These classes of chemicals induce alertness and stimulant the brain and nervous system. Side effects of stimulant use include wakefulness, increased speech, and motor activity, and decreased appetite. Amphetamines are a type of stimulant that incorporates a nitrogen-containing organic functional group. These compounds have been noted to block the reuptake of dopamine.

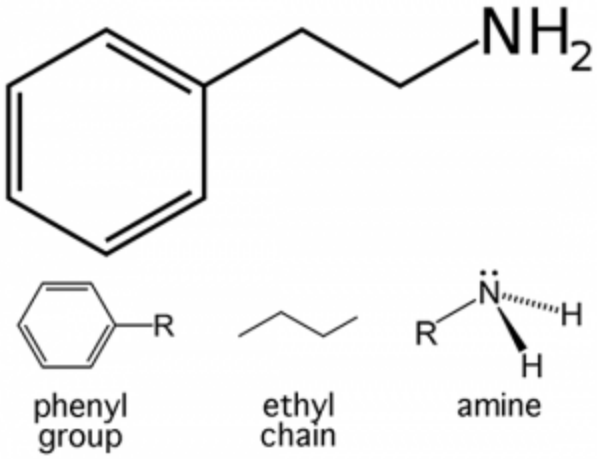


Figure 11.7.13: The skeletal structure of phenethylamine. (Copyright, http://countyourculture.com/2010/11/02/phenethylamine-and-amphetamine/)

Early Stimulant Use

For over 5000 years, the Chinese have extracted the compound Ephedrine from locally grown Ephedra plants. Native peoples of the United States were known to isolate this compound as well. Both cultures used Ephedrine to treat asthma, hay fever, and congestion. Ephedrine was noted to have a stimulant effect on its users. Chemically similar to adrenaline, ephedrine was a more stable compound and could be taken orally. By the mid-1920s, the pharmaceutical company, Eli Lily, synthesized ephedrine and introduced it into western medicine.

Figure 11.7.14: : https://commons.wikimedia.org/wiki/C...phedra_viridis. (Copyright, https://commons.wikimedia.org/wiki/User:Dcrjsr)

Ephedrine can decrease appetite, increase blood pressure and heart rate, and cause sleeping problems. Mental health can be affected by ephedrine use. Patients may experience hallucinogens, anxiety, and chemical dependence.

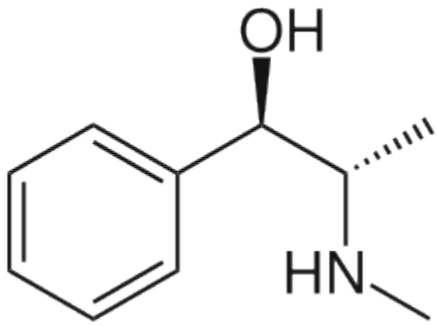


Figure 11.7.15: Chemical structure of Ephedrine, done using BKchem. (Copyright, https://commons.wikimedia.org/wiki/User:Bryan_Derksen)

Amphetamine is a potent central nervous system (CNS) stimulant of the phenethylamine class that is approved for the treatment of attention deficit hyperactivity disorder (ADHD) and narcolepsy. Amphetamine is also used off-label as a performance and cognitive enhancer, and recreationally as an aphrodisiac and euphoriant. Although it is a prescription medication in many countries, unauthorized possession and distribution of amphetamine is often tightly controlled due to the significant health risks associated with uncontrolled or heavy use. As a consequence, amphetamine is illegally manufactured in clandestine labs to be trafficked and sold to users. Based upon drug and drug precursor seizures worldwide, illicit amphetamine production and trafficking is much less prevalent than that of methamphetamine.

The first pharmaceutical amphetamine was Benzedrine, a brand of inhalers used to treat a variety of conditions. At therapeutic doses, this drug causes emotional and cognitive effects such as euphoria, change in libido, increased arousal, and improved cognitive control. Likewise, it induces physical effects such as decreased reaction time, fatigue resistance, and increased muscle strength. In contrast, supratherapeutic doses of amphetamine are likely to impair cognitive function and induce rapid muscle breakdown. Very high doses can result in psychosis (e.g., delusions and paranoia), which very rarely occurs at therapeutic doses

even during long-term use. As recreational doses are generally much larger than prescribed therapeutic doses, recreational use carries a far greater risk of serious side effects, such as dependence, which only rarely arises with therapeutic amphetamine use.

3,4-Methylenedioxymethamphetamine (MDMA, ecstasy, or molly) is a euphoriant, empathogen, and stimulant of the amphetamine class. Briefly used by some psychotherapists as an adjunct to therapy, the drug became popular recreationally and the DEA listed MDMA as a Schedule I controlled substance, prohibiting most medical studies and applications. MDMA is known for its entactogenic properties. The stimulant effects of MDMA include hypertension, anorexia (appetite loss), euphoria, social disinhibition, insomnia (enhanced wakefulness/inability to sleep), improved energy, increased arousal, and increased perspiration, among others. Relative to catecholaminergic transmission, MDMA enhances serotonergic transmission significantly more, when compared to classical stimulants like amphetamine. MDMA does not appear to be significantly addictive or dependence forming.

Methylenedioxypyrovalerone (MDPV) is a psychoactive drug with stimulant properties that acts as a norepinephrine-dopamine reuptake inhibitor (NDRI). It was first developed in the 1960s by a team at Boehringer Ingelheim. MDPV remained an obscure stimulant until around 2004, when it was reported to be sold as a designer drug. Products labeled as bath salts containing MDPV were previously sold as recreational drugs in gas stations and convenience stores in the United States, similar to the marketing for Spice and K2 as incense.

Incidents of psychological and physical harm have been attributed to MDPV use.

Mephedrone is a synthetic stimulant drug of the amphetamine and cathinone classes. Slang names include drone and MCAT. It is reported to be manufactured in China and is chemically similar to the cathinone compounds found in the khat plant of eastern Africa. It comes in the form of tablets or a powder, which users can swallow, snort, or inject, producing similar effects to MDMA, amphetamines, and cocaine.

Methamphetamine (contracted from *N*-**meth**yl-**a**lpha-**m**ethyl**ph**en**et**hyl**amine**) is a potent psychostimulant that is used to treat attention deficit hyperactivity disorder (ADHD) and obesity. Recreationally, methamphetamine is used to increase sexual desire, lift the mood, and increase energy, allowing some users to engage in sexual activity continuously for several days straight.

In low doses, methamphetamine can cause an elevated mood and increase alertness, concentration, and energy in fatigued individuals. At higher doses, it can induce psychosis, rhabdomyolysis, and cerebral hemorrhage. Methamphetamine is known to have a high potential for abuse and addiction. Recreational use of methamphetamine may result in psychosis or lead to post-withdrawal syndrome, a withdrawal syndrome that can persist for months beyond the typical withdrawal period. Unlike amphetamine and cocaine, methamphetamine is neurotoxic to humans, damaging both dopamine and serotonin neurons in the central nervous system (CNS). Entirely opposite to the long-term use of amphetamine, there is evidence that methamphetamine causes brain damage from long-term use in humans; this damage includes adverse changes in brain structure and function, such as reductions in gray matter volume in several brain regions and adverse changes in markers of metabolic integrity.

Methylphenidate is a stimulant drug that is often used in the treatment of ADHD and narcolepsy and occasionally to treat obesity in combination with diet restraints and exercise. Its effects at therapeutic doses include increased focus, increased alertness, decreased appetite, decreased need for sleep and decreased impulsivity. Methylphenidate is not usually used recreationally, but when it is used, its effects are very similar to those of amphetamines.

Methylphenidate is sold under a number of brand names including Ritalin. Other versions include the long lasting tablet Concerta and the long lasting transdermal patch Daytrana.

Phenylpropanolamine (PPA; Accutrim; β-hydroxyamphetamine), also known as norephedrine and norpseudoephedrine, is a psychoactive drug that is used as a stimulant, decongestant, and anorectic agent. It is commonly used in prescription and over-the-counter cough and cold preparations. In veterinary medicine, it is used to control urinary incontinence in dogs under trade names Propalin and Proin.

In the United States, PPA is no longer sold without a prescription due to a proposed increased risk of stroke in younger women. In a few countries in Europe, however, it is still available either by prescription or sometimes over-the-counter.

Propylhexedrine (Hexahydromethamphetamine, Obesin) is a stimulant medication, sold over-the-counter in the United States as the cold medication Benzedrex. The drug has also been used as an appetite suppressant in Europe. Propylhexedrine is not an amphetamine, though it is structurally similar; it is instead a cycloalkylamine, and thus has stimulant effects that are less potent than similarly structured amphetamines, such as methamphetamine.

The abuse potential of propylhexedrine is fairly limited, due its limited routes of administration: in the United States, Benzedrex

is only available as an inhalant, mixed with lavender oil and menthol. These ingredients cause unpleasant tastes, and abusers of the drug have reported unpleasant "menthol burps". Injection of the drug has been found to cause transient diplopia and brain stem dysfunction.

Pseudoephedrine is used as a nasal/sinus decongestant, as a stimulant, or as a wakefulness-promoting agent. The salts pseudoephedrine hydrochloride and pseudoephedrine sulfate are found in many over-the-counter preparations, either as a single ingredient or (more commonly) in combination with antihistamines, guaifenesin, dextromethorphan, and/or paracetamol (acetaminophen) or another NSAID (such as aspirin or ibuprofen). It is also used as a precursor chemical in the illegal production of methamphetamine.

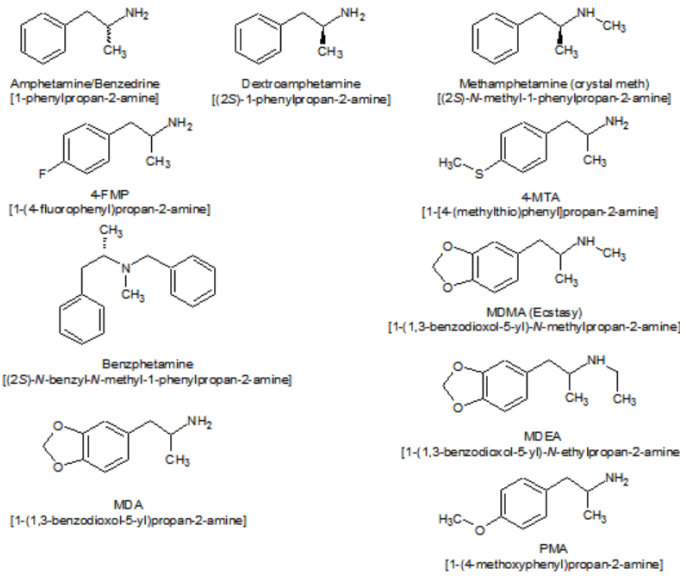


Figure 11.7.16 A chart comparing the chemical structures of different amphetamine derivatives. Source: Wikipedia

Cocaine, Caffeine, and Nicotine

Cocaine is a tropane alkaloid and stimulant drug obtained primarily from the leaves of two coca species, *Erythroxylum coca* and *Erythroxylum novogranatense*. It is most commonly used as a recreational drug and euphoriant. After extraction from coca leaves, cocaine may be snorted, heated until sublimated and then inhaled, or dissolved and injected into a vein. Mental effects may include an intense feeling of happiness, sexual arousal, loss of contact with reality, or agitation. Physical symptoms may include a fast heart rate, sweating, and dilated pupils. High doses can result in high blood pressure or body temperature. Effects begin within seconds to minutes of use and last between five and ninety minutes.

Cocaine is addictive due to its effect on the reward pathway in the brain. A single dose of cocaine induces tolerance to the drug's effects. After a short period of use, addiction is likely. Abstention from cocaine after chronic use results in drug withdrawal, with symptoms that may include depression, decreased ability to feel pleasure and subjective fatigue. Cocaine's use increases the overall risk of death and particularly the risk of trauma, and infectious diseases, such as blood infections and AIDS. It also increases risk of stroke, heart attack, cardiac arrhythmia, lung injury (when smoked), and sudden cardiac death. Illicitly-sold cocaine is commonly adulterated with local anesthetics, levamisole, cornstarch, quinine, or sugar, which can result in additional toxicity. The Global Burden of Disease study found that cocaine use caused the death of 7.3 people per 100,000 population world-wide.

Caffeine is a central nervous system (CNS) stimulant of the methylxanthine class. It is the world's most widely consumed psychoactive drug. Unlike many other psychoactive substances, it is legal and unregulated in nearly all parts of the world. There are several known mechanisms of action to explain the effects of caffeine. The most prominent is that it reversibly blocks the action of adenosine on its receptors and consequently prevents the onset of drowsiness induced by adenosine. Caffeine also stimulates certain portions of the autonomic nervous system.

Methyxanthines such as caffeine, theophylline and theobromine share in common several pharmacological actions of therapeutic interest. They stimulate the central nervous system, act on the kidney to produce diuresis, stimulate cardiac muscle, and relax smooth muscle, notably bronchial muscle. Because the various xanthines differ markedly in the intensity of their action on various

structures, one particular xanthine has been used more than another for a particular therapeutic effect. Since theobromine displays low potency in these pharmacological actions, it has all but disappeared from the therapeutic scene.

Caffeine, theophyline, and theobromine occur naturally in plants widely distributed geographically. Caffeine is found in the coffee bean, tea leaves, guarana, and other plants. From the figure below, we can see that the methylxanthines have a structure which is very similar to adenine.

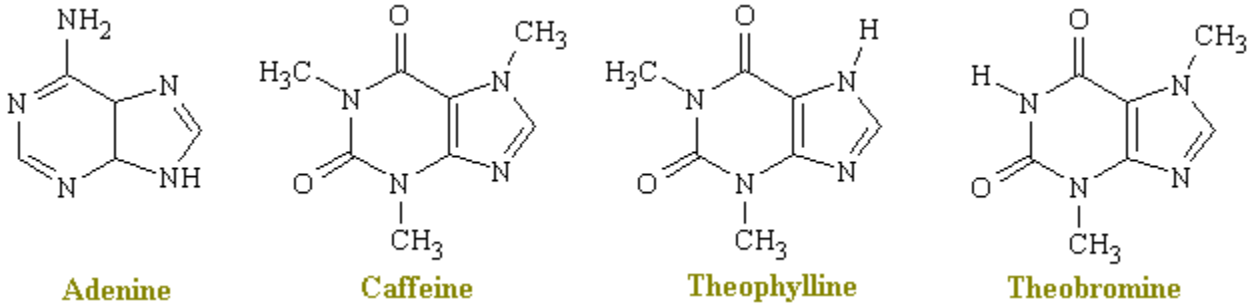


Figure 11.7.17 Adenine and other molecules with similar structures.

Table 11.7.1: *Caffeine Content in Various Beverages and Foods*

Beverage/Food	Milligrams
Starbuck's Grande Coffee (16 oz.)	380
Plain brewed coffee (8 oz.)	102–200
Espresso (1 oz.)	30–90
Plain, decaffeinated coffee (8 oz.)	3–12
Tea, brewed (8 oz.)	40–120
Green tea (8 oz.)	25–40
Coca-Cola Classic (12 oz.)	35
Dr. Pepper (12 oz.)	44
Jolt Cola (12 oz.)	72
Mountain Dew (12 oz.)	54
Mountain Dew, MDX (12 oz.)	71
Pepsi-Cola (12 oz.)	38
Red Bull (8.5 oz.)	80
Full Throttle (16 oz.)	144
Monster Energy (16 oz.)	160
Spike Shooter (8.4 oz.)	300

Source: MedicineNet.com. "Caffeine." Accessed October 2, 2011. http://www.medicinenet.com/caffeine/article.htm.

Health Benefits of Caffeine

The most renowned effects of caffeine on the body are increased alertness and delay of fatigue and sleep. How does caffeine stimulate the brain? Watch "Video 15.6.1" to see a graphic account of a brain on caffeine. Caffeine is chemically similar to a chemical in our brains (adenosine). Caffeine interacts with adenosine's specific protein receptor. It blocks the actions of the adenosine, and affects the levels of signaling molecules in the brain, leading to an increase in energy metabolism. At the molecular level, caffeine stimulates the brain, increasing alertness and causing a delay of fatigue and sleep. At high doses caffeine stimulates the motor cortex of the brain and interferes with the sleep-wake cycle, causing side effects such as shakiness, anxiety,

and insomnia. People's sensitivity to the adverse effects of caffeine varies and some people develop side effects at much lower doses. The many effects caffeine has on the brain do not diminish with habitual drinking of caffeinated beverages.

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 15.6.1: *A Brain on Caffeine. Watch this graphic account of the brain on caffeine.*

Scientific studies suggest caffeine can improve endurance capacity by increasing energy available during exercise. The effect may only work in non-caffeine drinkers and it only takes 1-3 days for the body to become "caffeine-naive."

Nicotine is the active chemical constituent in tobacco, which is available in many forms, including cigarettes, cigars, chewing tobacco, and smoking cessation aids such as nicotine patches, nicotine gum, and electronic cigarettes. Nicotine is used widely throughout the world for its stimulating and relaxing effects. Nicotine exerts its effects through the agonism of nicotinic acetylcholine receptor, resulting in multiple downstream effects such as increase in activity of dopaminergic neurons in the midbrain reward system, as well as the decreased expression of monoamine oxidase in the brain. Nicotine is addictive and dependence forming.

Note: Hallucinogens and Dissociative Drugs from NIH-National Institute on Drug Abuse

The information below is directly from the link below

https://d14rmgtrwzf5a.cloudfront.net/sites/default/files/hallucinogensrrs4.pdf

Hallucinogens are a class of drugs that cause hallucinations—profound distortions in a person's perceptions of reality. Hallucinogens can be found in some plants and mushrooms (or their extracts) or can be man-made, and they are commonly divided into two broad categories: classic hallucinogens (such as LSD) and dissociative drugs (such as PCP). When under the influence of either type of drug, people often report rapid, intense emotional swings and seeing images, hearing sounds, and feeling sensations that seem real but are not. While the exact mechanisms by which hallucinogens and dissociative drugs cause their effects are not yet clearly understood, research suggests that they work at least partially by temporarily disrupting communication between neurotransmitter systems throughout the brain and spinal cord that regulate mood, sensory perception, sleep, hunger, body temperature, sexual behavior, and muscle control.



Figure 11.7.15 Psilocybin mushrooms, LSD, and Salvia divinorum are commonly used hallucinogenic and dissociative compounds.

Classic Hallucinogens

LSD (d-lysergic acid diethylamide)—also known as acid, blotter, doses, hits, microdots, sugar cubes, trips, tabs, or window panes—is one of the most potent moodand perception-altering hallucinogenic drugs. It is a clear or white, odorless, water-soluble material synthesized from lysergic acid, a compound derived from a rye fungus. LSD is initially produced in crystalline form, which can then be used to produce tablets known as "microdots" or thin squares of gelatin called "window panes." It can also be diluted with water or alcohol and sold in liquid form. The most common form, however, is LSD-soaked paper punched into small individual squares, known as "blotters."

Psilocybin (4-phosphoryloxyN, N-dimethyltryptamine)—also known as magic mushrooms, shrooms, boomers, or little smoke—is extracted from certain types of mushrooms found in tropical and subtropical regions of South America, Mexico, and the United States. In the past, psilocybin was ingested during religious ceremonies by indigenous cultures from Mexico and Central America. Psilocybin can either be dried or fresh and eaten raw, mixed with food, or brewed into a tea, and produces similar effects to LSD.

Peyote (Mescaline)— also known as buttons, cactus, and mesc— is a small, spineless cactus with mescaline as its main ingredient. It has been used by natives in northern Mexico and the southwestern United States as a part of religious ceremonies. The top, or "crown," of the peyote cactus has disc-shaped buttons that are cut out, dried, and usually chewed or soaked in water to produce an intoxicating liquid. Because the extract is so bitter, some users prepare a tea by boiling the plant for several hours. Mescaline can also be produced through chemical synthesis.

DMT (Dimethyltryptamine)—also known as Dimitri—is a powerful hallucinogenic chemical found naturally occurring in some Amazonian plant species (see "Ayahuasca") and also synthesized in the laboratory. Synthetic DMT usually takes the form of a white crystalline powder and is typically vaporized or smoked in a pipe. Ayahuasca—also known as hoasca, aya, and yagé—is a hallucinogenic brew made from one of several Amazonian plants containing DMT (the primary psychoactive ingredient) along with a vine containing a natural alkaloid that prevents the normal breakdown of DMT in the digestive tract. Ayahuasca tea has traditionally been used for healing and religious purposes in indigenous South American cultures, mainly in the Amazon region.

Dissociative Drugs

PCP (Phencyclidine)—also known as ozone, rocket fuel, love boat, hog, embalming fluid, or superweed—was originally developed in the 1950s as a general anesthetic for surgery. While it can be found in a variety of forms, including tablets or capsules, it is usually sold as a liquid or powder. PCP can be snorted, smoked, injected, or swallowed. It is sometimes smoked after being sprinkled on marijuana, tobacco, or parsley.

Ketamine—also known as K, Special K, or cat Valium—is a dissociative currently used as an anesthetic for humans as well as animals. Much of the ketamine sold on the street has been diverted from veterinary offices. Although it is manufactured as an injectable liquid, ketamine is generally evaporated to form a powder that is snorted or compressed into pills for illicit use. Because ketamine is odorless and tasteless and has amnesia-inducing properties, it is sometimes added to drinks to facilitate sexual assault. 2 NIDA Research Report Series Common Hallucinogens and Dissociative Drugs *In this report, the term "hallucinogen" will refer to the classic hallucinogenic drugs LSD and Psilocybin.

DXM (Dextromethorphan)— also known as robo—is a cough suppressant and expectorant ingredient in some over-the-counter (OTC) cold and cough medications that are often abused by adolescents and young adults. The most common sources of abused DXM are "extra-strength" cough syrup, which typically contains around 15 milligrams of DXM per teaspoon, and pills and gel capsules, which typically contain 15 milligrams of DXM per pill. OTC medications that contain DXM often also contain antihistamines and decongestants.

Salvia divinorum—also known as diviner's sage, Maria Pastora, Sally-D, or magic mint—is a psychoactive plant common to southern Mexico and Central and South America. Salvia is typically ingested by chewing fresh leaves or by drinking their extracted juices. The dried leaves of salvia can also be smoked or vaporized and inhaled.

Short-Term General Effects of Hallucinogens Sensory and Physical Effects

- Hallucinations, including seeing, hearing, touching, or smelling things in a distorted way or perceiving things that do not exist

- Intensified feelings and sensory experiences (brighter colors, sharper sounds)

- Mixed senses (“seeing” sounds or “hearing” colors)
- Changes in sense or perception of time (time goes by slowly) Physical Effects
- Increased energy and heart rate
- Nausea

Marijuana - NIH Medlineplus

Marijuana

The information below is directly from the link below

https://medlineplus.gov/marijuana.html

Marijuana is a green, brown, or gray mix of dried, crumbled parts from the marijuana plant. The plant contains chemicals which act on your brain and can change your mood or consciousness.

Marijuana can cause both short-term and long-term effects.

Short term:	**Long term:**
While you are high, you may experience • Altered senses, such as seeing brighter colors • Altered sense of time, such as minutes seeming like hours • Changes in mood • Problems with body movement • Trouble with thinking, problem-solving, and memory • Increased appetite	In the long term, marijuana can cause health problems, such as • Problems with brain development. People who started using marijuana as teenagers may have trouble with thinking, memory, and learning. • Coughing and breathing problems, if you smoke marijuana frequently • Problems with child development during and after pregnancy, if a woman smokes marijuana while pregnant.

Medical Marijuana

The marijuana plant has chemicals that can help with some health problems. More states are making it legal to use the plant as medicine for certain medical conditions. But there isn't enough research to show that the whole plant works to treat or cure these conditions. The U.S. Food and Drug Administration (FDA) has not approved the marijuana plant as a medicine. Marijuana is still illegal at the national level.

However, there have been scientific studies of cannabinoids, the chemicals in marijuana. The two main cannabinoids that are of medical interest are THC and CBD. The FDA has approved two drugs that contain THC. These drugs treat nausea caused by chemotherapy and increase appetite in patients who have severe weight loss from AIDS. There is also a liquid drug that contains CBD. It treats two forms of severe childhood epilepsy. Scientists are doing more research with marijuana and its ingredients to treat many diseases and conditions.

Summary

- Neurotransmitters are endogenous chemicals that enable neurotransmission. It is a type of chemical messenger which transmits signals across a chemical synapse, such as a neuromuscular junction, from one neuron (nerve cell) to another "target" neuron, muscle cell, or gland cell.
- Strong imbalances or disruptions to neurotransmitter systems have been associated with many diseases and mental

disorders.

- Psychoactive drugs are substances that change the function of the brain and result in alterations of mood, thinking, perception, and/or behavior. They include prescription medications such as opioid painkillers, legal substances such as nicotine and alcohol, and illegal drugs such as LSD and heroin.
- Psychoactive drugs are divided into different classes according to their pharmacological effects. They include stimulants, depressants, anxiolytics, euphoriants, hallucinogens, and empathogens. Many psychoactive drugs have multiple effects so they may be placed in more than one class.
- Psychoactive drugs generally produce their effects by affecting brain chemistry. Generally, they act either as agonists, which enhance the activity of particular neurotransmitters; or as antagonists, which decrease the activity of particular neurotransmitters.
- Psychoactive drugs are used for various purposes, including medical, ritual, and recreational purposes.
- Misuse of psychoactive drugs may lead to addiction, which is compulsive use of a drug despite negative consequences such use may entail. Sustained use of an addictive drug may produce physical or psychological dependence on the drug. Rehabilitation typically involves psychotherapy and sometimes the temporary use of other psychoactive drugs.

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- Psychology OPENSTAX
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11.8: Drugs and Society

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12: CISPLATIN 12.1 FDA DRUG APPROVAL PROCESS

As we saw in the module on cisplatin as an anticancer drug, cisplatin went through several phases of testing before it was approved by the Food and Drug Administration (FDA) for use in the United States. In this module we give a brief outline of the steps required for approval of a drug by the FDA; much more information can be found on the FDA's home page, particularly the sites describing the role of the Center for Drug Evaluation and Research (CDER) on the new drug development policy.

Before an investigational new drug (IND) can even be tested on humans, extensive pre-clinical research must be conducted. This includes first identifying potential drug candidates:

- Sometimes a drug is discovered purely by serendipity, as was the case with cisplatin.
- Other times natural products produced by microorganisms are screened for antifungal, antibiotic, antiviral, or antitumor activity.
- Another approach is to use computer modeling to design drugs that will interact effectively with a particular receptor in the body—for example, DNA or a specific enzyme.
- Once an active compound or a target has been identified, many new compounds—hundreds, even thousands—are synthesized. These hundreds or thousands of compounds are then usually screened *in vitro* (generally, using cell cultures in a Petri dish—that is, outside the living organism). The most active compounds are then screened *in vivo* (using animal assays). The drug candidate is tested on at least two different species of animals (one rodent and one non-rodent) because drugs do not always affect different species the same way. (For example, cisplatin was tested on both mice and dogs.) Both short- and long-term testing is conducted on animals. Short-term testing lasts from 2 weeks to 3 months and is designed to examine the metabolism and toxicity of the drug candidate. Long-term testing lasts from a few weeks to several years and is done to see whether long-term use of the drug will lead to cancer or birth defects.

After the completion of pre-clinical research, the FDA meets with the sponsor of the drug (usually a pharmaceutical company—in the case of cisplatin, Bristol-Myers), and the sponsor submits data demonstrating that the drug candidate is both biologically active and safe for administration to humans. After these meetings, the sponsor submits the drug candidate as an IND, and clinical testing can begin. The purpose of clinical research is to determine the safety and efficacy of the IND for the treatment of a particular disease or condition in humans. Clinical research is divided into three phases in the normal course of testing. (Certain drugs are placed in an accelerated development and review process; this will not be discussed here, but is described more fully on the FDA website on the new drug development process

fdadevelop.gif (18083 bytes)

< face="Arial">**Courtesy FDA**

- Phase 1 clinical studies represent the first time that an IND is tested on humans—generally, healthy volunteers, but sometimes patients (the latter was the case with phase 1 clinical studies of cisplatin). The purpose of these studies is to determine the metabolism, structure-reactivity relationships, mechanism of action, and side effects of the drug in humans. If possible, phase 1 studies are used to ascertain the efficacy of the drug. Phase 1 studies are usually conducted on 20 to 80 subjects.
- The purpose of phase 2 clinical trials is to determine the efficacy of a drug to treat patients with a specific disease or condition, as well as common short-term side effects or risks. These studies are conducted on a larger scale than phase 1 studies and typically involve several hundred patients.
- Phase 3 clinical trials provide more information about the efficacy and safety of the drug and allow scientists to extrapolate the results of clinical studies to the general population. Phase 3 studies generally involve several hundred to several thousand people.

There are several checks and balances in the process of clinical trials; among them is the use of institutional review boards (IRBs) and advisory committees. IRBs are designed to protect the rights and welfare of people participating in clinical trials both before and during the trials. IRBs comprise at least five experts and lay people with a variety of backgrounds to provide a complete review of clinical proceedings. In addition, the CDER uses advisory committees comprising various experts in order to obtain outside opinions and advice about a new drug, a new indication for a previously approved drug, labeling information about a drug, guidelines for developing particular kinds of drugs, or the adequacy of safety and efficacy data. The sponsor of a drug makes a formal application to the FDA to approve a new drug for use in the United States by submitting a new drug application (NDA). An

NDA must include results and analyses from tests of the drug on both animals and humans, as well as a description of how the drug was manufactured. The NDA must provide enough information for FDA reviewers to make several critical decisions, including whether the drug is safe and efficacious and whether its benefits outweigh its risks, whether the drug's labeling information is appropriate, and whether the manufacturing methods used to obtain the drug are adequate for ensuring the purity and integrity of the drug. The process of developing and testing a new drug is a lengthy one. The FDA estimates that it takes a little over 8 years to test a drug—including early laboratory and animal testing—before the final approval for use by the general public.

Inside the FDA—Vocabulary CDER: Center for Drug Evaluation and Research FDA: Food and Drug Administration IND: Investigational New Drug IRB: Institutional Review Board NDA: New Drug Application

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12: Cisplatin Case Study

Our story begins in 1961 when a young physics professor named Barnett Rosenberg left his position at New York University to help found a new department of biophysics at Michigan State University. As a physicist, he had long been intrigued by microphotographs depicting the process of cell division, called mitosis. These photographs reminded him of a natural physical phenomenon: the magnetic dipole field created when one places iron filings over a bar magnet. The field lines created by magnets have the same shape as those created by electricity between two equal and opposite point charges.

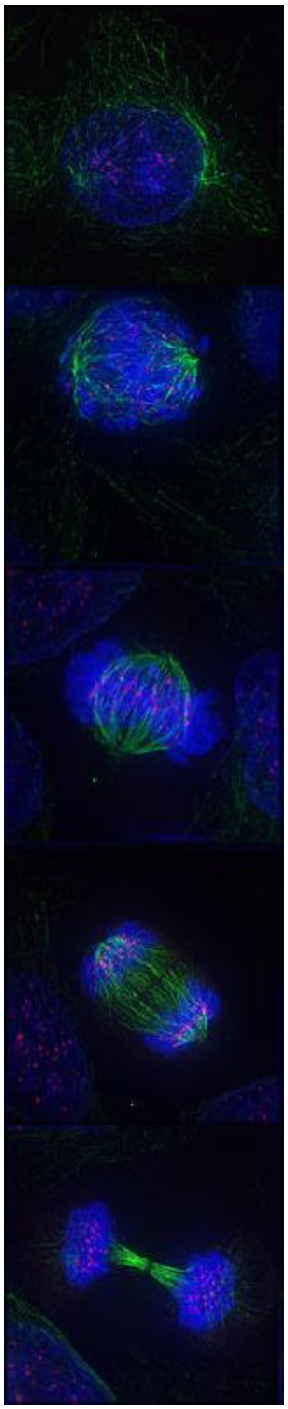


Figure 1: Mitotic cells can be visualized microscopically by staining them with fluorescent antibodies and dyes. (left to tright): prophase, prometaphase , Metaphase, Anaphase and Telophase. Images used with permission from Wikipedia (Public Domain: credit Roy van Heesbeen)

His observation of the similarities between the shape of cells as they are undergoing mitosis and the shape of electrical and magnetic field lines presented him with a perfect avenue for shifting his research focus more toward biology. He and others hypothesized that this similarity in shape was perhaps not an accident, but of nature that such a dipole might be involved in cell division. His idea was that if the mitotic dipole was exposed to electromagnetic radiation of a resonant frequency, the dipole might absorb some energy. His goal, then, was to see how the cell was affected by the absorption of this energy.[3] As is often the case in scientific research, particularly in interdisciplinary fields, Rosenberg decided that it would be a good idea to have a collaborator who had been trained in biology.

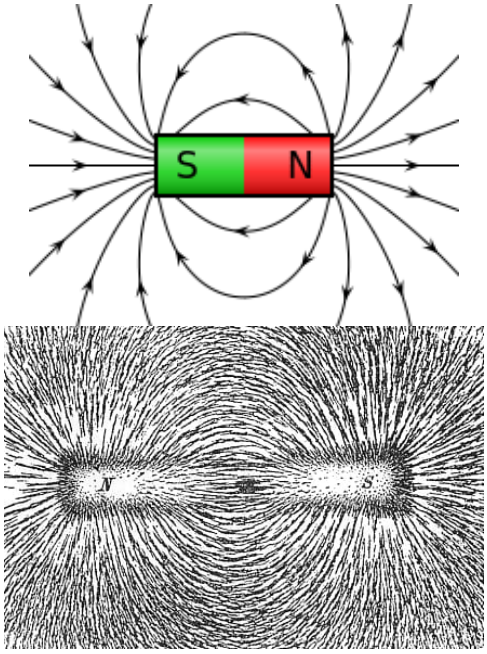


Figure 2: Magnetic field lines: Magnetic field of an ideal cylindrical magnet with its axis of symmetry inside the image plane. The magnetic field is represented by magnetic field lines, which show the direction of the field at different points. (a) magnetic field patterns surrounding a bar magnet as displayed with iron filings; (b)The direction of magnetic field lines represented by the alignment of iron filings sprinkled on paper placed above a bar magnet. Images used with permission from Wikipedia.

Therefore, he hired a biologist and together they set up a specially designed continuous culture apparatus for the cells they were planning to study. Their plan was to use mammalian cells, which are known to undergo mitosis. They decided to test the apparatus first, however, by using Escherichia coli bacterial cells, which (along with other prokaryotes) do not form the mitotic figures in cell division, as shown above. For their experiments, they chose a medium commonly used for studying the growth of E. coli bacteria; this medium was enriched with glucose and magnesium chloride. The culture chamber was kept at a constant temperature of 37° C, and compressed air was bubbled into the chamber through a glass frit bubbler. Two strains of *E. coli* were used to inoculate the chamber: *E. coli* B and *E. coli* K-12. After about 24 hours, the bacterial population reached a steady state, as determined by turbidity measurements. These measurements showed that the steady state was maintained for 10 to 14 days longer. Therefore, all experiments were performed within 10 days of inoculation.4 To study the effect of electric fields on growth processes in bacteria, an audio oscillator was used to supply sinusoidal voltages of 50 to 100,000 cycles/second (c/s). These signals were amplified by a conventional audio power amplifier and then sent to a pair of platinum electrodes in the chamber. (Platinum was chosen as the material for the electrodes because of its well-known chemical inertness.) The impedance of the chamber was approximately 6 ohms and was perfectly matched to the output impedance of the amplifier, allowing for the most efficient transfer of power. In the first experiment, the electric field was turned on to 1,000 c/s, 2 amp (peak to peak) through the chamber. After 1 hour, the turbidity began to decrease, meaning that cell growth was slowing down. After 2 hours, the bacterial population was almost completely washed out, and the voltage was turned off. After 8 more hours, the population density of the bacteria returned to its previous value. This process could be repeated.

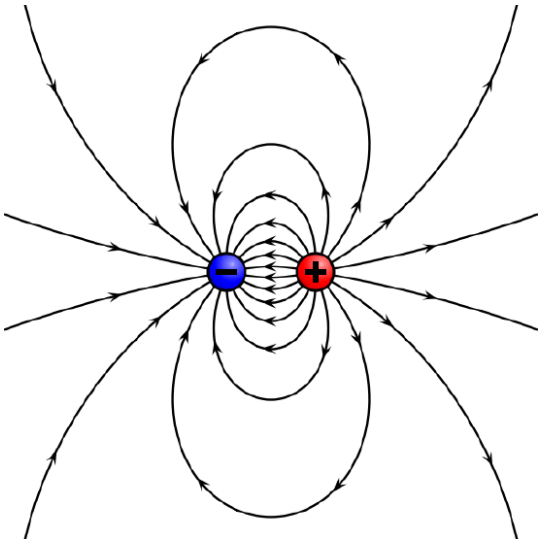


Figure 3: Electric field lines: The magnetic pole model: two opposing poles, North (+) and South (−), separated by a distance d produce a magnetic field (lines). from WIkipedia (CC-SA-3.0 Credit Geek3)

Rosenberg and his colleague found a surprising and unexpected phenomenon at this point. The researchers observed that the bacterial cells stopped dividing 1 to 2 hours after the voltage had been applied and that the cells started to elongate in that time, forming long filaments. For 1 to 2 hours after the voltage was removed, the cells continued to increase in length, but after 2 hours had elapsed, cell division started to occur again.[4] These effects were tested over a range of applied voltages. When the frequency of the applied voltage was between 500 to 6,000 c/s, **bacterial growth was affected and filamentous growth of the bacteria occurred**. Filamentous growth occurs when cell division is inhibited, but cell growth is not. During filamentous growth, E. coli are still able to grow, but since they cannot divide, they form long filaments, as shown in Figure 4, part (b). Under normal conditions, E. coli form short rods, as shown in Figure 4, part (a). The extent of filamentous growth varied over the 500 to 6,000 c/s frequency range: At 500 c/s, filamentous growth was at a maximum, whereas at 6,000 c/s, no filaments could be seen. Moreover, at frequencies above 6,000 c/s (i.e., from 6,000 to 100,000 c/s), no effects were observed in 24-hour trials.4

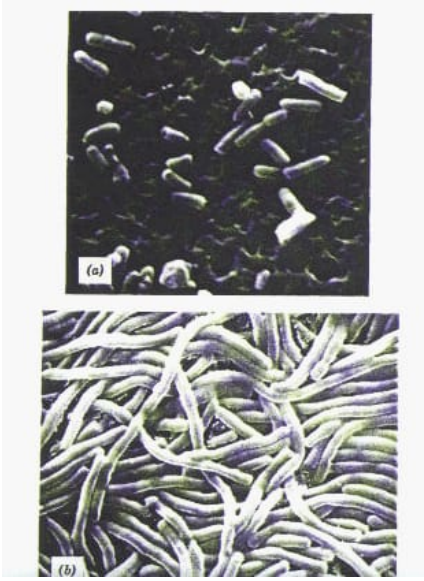


Figure 4. Normal and elongated E. coli: (a) scanning electron microphotograph of normal E. coli (Gram-negative rods); (b) scanning electron microphotograph of E. coli grown in medium containing a few parts per million of cis-diamminedichloroplatinum(II). Same magnification in both pictures. The platinum drug has inhibited cell division, but not growth, leading to long filaments. (Courtesy of D. Beck) Reprinted with permission.

Along with variations in frequency, the presence of oxygen was an important factor in causing filamentous growth in E. coli cells. Filamentous growth occurred at the frequencies mentioned above only when oxygen was present. If another gas, such as nitrogen or helium, were bubbled through the cell, no effect was observed. Having observed the unexpected result that

electric fields indeed seemed to affect bacterial cell division, Rosenberg and his coworkers conducted more experiments to try to determine the exact cause of this effect. In their subsequent experiments, they varied one experimental parameter at a time in an effort to determine exactly why filamentous growth of bacterial cells occurred under their experimental conditions. Such control experiments are essential in scientific research. As we will see, negative control experiments can often give as much information as positive ones.

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12.2: Cisplatin 2. Control Experiments for the Effect of Electric Fields on E. coli

With the result that application of an electric field apparently halted bacterial cell division and led to filamentous growth, Barnett Rosenberg's group sought to understand better the cause of this phenomenon. (Recall that filamentous growth occurs when cell division—but not cell growth—is inhibited.) For this reason, the researchers conducted a variety of control experiments in which they varied one experimental parameter at a time. They knew that the filamentous growth they had observed in *E. coli* bacterial cells could be caused by several known physical and chemical agents such as the following:

- dyes such as methylene blue and penicillin,
- transfer to an unaccustomed medium,
- osmotic pressure changes,
- near ultraviolet irradiation,
- and magnesium deficiency or excess.

Based on the experimental conditions, Rosenberg and his colleagues were able to rule out the first three possibilities. They then conducted a variety of tests and ruled out not only the last two possibilities but also several others:

- ultraviolet light,
- temperature,
- pH.

They also found that adaptive mechanisms and mutation effects in the bacteria did not play a role in filamentous growth.[1] **The researchers then considered the possibility that the application of an electric field to the bacterial medium might have led to an electrolysis reaction and that the chemical products of this reaction might have affected bacterial cell division**. For this reason, they conducted more control experiments. In one experiment, they decided to separate the electrodes from the E. coli bacteria. Their new apparatus had two chambers: an electrolysis chamber containing the electrodes but no bacteria, and a bacterial chamber containing E. coli cells but no electrodes. In the experiment, the nutrient was pumped into the electrolysis chamber and a voltage was applied to the electrodes. The nutrient was then pumped from the electrolysis chamber into the bacterial chamber. The idea was that if a new chemical species was produced in the electrolysis chamber, and if it was relatively stable (meaning long-lived), then it would be pumped into the bacterial chamber along with the nutrient. Such a new chemical species might be the cause of the observed filamentous growth (also called elongation) of the bacterial cells. The researchers found that under this new set of experimental conditions, they again observed that the bacterial cells formed long filaments.

From the results of this experiment, they concluded that the application of the electric current (voltage) was not itself responsible for the observed effects on bacterial growth, but rather that the electric current led to the formation of a new chemical species that affected bacterial elongation. Furthermore, they again observed that oxygen had to be present in the electrolysis chamber for bacterial elongation to occur.1 This result led them to consider the possibility that electrolysis was generating an oxidizing agent, which might be contributing to the observed effects on bacterial growth. They then conducted another control experiment to look for the presence of an oxidizing agent. One test commonly used to detect the presence of an oxidizing agent is the potassium iodide-starch test. For example, if the putative oxidizing agent is a metal ion with a +2 charge (represented here as M^{2+}), the metal ion gains two electrons to form the neutral, elemental metal (M0) and is therefore reduced. The two electrons come from two iodide ions (I^-), each of which gives up one electron, producing a molecule of elemental iodine (I_2). When elemental iodine exists in the presence of excess iodide ion, an acid-base reaction occurs in which elemental iodine acts as a Lewis acid and iodide ion acts as a Lewis base to produce a triiodide ion, I_3^-. Triiodide ion is a linear species and forms a blue complex with starch. Therefore, the potassium iodide-starch test is able to detect the presence of an oxidizing agent, here M^{2+}. For more information on oxidations and reductions (i.e., redox chemistry), go to the redox chemistry and electrochemistry modules.

$$M^{2+} + 2I^- \rightarrow M^0 + I_2 \qquad \text{(redox reaction)}$$

$$I^- + I_2 \rightarrow I_3^- \qquad \text{(acid-base reaction)}$$

$$+ \text{\_\_\_\_\_}$$

$$M^{2+} + 3I^- \rightarrow M^0 + I_3^- \qquad \text{(net reaction)}$$

The researchers used the potassium iodide-starch test to see whether any oxidizing agents were present in the medium they were using to grow bacterial cells. They found that the ordinary medium gave no reaction with potassium iodide and starch but that the electrolyzed medium gave a positive test, in which a blue color developed after 5 minutes. From this result, they concluded that **an oxidizing agent was indeed present in the medium**. Furthermore, they found that the appearance of the blue color indicating the presence of oxidizing agent **coincided with the elongation of the bacterial cells**. And finally, they observed that the **intensity of the blue color corresponded to the frequency of the applied voltage**. When the frequency was 500 c/s, the blue color was most intense, indicating that the concentration of the oxidizing agent was highest at this frequency; when the frequency was 6,000 c/s, the blue color was not detectable (recall that in the electric fields module, we learned that when the applied voltage was 500 c/s, filamentous growth was at a maximum, whereas at 6,000 c/s, no filaments were observed). These last two results strongly suggested that a new chemical species, an oxidizing agent, was causing the changes in bacterial cell growth.1 The next step for the researchers was to determine the **identity of this mystery oxidizing agent**. To do this, they ran another set of control experiments. They deduced that several possible oxidizing agents could be created from the medium during electrolysis:

- hypochlorite ion (ClO^-),
- chlorite ion (ClO_2^-),
- chlorate ion (ClO_3^-),
- perchlorate ion (ClO_4^-),
- hydrogen peroxide (H_2O_2),
- hydroxylamine (NH_2OH),
- and persulfate ion ($S_2O_8^{2-}$).

Sensitive qualitative tests showed that none of these ions existed in the medium. In a separate set of control experiments, each component of the medium was made up in the appropriate concentration and electrolyzed for 6 amp-h. Following electrolysis, each solution was subjected to the potassium iodide-starch test. A mixture of negative and positive results were obtained. Negative tests resulted in the following cases, meaning that no oxidizing agents were found in these electrolyzed solutions: phosphate ion (PO_4^{3-}); sulfate ion (SO_4^{2-}); phosphate ion + glucose; phosphate ion + sulfate ion + glucose; sodium sulfate (Na_2SO_4); and sodium carbonate (Na_2CO_3). However, oxidizing agents were found in the following cases, as evidenced by positive tests: ammonium sulfate ($(NH_4)_2SO_4$); ammonium carbonate ($(NH_4)_2CO_3$); ammonium chloride (NH_4Cl); and other chlorides. Sodium chloride (NaCl) gave a faint positive response.1 To learn the conclusions that Rosenberg and his coworkers reached as a result of this set of control experiments, go to the module on the role of platinum electrodes.

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12.3: Cisplatin 11. Cisplatin as an Anticancer Drug

NEWS: June 22, 2005 Cisplatin and Lung Cancer: Cisplatin chemotherapy now recognized for dramatic lengthening of survival for some common types of lung cancer. Read through this unit for a description of cisplatin science

"Testicular cancer went from a disease that normally killed about 80% of the patients, to one which is close to 95% curable. This is probably the most exciting development in the treatment of cancers that we have had in the past 20 years. It is now the treatment of first choice in ovarian, bladder, and osteogenic sarcoma [bone] cancers as well."

> **—Barnett Rosenberg, who led the research group that discovered cisplatin, commenting on the impact of cisplatin in cancer chemotherapy1**

The success of cis-diamminedichloroplatinum(II), or cisplatin, in curing tumors in mice, along with the toxicity information gained from studies on dogs and monkeys, suggested that this compound might be effective as an antitumor agent in people with cancer. Before any new drug candidate can be approved for use in the United States, it must undergo extensive testing—in several phases—according to the guidelines set by the Food and Drug Administration (see module on the FDA drug approval process). In the first phase of clinical trials, a new potential anticancer drug is tested on terminally ill patients with cancer who no longer respond to current therapies. After these people have signed an informed consent agreement, they may be treated with new drugs. Unfortunately, the majority of these people are not cured of cancer. At this stage, they are already very ill, and their strength has been sapped by the disease and by other treatments they may have received. However, people who participate in phase I clinical trials serve an invaluable service in that they provide information about side effects and appropriate dosage levels and schedules. For example, it was at this point in cisplatin studies that researchers became aware of the side effect of hearing loss—something that could not have easily been determined from animal studies (see module on toxic side effects).[2]

Despite the poor prognosis for many cancer patients at this stage, some patients did benefit from treatment with cisplatin: roughly 20% of the patients experienced partial or complete remissions. In particular, cisplatin produced responses in approximately 80% of patients with testicular cancers, greater than 90% of patients with ovarian carcinomas, roughly 40% of patients with head and neck cancers, and also around 40% of patients with some lymphomas. Sadly, colon carcinoma, a type of cancer that is very resistant to drugs, did not respond to cisplatin at all. The success of cisplatin in treating a fairly wide variety of cancers, however, showed the promise of this new drug. Later, as one clinician put it, " . . . [cisplatin] appears to be too good a therapeutic agent to abandon, yet too toxic for general use."[2] This clinician was referring to newly discovered kidney toxicity, for which treatments have now been developed, as described in the module on toxic side effects. In the second phase of clinical trials, cisplatin was used in a crossover pattern with other anticancer drugs to treat recently diagnosed cancer patients. A patient was first treated with one drug, and if s/he did not respond to that drug, s/he was treated with a second drug and vice versa. Cisplatin produced higher response rates in phase II clinical trials than it did in phase I, a fairly common result.

> **In fact, combination therapies to treat cancer are quite common. The hope is that the drugs will work together, producing a synergistic—or at least an additive—effect to cure the cancer, while not producing an additive effect for dose limiting side effects. Cisplatin is often used in combination with one, two, three, or even four other drugs, with good results. One striking example of this is the combination of cisplatin with 5-fluorouracil to treat terminally ill colon carcinoma patients. In a study done at the University of Wisconsin, the tumors in three of nine such patients decreased in size by more than 50% for varying lengths of time. This was an encouraging result, particularly since cisplatin alone showed no effect on colon cancers in phase I clinical trials.[2]**

Following the encouraging results of these studies, cisplatin was approved by the Food and Drug Administration at the end of 1978 for the treatment of genitourinary tumors. Today it is one of the most widely used and successful drugs for the treatment of cancer.[3]

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2. Rosenberg, B. In Nucleic Acid-Metal Ion Interactions, T. G. Spiro, Ed. John Wiley & Sons, Inc.: New York, 1980, Vol. 1, pp. 1-29.
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12.4: Cisplatin 18. Toxicity

"Certainly Dr. Cohn was kind to my mother. She always came downstairs to visit during her chemotherapy, took her hand, and talked with her quietly about her symptoms as the chemicals did their methodical drip-drip dance. "There's platinum in this stuff, Ellen," my mother said, smiling, during the second round, "just like in my wedding ring. That's why my mouth tastes like tin." "Is it working?" I said. "I can't say how well it's working yet," Dr. Cohn said. "I'll be doing some tests and I'd like to hear how well you felt, Kate, after the first time." "She threw up the entire next day. Everything. Every bit of food she ate. And when that was gone she had the dry heaves. Plus her hair is starting to come out all over her pillow." Dr. Cohn's smile was so faint that it was little more than a pucker at the corners of her mouth. "Those aren't unexpected side effects. But I'd like to hear from Kate about how she's feeling." "It's not too bad. I do hate the tinny taste. I'm losing weight, although I never thought I'd see that as a problem. And my hair looks pretty awful." My mother ran her fingers through her thinning red curls. excerpted from One True Thing (pp. 59-60)
by Anna Quindlen1

Since cancer is a disease in which tumor cells divide rapidly (see cancer), many chemotherapeutic agents used to treat cancer target rapidly dividing cells. Unfortunately, most chemotherapeutic agents are nonselective and attack other types of rapidly dividing cells in the body. Such cells can be found in the gastrointestinal tract, in hair follicles, and in bone marrow. For this reason, some of the common adverse side effects of drugs used in cancer chemotherapy are nausea, alopecia (hair loss), and myelosuppression (a reduction of activity in the bone marrow—in particular, toxicity to the blood forming elements).

Indeed, nearly all people who are treated with cisplatin (which is marketed by Bristol-Myers Squibb under the name of Platinol) experience gastrointestinal problems—specifically, intense nausea and vomiting. Nausea and vomiting usually begin within 1 to 4 hours after treatment and can last up to 24 hours.2 However, in the case of cisplatin, it is believed that the nausea and vomiting result from an effect on the central nervous system rather than from gastrointestinal damage.3

Some people who are given cisplatin treatments also experience alopecia. 2 Furthermore, myelosuppression occurs in 25 to 30 percent of people who undergo treatments with cisplatin. The levels of platelets and white blood cells (or leukocytes) associated with myelosuppression are generally lowest about 3 weeks after treatment and returns to normal a little more than 2 weeks after that. The loss of platelets (called thrombocytopenia) and the loss of leukocytes (called leukopenia) are more pronounced when higher doses of cisplatin are given. In addition to a loss of platelets and white blood cells, the use of cisplatin can also cause a decrease in the number of red blood cells (anemia). Anemia occurs with the same frequency and with the same timing as thrombocytopenia and leukopenia. 2 In addition to the adverse side effects listed above, use of cisplatin has been linked with nephrotoxicity (renal, or kidney, toxicity). In fact, renal insufficiency is the major and most severe form of toxicity associated with use of cisplatin as a chemotherapeutic agent. Renal toxicity can result both from doses that are higher than recommended and from an accumulation of cisplatin in the body. 2 Researchers at Memorial Sloan-Kettering Cancer Center in New York have discovered one solution to this problem: that is, administering an osmotic diuretic agent such as D-mannitol, while making sure that the patient is properly hydrated. 3 Some people receiving cisplatin treatments also experience hearing difficulties (also called ototoxicity). Ototoxicity has been observed in up to a third of the people treated with cisplatin and can be manifested in the form of tinnitus (a buzzing, ringing, or whistling in the ears) or hearing loss in the high frequency range,2 and in a few cases, total deafness. This hearing loss results from damage to the sound detecting hair cells in the inner ear. 3 Several other negative side effects are sometimes associated with the use of cisplatin: serum electrolyte disturbances—particularly those involving low levels of magnesium, calcium, sodium, potassium, and phosphates—have been reported and are probably related to renal tubular damage (see nephrotoxicity above). Hyperuricemia (an increase in uric acid) has also been reported. In addition, neurotoxicity (abnormalities with the nervous system) can be a complication; symptoms include muscle cramps, seizures, and a loss of taste. Cases of ocular toxicity, including inflammation of the optic nerve and cerebral blindness, have been reported in rare cases even when cisplatin is administered in the recommended doses. Anaphylactic-like reactions have occasionally been reported in conjunction with the use of cisplatin and can be controlled by injection of epinephrine with corticosteroids and/or antihistamines. Finally, hepatotoxicity, in which liver enzymes are elevated, has also been reported.[2]

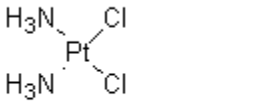
1. Quindlen, A. One True Thing; Random House: New York, 1994.
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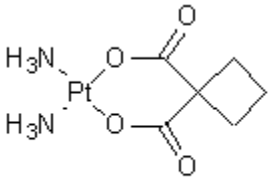
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12.5: Cisplatin 16. Profits from Cisplatin

Following the discovery by Barnett Rosenberg and coworkers at Michigan State University (see modules on electric fields, platinum electrodes, one scientist's story) that certain platinum-containing compounds inhibited cell division and cured solid tumors resulting from cancer, cisplatin was approved by the Food and Drug Administration for the treatment of genitourinary tumors in 1978.1 Since then, Michigan State has collected over $160 million in royalties from cisplatin and a related drug, carboplatin, which was approved by the FDA in 1989 for the treatment of ovarian cancers.



cisplatin
(*cis*-diamminedichloroplatinum(II))



carboplatin

The distribution of the profits from these drugs has generated a great deal of controversy, resulting in two (unrelated) lawsuits. Let's provide some background, as well as some detail about these two lawsuits because they illustrate difficulties that can arise when universities and businesses cooperate.

After cisplatin was discovered by researchers at Michigan State, members of the university made an agreement with a pharmaceutical company, Bristol-Myers, to develop and produce the drug. This agreement was made possible with the help of a technology-licensing organization known today as Research Corporation Technology (RCT). RCT is a daughter company of Research Corporation, a charitable foundation founded in 1912, whose stated mission is to promote "the advancement and extension of technical and scientific investigation, research, experimentation, and education" by making grants.

In 1950, well before Barnett Rosenberg and his coworkers discovered the medical uses of cisplatin, Michigan State University and Research Corporation signed an agreement that stated that for any discoveries patented at Michigan State, the inventors would receive the first 15 percent of all royalties, and the remaining royalties would then be split evenly between the university and Research Corporation.2 In 1987, Research Corporation decided that investing and business activities related to "technology transfer" (such as the licensing agreements between universities where discoveries were made and companies where these inventions could be developed and marketed) could not be carried out by a charitable organization such as itself. For this reason, Research Corporation created a new daughter company to handle all its rights to patents and licenses. This new company was called Research Corporation Technology (RCT), an independent, nonprofit company that pays taxes and has no shareholders; the formation of this new company was made possible by the Tax Reform Act of 1986. Unlike its parent company, the new company (RCT) does not make grants; rather, revenues from RCT are used to finance licensing activities, patent litigation, and creation of new companies to commercialize inventions. RCT has agreements—like the one with Michigan State—with over 100 universities.2

In 1995, Michigan State University attempted to terminate its contract with RCT by suing them. There were several reasons that the university wanted to become independent of RCT. First, representatives for Michigan State accused RCT of using its share of the profits of cisplatin and carboplatin to "unconscionably" enrich itself—particularly because RCT was not using these profits to promote scientific research at universities; rather, RCT was using the funds to give salary bonuses to high-ranking officials at the corporation. In short, Michigan State contended that RCT was acting in a manner that was inconsistent with the charitable mission of its parent company, Research Corporation. Michigan State hoped to rectify the wrongs by seeking damages of $50 million each from Research Corporation and RCT. Second, representatives for the university stated that the services provided by RCT were no longer needed because there was enough in-house expertise at Michigan State. Finally, along with the damages sought by the university, there was a significant financial incentive for Michigan State to end its relationship with RCT.

Under the original agreement, the university and RCT evenly split the profits remaining after the inventors received their share. Under a termination clause in the original agreement, however, the university would be entitled to 70 percent of the remaining profits if its agreement with RCT ended. Both Michigan State and RCT profit handsomely from the royalty income generated by cisplatin and carboplatin: Nearly all of the university's royalty income comes from cisplatin and carboplatin, and nearly 20 percent of RCT revenues comes from the two drugs.2 Because of this monetary component to the lawsuit, many people—including officials at other universities having contracts with RCT, as well as Barnett Rosenberg, whose research group discovered cisplatin and carboplatin—perceived that greed, not moral principle, was the primary driving force for the lawsuit. One official at another

midwestern university having a similar agreement with RCT stated that "a contract is a contract" and that "it was good enough when they signed up." Likewise, Barnett Rosenberg, who has retired from Michigan State and now conducts research at a private research institute in Michigan, said that he didn't think that he "would have acted the way the university did." He is proud of his role in the development of these important anticancer drugs but has stated that his efforts have "led to the creation of a lot of selfish, money-hungry university personnel." He said that he puts most of the money he receives from royalties of the drugs toward research projects at the Barros Institute in Lansing Michigan of which he is president.2 Now, a settlement has been reached between Michigan State and Research Corporation Technology.

Exact details of the settlement are not available; however, it is known that RCT will continue to manage the patents and licensing for cisplatin and carboplatin—but under new terms. Furthermore, the allocation of royalty income from the two drugs has changed: Michigan State will now receive more of the royalties from carboplatin, which now earns nearly six times more money in royalties than cisplatin. (See Figure 1.)

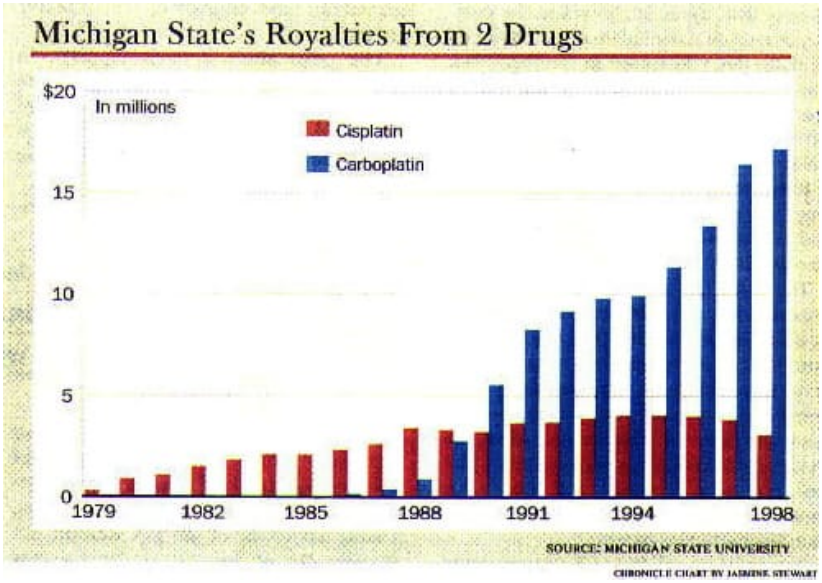


Figure 1. After FDA approval of carboplatin for the treatment of ovarian cancers in 1989, it outpaced cisplatin in earning royalties. Reprinted with permission. 3

Finally, RCT will pay Michigan State an additional $4.5 million over the next 2 years. Neither Michigan State University nor Research Corporation Technology commented extensively on the settlement, but each organization issued written statements that it had resolved its differences with the other.2 An additional benefit of the settlement for Michigan State is that the university will be represented by RCT in a second lawsuit arising from disagreements about profits from the anticancer drugs. In October 1996, a new patent was issued that allowed Michigan State to continue to receive royalties after the original cisplatin patent (issued in 1979) expired in 1996. In addition, Bristol-Myers Squibb will retain exclusive rights to the drug, as well as the obligation to pay royalties to Michigan State, until 2012. The second lawsuit has been brought by four generic-drug companies that want to manufacture their own, presumably less expensive, versions of cisplatin. Manufacture of these generic drugs would not require royalty payments to Michigan State. Lawyers for the four generic-drug companies contend that the 1996 patent should never have been issued because its fundamental claim is nearly the same as that made in the original patents. Double patenting is illegal. Because of their recent settlement with RCT, Michigan State will not be directly involved in this lawsuit because RCT is responsible for protecting the cisplatin patent.2

In short, cisplatin has produced not only physical, but also monetary, side effects. In the past few years, there have been two lawsuits contesting how the profits of cisplatin should be distributed.

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