CH100: EVERYDAY CHEMISTRY



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CH100: Everyday Chemistry (Southeast)

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Licensing

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CHAPTER OVERVIEW

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1.1: Basic Definitions

Learning Objective

- Learn the basic terms used to describe matter

The definition of chemistry—the study of the interactions of matter with other matter and with energy—uses some terms that should also be defined. We start the study of chemistry by defining basic terms.

Matter

Matter is anything that has mass and takes up space. A book is matter, a computer is matter, food is matter, and dirt in the ground is matter. Sometimes matter may be difficult to identify. For example, air is matter, but because it is so thin compared to other matter (e.g., a book, a computer, food, and dirt), we sometimes forget that air has mass and takes up space. Things that are not matter include thoughts, ideas, emotions, and hopes.

Example 1.1.1

Which of the following is matter and not matter?

a. a hot dog
b. love
c. a tree

Solution

a. A hot dog has mass and takes up space, so it is matter.
b. Love is an emotion, and emotions are not matter.
c. A tree has mass and takes up space, so it is matter.

Exercise 1.1.1

Which of the following is matter and not matter?

a. the moon
b. an idea for a new invention

Answer a

The moon is matter.

Answer b

The invention itself may be matter, but the idea for it is not.

To understand matter and how it changes, we need to be able to describe matter. There are two basic ways to describe matter: physical properties and chemical properties.

Physical properties

Physical properties are characteristics that describe matter as it exists. Some physical characteristics of matter are shape, color, size, and temperature. An important physical property is the **phase** (or **state**) of matter. The three fundamental phases of matter are solid, liquid, and gas (Figure **1.1.1**).



Figure 1.1.1: The Phases of Matter. Chemistry recognizes three fundamental phases of matter: solid (left), liquid (middle), and gas (right). (CC BY-SA 3.0; Spirit469). The solid state depicted is an ice sculpture of an angel, the liquid state is a drop of water, and the gas is clouds made of water vapor.

Chemical Properties

Chemical properties are characteristics of matter that describe how matter changes form in the presence of other matter. Does a sample of matter burn? Burning is a chemical property. Does it behave violently when put in water? This reaction is a chemical property as well (Figure 1.1.2). In the following chapters, we will see how descriptions of physical and chemical properties are important aspects of chemistry.



Figure 1.1.2: Chemical Properties. The fact that this match burns is a chemical property of the match. (Sebastian Ritter (Rise0011)).

Physical Change

A physical change occurs when a sample of matter changes one or more of its physical properties. For example, a solid may melt (Figure 1.1.3), or alcohol in a thermometer may change volume as the temperature changes. A physical change does not affect the chemical composition of matter.



Figure 1.1.2: Physical Changes: The solid ice melts into liquid water—a physical change. A time-lapse animation of ice cubes melting in a glass over 50 minutes. (Public Domain; Moussa).

Chemical Change

Chemical change is the process of demonstrating a chemical property, such as the burning match in Figure 1.1.2 "Chemical Properties". As the matter in the match burns, its chemical composition changes, and new forms of matter with new physical properties are created. Note that chemical changes are frequently accompanied by physical changes, as the new matter will likely have different physical properties from the original matter.

Example 1.1.2

Describe each process as a physical change or a chemical change.

a. Water in the air turns into snow.
b. A person's hair is cut.
c. Bread dough becomes fresh bread in an oven.

Solution

a. Because the water is going from a gas phase to a solid phase, this is a physical change.
b. Your long hair is being shortened. This is a physical change.
c. Because of the oven's temperature, chemical changes are occurring in the bread dough to make fresh bread. These are chemical changes. (In fact, a lot of cooking involves chemical changes.)

Exercise 1.1.2

Identify each process as a physical change or a chemical change.

a. A fire is raging in a fireplace.
b. Water is warmed to make a cup of coffee.

Answer a

chemical change

Answer b

physical change

Substance

A sample of matter that has the same physical and chemical properties throughout is called a substance. Sometimes the phrase *pure substance* is used, but the word *pure* isn't needed. The definition of the term *substance* is an example of how chemistry has a specific definition for a word that is used in everyday language with a different, vaguer definition. Here, we will use the term *substance* with its strict chemical definition.

> *Chemistry recognizes two different types of substances: elements and compounds.*

Element

An element is the simplest type of chemical substance; it cannot be broken down into simpler chemical substances by ordinary chemical means. There are 118 elements known to science, of which 80 are stable. (The other elements are radioactive, a condition we will consider in Chapter 15.) Each element has its own unique set of physical and chemical properties. Examples of elements include iron, carbon, and gold.

Compound

A compound is a combination of more than one element. The physical and chemical properties of a compound are different from the physical and chemical properties of its constituent elements; that is, it behaves as a completely different substance. There are over 50 million compounds known, and more are being discovered daily. Examples of compounds include water, penicillin, and sodium chloride (the chemical name for common table salt).

Mixtures

Physical combinations of more than one substance are called mixtures. Elements and compounds are not the only ways in which matter can be present. We frequently encounter objects that are physical combinations of more than one element or compound—mixtures. There are two types of mixtures.

Heterogeneous Mixture

A heterogeneous mixture is a mixture composed of two or more substances. It is easy to tell, sometimes by the naked eye, that more than one substance is present.

Homogeneous Mixture/ Solution

A homogeneous mixture is a combination of two or more substances that is so intimately mixed, that the mixture behaves as a single substance. Another word for a homogeneous mixture is a solution. Thus, a combination of salt and steel wool is a heterogeneous mixture because it is easy to see which particles of the matter are salt crystals and which are steel wool. On the other hand, if you take salt crystals and dissolve them in water, it is very difficult to tell that you have more than one substance present just by looking—even if you use a powerful microscope. The salt dissolved in water is a homogeneous mixture, or a solution (Figure **1.1.3**).



Figure **1.1.3**: Types of Mixtures © Thinkstock. On the left, the combination of two substances is a heterogeneous mixture because the particles of the two components look different. On the right, the salt crystals have dissolved in the water so finely that you cannot tell that salt is present. The homogeneous mixture appears like a single substance.

✔ Example 1.1.3

Identify the following combinations as heterogeneous mixtures or homogenous mixtures.

a. soda water (carbon dioxide is dissolved in water)
b. a mixture of iron metal filings and sulfur powder (both iron and sulfur are elements)

Solution

a. Because carbon dioxide is dissolved in water, we can infer from the behavior of salt crystals dissolved in water that carbon dioxide dissolved in water is (also) a homogeneous mixture.
b. Assuming that the iron and sulfur are simply mixed together, it should be easy to see what is iron and what is sulfur, so this is a heterogeneous mixture.

? Exercise 1.1.3

a. the human body
b. an amalgam, a combination of some other metals dissolved in a small amount of mercury

Answer a

heterogeneous mixture

Answer b

homogeneous mixture

There are other descriptors that we can use to describe matter, especially elements. We can usually divide elements into metals and nonmetals, and each set shares certain (but not always all) properties.

Metal

A metal is an element that conducts electricity and heat well and is shiny, silvery, solid, ductile, and malleable. At room temperature, metals are solid (although mercury is a well-known exception). A metal is ductile because it can be drawn into thin wires (a property called *ductility*); and malleable because it can be pounded into thin sheets (a property called *malleability*).

Nonmetal

A non-metal is an element that is brittle when solid, and does not conduct electricity or heat very well. Non-metals cannot be made into thin sheets or wires (Figure **1.1.4**). Nonmetals also exist in a variety of phases and colors at room temperature.

Semi-metals

Some elements have properties of both metals and nonmetals and are called semi-metals (or metalloids). We will see later how these descriptions can be assigned rather easily to various elements.



Figure **1.1.4**: Semimetals © Thinkstock. On the left is some elemental mercury, the only metal that exists as a liquid at room temperature. It has all the other expected properties of a metal. On the right, elemental sulfur is a yellow nonmetal that usually is found as a powder.

Describing Matter Flowchart

"Describing Matter" is a flowchart of the relationships among the different ways of describing matter.

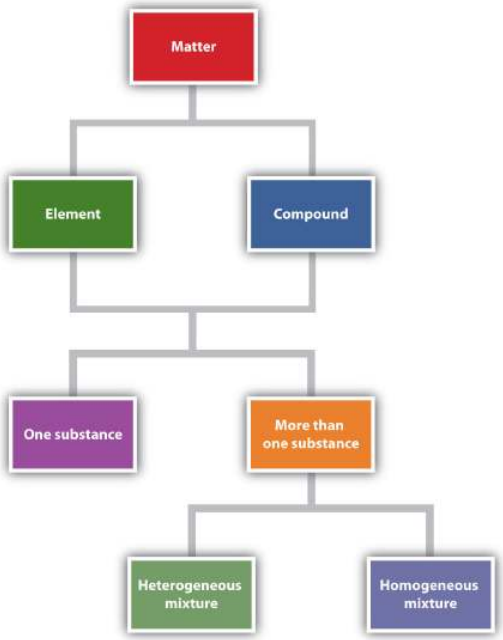


Figure 1.1.5: Describing Matter. This flowchart shows how matter can be described. Matter forks into element & compound which fork into 1 substance & more than 1 substance. More than 1 substance is homogenous or heterogenous.

Example 1.1.1: Chemistry is Everywhere: In the Morning

Most people have a morning ritual, a process that they go through every morning to get ready for the day. Chemistry appears in many of these activities.

- If you take a shower or bath in the morning, you probably use soap, shampoo, or both. These items contain chemicals that interact with the oil and dirt on your body and hair to remove them and wash them away. Many of these products also contain chemicals that make you smell good; they are called *fragrances*.
- When you brush your teeth in the morning, you usually use toothpaste, a form of soap, to clean your teeth. Toothpastes typically contain tiny, hard particles called *abrasives* that physically scrub your teeth. Many toothpastes also contain fluoride, a substance that chemically interacts with the surface of the teeth to help prevent cavities.
- Perhaps you take vitamins, supplements, or medicines every morning. Vitamins and other supplements contain chemicals your body needs in small amounts to function properly. Medicines are chemicals that help combat diseases and promote health.
- Perhaps you make some fried eggs for breakfast. Frying eggs involves heating them enough so that a chemical reaction occurs to cook the eggs.
- After you eat, the food in your stomach is chemically reacted so that the body (mostly the intestines) can absorb food, water, and other nutrients.
- If you drive or take the bus to school or work, you are using a vehicle that probably burns gasoline, a material that burns fairly easily and provides energy to power the vehicle. Recall that burning is a chemical change.

These are just a few examples of how chemistry impacts your everyday life. And we haven't even made it to lunch yet!



Figure 1.1.6: Chemistry in Real Life © Thinkstock. Examples of chemistry can be found everywhere—in personal hygiene products, food, and motor vehicles. Personal hygiene products (left), food (middle), and motor vehicles (right) images next to one another.

Key Takeaways

- Chemistry is the study of matter and its interactions with other matter and energy.
- Matter is anything that has mass and takes up space.
- Matter can be described in terms of physical properties and chemical properties.
- Physical properties and chemical properties of matter can change.
- Matter is composed of elements and compounds.
- Combinations of different substances are called mixtures.
- Elements can be described as metals, nonmetals, and semi-metals.

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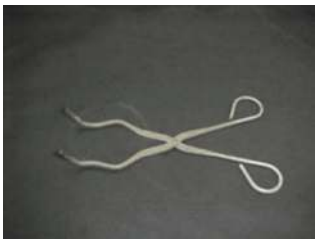
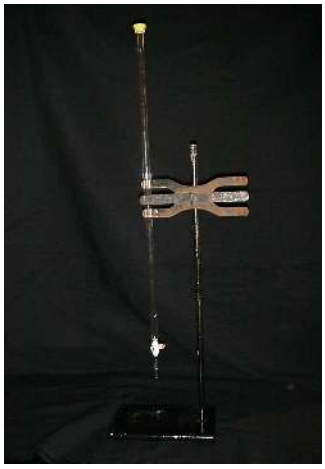
1.2: General Glassware

Beaker	Erlenmeyer Flask	Powder Funnel	Graduated Cylinder
			
Used by all sciences.	to contain solutions; can be easily swirled.	for pouring solids into a beaker or any container. Other kind of funnels can be used to pour liquid.	Used mainly for measuring liquid volume.

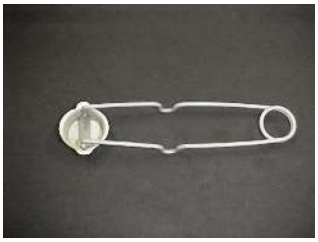
Side Arm Flask	Test Tubes	Watch glass
		
Commonly used in filtering processes involving vacuum filtration, often with a buchner funnel or crucible funnel. The inlet may be connected to a vacuum source, such as a water aspirator. It is also known as Suction flask.	Classified as TC (to contain), test tubes are used to hold samples. Shown at left are several placed in a test tube rack.	Commonly used to cover containers; also can be used for miscellaneous tasks

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1.3: General Chemistry Equipment

Mortar and Pestle	Beaker	Tongs	Buret
			
Good for grinding compounds.	Used by all sciences.	Tongs are often used to lift/move hot items. These are used in all of the various chemistries.	Used in titrations.

Rubber Policeman	Spatulas	Stir Bars	Evaporating Dish (Crucibles)
			
Used to scrape wet crystals from a flask. The rubber tip prevents scratching of the glass.	scoop substances in small amounts	used for stirring solution with the help of magnetic stir plate	These are used for heating things over a bunsen burner. They are ceramic, and so are relatively flame resistant. Often placed on a ring (on a ring stand) or on a wire mesh on a ring.

Cork Ring	Flint	Ring	Disposable Pipet
			
Used to hold round bottom flasks, to prevent them from rolling around and falling, and to keep them from touching the benchtop.	This is used for starting a flame or often a bunsen burner.	A ring is used to hold things. In organic chemistry this is oftem used to hold a separatory funnel. In general chemistry this is used to hold wire mesh to heat things on.	This is used in all of the chemistries for measuring out small quantities of a substance relatively imprecisely. Also often used as a dropper.

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1.4: Chemistry as a Science

Learning Objective

- Learn what science is and how it works.

Chemistry is a branch of science. Although science itself is difficult to define exactly, the following definition can serve as a starting point. Science is the process of knowing about the natural universe through observation and experiment. Science is not the only process of knowing (e.g., the ancient Greeks simply sat and *thought*), but it has evolved over more than 350 years into the best process that humanity has devised, to date, to learn about the universe around us.

The process of science is usually stated as the *scientific method*, which is rather naively described as follows:

1. state a hypothesis,
2. test the hypothesis, and
3. refine the hypothesis

In actuality, the process is not that simple. (For example, a scientist does not go into their lab every day and exclaim, "I am going to state a hypothesis today and spend the day testing it!") The process is not that simple because science and scientists have a body of knowledge that has already been identified as coming from the highest level of understanding, and most scientists build from that body of knowledge.

An educated guess about how the natural universe works is called a **hypothesis**. A scientist who is familiar with how part of the natural universe works—say, a chemist—is interested in furthering that knowledge. That person makes a reasonable guess—a hypothesis—that is designed to see if the universe works in a new way as well. Here's an example of a hypothesis: "if I mix one part of hydrogen with one part of oxygen, I can make a substance that contains both elements."

> *For a hypothesis to be termed a scientific hypothesis, it has to be something that can be supported or refuted through carefully crafted experimentation or observation.*

Most good hypotheses are grounded in previously understood knowledge and represent a testable extension of that knowledge. The scientist then devises ways to test if that hypothesis is correct or not. That is, the scientist plans experiments. **Experiments are tests of the natural universe to see if a guess (hypothesis) is correct**. An experiment to test our previous hypothesis would be to actually mix hydrogen and oxygen and see what happens. Most experiments include observations of small, well-defined parts of the natural universe designed to see results of the experiments.

A Scientific Hypothesis

A hypothesis is often written in the form of an if/then statement that gives a possibility (if) and explains what may happen because of the possibility (then). For example, if eating elemental sulfur repels ticks, then someone that is eating sulfur every day will not get ticks.



Why do we have to do experiments? Why do we have to test? Because the natural universe is not always so obvious, experiments are necessary. For example, it is fairly obvious that if you drop an object from a height, it will fall. Several hundred years ago (coincidentally, near the inception of modern science), the concept of gravity explained that test. However, is it obvious that the

entire natural universe is composed of only about 115 fundamental chemical building blocks called elements? This wouldn't seem true if you looked at the world around you and saw all the different forms matter can take. In fact, the concept of *the element* is only about 200 years old, and the last naturally occurring element was identified about 80 years ago. It took decades of tests and millions of experiments to establish what the elements actually are. These are just two examples; a myriad of such examples exists in chemistry and science in general.

When enough evidence has been collected to establish a general principle of how the natural universe works, the evidence is summarized in a theory. A theory is a general statement that explains a large number of observations. "All matter is composed of atoms" is a general statement, a theory, that explains many observations in chemistry. A theory is a very powerful statement in science. There are many statements referred to as "the theory of _____" or the "_____ theory" in science (where the blanks represent a word or concept). When written in this way, theories indicate that science has an overwhelming amount of evidence of its correctness. We will see several theories in the course of this text.

A specific statement that is thought to never be violated by the entire natural universe is called a law. A scientific law is the highest understanding of the natural universe that science has and is thought to be inviolate. The fact that all matter attracts all other matter —the law of gravitation—is one such law. Note that the terms *theory* and *law* used in science have slightly different meanings from those in common usage; where theory is often used to mean hypothesis ("I have a theory…"), and a law is an arbitrary limitation that can be broken but with potential consequences (such as speed limits). Here again, science uses these terms differently, and it is important to apply their proper definitions when you use these words in science. (Figure **1.4.1**)

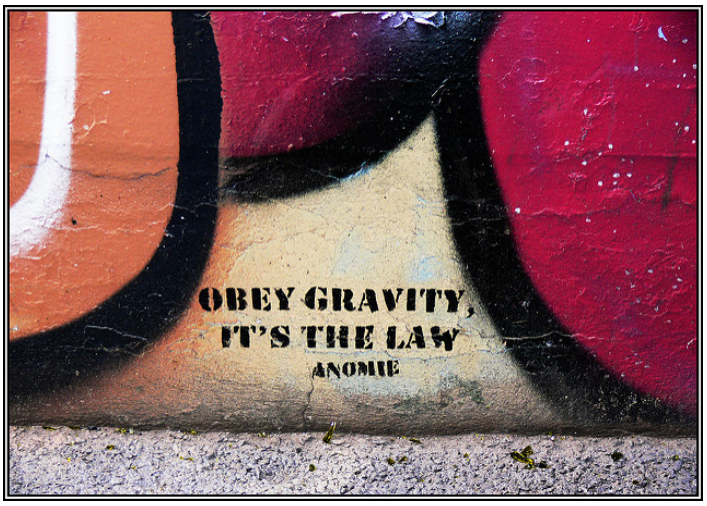


Figure **1.4.1**: Defining a law. Does this graffiti mean "law" the way science defines "law"? (CC BY-SA-NC-ND; Koppenbadger).

There is an additional phrase in our definition of science: "the natural universe." Science is concerned *only* with the natural universe. What is the natural universe? It's anything that occurs around us, well, naturally. Stars, planets, the appearance of life on earth; as well as how animals, plants, and other matter function are all part of the natural universe. Science is concerned with that—and *only* that.

Of course, there are other things that concern us. For example, is the English language part of science? Most of us can easily answer no; English is not science. English is certainly worth knowing (at least for people in predominantly English-speaking countries), but why isn't it science? English, or any human language, is not science because ultimately it is *contrived*; it is made up. Think of it: the word spelled b-l-u-e represents a certain color, and we all agree what color that is. But what if we used the word h-a-r-d to describe that color? (Figure **1.4.2**) That would be fine—as long as everyone agreed. Anyone who has learned a second language must initially wonder why a certain word is used to describe a certain concept; ultimately, the speakers of that language agreed that a particular word would represent a particular concept. It was contrived.

That doesn't mean language isn't worth knowing. It is very important in society. But it's not *science*. Science deals only with what occurs naturally.

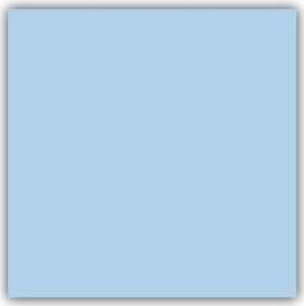


Figure 1.4.2: English Is Not Science. How would you describe this color? Blue or hard? Either way, you're not doing science.

Example 1.4.1: Identifying Science

Which of the following fields would be considered science?

a. geology, the study of the earth
b. ethics, the study of morality
c. political science, the study of governance
d. biology, the study of living organisms

Solution

a. Because the earth is a natural object, the study of it is indeed considered part of science.
b. Ethics is a branch of philosophy that deals with right and wrong. Although these are useful concepts, they are not science.
c. There are many forms of government, but all are created by humans. Despite the fact that the word *science* appears in its name, political science is not true science.
d. Living organisms are part of the natural universe, so the study of them is part of science.

Exercise 1.4.1

Which is part of science, and which is not?

a. dynamics, the study of systems that change over time
b. aesthetics, the concept of beauty

Answer A

science

Answer B

not science

The field of science has gotten so big that it is common to separate it into more specific fields. First, there is mathematics, the language of science. All scientific fields use mathematics to express themselves—some more than others. Physics and astronomy are scientific fields concerned with the fundamental interactions between matter and energy. Chemistry, as defined previously, is the study of the interactions of matter with other matter and with energy. Biology is the study of living organisms, while geology is the study of the earth. Other sciences can be named as well. Understand that these fields are not always completely separate; the boundaries between scientific fields are not always readily apparent. A scientist may be labeled a biochemist if he or she studies the chemistry of biological organisms.

Finally, understand that science can be either qualitative or quantitative. Qualitative implies a description of the quality of an object. For example, physical properties are generally qualitative descriptions: sulfur is yellow, your math book is heavy, or that statue is pretty. A quantitative description represents the specific amount of something; it means knowing how much of something is present, usually by counting or measuring it. Some quantitative descriptions include: 25 students in a class, 650 pages in a book, or a velocity of 66 miles per hour. Quantitative expressions are very important in science; they are also very important in chemistry.

Example 1.4.2: qualitative vs. quantitative Descriptions

Identify each statement as either a qualitative description or a quantitative description.

a. Gold metal is yellow.
b. A ream of paper has 500 sheets in it.
c. The weather outside is snowy.
d. The temperature outside is 24 degrees Fahrenheit.

Solution

a. Because we are describing a physical property of gold, this statement is qualitative.
b. This statement mentions a specific amount, so it is quantitative.
c. The word *snowy* is a description of how the day is; therefore, it is a qualitative statement.
d. In this case, the weather is described with a specific quantity—the temperature. Therefore, it is quantitative.

Exercise 1.4.2

Are these qualitative or quantitative statements?

a. Roses are red, and violets are blue.
b. Four score and seven years ago….

Answer A

qualitative

Answer B

quantitative

Food and Drink Application: Carbonated Beverages

Some of the simple chemical principles discussed in this chapter can be illustrated with carbonated beverages: sodas, beer, and sparkling wines. Each product is produced in a different way, but they all have one thing in common: they are solutions of carbon dioxide dissolved in water.

Carbon dioxide is a compound composed of carbon and oxygen. Under normal conditions, it is a gas. If you cool it down enough, it becomes a solid known as dry ice. Carbon dioxide is an important compound in the cycle of life on earth.

Even though it is a gas, carbon dioxide can dissolve in water, just like sugar or salt can dissolve in water. When that occurs, we have a homogeneous mixture, or a solution, of carbon dioxide in water. However, very little carbon dioxide can dissolve in water. If the atmosphere were pure carbon dioxide, the solution would be only about 0.07% carbon dioxide. In reality, the air is only about 0.03% carbon dioxide, so the amount of carbon dioxide in water is reduced proportionally.

However, when soda and beer are made, manufacturers do two important things: they use pure carbon dioxide gas, and they use it at very high pressures. With higher pressures, more carbon dioxide can dissolve in the water. When the soda or beer container is sealed, the high pressure of carbon dioxide gas remains inside the package. (Of course, there are more ingredients in soda and beer besides carbon dioxide and water.)

When you open a container of soda or beer, you hear a distinctive *hiss* as the excess carbon dioxide gas escapes. But something else happens as well. The carbon dioxide in the solution comes out of solution as a bunch of tiny bubbles. These bubbles impart a pleasing sensation in the mouth, so much so that the soda industry sold over *225 billion* servings of soda in the United States alone in 2009.

Some sparkling wines are made in the same way—by forcing carbon dioxide into regular wine. Some sparkling wines (including champagne) are made by sealing a bottle of wine with some yeast in it. The yeast *ferments*, a process by which the yeast converts sugars into energy and excess carbon dioxide. The carbon dioxide produced by the yeast dissolves in the wine.

Then, when the champagne bottle is opened, the increased pressure of carbon dioxide is released, and the drink bubbles just like an expensive glass of soda.



Figure **1.4.3**: Carbonated Beverages © Thinkstock. Soda, beer, and sparkling wine take advantage of the properties of a solution of carbon dioxide in water.

Soda (left), beer (middle) and sparkling wine (right) images next to one another.

Key Takeaways

- Science is a process of knowing about the natural universe through observation and experiment.
- Scientists go through a rigorous process to determine new knowledge about the universe; this process is generally referred to as the scientific method.
- Science is broken down into various fields, of which chemistry is one.
- Science, including chemistry, is both qualitative and quantitative.

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1.5: Expressing Numbers

Learning Objective

- Learn to express numbers properly

Quantities have two parts: the number and the unit. The number tells "how many." It is important to be able to express numbers properly so that the quantities can be communicated properly.

Standard Notation

Standard notation is the straightforward expression of a number. Numbers such as 17, 101.5, and 0.00446 are expressed in standard notation. For relatively small numbers, standard notation is fine. However, for very large numbers, such as 306,000,000, or for very small numbers, such as 0.000000419, standard notation can be cumbersome because of the number of zeros needed to place nonzero numbers in the proper position.

Scientific Notation

Scientific notation is an expression of a number using powers of 10. Powers of 10 are used to express numbers that have many zeros:

Table shows the powers of 10. Left side has 10 to a specific power and right side has the equation for the specified power.

10^0	= 1
10^1	= 10
10^2	= 100 = 10 × 10
10^3	= 1,000 = 10 × 10 × 10
10^4	= 10,000 = 10 × 10 × 10 × 10

and so forth.

The raised number to the right of the 10 indicates the number of factors of 10 in the original number. (Scientific notation is sometimes called *exponential notation*.) The exponent's value is equal to the number of zeros in the number expressed in standard notation.

Small numbers can also be expressed in scientific notation but with negative exponents:

Table shows the powers of 10 with negative exponents. Left side has 10 to a specific power and right side has the equation for the specified power.

10^{-1}	= 0.1 = 1/10
10^{-2}	= 0.01 = 1/100
10^{-3}	= 0.001 = 1/1,000
10^{-4}	= 0.0001 = 1/10,000

and so forth. Again, the value of the exponent is equal to the number of zeros in the denominator of the associated fraction. A negative exponent implies a decimal number less than one.

A number is expressed in scientific notation by writing the first nonzero digit, then a decimal point, and then the rest of the digits. The part of a number in scientific notation that is multiplied by a power of 10 is called the *coefficient*. We determine the power of 10 needed to make that number into the original number and multiply the written number by the proper power of 10. For example, to write 79,345 in scientific notation,

$$79,345 = 7.9345 \times 10,000 = 7.9345 \times 10^4$$

Thus, the number in scientific notation is 7.9345×10^4. For small numbers, the same process is used, but the exponent for the power of 10 is negative:

$$0.000411 = 4.11 \times \frac{1}{10,000} = 4.11 \times 10^{-4}$$

Typically, the extra zero digits at the end or the beginning of a number are not included (Figure 1.5.1).



Figure 1.5.1: "The Blue Marble" is a famous photograph of the Earth taken on December 7, 1972, by the crew of the Apollo 17 spacecraft en route to the Moon at a distance of about 29,000 kilometers. This is 2.9×10^4 *km*. It shows Africa, Antarctica, and the Arabian Peninsula.The earth is about 93,000,000 miles from the sun. In scientific notation, this is 9.3×10^7 *miles*. (Public Domain; NASA/Apollo 17 crew; taken by either Harrison Schmitt or Ron Evans.)

Example 1.5.1: Expressing Numbers in Scientific Notation

Express these numbers in scientific notation.

a. 306,000
b. 0.00884
c. 2,760,000
d. 0.000000559

Solution

a. The number 306,000 is 3.06 times 100,000, or 3.06 times 10^5. In scientific notation, the number is 3.06×10^5.
b. The number 0.00884 is 8.84 times 1/1,000, which is 8.84 times 10^{-3}. In scientific notation, the number is 8.84×10^{-3}.
c. The number 2,760,000 is 2.76 times 1,000,000, which is the same as 2.76 times 10^6. In scientific notation, the number is written as 2.76×10^6. Note that we omit the zeros at the end of the original number.
d. The number 0.000000559 is 5.59 times 1/10,000,000, which is 5.59 times 10^{-7}. In scientific notation, the number is written as 5.59×10^{-7}.

Exercise 1.5.1

Express these numbers in scientific notation.

a. 23,070
b. 0.0009706

Answer a

2.307×10^4

Answer b

9.706×10^{-4}

Another way to determine the power of 10 in scientific notation is to count the number of places you need to move the decimal point to get a numerical value between 1 and 10. The number of places equals the power of 10. This number is positive if you move the decimal point to the right and negative if you move the decimal point to the left:

$$56,900 = 5.69 \times 10^{4} \qquad 0.000028 = 2.8 \times 10^{-5}$$

4 3 2 1 (for 56,900) 1 2 3 4 5 (for 0.000028)

56900 can be written as 5.69 times 10 to the fourth power. 0.000028 can be written as 2.8 times 10 to the negative fifth power.

Many quantities in chemistry are expressed in scientific notation. When performing calculations, you may have to enter a number in scientific notation into a calculator. Be sure you know how to correctly enter a number in scientific notation into your calculator. Different models of calculators require different actions for properly entering scientific notation. If in doubt, consult your instructor immediately (Figure **1.5.2**).

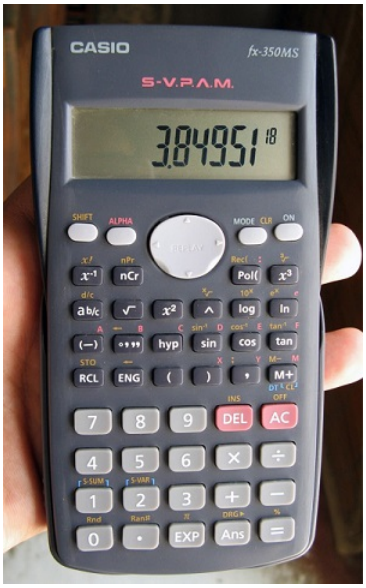


Figure **1.5.2**: This calculator shows only the coefficient and the power of 10 to represent the number in scientific notation. Thus, the number being displayed is $\mathbf{3.84951 \times 10^{18}}$, or $\mathbf{3,849,510,000,000,000,000}$. Source: "Casio" Asim Bijarani is licensed under Creative Commons Attribution 2.0 Generic.

Key Takeaways

- Standard notation expresses a number normally.
- Scientific notation expresses a number as a coefficient times a power of 10.
- The power of 10 is positive for numbers greater than 1 and negative for numbers between 0 and 1.

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1.6: Significant Figures - Writing Numbers to Reflect Precision

Learning Objectives

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

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

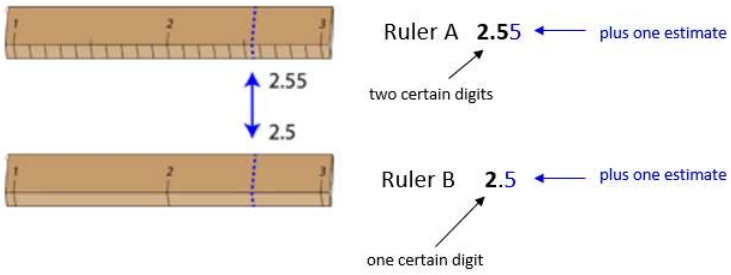


Figure 1.6.1: Measurement with two different rulers.
Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

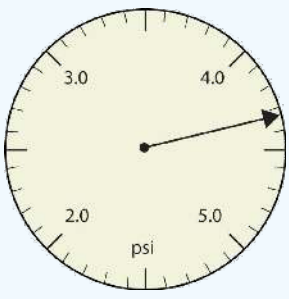
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest $0.1\ \text{g}$, other highly sensitive balances are capable of measuring to the nearest $0.001\ \text{g}$ or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 1.6.1 shows two rulers making the same measurement of an object (indicated by the blue arrow).

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

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

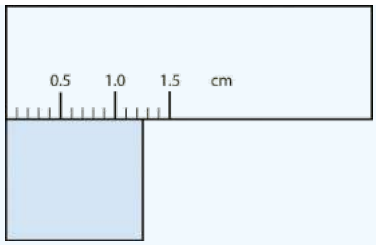
Example 1.6.1: Reporting Measurements to the Proper Number of Significant Figures

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



a.

b.



Ruler measuring a rectangle in units of centimeters, with the rectangle's edge between 1.2 and 1.3 cm marks

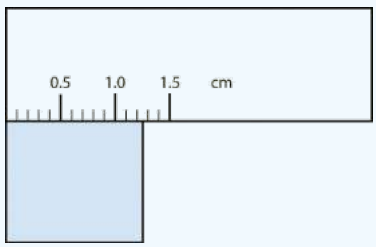
Solutions

Solutions to Example 2.3.1

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

? Exercise 1.6.1

What would be the reported width of this rectangle?



Answer

1.25 cm

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

Table **1.6.1**: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• 237 has three significant figures. • 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• 1400 is ambiguous. ◦ 1.4×10^3 has two significant figures. ◦ 1.40×10^3 three significant figures. ◦ 1.400×10^3 has four significant figures.
5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• 620.0 has four significant figures. • 19.000 has five significant figures.

Exact Numbers

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

Example 1.6.2

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

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

Solution

Solution to Example 2.3.2

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

? Exercise 1.6.2

Give the number of significant figures in each.

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

Answer a

three significant figures

Answer b

four significant figures

Answer c

four significant figures

Answer d

infinite (exact number)

Answer e

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

Accuracy and Precision

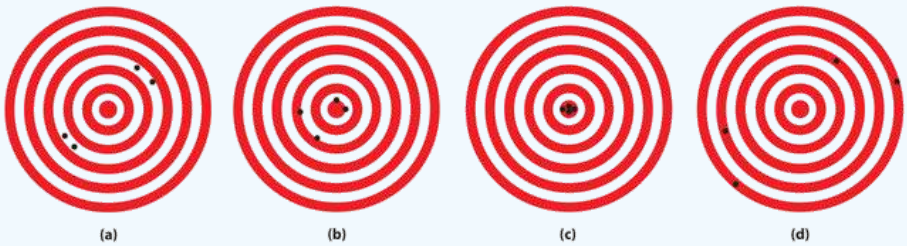
Measurements may be accurate, meaning that the measured value is the same as the true value; they may be precise, meaning that multiple measurements give nearly identical values (i.e., reproducible results); they may be both accurate and precise; or they may be neither accurate nor precise. The goal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



Video **1.6.1**: Difference between precision and accuracy.

Example 1.6.3

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



Which target shows

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

Solution

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

Summary

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

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1.7: The Basic Units of Measurement

Learning Objectives

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

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

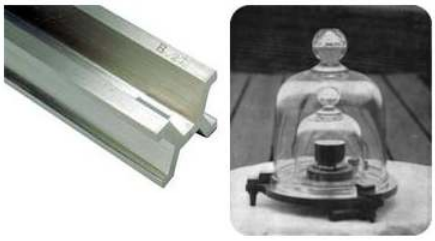


Figure **1.7.1**: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table **1.7.1**.

Table **1.7.1**: SI Base Units of Measurement

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

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

Unfortunately, the Metric System is Not Ubiquitous

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

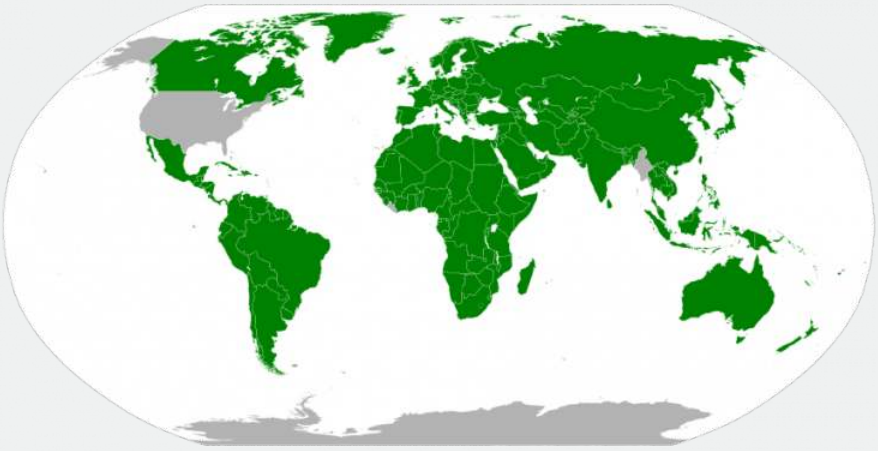


Figure 1.7.1: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word μεγας, meaning "great". Table 1.7.2 lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 1.7.2: SI Prefixes

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

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

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

Example 1.7.1: Unit Abbreviations

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

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

Solutions

Solutions to Example 2.5.1

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

Exercise 1.7.1

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

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

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Summary

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

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1.8: Problem Solving and Unit Conversions

Learning Objectives

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

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

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can convert meters into centimeters, or a more complex one can convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you are working with (Table 1.8.1).

Table 1.8.1: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	29.6 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume

*Pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.

Of course, there are other ratios which are not listed in Table 1.8.1. They may include:

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

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

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

or

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

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

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

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

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

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

Performing Dimensional Analysis

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

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

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

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:

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

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

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

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

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

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

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 1.8.1 shows a concept map for constructing a proper conversion.

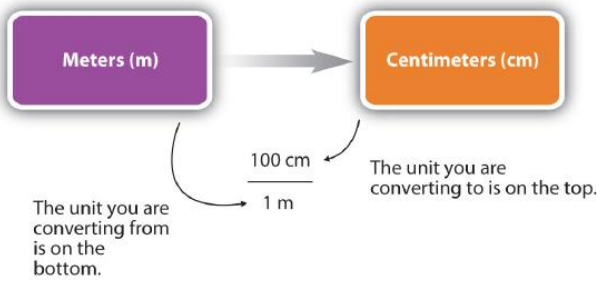


Figure 1.8.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Converting meters to centimeters: the unit you are converting is on the bottom of the fraction, and the unit to convert to is on top

General Steps in Performing Dimensional Analysis

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

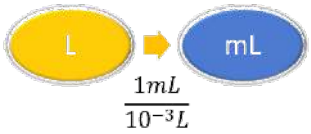
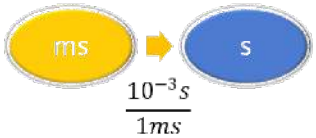
Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

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

Example 1.8.1

Steps for Problem Solving for Example 2.6.1 and 2.6.2

	Example 1.8.1	Example 1.8.2
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities.	$1\,mL = 10^{-3} L$	$1\,ms = 10^{-3} s$
Prepare a concept map and use the proper conversion factor.	 $\frac{1mL}{10^{-3}L}$	 $\frac{10^{-3}s}{1ms}$
Cancel units and calculate.	$4.7 \cancel{L} \times \frac{1\ \text{mL}}{10^{-3}\ \cancel{L}} = 4,700\ \text{mL}$ or $4.7 \cancel{L} \times \frac{1,000\ \text{mL}}{1\ \cancel{L}} = 4,700\ \text{mL}$ or 4.7 x 10^3 2SF, not ambiguous	$18\ \cancel{\text{ms}} \times \frac{10^{-3}\ \text{s}}{1\ \cancel{\text{ms}}} = 0.018\ \text{s}$ or $18\ \cancel{\text{ms}} \times \frac{1\ \text{s}}{1,000\ \cancel{\text{ms}}} = 0.018\ \text{s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

? Exercise 1.8.1

Perform each conversion.

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

Answer a:

$1.01000x10^{-4}s$

Answer b:

$3.208x10^{4}g$

Answer c:

$1.53x10^{2}cg$

Summary

- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce the appropriate units.

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1.9: Temperature and Density

Learning Objective

- Learn about the various temperature scales that are commonly used in chemistry.
- Define density and use it as a conversion factor.

There are other units in chemistry that are important, and we will cover others over the course of the entire book. One of the fundamental quantities in science is temperature. Temperature is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called degrees, and there are several temperature scales in use. In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as "degrees Fahrenheit"). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32°F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212°F.

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

$$^{\circ}\mathrm{C} = (^{\circ}\mathrm{F}\text{-}32) \times \tfrac{5}{9} \tag{1.9.1}$$

$$^{\circ}\mathrm{F} = (^{\circ}\mathrm{C} \times \tfrac{9}{5}) + 32 \tag{1.9.2}$$

Example 1.9.1: Conversions

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

Solution

a. Using Equation 1.9.1, we have

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

b. Using Equation 1.9.2, we have

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

Exercise 1.9.1

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

Answer a

−17.8 °C

Answer b

414 °F

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

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

For most purposes, it is acceptable to use 273 instead of 273.15. Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as "two hundred ninety-five kelvins" and not "two hundred ninety-five degrees Kelvin."

The reason that the Kelvin scale is defined this way is because there exists a minimum possible temperature called absolute zero (zero kelvins). The Kelvin temperature scale is set so that 0 K is absolute zero, and temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example.

✔ Example 1.9.2: Room Temperature

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

Solution

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

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

Then we use the appropriate formula above to determine the temperature in the Kelvin scale:

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

So, room temperature is about 295 K.

? Exercise 1.9.2

What is 98.6°F on the Kelvin scale?

Answer

310.2 K

Figure **1.9.1** compares the three temperature scales. Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. The other two countries are Liberia and Myanmar (formerly Burma). People driving near the borders of Canada or Mexico may pick up local radio stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.

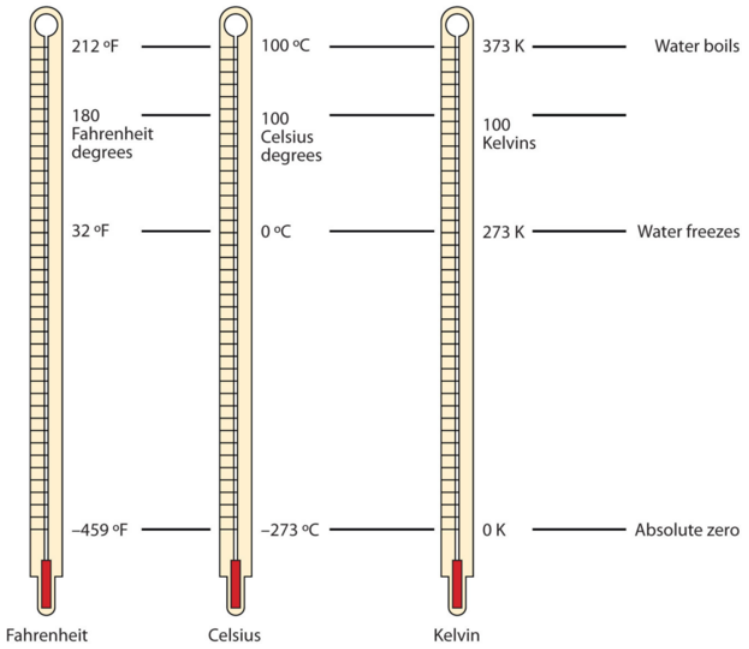


Figure 1.9.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales.

Density

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

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

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has derived units, just like velocity. Common units for density include g/mL, g/cm^3, g/L, kg/L, and even kg/m^3. Densities for some common substances are listed in Table 1.9.1.

Table 1.9.1: Densities of Some Common Substances

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

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (from Table 1.9.1), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

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

where we have limited our answer to two significant figures.

Example 1.9.3: Mercury

What is the mass of 44.6 mL of mercury?

Solution

Use the density from Table 1.9.1 "Densities of Some Common Substances" as a conversion factor to go from volume to mass:

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

The mass of the mercury is 607 g.

Exercise 1.9.3

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

Answer

197 g

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

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

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

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

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

$$\frac{1}{d} = \frac{1\,mL}{19.3\,g}$$

This reciprocal density is still a useful conversion factor, but now the mass unit will cancel and the volume unit will be introduced. Thus, if we want to know the volume of 45.9 g of gold, we would set up the conversion as follows:

$$45.9 \cancel{g} \times \frac{1\,mL}{19.3\,\cancel{g}} = 2.38\,mL$$

Note how the mass units cancel, leaving the volume unit, which is what we are looking for.

Example 1.9.4: Calculating Volume from Density

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/cm^3, what is the volume of the cork?

Solution

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

$$\frac{1}{d} = \frac{1\,cm^3}{0.22\,g}$$

We can use this expression as the conversion factor. So

$$3.78 \cancel{g} \times \frac{1\,cm^3}{0.22\,\cancel{g}} = 17.2\,cm^3$$

? Exercise 1.9.4

What is the volume of 3.78 g of gold?

Answer

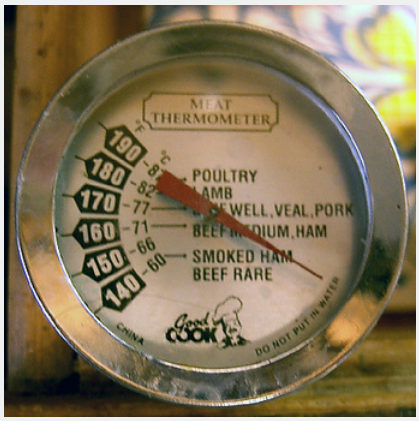
0.196 cm^3

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

Food and Drink Application: Cooking Temperatures

Because degrees Fahrenheit is the common temperature scale in the United States, kitchen appliances, such as ovens, are calibrated in that scale. A cool oven may be only 150°F, while a cake may be baked at 350°F and a chicken roasted at 400°F. The broil setting on many ovens is 500°F, which is typically the highest temperature setting on a household oven.

People who live at high altitudes, typically 2,000 ft above sea level or higher, are sometimes urged to use slightly different cooking instructions on some products, such as cakes and bread, because water boils at a lower temperature the higher in altitude you go, meaning that foods cook slower. For example, in Cleveland water typically boils at 212°F (100°C), but in Denver, the Mile-High City, water boils at about 200°F (93.3°C), which can significantly lengthen cooking times. Good cooks need to be aware of this.



A meat thermometer with a dial. Notice the markings for Fahrenheit (outer scale) and Celsius (inner scale) temperatures. Recipes for cooking food in an oven can use very different numbers, depending on the country you're in. (CC BY2.0 Bev Sykes)

At the other end is pressure cooking. A pressure cooker is a closed vessel that allows steam to build up additional pressure, which increases the temperature at which water boils. A good pressure cooker can get to temperatures as high as 252°F (122°C); at these temperatures, food cooks much faster than it normally would. Great care must be used with pressure cookers because of the high pressure and high temperature. (When a pressure cooker is used to sterilize medical instruments, it is called an *autoclave*.)

Other countries use the Celsius scale for everyday purposes. Therefore, oven dials in their kitchens are marked in degrees Celsius. It can be confusing for US cooks to use ovens abroad—a 425°F oven in the United States is equivalent to a 220°C oven in other countries. These days, many oven thermometers are marked with both temperature scales.

Key Takeaways

- Chemistry uses the Celsius and Kelvin scales to express temperatures.
- A temperature on the Kelvin scale is the Celsius temperature plus 273.15.
- The minimum possible temperature is absolute zero and is assigned 0 K on the Kelvin scale.
- Density relates the mass and volume of a substance.
- Density can be used to calculate volume from a given mass or mass from a given volume.

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1.10: Hypothesis, Theories, and Laws

Learning Objectives

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

Although many have taken science classes throughout the course of their studies, people often have incorrect or misleading ideas about some of the most important and basic principles in science. Most students have heard of hypotheses, theories, and laws, but what do these terms really mean? Prior to reading this section, consider what you have learned about these terms before. What do these terms mean to you? What do you read that contradicts or 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, depending 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 when you predicted the outcome of your experiment.
- Tentative or proposed explanation: hypotheses can be suggestions about why something is observed. In order for it to be scientific, however, 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, we should see ___ result when we perform ___ test."

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

What is a Theory?

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

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

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

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

> *Theories aren't likely to change.*

What is a Law?

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

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

What is a Belief?

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

Laws vs. Theories

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



Video 1.10.1: What's the difference between a scientific law and theory?

Summary

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

Contributions & Attributions

-
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1.11: Classification of Matter

Learning Objectives

- To classify matter.

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

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

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

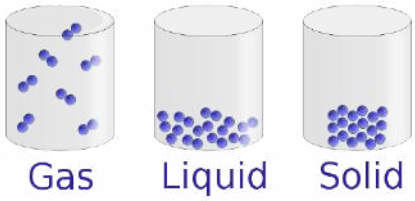


Figure **1.11.1**: The Three States of Matter. Solids have a defined shape and volume. Liquids have a fixed volume but flow to assume the shape of their containers. Gases completely fill their containers, regardless of volume. Figure used with permission from Wikipedia

Pure Substances and Mixtures

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

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



Figure 1.11.2: A Heterogeneous Mixture. Under a microscope, whole milk is actually a heterogeneous mixture composed of globules of fat and protein dispersed in water. Figure used with permission from Wikipedia

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

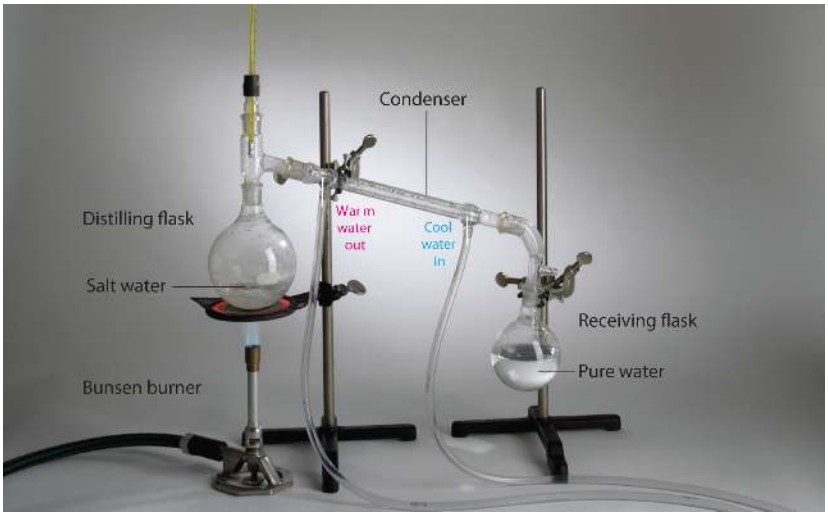


Figure 1.11.3: The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.

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

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

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



Figure **1.11.4**: The Crystallization of Sodium Acetate from a Concentrated Solution of Sodium Acetate in Water. The addition of a small “seed” crystal (a) causes the compound to form white crystals, which grow and eventually occupy most of the flask. Video can be found here: www.youtube.com/watch?v=BLq5NibwV5g

closeup of bulb flask containing liquid with many thin spike crystals radially emerging from the center.

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

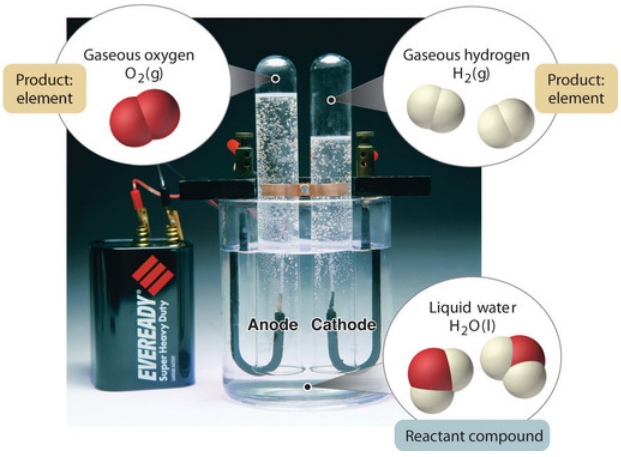


Figure **1.11.5**: The Decomposition of Water to Hydrogen and Oxygen by Electrolysis. Water is a chemical compound; hydrogen and oxygen are elements.

a battery wired to an anode and cathode placed in a beaker filled with water. Two inverted test tubes are submerged in the water and placed over each of the electrodes to collect the gaseous products. Magnifying pointers show the molecular structure of water in the beaker as well as the hydrogen gas collected on the anode side and oxygen gas on the cathode side.



Different Definitions of Matter: Different Definitions of Matter, YouTube (opens in new window) [youtu.be]

In general, a reverse chemical process breaks down compounds into their elements. For example, water (a compound) can be decomposed into hydrogen and oxygen (both elements) by a process called electrolysis. In electrolysis, electricity provides the energy needed to separate a compound into its constituent elements (Figure 1.11.5). A similar technique is used on a vast scale to obtain pure aluminum, an element, from its ores, which are mixtures of compounds. Because a great deal of energy is required for electrolysis, the cost of electricity is by far the greatest expense incurred in manufacturing pure aluminum. Thus recycling aluminum is both cost-effective and ecologically sound.

The overall organization of matter and the methods used to separate mixtures are summarized in Figure 1.11.6.

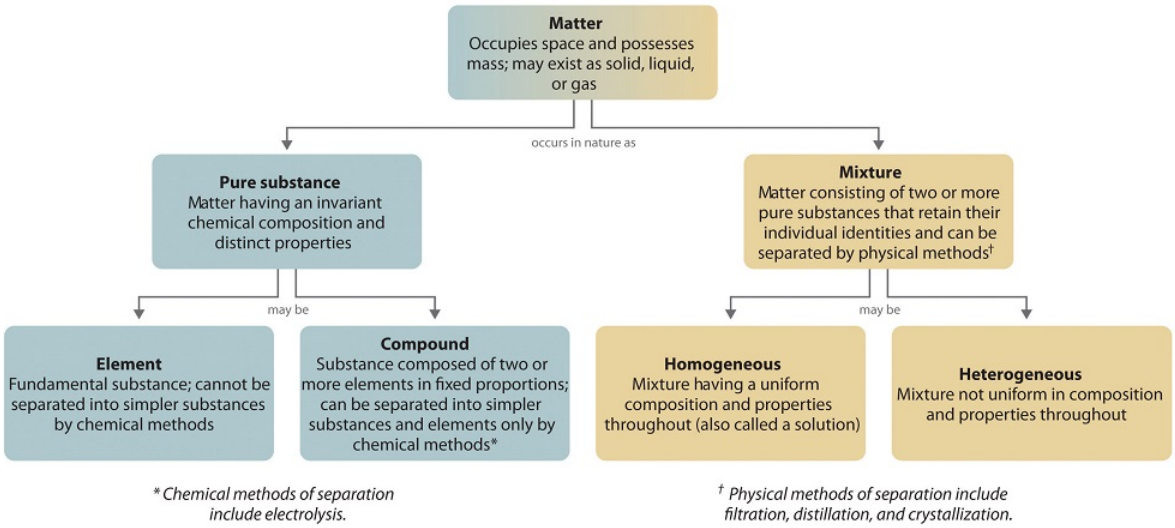


Figure 1.11.6: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Example 1.11.1

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

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

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.

B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

a. **A** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration. **B** Because the composition of the solution is uniform throughout, it is a homogeneous mixture.

b. **A** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure. B Because its composition is not uniform throughout, orange juice is a heterogeneous mixture.

c. **A** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence a compact disc is not chemically pure. **B** The regions of different composition indicate that a compact disc is a heterogeneous mixture.

d. **A** Aluminum oxide is a single, chemically pure compound.

e. **A** Selenium is one of the known elements.

? Exercise 1.11.1

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

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

Answer A

solution

Answer B

element

Answer C

heterogeneous mixture

Answer D

compound



Different Definitions of Changes: Different Definitions of Changes, YouTube(opens in new window) [youtu.be] (opens in new window)

Summary

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

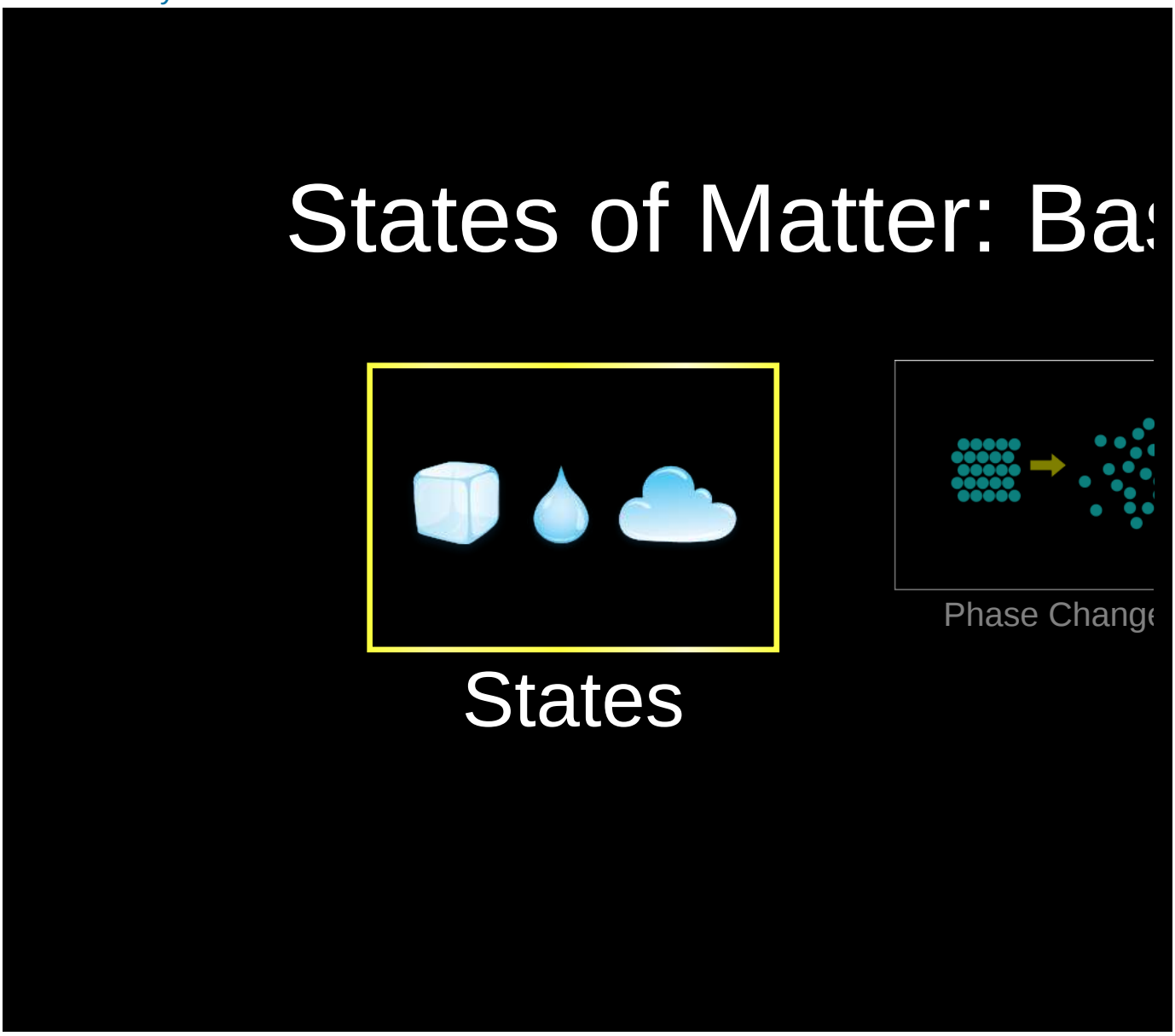
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1.12: Activity: States of Matter



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1.13: Compounds Display Constant Composition

When building a house, the starting point is a blueprint of what the house will look like. The plan states how many windows and what kind, how many doors and what style, how many rooms and what type (bedroom, kitchen, other). The blueprint shows how the different pieces will go together to make the house. As long as the blueprint is followed and exactly the same items are used, the house will be identical to its blueprint.

Compounds

A **compound** is a substance that contains two or more elements chemically combined in a fixed proportion. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen particles as carbon particles. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances—carbon and hydrogen.

Recall that the components of a mixture can be separated from one another by physical means. This is not true for a compound. Table salt is a compound consisting of equal parts of the elements sodium and chlorine. Salt cannot be separated into its two elements by filtering, distillation, or any other physical process. Salt and other compounds can only be decomposed into their elements by a chemical process. A **chemical change** is a change that produces matter with a different composition. Many compounds can be decomposed into their elements by heating. When sugar is heated, it decomposes into carbon and water. Water is still a compound, but one which cannot be broken down into hydrogen and oxygen by heating. Instead, the passage of an electrical current through water will produce hydrogen and oxygen gases.

The properties of compounds are generally very different than the properties of the elements from which the compound is formed. Sodium is an extremely reactive soft metal that cannot be exposed to air or water. Chlorine is a deadly gas. The compound sodium chloride is a white solid which is essential for all living things (see below).



Figure **1.13.1**: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine.

Summary

- A compound is a substance that contains two or more elements chemically combined in a fixed proportion.
- A chemical change is a change that produces matter with a different composition.

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1.14: Chemical Formulas: How to Represent Compounds

Learning Objectives

- Determine the number of different atoms in a formula.
- Define chemical formula, molecular formula, and empirical formula.

A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a 2:1 ratio. The chemical formula for water is H_2O. Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is H_2SO_4.

Certain groups of atoms are bonded together to form what is called a polyatomic ion that acts as a single unit. Polyatomic ions are discussed in more detail in Section 5.5. Polyatomic ions are enclosed in parenthesis followed by a subscript if more than one of the same ion exist in a chemical formula. The formula $Ca_3(PO_4)_2$ represents a compound with the following:

$$3 \text{ Ca atoms} + 2 \ PO_4^{3-} \text{ ions}$$

To count the total number of atoms for formulas with polyatomic ions enclosed in parenthesis, use the subscript as a multiplier for each atom or number of atoms.

$$Ca_3(PO_4)_2$$

$$3 \text{ Ca} + 2 \text{ x1 P} + 2 \text{ x 4 O} = 3 \text{ Ca atoms} + 2 \text{ P atoms} + 8 \text{ O atoms}$$

Molecular Formula

A **molecular formula** is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

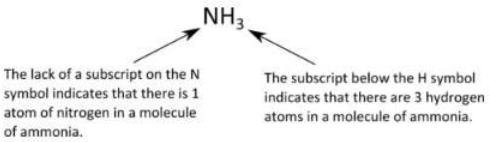


Figure 1.14.1: The molecular formula for ammonia. NH3. There is one atom of nitrogen and 3 atoms of hydrogen in a molecule of ammonia.

Note from the example that there are some standard rules to follow in writing molecular formulas. The arrangements of the elements depend on the particular structure, which is not of concern at this point. The number of atoms of each kind is indicated by a subscript following the atom. If there is only one atom, no number is written. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. We would not write N_3H for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O. When chemists analyze an unknown compound, often the first step is to determine its empirical formula.

- molecular formula: $C_6H_{12}O_6$
- empirical formula: CH_2O

There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.

Summary

- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- If only one atom of a specific type is present, no subscript is used.
- For atoms that have two or more of a specific type of atom present, a subscript is written after the symbol for that atom.
- Polyatomic ions in chemical formulas are enclosed in parentheses followed by a subscript if more than one of the same type of polyatomic ion exist.
- Molecular formulas do not indicate how the atoms are arranged in the molecule.
- The empirical formula tells the lowest whole-number ratio of elements in a compound. The empirical formula does not show the actual number of atoms.

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1.15: A Molecular View of Elements and Compounds

Learning Objectives

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

Atomic Elements

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

Molecular Elements

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

Table **1.15.1**: Elements That Exist as Diatomic Molecules

Hydrogen, H	Oxygen	Nitrogen	Fluorine	Chlorine	Bromine	Iodine

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules—a molecule with only two atoms (Table **1.15.1**). As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2, N_2, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8, while phosphorus exists as a four-atom molecule, P_4 (Figure **1.15.1**).

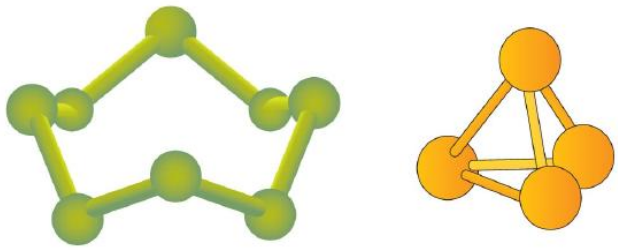


Figure **1.15.1**: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure **1.15.1** shows two examples of how molecules will be represented in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond.

Ionic Compounds

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

Molecular Compounds

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

Example 1.15.1

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

a. Fe
b. PCl_3
c. LiBr
d. P_4
e. oxygen gas

Solution

a. **Fe** (iron) is an element that is represented with no subscript, so it is an **atomic element**.
b. **PCl_3** is made up of two nonmetals, so it is a **molecular compound**.
c. **LiBr** is made up of lithium, a metal, and bromine, a nonmetal, so it is an **ionic compound**.
d. **P_4** is a substance that is made up of four atoms of the same element, so it is a **molecular element**.
e. The formula for **oxygen gas** is **O_2** so it is a molecular element.

Exercise 1.15.1

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

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

Answer a:

molecular element

Answer b:

atomic element

Answer c:

molecular compound

Answer d:

atomic element

Answer e:

ionic compound

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1.16: Methods for Separating Mixtures



Figure **1.16.1** (CC by 2.0; Marcin Chady via Flicker photo, Marcin Chady [www.flickr.com])

How did gold miners search for gold?

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

Separation of Mixtures

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

Here are a number of common separation techniques:

Chromatography

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

Distillation

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

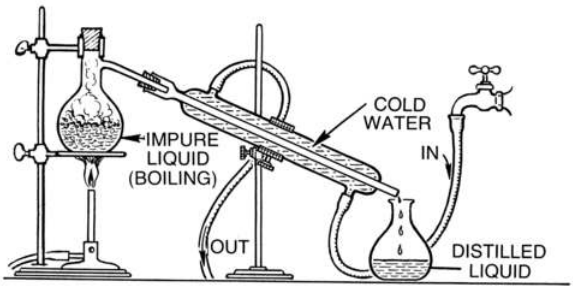


Figure **1.16.2**: Distillation apparatus. (Public Domain; Pearson Scott Foresman via Commons Wikimedia, Distillation Apparatus [commons.wikimedia.org])

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

Evaporation

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

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



Figure **1.16.3**: Once the sea water in these evaporation ponds has evaporated, the salt can be harvested. (CC BY 3.0; Sandip Dey (Wikimedia: Dey.sandip) via Commons Wikimedia, Salt pans in Marakkanam, India [commons.wikimedia.org])

Filtration

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

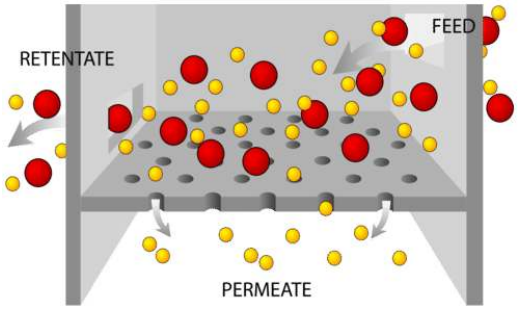


Figure **1.16.3**: Filtration. (Public Domain; Mariana Ruiz Villarreal (Wikimedia: LadyofHats) via Commons Wikimedia, Filtration Diagram [commons.wikimedia.org])



Summary

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

Review

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

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1.17: Properties of Matter

Learning Objectives

- To separate physical from chemical properties and changes

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 properties can be extensive or intensive. **Extensive properties** vary with the amount of the substance and include mass, weight, and volume. **Intensive properties**, in contrast, do not depend on the amount of the substance; they include color, melting point, boiling point, electrical conductivity, and physical state at a given temperature. For example, elemental sulfur is a yellow crystalline solid that does not conduct electricity and has a melting point of 115.2 °C, no matter what amount is examined (Figure 1.17.1). Scientists commonly measure intensive properties to determine a substance's identity, whereas extensive properties convey information about the amount of the substance in a sample.

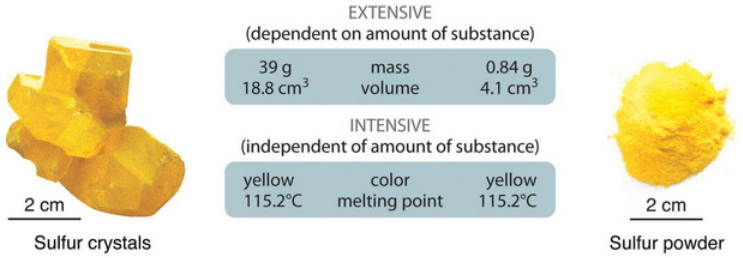


Figure 1.17.1: The Difference between Extensive and Intensive Properties of Matter. Because they differ in size, the two samples of sulfur have different extensive properties, such as mass and volume. In contrast, their intensive properties, including color, melting point, and electrical conductivity, are identical.

Although mass and volume are both extensive properties, their ratio is an important intensive property called **density** (ρ). Density is defined as mass per unit volume and is usually expressed in grams per cubic centimeter (g/cm^3). As mass increases in a given volume, density also increases. For example, lead, with its greater mass, has a far greater density than the same volume of air, just as a brick has a greater density than the same volume of Styrofoam. At a given temperature and pressure, the density of a pure substance is a constant:

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

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

Pure water, for example, has a density of 0.998 g/cm^3 at 25 °C. The average densities of some common substances are in Table 1.17.1. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float" (Figure 1.17.2).

Table 1.17.1: Densities of Common Substances

Substance	Density at 25 °C (g/cm^3)	Substance	Density at 25 °C (g/cm^3)
blood	1.035	corn oil	0.922
body fat	0.918	mayonnaise	0.910
whole milk	1.030	honey	1.420



Figure 1.17.2: Water and oil. Since the oil has a lower density than water, it floats on top. (CC-BY SA 3.0; Victor Blacus).

Physical Property and Change

Physical changes are changes in which no chemical 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** (Figure 1.17.3), since some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.



Figure 1.17.3: Ice Melting is a physical change. When solid water (H_2O) as ice melts into a liquid (water), 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.

Chemical Properties and 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 paper. In contrast to the act of ripping paper, the act of burning paper actually results in the formation of new chemicals (carbon dioxide and water, to be exact). Another example of chemical change occurs when water is formed. Each molecule contains two atoms of hydrogen and one atom of oxygen chemically bonded.

Another example of a chemical change is what occurs when natural gas is burned in your furnace. This time, before the reaction we have a molecule of methane, CH_4, and two molecules of oxygen, O_2, while after the reaction we have two molecules of water, H_2O, and one molecule of carbon dioxide, CO_2. In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.

The combustion of magnesium metal is also chemical change (Magnesium + Oxygen → Magnesium Oxide):

$$2\,Mg + O_2 \rightarrow 2\,MgO$$

as is the rusting of iron (Iron + Oxygen → Iron Oxide/ Rust):

$$4\,Fe + 3\,O_2 \rightarrow 2\,Fe_2O_3$$

Using the components of composition and properties, we have the ability to distinguish one sample of matter from the others.



Different Definitions of Changes: Different Definitions of Changes, YouTube(opens in new window) [youtu.be]



Different Definitions of Properties: Different Definitions of Properties, YouTube(opens in new window) [youtu.be]

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1.18: What is Energy?

Learning Objectives

By the end of this chapter, students will be able to

- List specific examples of non-renewable energy sources
- Explain what makes an energy source non-renewable
- Describe the main types of fossil fuels and how they formed
- Explain the environmental impacts associated with exploration, extraction and use of the different types of fossil fuels
- Explain nuclear energy, how it works, its benefits and risks



Watts Bar Nuclear Power Plant in East Tennessee. Image from Tennessee Valley Authority (Public Domain).

Energy is the ability of a system to do work. A system has done work if it has exerted a force on another system over some distance. When this happens, energy is transferred from one system to another. At least some of the energy is also transformed from one type to another during this process. One can keep track of how much energy transfers into or out of a system. There are two categories that all energy falls into: kinetic and potential. **Kinetic energy** refers to types of energy associated with motion (Figure 1.18.1, top). For example, a rock rolling down a hill, the wind blowing through trees, water flowing over a dam, and a cyclist riding a bicycle are just a few examples of kinetic energy. **Potential energy** is energy possessed by an object or system due to its position in space relative to another object or system and forces between the two (Figure 1.18.1, bottom). Examples include a rock poised at the top of a hill and water stored behind a dam. Some forms of energy are part kinetic and part potential energy. **Chemical energy** describes the potential of a chemical substance to undergo a chemical reaction and transform other chemical substances; hence it is a form of potential energy. Examples include energy stored in the food you eat and the gasoline that you put in your car.



Figure 1.18.1: *Examples kinetic (top) and potential (bottom) forms of energy. All images were obtained from Wikimedia commons (public domain).*

Living organisms need energy to perform life-sustaining “work” in order to survive. For nearly all living systems on Earth, the sun is the ultimate source of that energy. Over time, we humans have developed an understanding of energy that has allowed us to harness it for uses well beyond basic survival. The development and evolution of human society is largely attributed to our relationship with energy. The first major advancement in human understanding of energy was the mastery of fire for cooking and heating. Modern civilization is especially dependent on energy and some of its most distinct characteristics such as population growth, environmental impact and climate change are all a consequence of energy use. We use energy to heat and light our homes; power our machinery; fuel our vehicles; produce plastics, pharmaceuticals, and synthetic fibers; and provide the comforts and conveniences to which we have grown accustomed in the industrial age. Societal complexity, affluence, and the gap between poor and rich peoples are all related to our level of energy consumption.

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1.19: Energy and Chemical and Physical Change

Learning Objectives

- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.
- Determine whether a reaction is endothermic or exothermic through observations, temperature changes, or an energy diagram.

So far, we have talked about how energy exists as either *kinetic energy* or *potential energy* and how energy can be transferred as either *heat* or *work*. While it's important to understand the difference between kinetic energy and potential energy and the difference between heat and work, the truth is, energy is constantly changing. Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy. Likewise, energy that is transferred as work might later end up transferred as heat, while energy that is transferred as heat might later end up being used to do work.

Even though energy can change form, it must still follow one fundamental law: *Energy cannot be created or destroyed, it can only be changed from one form to another*. This law is known as the **Law of Conservation of Energy**. In a lot of ways, energy is like money. You can exchange quarters for dollar bills and dollar bills for quarters, but no matter how often you convert between the two, you will not end up with any more or any less money than you started with. Similarly, you can transfer (or spend) money using cash, or transfer money using a credit card; but you still spend the same amount of money, and the store still makes the same amount of money.

A campfire is an example of basic thermochemistry. The reaction is initiated by the application of heat from a match. The reaction converting wood to carbon dioxide and water (among other things) continues, releasing heat energy in the process. This heat energy can then be used to cook food, roast marshmallows, or just keep warm when it's cold outside.



An image of a campfire with colored flames, made by the burning of a garden hose in a copper pipe. (CC SA-BY 3.0; Jared)

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation of energy** states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe: *the system* and *the surroundings*. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** are everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted, while the surroundings are the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat.

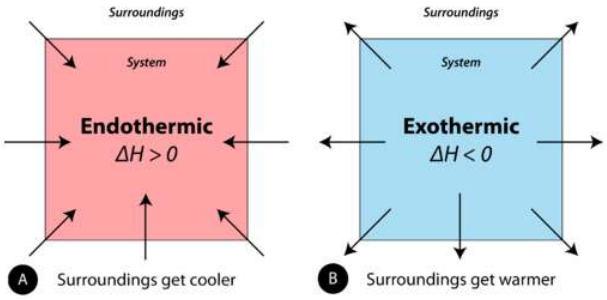


Figure 1.19.1: (A) Endothermic reaction. (B) Exothermic reaction.

Endothermic reaction: surroundings get cooler and delta H is greater than 0, Exothermic reaction: surroundings get warmer and delta H is less than 0

During phase changes, energy changes are usually involved. For example, when solid dry ice vaporizes (physical change), carbon dioxide molecules absorb energy. When liquid water becomes ice, energy is released. Remember that all chemical reactions involve a change in the bonds of the reactants. The bonds in the reactants are broken and the bonds of the products are formed. Chemical bonds have **potential energy** or "stored energy". Because we are changing the bonding, this means we are also changing how much of this "stored energy" there is in a reaction.

Energy changes are frequently shown by drawing an energy diagram. Energy diagrams show the stored/hidden energy of the reactants and products as well as the activation energy. If, on an energy diagram, the products have more stored energy than the reactants started with, the reaction is endothermic. You had to give the reaction energy. If, on the energy diagram, the products have less stored energy than the reactants started with, the reaction is exothermic.

✔ Example 1.19.1

Label each of the following processes as endothermic or exothermic.

a. water boiling
b. gasoline burning
c. ice forming on a pond

Solution

a. Endothermic—you must put a pan of water on the stove and give it heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
b. Exothermic—when you burn something, it feels hot to you because it is giving off heat into the surroundings.
c. Exothermic—think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.

? Exercise 1.19.1

Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

Answer (a)

exothermic

Answer (b)

endothermic

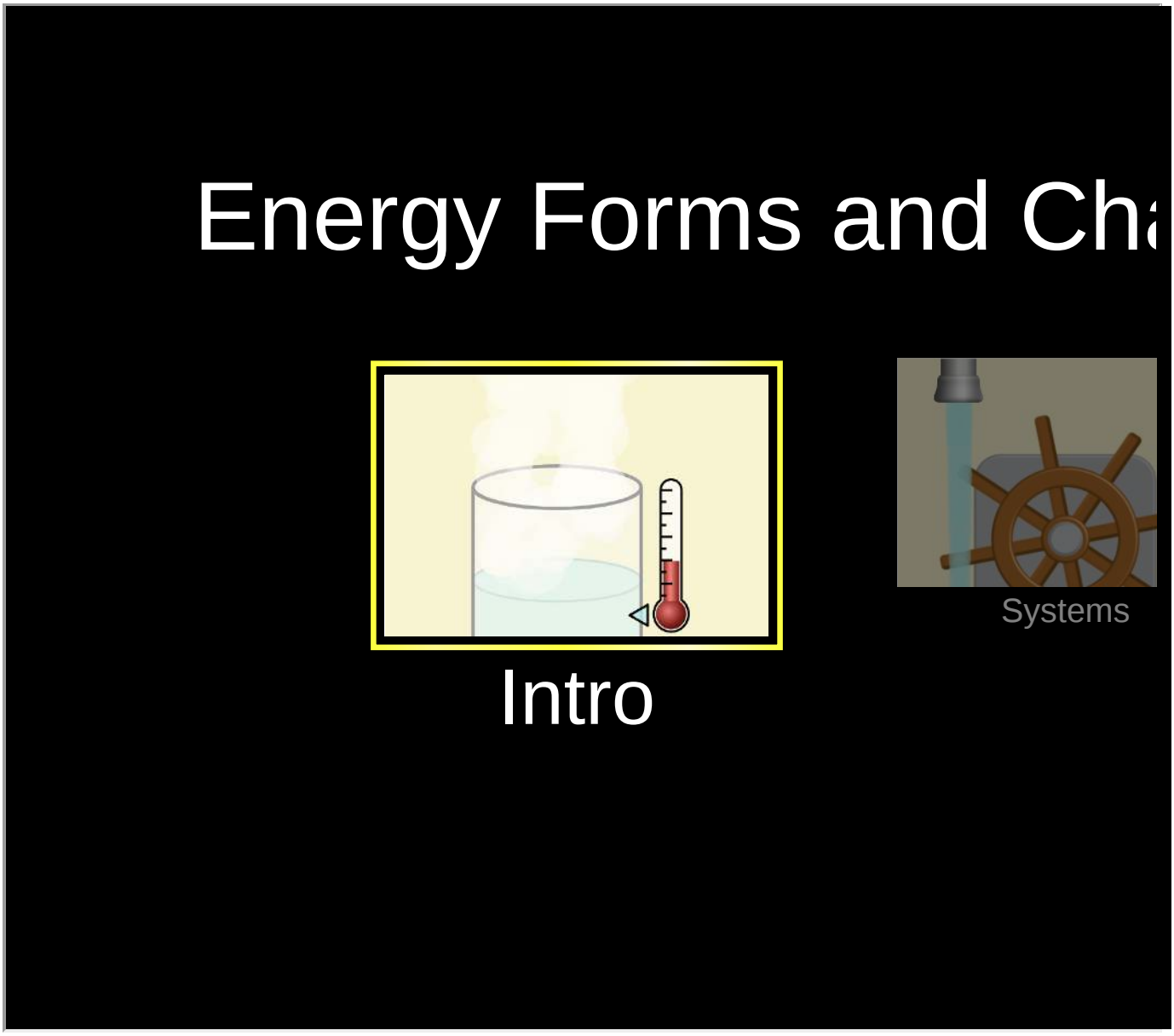
Summary

Phase changes involve changes in energy. All chemical reactions involve changes in energy. This may be a change in heat, electricity, light, or other forms of energy. Reactions that absorb energy are endothermic. Reactions that release energy are exothermic.

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1.20: Activity: Energy Forms and Changes



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CHAPTER OVERVIEW

2: Chemical Reactions

2.1: The Atomic Theory of Matter
2.2: Masses of Atoms and Molecules
2.3: Stoichiometry
2.4: The Mole
2.5: Mole-Mass and Mass-Mass Calculations
2.6: Chemical Equations
2.7: The Mole in Chemical Reactions
2.8: Activity: Balancing Chemical Equations
2.9: Some Simple Patterns of Chemical Reactivity

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2.1: The Atomic Theory of Matter

Learning Objectives

- Explain how all matter is composed of atoms.
- Describe the modern atomic theory.

Take some aluminum foil. Cut it in half. Now there are 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 this exercise be taken, at least in theory? Can one 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.

John Dalton (1766-1844) is the scientist credited for proposing the atomic theory. The theory explains several concepts that are relevant in the observable world: the composition of a pure gold necklace, what makes the pure gold necklace different than a pure silver necklace, and what occurs when pure gold is mixed with pure copper. This section explains the theories that Dalton used as a basis for his theory: (1) Law of Conservation of Mass, (2) Law of Definite Proportions, and (3) Law of Multiple Proportions

Law 1: The Conservation of Mass

"Nothing comes from nothing" is an important idea in ancient Greek philosophy that argues that what exists **now** has always **existed**, since no new matter can come into existence where there was none before. Antoine Lavoisier (1743-1794) restated this principle for chemistry with the law of conservation of mass, which "means that the atoms of an object cannot be created or destroyed, but can be moved around and be changed into different particles." This law says that when a chemical reaction rearranges atoms into a new product, the mass of the reactants (chemicals before the chemical reaction) is the same as the mass of the products (the new chemicals made). More simply, whatever you do, you will still have the same amount of stuff (however, certain nuclear reactions like fusion and fission can convert a small part of the mass into energy.

The law of conservation of mass states that the total mass present before a chemical reaction is the same as the total mass present after the chemical reaction; in other words, **mass is conserved**. The law of conservation of mass was formulated by Lavoisier as a result of his combustion experiment, in which he observed that the mass of his original substance—a glass vessel, tin, and air—was equal to the mass of the produced substance—the glass vessel, “tin calx”, and the remaining air.

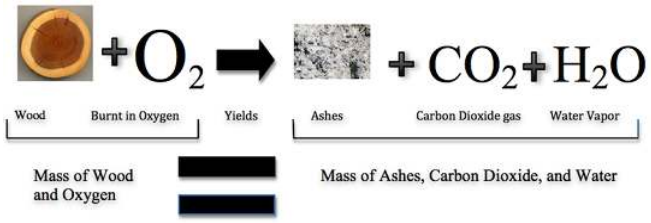


Figure **2.1.1**: Image of the wood courtesy of Ehamberg and Stannered on Wikimedia Commons, available under Creative Commons Attribution 2.5 Generic license. Image of ashes courtesy of Walter Siegmund. Image as a whole constructed by Jessica Thornton (UCD).

Chemical equation; wood + oxygen yields ashes + carbon dioxide gas + water vapor. The mass of wood and oxygen equals the mass of ashes, carbon dioxide, and water.

Historically, this was a difficult concept for scientists to grasp. If this law was true, then how could a large piece of wood be reduced to a small pile of ashes? The wood clearly has a greater mass than the ashes. From this observation scientists concluded that mass had been lost. However, Figure **2.1.1** shows that the burning of word does follow the law of conservation of mass. Scientists did not account for the gases that play a critical role in this reaction.

> *The law of conservation of mass states that the total mass present before a chemical reaction is the same as the total mass present after the chemical reaction.*

Law 2: Definite Proportions

Joseph Proust (1754-1826) formulated the *law of definite proportions* (also called the **Law of Constant *Composition*** or **Proust's Law**). This law states that if a compound is broken down into its constituent elements, the masses of the constituents will always have the same proportions, regardless of the quantity or source of the original substance. Joseph Proust based this law primarily on his experiments with basic copper carbonate. The illustration below depicts this law in action.

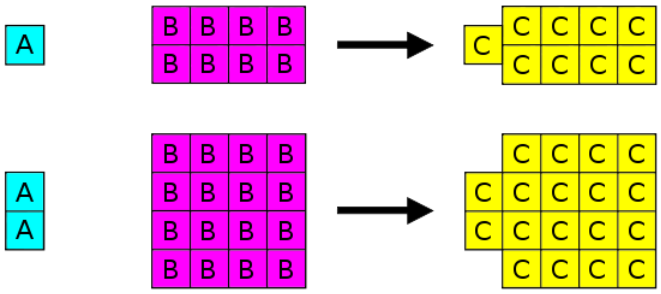


Figure **2.1.2**: If 1 gram of A reacts with 8 grams of B, then by the Law of Definite Proportions, 2 grams of A must react with 16 grams of B. If 1 gram of A reacts with 8 grams of B, then by the Law of Conservation of Mass, they must produce 9 grams of C. Similarly, when 2 grams of A react with 16 grams of B, they must produce 18 grams of C.

Top: Single block labelled A. Eight blocks labelled B with arrow pointing to nine blocks labelled C. Bottom: Two blocks labelled A. Sixteen blocks labelled B with arrow pointing to eighteen blocks labelled C.

> *Law of Definite Proportions states that in a given type of chemical substance, the elements are always combined in the same proportions by mass.*

The Law of Definite Proportions applies when elements are reacted together to form *the same* product. Therefore, while the Law of Definite Proportions can be used to compare two experiments in which hydrogen and oxygen react to form water, the Law of Definite Proportions can *not* be used to compare one experiment in which hydrogen and oxygen react to form water, and another experiment in which hydrogen and oxygen react to form hydrogen peroxide (peroxide is another material that can be made from hydrogen and oxygen).

Example 2.1.1: water

Oxygen makes up 88.8% of the mass of any sample of pure water, while hydrogen makes up the remaining 11.2% of the mass. You can get water by melting ice or snow, by condensing steam, from river, sea, pond, etc. It can be from different places: USA, UK, Australia, or anywhere. It can be made by chemical reactions like burning hydrogen in oxygen.

However, if the water is **pure**, it will **always** consist of 88.8 % oxygen by mass and 11.2 % hydrogen by mass, irrespective of its source or method of preparation.

Law 3: Multiple Proportions

Many combinations of elements can react to form more than one compound. In such cases, this law states that the weights of one element that combine with a fixed weight of another of these elements are integer multiples of one another. It's easy to say this, but please make sure that you understand how it works. Nitrogen forms a very large number of oxides, five of which are shown here.

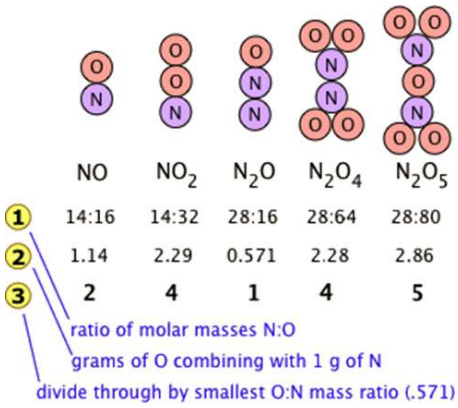


Figure **2.1.3**: Law of Multiple Proportions applied to nitrogen oxides (NO_x) compounds. Table of several nitrogen-oxygen compounds. Column headings, left to right: N O, N O 2, N 2 O, N 2 O 4, N 2 O 5. Second row heading: ratio of molar masses N:O. Values, left to right: 14:16, 14:32, 28:16, 28:64, 28:80. Second row heading: grams of O combining with 1 gram of N. Values: 1.14, 2.29, 0.571, 2.28, 2.86. Third row heading: divide through by smallest O:N mass ratio (.571). Values: 2, 4, 1, 4, 5. (CC-BY; Stephen Lower)

- Line ① shows the ratio of the relative weights of the two elements in each compound. These ratios were calculated by simply taking the molar mass of each element, and multiplying by the number of atoms of that element per mole of the compound. Thus for NO_2, we have $(1 \times 14) : (2 \times 16) = 14:32$. (These numbers were not known in the early days of Chemistry because atomic weights (i.e., molar masses) of most elements were not reliably known.)
- The numbers in Line ② are just the mass ratios of O:N, found by dividing the corresponding ratios in line 1. But someone who depends solely on experiment would work these out by finding the mass of O that combines with unit mass (1 g) of nitrogen.
- Line is obtained by dividing the figures the previous line by the smallest O:N ratio in the line above, which is the one for N_2O. Note that just as the law of multiple proportions says, the weight of oxygen that combines with unit weight of nitrogen work out to small integers.
- Of course we just as easily could have illustrated the law by considering the mass of nitrogen that combines with one gram of oxygen; it works both ways!

> *The law of multiple proportions states that if two elements form more than one compound between them, the masses of one element combined with a fixed mass of the second element form in ratios of small integers.*

Example 2.1.2: Oxides of Carbon

Consider two separate compounds are formed by only carbon and oxygen. The first compound contains 42.9% carbon and 57.1% oxygen (by mass) and the second compound contains 27.3% carbon and 72.7% oxygen (again by mass). Is this consistent with the law of multiple proportions?

Solution

The *Law of Multiple Proportions* states that the masses of one element which combine with a fixed mass of the second element are in a ratio of **whole** numbers. Hence, the masses of oxygen in the two compounds that combine with a fixed mass of carbon should be in a whole-number ratio.

Thus for every 1 g of the first compound there are 0.57 g of oxygen and 0.429 g of carbon. The mass of oxygen per gram carbon is:

$$\frac{0.571 \text{ g oxygen}}{0.429 \text{ g carbon}} = 1.33 \frac{\text{g oxygen}}{\text{g carbon}}$$

Similarly, for 1 g of the second compound, there are 0.727 g oxygen and 0.273 g of carbon. The ration of mass of oxygen per gram of carbon is

$$\frac{0.727 \text{ g oxygen}}{0.273 \text{ g carbon}} = 2.66 \frac{\text{g oxygen}}{\text{g carbon}}$$

Dividing the mass of oxygen per g of carbon of the second compound:

$$\frac{2.66}{1.33} = 2$$

Hence the masses of oxygen combine with carbon in a 2:1 ratio which s consistent with the Law of Multiple Proportions since they are whole numbers.

Dalton's Atomic Theory

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure **2.1.4**), is a fundamental concept that states that all elements are composed of atoms. Previously, an atom was defined 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 the human pinkie (about 1 cm).

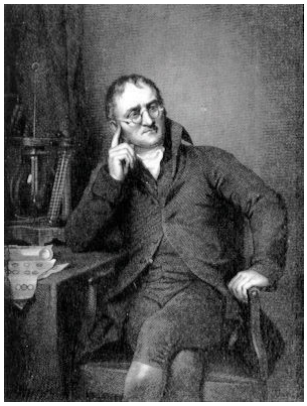


Figure **2.1.4**: John Dalton was an English scientist who enunciated the modern atomic theory.

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. 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.

Dalton's Theory was a powerful development as it explained the three laws of chemical combination (above) and recognized a workable distinction between the fundamental particle of an element (atom) and that of a compound (molecule). Six postulates are involved in Dalton's Atomic Theory:

1. All matter consists of indivisible particles called atoms.
2. Atoms of the same element are similar in shape and mass, but differ from the atoms of other elements.
3. Atoms cannot be created or destroyed.
4. Atoms of different elements may combine with each other in a fixed, simple, whole number ratios to form compound atoms.
5. Atoms of same element can combine in more than one ratio to form two or more compounds.
6. The atom is the smallest unit of matter that can take part in a chemical reaction.

In light of the current state of knowledge in the field of Chemistry, Dalton's theory had a few drawbacks. According to Dalton's postulates,

1. The indivisibility of an atom was proved wrong: an atom can be further subdivided into protons, neutrons and electrons. However an atom is the smallest particle that takes part in chemical reactions.

2. According to Dalton, the atoms of same element are similar in all respects. However, atoms of some elements vary in their masses and densities. These atoms of different masses are called isotopes. For example, chlorine has two isotopes with mass numbers 35 and 37.
3. Dalton also claimed that atoms of different elements are different in all respects. This has been proven wrong in certain cases: argon and calcium atoms each have an same atomic mass (40 amu).
4. According to Dalton, atoms of different elements combine in simple whole number ratios to form compounds. This is not observed in complex organic compounds like sugar ($C_{12}H_{22}O_{11}$).
5. The theory fails to explain the existence of allotropes (different forms of pure elements); it does not account for differences in properties of charcoal, graphite, diamond.

Despite these drawbacks, the importance of Dalton's theory should not be underestimated. He displayed exceptional insight into the nature of matter. and his ideas provided a framework that was later modified and expanded by other. Consequentiually, John Dalton is often considered to be the father of modern atomic theory.



Fundamental Experiments in Chemistry: Fundamental Experiments in Chemistry, YouTube(opens in new window) [youtu.be]

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Summary

This article explains the theories that Dalton used as a basis for his theory: (1) the Law of Conservation of Mass, (2) the Law of Constant Composition, (3) the Law of Multiple Proportions.

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2.2: Masses of Atoms and Molecules

Learning Objective

- Express the masses of atoms and molecules.

Because matter is defined as anything that has mass and takes up space, it should not be surprising to learn that atoms and molecules have mass.

Individual atoms and molecules, however, are very small, and the masses of individual atoms and molecules are also very small. For macroscopic objects, we use units such as grams and kilograms to state their masses, but these units are much too big to comfortably describe the masses of individual atoms and molecules. Another scale is needed.

Atomic Mass Unit

The atomic mass unit (u; some texts use amu, but this older style is no longer accepted) is defined as one-twelfth of the mass of a carbon-12 atom, an isotope of carbon that has six protons and six neutrons in its nucleus. By this scale, the mass of a proton is 1.00728 u, the mass of a neutron is 1.00866 u, and the mass of an electron is 0.000549 u. There will not be much error if you estimate the mass of an atom by simply counting the total number of protons and neutrons in the nucleus (i.e., identify its mass number) and ignore the electrons. Thus, the mass of carbon-12 is about 12 u, the mass of oxygen-16 is about 16 u, and the mass of uranium-238 is about 238 u. More exact masses are found in scientific references—for example, the exact mass of uranium-238 is 238.050788 u, so you can see that we are not far off by using the whole-number value as the mass of the atom.

What is the mass of an element? This is somewhat more complicated because most elements exist as a mixture of isotopes, each of which has its own mass. Thus, although it is easy to speak of the mass of an atom, when talking about the mass of an element, we must take the isotopic mixture into account.

Atomic Mass

The atomic mass of an element is a weighted average of the masses of the isotopes that compose an element. What do we mean by a weighted average? Well, consider an element that consists of two isotopes, 50% with mass 10 u and 50% with mass 11 u. A weighted average is found by multiplying each mass by its fractional occurrence (in decimal form) and then adding all the products. The sum is the weighted average and serves as the formal atomic mass of the element. In this example, we have the following:

0.50 × 10 u	= 5.0 u
0.50 × 11 u	= 5.5 u
Sum	= 10.5 u = the atomic mass of our element

Note that no atom in our hypothetical element has a mass of 10.5 u; rather, that is the average mass of the atoms, weighted by their percent occurrence.

This example is similar to a real element. Boron exists as about 20% boron-10 (five protons and five neutrons in the nuclei) and about 80% boron-11 (five protons and six neutrons in the nuclei). The atomic mass of boron is calculated similarly to what we did for our hypothetical example, but the percentages are different:

0.20 × 10 u	= 2.0 u
0.80 × 11 u	= 8.8 u
Sum	= 10.8 u = the atomic mass of boron

Thus, we use 10.8 u for the atomic mass of boron.

Virtually all elements exist as mixtures of isotopes, so atomic masses may vary significantly from whole numbers. Table **2.2.1** lists the atomic masses of some elements. The atomic masses in Table **2.2.1** are listed to three decimal places where possible, but in most cases, only one or two decimal places are needed. Note that many of the atomic masses, especially the larger ones, are not very close to whole numbers. This is, in part, the effect of an increasing number of isotopes as the atoms increase in size. (The record number is 10 isotopes for tin.)

Table **2.2.1**: Selected Atomic Masses of Some Elements

Element Name	Atomic Mass (u)	Element Name	Atomic Mass (u)
Aluminum	26.981	Molybdenum	95.94
Argon	39.948	Neon	20.180
Arsenic	74.922	Nickel	58.693
Barium	137.327	Nitrogen	14.007
Beryllium	9.012	Oxygen	15.999
Bismuth	208.980	Palladium	106.42
Boron	10.811	Phosphorus	30.974
Bromine	79.904	Platinum	195.084
Calcium	40.078	Potassium	39.098
Carbon	12.011	Radium	n/a
Chlorine	35.453	Radon	n/a
Cobalt	58.933	Rubidium	85.468
Copper	63.546	Scandium	44.956
Fluorine	18.998	Selenium	78.96
Gallium	69.723	Silicon	28.086
Germanium	72.64	Silver	107.868
Gold	196.967	Sodium	22.990
Helium	4.003	Strontium	87.62
Hydrogen	1.008	Sulfur	32.065
Iodine	126.904	Tantalum	180.948
Iridium	192.217	Tin	118.710
Iron	55.845	Titanium	47.867
Krypton	83.798	Tungsten	183.84
Lead	207.2	Uranium	238.029
Lithium	6.941	Xenon	131.293
Magnesium	24.305	Zinc	65.409
Manganese	54.938	Zirconium	91.224
Mercury	200.59	Molybdenum	95.94

Note: Atomic mass is given to three decimal places, if known.

Molecular Mass

Now that we understand that atoms have mass, it is easy to extend the concept to the mass of molecules. The molecular mass is the sum of the masses of the atoms in a molecule. This may seem like a trivial extension of the concept, but it is important to count the number of each type of atom in the molecular formula. Also, although each atom in a molecule is a particular isotope, we use the weighted average, or atomic mass, for each atom in the molecule.

For example, if we were to determine the molecular mass of dinitrogen trioxide, N_2O_3, we would need to add the atomic mass of nitrogen two times with the atomic mass of oxygen three times:

2 N masses = 2 × 14.007 u	= 28.014 u
3 O masses = 3 × 15.999 u	= 47.997 u
Total	= 76.011 u = the molecular mass of N_2O_3

We would not be far off if we limited our numbers to one or even two decimal places.

Example 2.2.1

What is the molecular mass of each substance?

a. NBr_3
b. C_2H_6

Solution

Add one atomic mass of nitrogen and three atomic masses of bromine:

Solutions to Example 3.4.1

1 N mass	= 14.007 u
3 Br masses = 3 × 79.904 u	= 239.712 u
Total	= 253.719 u = the molecular mass of NBr_3

Add two atomic masses of carbon and six atomic masses of hydrogen:

Solutions to Example 3.4.1

2 C masses = 2 × 12.011 u	= 24.022 u
6 H masses = 6 × 1.008 u	= 6.048 u
Total	= 30.070 u = the molecular mass of C_2H_6

The compound C_2H_6 also has a common name—ethane.

Exercise 2.2.1

What is the molecular mass of each substance?

a. SO_2
b. Extra close brace or missing open brace

Answer a

64.063 u

Answer b

87.968 u

Chemistry is Everywhere: Sulfur Hexafluoride

On March 20, 1995, the Japanese terrorist group Aum Shinrikyo (Sanskrit for "Supreme Truth") released some sarin gas in the Tokyo subway system; twelve people were killed, and thousands were injured (Figure **2.2.2*a***). Sarin (molecular formula $C_4H_{10}FPO_2$) is a nerve toxin that was first synthesized in 1938. It is regarded as one of the most deadly toxins known, estimated to be about 500 times more potent than cyanide. Scientists and engineers who study the spread of chemical weapons such as sarin (yes, there are such scientists) would like to have a similar less dangerous chemical to study, indeed one that is nontoxic, so they are not at risk themselves.

Sulfur hexafluoride is used as a model compound for sarin. SF_6 (a molecular model of which is shown Figure **2.2.2*b***) has a similar molecular mass (about 146 u) as sarin (about 140 u), so it has similar physical properties in the vapor phase. Sulfur hexafluoride is also very easy to accurately detect, even at low levels, and it is not a normal part of the atmosphere, so there is little potential for contamination from natural sources. Consequently, SF_6 is also used as an aerial tracer for ventilation systems in buildings. It is nontoxic and very chemically inert, so workers do not have to take special precautions other than watching for asphyxiation.

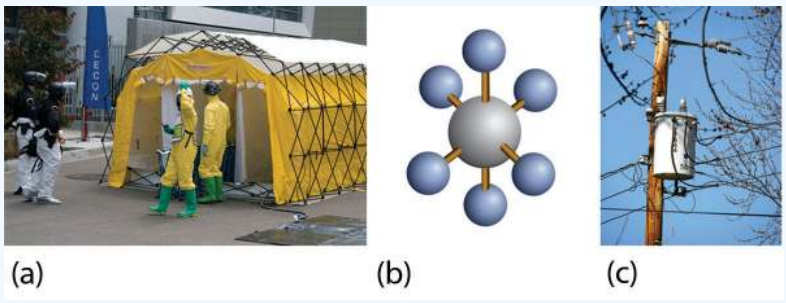


Figure **2.2.2**: Sarin and Sulfur Hexafluoride © Thinkstock. (a) Properly protected workers clear out the Tokyo subway after the nerve toxin sarin was released. (b) A molecular model of SF_6. (c) A high-voltage electrical switchgear assembly that would be filled with SF_6 as a spark suppressant.

Sulfur hexafluoride also has another interesting use: a spark suppressant in high-voltage electrical equipment. High-pressure SF_6 gas is used in place of older oils that may have contaminants that are not environmentally friendly (Figure **2.2.2*c***).

Key Takeaways

- The atomic mass unit (u) is a unit that describes the masses of individual atoms and molecules.
- The atomic mass is the weighted average of the masses of all isotopes of an element.
- The molecular mass is the sum of the masses of the atoms in a molecule.

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2.3: 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,

$$9\,\cancel{c\,pancake\,mix} \times \frac{1\,egg}{\cancel{2\,c\,pancake\,mix}} = 4.5\,eggs$$

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:

$$9\,\cancel{c\,pancake\,mix} \times \frac{\frac{1}{2}\,c\,milk}{\cancel{2\,c\,pancake\,mix}} = 2.25\,c\,milk$$

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:

$$2H_2 \text{ molecules} \Leftrightarrow 1O_2 \text{ molecule} \Leftrightarrow 2H_2O \text{ molecules}$$

These equivalences allow us to construct conversion factors:

$$\frac{2\,molecules\,H_2}{1\,molecule\,O_2} \quad \frac{2\,molecules\,H_2}{2\,molecules\,H_2O} \quad \frac{1\,molecule\,H_2}{2\,molecules\,H_2O}$$

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 H_2. 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* H_2 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 2.3.1

How many molecules of SO_3 are needed to react with 144 molecules of Fe_2O_3 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 Fe_2O_3 and SO_3. The number of molecules of Fe_2O_3 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO_3 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 SO_3 to react with 144 molecules of Fe_2O_3.

Exercise 2.3.1

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

87 molecules

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 2.3.2

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 2.3.2

How many molecules of $Fe_2(SO_4)_3$ can you make from 777 atoms of S?

Answer

259 molecules

Summary

Quantities of substances can be related to each other using balanced chemical equations.

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2.4: 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 2.4.1

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 2.4.1

How many molecules are present in 4.61×10^{-2} mol of O_2?

Answer

2.78×10^{22} molecules

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:

$$1.50 \times 10^{25} \cancel{\text{ atoms C}} \times \frac{12.0000\, g\, C}{6.022 \times 10^{23} \cancel{\text{ atoms C}}} = 299\, g\, C$$

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 2.4.2: 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 2.4.2

What is the molar mass of $AgNO_3$?

Answer

169.87 g/mol

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 2.4.3

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 2.4.3

What is the mass of 33.7 mol of $\mathbf{H_2O}$?

Answer

607 g

Example 2.4.4

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

$$(1.0079 + 1.0079 + 15.999) = 18.015\ g/mol.$$

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 2.4.4

How many moles are present in 35.6 g of H_2SO_4 (molar mass = 98.08 g/mol)?

Answer

0.363 mol

Other conversion factors can be combined with the definition of mole—density, for example.

Example 2.4.5

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 2.4.5

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

0.201 mol

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.

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2.5: 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 2.5.1

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

We will do this in two steps: convert the mass of $AlCl_3$ to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of $AlCl_3$ 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 $AlCl_3$ and the number of moles of HCl:

$$frac6\ mol\ HCl2\ \cancel{mol\ AlCl_3}$$

Applying this conversion factor to the quantity of $AlCl_3$, 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 2.5.1

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer

0.442 mol

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 2.5.2

How many grams of NH_3 will be produced when 33.9 mol of H_2 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 2.5.2

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

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.)

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):

$$45.3\ g\,SO_2 \times \frac{1\ mol\ SO_2}{64.06\ g\,SO_2} = 0.707\ mol\ SO_2$$

Second, we use the balanced chemical reaction to convert from moles of SO_2 to moles of SO_3:

$$0.707\ mol\,SO_2 \times \frac{2\ mol\ SO_3}{2\ mol\,SO_2} = 0.707\ mol\ SO_3$$

Finally, we use the molar mass of SO_3 (80.06 g/mol) to convert to the mass of SO_3:

$$0.707\ mol\,SO_3 \times \frac{80.06\ g\ SO_3}{1\ mol\,SO_3} = 56.6\ g\ SO_3$$

We can also perform all three steps sequentially, writing them on one line as

$$45.3\ g\,SO_2 \times \frac{1\ mol\ SO_2}{64.06\ g\,SO_2} \times \frac{2\ mol\ SO_3}{2\ mol\,SO_2} \times \frac{80.06\ g\ SO_3}{1\ mol\,SO_3} = 56.6\ g\ SO_3$$

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 2.5.3

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

mass K → mol K → mol Mg → 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,

$$86.4\ g\,K \times \frac{1\ mol\ K}{39.09\ g\,K} \times \frac{1\ mol\,Mg}{2\ mol\,K} \times \frac{24.31\ g\ Mg}{1\ mol\,Mg} = 26.87\ g\ Mg$$

Exercise 2.5.3

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

3.77 g

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.

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2.6: Chemical Equations

Learning Objectives

- To describe a chemical reaction.
- To calculate the quantities of compounds produced or consumed in a chemical reaction

What happens to matter when it undergoes chemical changes? The Law of conservation of mass says that "***Atoms are neither created, nor destroyed, during any chemical reaction***." Thus, the same collection of atoms is present after a reaction as before the reaction. The changes that occur during a reaction just involve the *rearrangement* of atoms. In this section we will discuss ***stoichiometry*** (the "measurement of elements").

Chemical Equations

As shown in Figure **2.6.1**, applying a small amount of heat to a pile of orange ammonium dichromate powder results in a vigorous reaction known as the ammonium dichromate volcano. Heat, light, and gas are produced as a large pile of fluffy green chromium(III) oxide forms. This reaction is described with a chemical equation, an expression that gives the identities and quantities of the substances in a chemical reaction.



Figure **2.6.1**: An Ammonium Dichromate Volcano: Change during a Chemical Reaction. The starting material is solid ammonium dichromate. A chemical reaction transforms it to solid chromium(III) oxide, depicted showing a portion of its chained structure, nitrogen gas, and water vapor (in addition, energy in the form of heat and light is released). During the reaction, the distribution of atoms changes, but the number of atoms of each element does not change. Because the numbers of each type of atom are the same in the reactants and the products, the chemical equation is balanced. (CC BY-SA 3.0; Mikk Mihkel Vaabel via Wikipedia). See video here: www.youtube.com/watch?v=CW4hN0dYnkM

Chemical reactions are represented on paper by ***chemical equations***. For example, hydrogen gas (H_2) can react (burn) with oxygen gas (O_2) to form water (H_2O). The *chemical equation* for this *reaction* is written as:

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

Chemical formulas and other symbols are used to indicate the starting materials, or reactants, which by convention are written on the left side of the equation, and the final compounds, or products, which are written on the right. An arrow points from the reactant to the products. The chemical reaction for the ammonium dichromate volcano in Figure **2.6.1** is

$$\underbrace{(NH_4)_2Cr_2O_7}_{reactant} \rightarrow \underbrace{Cr_2O_3 + N_2 + 4\,H_2O}_{products} \tag{2.6.1}$$

The arrow is read as “yields” or “reacts to form.” Equation 2.6.1 indicates that ammonium dichromate (the reactant) yields chromium(III) oxide, nitrogen, and water (the products). The equation for this reaction is even more informative when written as follows:

$$(NH_4)_2Cr_2O_7\,(s) \rightarrow Cr_2O_3\,(s) + N_2\,(g) + 4\,H_2O\,(g) \tag{2.6.2}$$

Equation 2.6.2 is identical to Equation 2.6.1 except for the addition of abbreviations in parentheses to indicate the physical state of each species. The abbreviations are (s) for solid, (l) for liquid, (g) for gas, and (aq) for an aqueous solution, a solution of the substance in water.

Consistent with the law of conservation of mass, the numbers of each type of atom are the same on both sides of Equations 2.6.1 and 2.6.2. Each side of the reaction has two chromium atoms, seven oxygen atoms, two nitrogen atoms, and eight hydrogen atoms.

In a balanced chemical equation, both the numbers of each type of atom and the total charge are the same on both sides. Equations 2.6.1 and 2.6.2 are balanced chemical equations. What is different on each side of the equation is how the atoms are arranged to make molecules or ions. A chemical reaction represents a change in the distribution of atoms, but not in the number of atoms. In this reaction, and in most chemical reactions, bonds are broken in the reactants (here, Cr–O and N–H bonds), and new bonds are formed to create the products (here, O–H and N≡N bonds). If the numbers of each type of atom are different on the two sides of a chemical equation, then the equation is unbalanced, and it cannot correctly describe what happens during the reaction. To proceed, the equation must first be balanced.

A chemical reaction changes only the distribution of atoms, not the number of atoms.



Introduction to Chemical Reaction Equations: Introduction to Chemical Reaction Equations, YouTube(opens in new window) [youtu.be]

Balancing Simple Chemical Equations

When a chemist encounters a new reaction, it does not usually come with a label that shows the balanced chemical equation. Instead, the chemist must identify the reactants and products and then write them in the form of a chemical equation that may or may not be balanced as first written. Consider, for example, the combustion of n-heptane (C_7H_{16}), an important component of gasoline:

$$\mathrm{C_7H_{16}\,(l) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)} \tag{2.6.3}$$

The complete combustion of any hydrocarbon with sufficient oxygen always yields carbon dioxide and water.



Figure 2.6.2: An Example of a Combustion Reaction. The wax in a candle is a high-molecular-mass hydrocarbon, which produces gaseous carbon dioxide and water vapor in a combustion reaction (Equation 2.6.3).

Equation 2.6.3 is not balanced: the numbers of each type of atom on the reactant side of the equation (7 carbon atoms, 16 hydrogen atoms, and 2 oxygen atoms) is not the same as the numbers of each type of atom on the product side (1 carbon atom, 2 hydrogen atoms, and 3 oxygen atoms). Consequently, the coefficients of the reactants and products must be adjusted to give the same numbers of atoms of each type on both sides of the equation. Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 2.6.3.

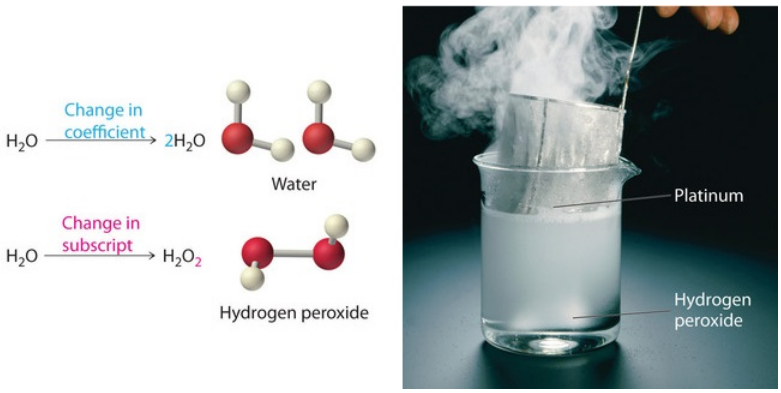


Figure **2.6.3**: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.

Left: Example of changing coefficients or subscripts. Right: Platinum dissolving in hydrogen peroxide.



Balancing Combustion Reactions: Balancing Combustions Reactions, YouTube(opens in new window) [youtu.be]

The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element on both sides.
3. Balance polyatomic ions (if present) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Check your work by counting the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

Example 2.6.1A: Combustion of Heptane

To demonstrate this approach, let’s use the combustion of n-heptane (Equation **2.6.3**) as an example.

1. **Identify the most complex substance**. The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.

2. **Adjust the coefficients**. Try to adjust the coefficients of the molecules on the other side of the equation to obtain the same numbers of atoms on both sides. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side:

$$\mathrm{C_7H_{16} + O_2 \rightarrow 7\,CO_2 + H_2O} \tag{2.6.4}$$

3. **Balance polyatomic ions as a unit**. There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms**. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H2O molecules, each of which contains 2 hydrogen atoms, on the right side:

$$\mathrm{C_7H_{16} + O_2 \rightarrow 7\,CO_2 + 8\,H_2O} \tag{2.6.5}$$

The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side:

$$\mathrm{C_7H_{16}\,(l) + 11\,O_2\,(g) \rightarrow 7\,CO_2\,(g) + 8\,H_2O\,(g)} \tag{2.6.6}$$

5. **Check your work**. The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start.

✔ Example 2.6.1*B*: Combustion of Isooctane

Consider, for example, a similar reaction, the combustion of isooctane ($\mathrm{C_8H_{18}}$). Because the combustion of any hydrocarbon with oxygen produces carbon dioxide and water, the unbalanced chemical equation is as follows:

$$\mathrm{C_8H_{18}\,(l) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)} \tag{2.6.7}$$

1. **Identify the most complex substance**. Begin the balancing process by assuming that the final balanced chemical equation contains a single molecule of isooctane.
2. **Adjust the coefficients**. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products:

$$\mathrm{C_8H_{18} + O_2 \rightarrow 8\,CO_2 + H_2O} \tag{2.6.8}$$

3. **Balance polyatomic ions as a unit**. This step does not apply to this equation.
4. **Balance the remaining atoms**. Eighteen hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products:

$$\mathrm{C_8H_{18} + O_2 \rightarrow 8\,CO_2 + 9\,H_2O} \tag{2.6.9}$$

The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient (25/2) to balance the oxygen atoms:

$$\mathrm{C_8H_{18} + \tfrac{25}{2}\,O_2 \rightarrow 8\,CO_2 + 9\,H_2O} \tag{2.6.10}$$

Equation 2.6.10 is now balanced, but we usually write equations with whole-number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by 2:

$$\mathrm{2\,C_8H_{18}\,(l) + 25\,O_2\,(g) \rightarrow 16\,CO_2\,(g) + 18\,H_2O\,(g)} \tag{2.6.11}$$

5. **Check your work**. The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side.



Balancing Complex Chemical Equations: Balancing Complex Chemical Equations, YouTube(opens in new window) [youtu.be]

Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

Example 2.6.1C: Hydroxyapatite

The reaction of the mineral hydroxyapatite ($Ca_5(PO_4)_3(OH)$) with phosphoric acid and water gives $Ca(H_2PO_4)_2 \cdot H_2O$ (calcium dihydrogen phosphate monohydrate). Write and balance the equation for this reaction.



Hydroxyapatite (Extra close brace or missing open brace) crystal

Given: reactants and product

Asked for: balanced chemical equation

Strategy:

A. Identify the product and the reactants and then write the unbalanced chemical equation.
B. Follow the steps for balancing a chemical equation.

Solution:

A We must first identify the product and reactants and write an equation for the reaction. The formulas for hydroxyapatite and calcium dihydrogen phosphate monohydrate are given in the problem (recall that phosphoric acid is H_3PO_4). The initial (unbalanced) equation is as follows:

$$Ca_5(PO_4)_3(OH)\,(s) + H_3PO_4\,(aq) + H_2O_{(l)} \rightarrow Ca(H_2PO_4)_2 \cdot H_2O_{(s)}$$

1. **B Identify the most complex substance**. We start by assuming that only one molecule or formula unit of the most complex substance, $Ca_5(PO_4)_3(OH)$, appears in the balanced chemical equation.

2. **Adjust the coefficients**. Because calcium is present in only one reactant and one product, we begin with it. One formula unit of $Ca_5(PO_4)_3(OH)$ contains 5 calcium atoms, so we need 5 $Ca(H_2PO_4)_2$•H_2O on the right side:

$$Ca_5(PO_4)_3(OH) + H_3PO_4 + H_2O \rightarrow 5\,Ca(H_2PO_4)_2 \cdot H_2O$$

3. **Balance polyatomic ions as a unit**. It is usually easier to balance an equation if we recognize that certain combinations of atoms occur on both sides. In this equation, the polyatomic phosphate ion (PO_4^{3-}), shows up in three places. In H_3PO_4, the phosphate ion is combined with three H^+ ions to make phosphoric acid (H_3PO_4), whereas in $Ca(H_2PO_4)_2 \cdot H_2O$ it is combined with two H^+ ions to give the dihydrogen phosphate ion. Thus it is easier to balance PO_4 as a unit rather than counting individual phosphorus and oxygen atoms. There are 10 PO_4 units on the right side but only 4 on the left. The simplest way to balance the PO_4 units is to place a coefficient of 7 in front of H_3PO_4:

$$\mathrm{Ca_5(PO_4)_3(OH) + 7\,H_3PO_4 + H_2O \rightarrow 5\,Ca(H_2PO_4)_2 \cdot H_2O}$$

Although OH^- is also a polyatomic ion, it does not appear on both sides of the equation. So oxygen and hydrogen must be balanced separately.

4. **Balance the remaining atoms**. We now have 30 hydrogen atoms on the right side but only 24 on the left. We can balance the hydrogen atoms using the least complex substance, H_2O, by placing a coefficient of 4 in front of H_2O on the left side, giving a total of 4 H_2O molecules:

$$\mathrm{Ca_5(PO_4)_3(OH)\,(s) + 7\,H_3PO_4\,(aq) + 4\,H_2O\,(l) \rightarrow 5\,Ca(H_2PO_4)_2 \cdot H_2O\,(s)}$$

The equation is now balanced. Even though we have not explicitly balanced the oxygen atoms, there are 41 oxygen atoms on each side.

5. **Check your work**. Both sides of the equation contain 5 calcium atoms, 10 phosphorus atoms, 30 hydrogen atoms, and 41 oxygen atoms.

? Exercise 2.6.1: Fermentation

Fermentation is a biochemical process that enables yeast cells to live in the absence of oxygen. Humans have exploited it for centuries to produce wine and beer and make bread rise. In fermentation, sugars such as glucose are converted to ethanol (CH_3CH_2OH and carbon dioxide CO_2. Write a balanced chemical reaction for the fermentation of glucose.



Commercial use of fermentation. *(a) Microbrewery vats are used to prepare beer. (b) The fermentation of glucose by yeast cells is the reaction that makes beer production possible.*

Answer

$$C_6H_{12}O_6(s) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$



Balancing Reactions Which Contain Polyatomics: Balancing Reactions Which Contain Polyatomics, YouTube(opens in new window) [youtu.be]

Interpreting Chemical Equations

In addition to providing qualitative information about the identities and physical states of the reactants and products, a balanced chemical equation provides quantitative information. Specifically, it gives the relative amounts of reactants and products consumed or produced in a reaction. The number of atoms, molecules, or formula units of a reactant or a product in a balanced chemical equation is the coefficient of that species (e.g., the 4 preceding H_2O in Equation 2.6.1). When no coefficient is written in front of a species, the coefficient is assumed to be 1. As illustrated in Figure **2.6.4**, the coefficients allow Equation 2.6.1 to be interpreted in any of the following ways:

- Two NH_4^+ ions and one $Cr_2O_7^{2-}$ ion yield 1 formula unit of Cr_2O_3, 1 N_2 molecule, and 4 H_2O molecules.
- One mole of $(NH_4)_2Cr_2O_7$ yields 1 mol of Cr_2O_3, 1 mol of N_2, and 4 mol of H_2O.
- A mass of 252 g of $(NH_4)_2Cr_2O_7$ yields 152 g of Cr_2O_3, 28 g of N_2, and 72 g of H_2O.
- A total of 6.022×10^{23} formula units of $(NH_4)_2Cr_2O_7$ yields 6.022×10^{23} formula units of Cr_2O_3, 6.022×10^{23} molecules of N_2, and 24.09×10^{23} molecules of H_2O.

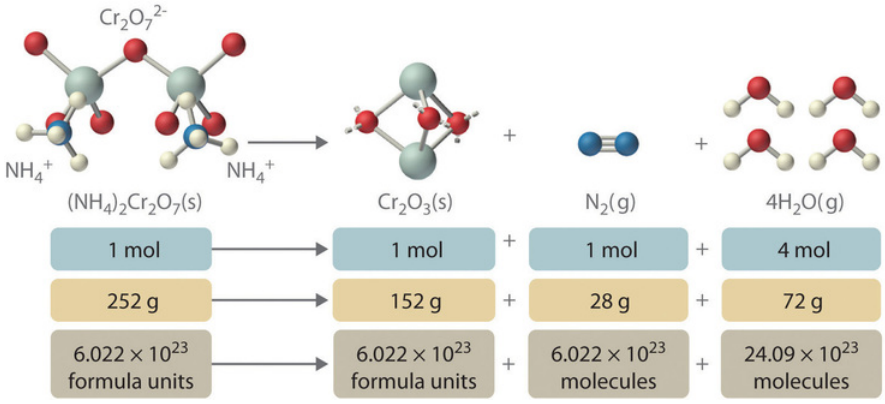


Figure **2.6.4**: The Relationships among Moles, Masses, and Formula Units of Compounds in the Balanced Chemical Reaction for the Ammonium Dichromate Volcano

Chemical equation: (N H 4) 2 C r 2 O 7 dissociates into C r 2 O 3, N 2, and H 2 O. Conversions are given between moles, mass, and molecules.

These are all chemically equivalent ways of stating the information given in the balanced chemical equation, using the concepts of the mole, molar or formula mass, and Avogadro's number. The ratio of the number of moles of one substance to the number of moles of another is called the mole ratio. For example, the mole ratio of H_2O to N_2 in Equation 2.6.1 is 4:1. The total mass of reactants equals the total mass of products, as predicted by Dalton's law of conservation of mass:

$$252\ g\ \text{of}\ (NH_4)_2Cr_2O_7$$

yield

$$152 + 28 + 72 = 252\ g \text{ of products.}$$

The chemical equation does not, however, show the rate of the reaction (rapidly, slowly, or not at all) or whether energy in the form of heat or light is given off. These issues are considered in more detail in later chapters.

An important chemical reaction was analyzed by Antoine Lavoisier, an 18th-century French chemist, who was interested in the chemistry of living organisms as well as simple chemical systems. In a classic series of experiments, he measured the carbon dioxide and heat produced by a guinea pig during respiration, in which organic compounds are used as fuel to produce energy, carbon dioxide, and water. Lavoisier found that the ratio of heat produced to carbon dioxide exhaled was similar to the ratio observed for the reaction of charcoal with oxygen in the air to produce carbon dioxide—a process chemists call combustion. Based on these experiments, he proposed that "Respiration is a combustion, slow it is true, but otherwise perfectly similar to that of charcoal." Lavoisier was correct, although the organic compounds consumed in respiration are substantially different from those found in charcoal. One of the most important fuels in the human body is glucose ($C_6H_{12}O_6$), which is virtually the only fuel used in the brain. Thus combustion and respiration are examples of chemical reactions.

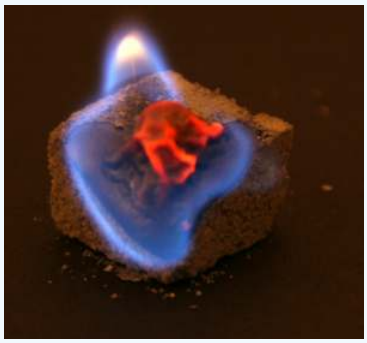
Example 2.6.2: Combustion of Glucose

The balanced chemical equation for the combustion of glucose in the laboratory (or in the brain) is as follows:

$$\mathrm{C_6H_{12}O_6\,(s) + 6\,O_2\,(g) \rightarrow 6\,CO_2\,(g) + 6\,H_2O\,(l)}$$

Construct a table showing how to interpret the information in this equation in terms of

a. a single molecule of glucose.
b. moles of reactants and products.
c. grams of reactants and products represented by 1 mol of glucose.
d. numbers of molecules of reactants and products represented by 1 mol of glucose.



The combustion of a sugar cube consisting of sucrose with a similar reaction to the combustion of glucose. from Wikipedia.

Given: balanced chemical equation

Asked for: molecule, mole, and mass relationships

Strategy:

A. Use the coefficients from the balanced chemical equation to determine both the molecular and mole ratios.
B. Use the molar masses of the reactants and products to convert from moles to grams.
C. Use Avogadro's number to convert from moles to the number of molecules.

Solution:

This equation is balanced as written: each side has 6 carbon atoms, 18 oxygen atoms, and 12 hydrogen atoms. We can therefore use the coefficients directly to obtain the desired information.

a. One molecule of glucose reacts with 6 molecules of O_2 to yield 6 molecules of CO_2 and 6 molecules of H_2O.
b. One mole of glucose reacts with 6 mol of O_2 to yield 6 mol of CO_2 and 6 mol of H_2O.

c. To interpret the equation in terms of masses of reactants and products, we need their molar masses and the mole ratios from part b. The molar masses in grams per mole are as follows: glucose, 180.16; O_2, 31.9988; CO_2, 44.010; and H_2O, 18.015.

$$\textbf{mass of reactants} = \textbf{mass of products}$$

$$g\,glucose + g\,O_2 = g\,CO_2 + g\,H_2O$$

$$1\,mol\,glucose\left(\frac{180.16\,g}{1\,mol\,glucose}\right) + 6\,mol\,O_2\left(\frac{31.9988\,g}{1\,mol\,O_2}\right)$$

$$= 6\,mol\,CO_2\left(\frac{44.010\,g}{1\,mol\,CO_2}\right) + 6\,mol\,H_2O\left(\frac{18.015\,g}{1\,mol\,H_2O}\right)$$

$$372.15\,g = 372.15\,g$$

C One mole of glucose contains Avogadro's number (6.022×10^{23}) of glucose molecules. Thus 6.022×10^{23} glucose molecules react with ($6 \times 6.022 \times 10^{23}$) = 3.613×10^{24} oxygen molecules to yield ($6 \times 6.022 \times 10^{23}$) = 3.613×10^{24} molecules each of CO_2 and H_2O.

In tabular form:

Solution to Example 3.1.2

	$C_6H_{12}O_{6\,(s)}$	+	$6O_{2\,(g)}$	→	$6CO_{2\,(g)}$		$6H_2O_{(l)}$
a.	1 molecule		6 molecules		6 molecules		6 molecules
b.	1 mol		6 mol		6 mol		6 mol
c.	180.16 g		191.9928 g		264.06 g		108.09 g
d.	6.022×10^{23} molecules		3.613×10^{24} molecules		3.613×10^{24} molecules		3.613×10^{24} molecule

? Exercise 2.6.2: Ammonium Nitrate Explosion

Ammonium nitrate is a common fertilizer, but under the wrong conditions it can be hazardous. In 1947, a ship loaded with ammonium nitrate caught fire during unloading and exploded, destroying the town of Texas City, Texas.



Ammonium nitrate can be hazardous. This aerial photograph of Texas City, Texas, shows the devastation caused by the explosion of a shipload of ammonium nitrate on April 16, 1947. For a video click here.

The explosion resulted from the following reaction:

$$2NH_4NO_{3\ (s)} \rightarrow 2N_{2\ (g)} + 4H_2O_{(g)} + O_{2\ (g)}$$

Construct a table showing how to interpret the information in the equation in terms of

a. individual molecules and ions.
b. moles of reactants and products.
c. grams of reactants and products given 2 mol of ammonium nitrate.
d. numbers of molecules or formula units of reactants and products given 2 mol of ammonium nitrate.

Answer:

Answer to Exercise 3.1.2

	$2NH_4NO_{3\ (s)}$	→	$2N_{2\ (g)}$	+	$4H_2O_{(g)}$	+	$O_{2\ (g)}$
a.	$2NH_4^+$ ions and $2NO_3^-$ ions		2 molecules		4 molecules		1 molecule
b.	2 mol		2 mol		4 mol		1 mol
c.	160.0864 g		56.0268 g		72.0608 g		31.9988 g
d.	1.204×10^{24} formula units		1.204×10^{24} molecules		2.409×10^{24} molecules		6.022×10^{23} molecules



Finding Mols and Masses of Reactants and Products Using Stoichiometric Factors (Mol Ratios): Finding Mols and Masses of Reactants and Products Using Stoichiometric Factors, YouTube(opens in new window) [youtu.be]

Summary

A chemical reaction is described by a chemical equation that gives the identities and quantities of the reactants and the products. In a chemical reaction, one or more substances are transformed to new substances. A chemical reaction is described by a **chemical equation**, an expression that gives the identities and quantities of the substances involved in a reaction. A chemical equation shows the starting compound(s)—the **reactants**—on the left and the final compound(s)—the **products**—on the right, separated by an arrow. In a balanced chemical equation, the numbers of atoms of each element and the total charge are the same on both sides of the equation. The number of atoms, molecules, or formula units of a reactant or product in a balanced chemical equation is the **coefficient** of that species. The **mole ratio** of two substances in a chemical reaction is the ratio of their coefficients in the balanced chemical equation.

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2.7: 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 2.7.1

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 2.7.1

Interpret this balanced chemical equation in terms of moles.

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$$

Answer

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

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 2.7.2

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 2.7.2

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer

118 mol

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.

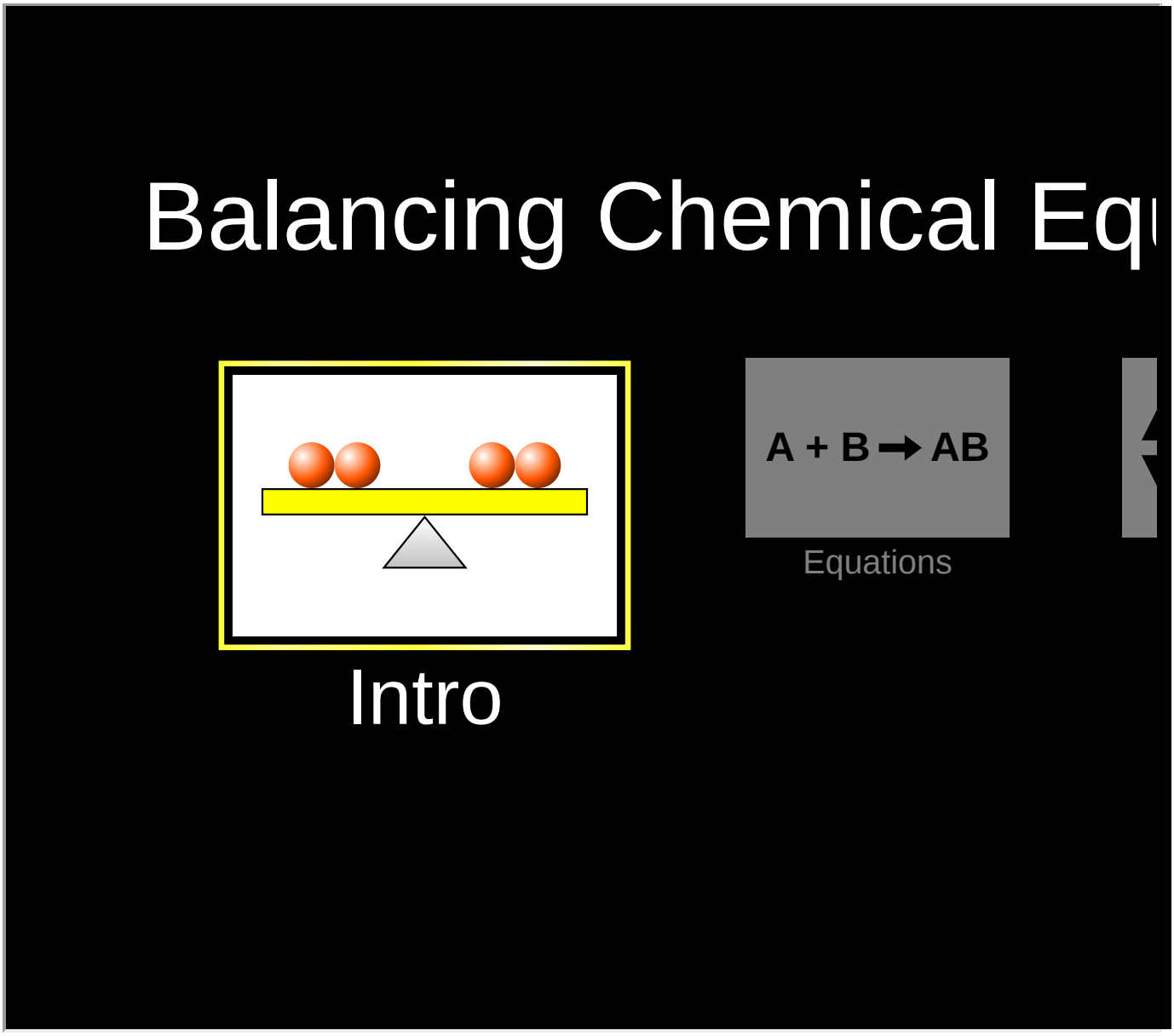
Summary

Balanced chemical reactions are balanced in terms of moles. A balanced chemical reaction gives equivalents in moles that allow stoichiometry calculations to be performed.

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2.8: Activity: Balancing Chemical Equations



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2.9: Some Simple Patterns of Chemical Reactivity

Learning Objectives

- To introduce the basic patterns of chemical reaction.

There is an underlying explanation of why a reaction takes place from an underlying chemical theory based on chemical reactivity. In this sections, we discussed various reactions to identify key chemical trends that allows chemists to predict the outcome of chemical reactions. By recognizing general patterns of chemical reactivity, you will be able to successfully predict the products formed by a given combination of reactants We can often predict a reaction if we have seen a similar reaction before. For example, sodium ($\mathbf{Na}$) reacts with water ($\mathbf{H_2O}$) to form sodium hydroxide ($\mathbf{NaOH}$) and $\mathbf{H_2}$ gas:

$$\mathbf{2\,Na\,(s) + 2\,H_2O\,(l) \longrightarrow 2\,NaOH\,(aq) + H_2\,(g)}$$

with $\mathbf{(aq)}$ indicating aqueous liquid. As discussed later, potassium ($\mathbf{K}$) is in the same family (column) of elements as sodium and exhibits similar chemistry. Therefore, one might *predict* that the reaction of $\mathbf{K}$ with $\mathbf{H_2O}$ would be similar to that of $\mathbf{Na}$:

$$\mathbf{2\,K\,(s) + 2\,H_2O\,(l) \longrightarrow 2\,KOH\,(aq) + H_2\,(g)}$$

In fact, all alkali metals react with water to form their hydroxide compounds and hydrogen.

Combination and Decomposition Reactions

In ***combination reactions,*** two or more compounds react to form a single, more complex compound. Many elements react with one another in this fashion to form compounds. The general chemical equation for combination reaction is:

$$\mathbf{A + B \longrightarrow C}$$

and an example includes the generation of ammonia ($\mathbf{NH_3}$) from nitrogen ($\mathbf{N_2}$), and hydrogen ($\mathbf{H_2}$):

$$\mathbf{N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g)}$$

In ***decomposition reactions*** one substance undergoes a reaction to form two or more simpler products. Such reactions often occur when compounds are heated or electricity is added. The general chemical equation for decomposition reaction is:

$$\mathbf{A \longrightarrow B + C}$$

For example, the thermal decomposition of limestone ($\mathbf{CaCO_3\,(s)}$) generates quicklime ($\mathbf{CaO\,(s)}$) and carbon dioxide $\mathbf{CO_2\,(g)}$

$$\mathbf{CaCO_3\,(s) \longrightarrow CaO\,(s) + CO_2\,(g)}$$

Combustion in Air

Combustion reactions are rapid reactions that produce a flame. Most common combustion reactions involve oxygen (O_2) from the air as a *reactant*. For example, wood will combust in the presence of oxygen in that air (with an ignition event) to generate a fire.



Figure **2.9.1**: The flames caused as a result of a fuel undergoing combustion (burning) (CC BY-SA 2.5; Et bål via Wikipedia)

A common class of compounds which can participate in combustion reactions are *hydrocarbons* (compounds that contain only carbon and hydrogen).

Table **2.9.1**: Examples of common hydrocarbons

Hydrocarbon	Molecular formula
methane	CH_4
propane	C_3H_8
butane	C_4H_{10}
octane	C_8H_{18}

When hydrocarbons are combusted they react with oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O). For example, when propane is burned the reaction is:

$$C_3H_8\,(g) + 5\,O_2\,(g) \longrightarrow 3\,CO_2\,(g) + 4\,H_2O\,(l) \tag{2.9.1}$$

Other compounds which contain carbon, hydrogen and oxygen (e.g. the alcohol *methanol* CH_3OH, and the sugar *glucose* $C_6H_{12}O_6$) also combust in the presence of oxygen (O_2) to produce CO_2 and H_2O.

Contributors and Attributions

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CHAPTER OVERVIEW

3: Atoms and the Periodic Table

3.1: Development of the Modern Periodic Table
3.2: Looking for Patterns: The Periodic Table
3.3: Indivisible - The Atomic Theory
3.4: The Properties of Protons, Neutrons, and Electrons
3.5: Activity: Build an Atom
3.6: Atomic Number
3.7: Mass Number
3.8: Isotopes
3.9: Atomic Models of the Twentieth Century
3.10: Bohr's Atomic Model
3.11: Understanding Atomic Spectra
3.12: Predicting Ion Charges

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3.1: Development of the Modern Periodic Table

Learning Objectives

- Appreciate how the modern periodic table was constructed.
- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.
- Distinguish between the A (atomic mass) value and the Z (atomic number) value when viewing the periodic table.
- Compare and contrast Mendeleev's and Moseley's ordering of the periodic table.

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, 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 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 regularly throughout the table (that is, they are periodic), it became known as the periodic table.

	4 werthig	3 werthig	2 werthig	1 werthig	1 werthig	2 werthig
	-	-	-	-	Li = 7,03	(Be =9,3?)
Differenz =	-	-	-	-	16,02	(14,7)
	C = 12,0	N = 14,04	O = 16,00	Fl = 19,0	Na = 23,05	Mg = 24,0
Differenz =	16,5	16,96	16,07	16,46	16,08	16,0
	Si = 28,5	P = 31,0	S = 32,07	Cl = 35,46	K = 39,13	Ca = 40,0
Differenz =	89,1 /2 = 44,55	44,0	46,7	44,51	46,3	47,6
	-	As = 75,0	Se = 78,8	Br = 79,97	Rb =85,4	Sr = 87,6
Differenz =	89,1 /2 = 44,55	45,6	49,5	46,8	47,6	49,5
	Sn = 117,6	Sb = 120,6	Te = 128,3	J =126,8	Cs = 133,0	Ba = 137,1
Differenz =	89,4 = 2*44,7	87,4 = 2*43,7	-	-	(71 = 2*35,5)	-
	Pb = 207,0	Bi = 208,0	-	-	(Tl = 204?)	-

Figure **3.1.1**: Copy of table on page 137 of (Julius) Lothar Meyer: "Die modernen Theorien der Chemie" (1864). (Copyright;By Kawarayaki - Own work, CC BY-SA 4.0, https://commons.wikimedia.org/w/inde...curid=75552705)

Dimitri Mendeleev (1834–1907)

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

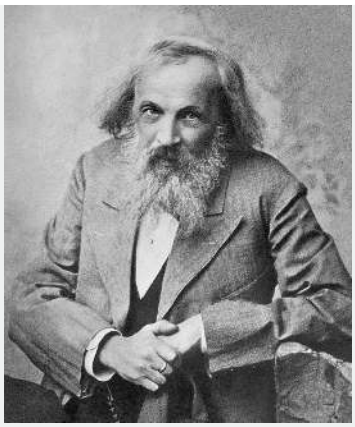


Figure **3.1.2**: Dmitri Mendeleev, the father of the modern periodic table.

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

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



Mendeleev's Periodic Table

Mendeleev, who first published his periodic table in 1869 (Figure **3.1.1**), is usually credited with the origin of the modern periodic table. The key difference between his arrangement of the elements and that of Meyer and others is that Mendeleev did not assume that all the elements had been discovered (actually, only about two-thirds of the naturally occurring elements were known at the time). Instead, he deliberately left blanks in his table at atomic masses 44, 68, 72, and 100, in the expectation that elements with those atomic masses would be discovered. Those blanks correspond to the elements we now know as scandium, gallium, germanium, and technetium.

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

Figure **3.1.3**: Mendeleev's Periodic Table, as Published in the German Journal Annalen der Chemie und Pharmacie in 1872. The column headings "Reihen" and "Gruppe" are German for "row" and "group." Formulas indicate the type of compounds formed by each group, with "R" standing for "any element" and superscripts used where we now use subscripts. Atomic masses are shown after equal signs and increase across each row from left to right.

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

The most convincing evidence in support of Mendeleev's arrangement of the elements was the discovery of two previously unknown elements whose properties closely corresponded with his predictions (Table **3.1.1**). Two of the blanks Mendeleev had left in his original table were below aluminum and silicon, awaiting the discovery of two as-yet-unknown elements, eka-aluminum and eka-silicon (from the Sanskrit eka, meaning "one," as in "one beyond aluminum"). The observed properties of gallium and germanium matched those of eka-aluminum and eka-silicon so well that once they were discovered, Mendeleev's periodic table rapidly gained acceptance.

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

Table **3.1.1**: Comparison of the Properties Predicted by Mendeleev in 1869 for eka-Aluminum and eka-Silicon with the Properties of Gallium (Discovered in 1875) and Germanium (Discovered in 1886)

Property	*eka*-Aluminum (predicted)	Gallium (observed)	*eka*-Silicon (predicted)	Germanium (observed)
atomic mass	68	69.723	72	72.64
element	metal	metal	dirty-gray metal	gray-white metal
	low mp*	mp = 29.8°C	high mp	mp = 98°C
	d = 5.9 g/cm^3	d = 5.91 g/cm^3	d = 5.5 g/cm^3	d = 5.323 g/cm^3
oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
	d = 5.5 g/cm^3	d = 6.0 g/cm^3	d = 4.7 g/cm^3	d = 4.25 g/cm^3
chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$
	volatile	mp = 78°C bp* = 201°C	bp < 100°C	bp = 87°C

*mp = melting point; bp = boiling point.

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

The Role of the Atomic Number in the Periodic Table

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

H. G. J. Moseley (1887–1915)

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

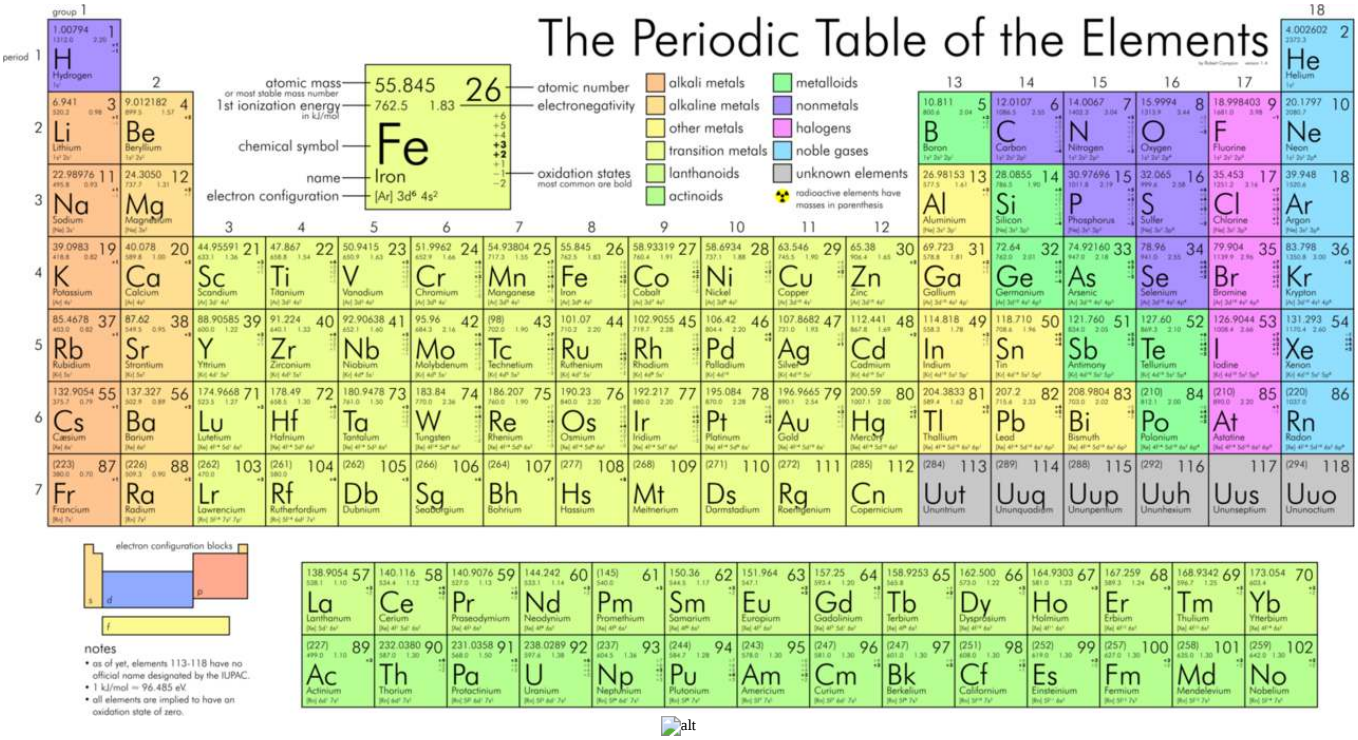


Figure **3.1.4**:G.J. Moseley

Features of the Periodic Table

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

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



alt

Figure **3.1.5**: Types of Elements. Elements are either metals, nonmetals, or metalloids. Each group is located in a different part of the periodic table. An interactive Periodic table can be found here. (Public Domain; 2012rc).

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi-metal), as shown in Figure **3.1.5**. 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 **3.1.5**, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids**. Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

Example 3.1.1

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

a. Se
b. Mg
c. Ge

Solution

a. In Figure **3.1.5**, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

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

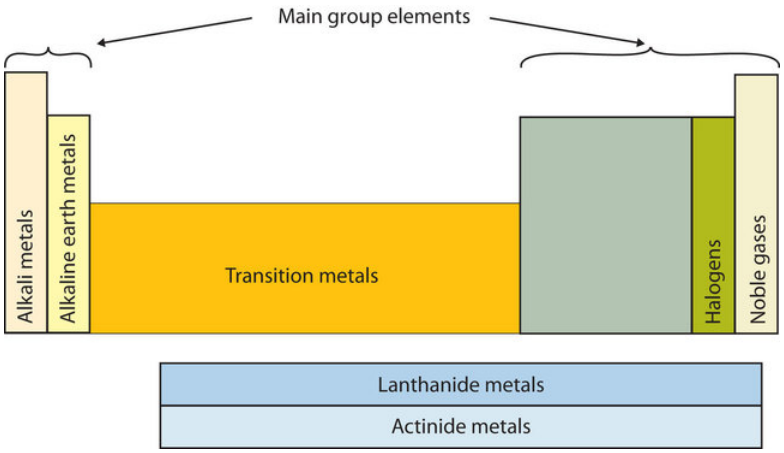


Figure **3.1.6**: Some sections of the periodic table have specific names. Be sure to know the section names for quizzes and tests.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins. To watch a video of your instructor presenting these groups and periods, access your Moodle class chapter 3 documents.

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

Example 3.1.2: Groups

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

a. Li
b. Ar
c. Ra

Solution:

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

Example 3.1.3: Classification of Elements

Provide elemental names for the following combinations:

a. The alkali metal in period three.
b. The halogen in period two
c. A metalloid in period four
d. A transition metal in period three

Solution:

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

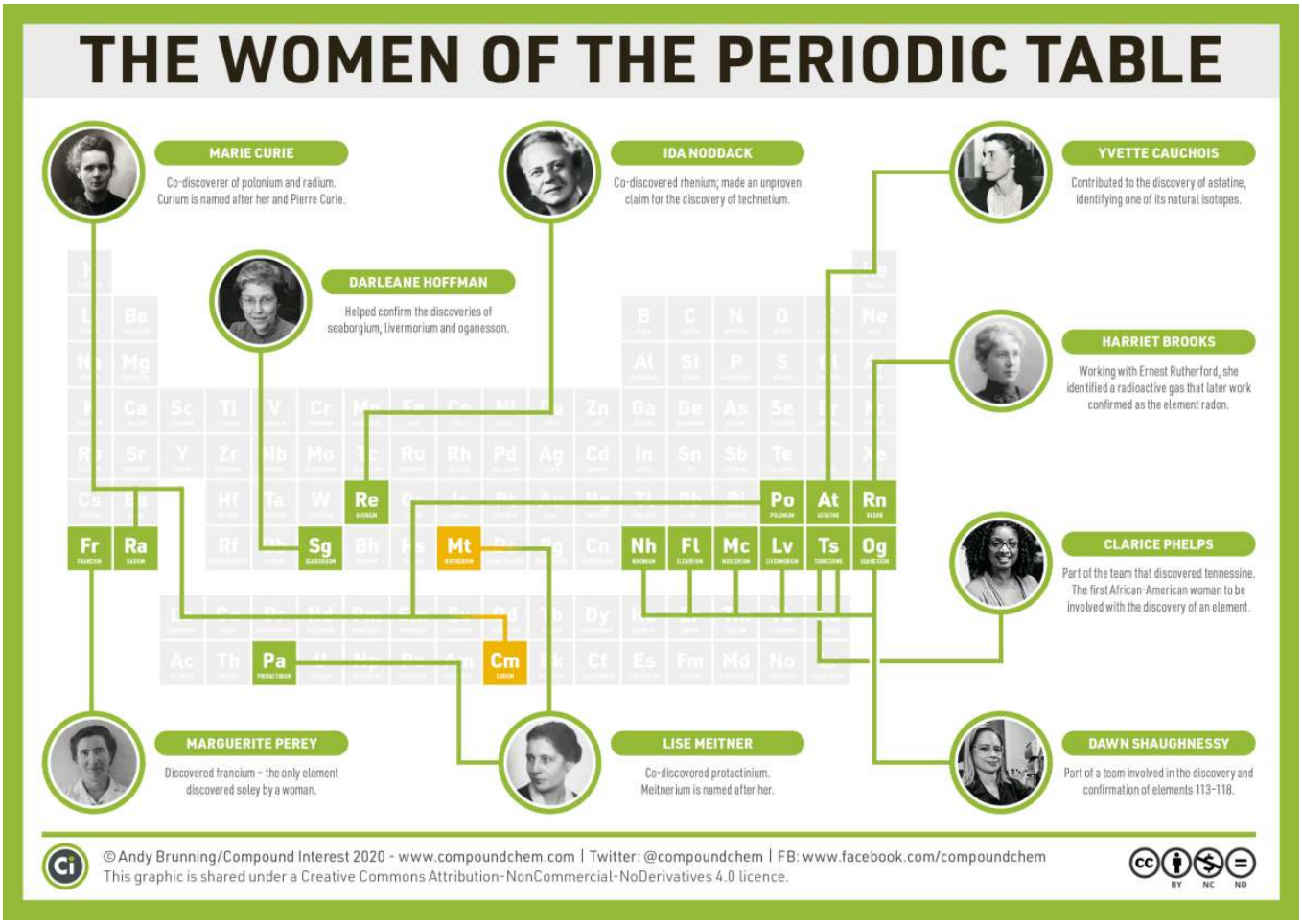


Figure **3.1.7**: The Women of the Periodic Table Graphic by Andy Brunning.

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
2. Sisler, Harry H. Electronic structure, properties, and the periodic law. New york; Reinhold publishing corporation, 1963.
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3.2: Looking for Patterns: 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 of the known elements on the basis of their chemical properties. A modern version is shown in Figure **3.2.1**. 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.

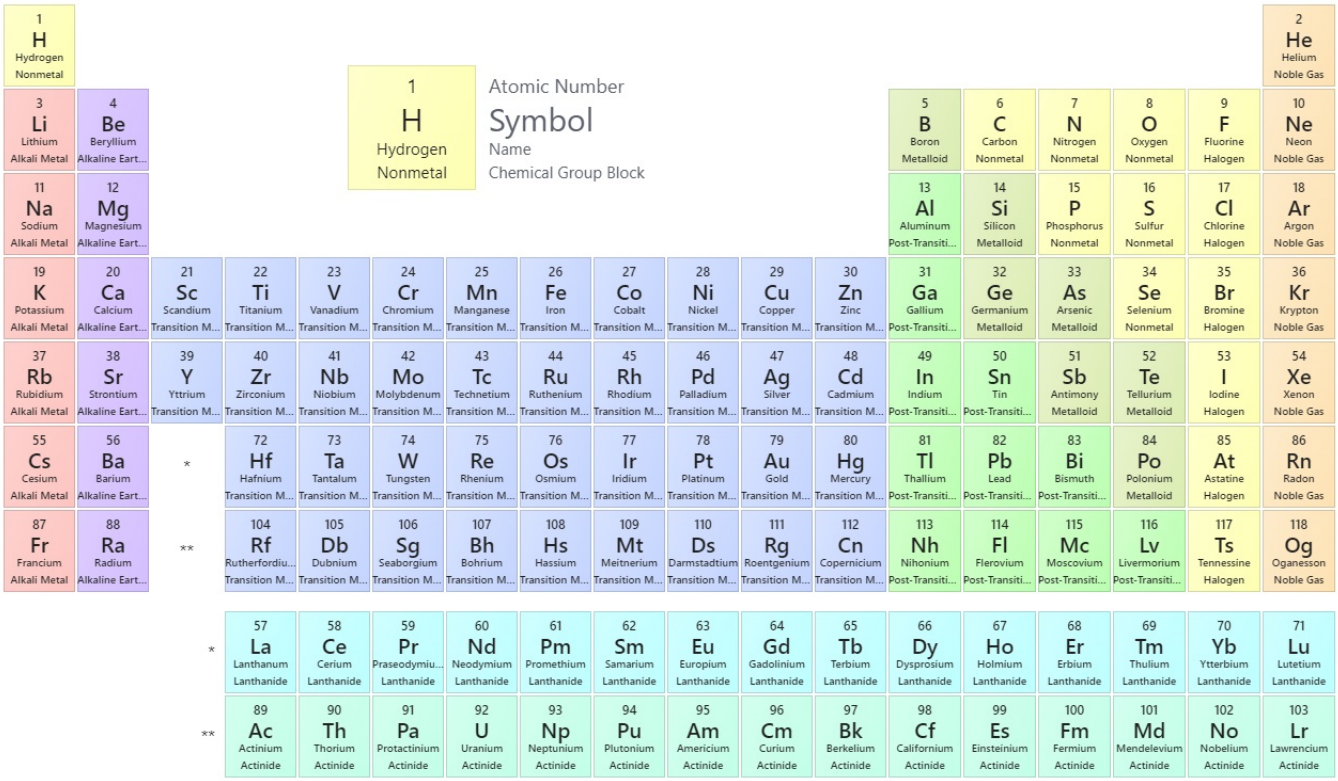


Figure **3.2.1**: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found Periodic Table of Elements. (Public Domain; PubChem via NIH)

Features of the Periodic Table

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

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

To Your Health: Radon

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

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

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

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

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure **3.2.2**. 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 **3.2.2**, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

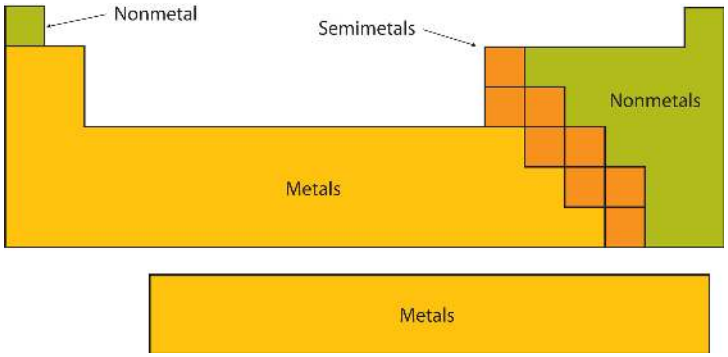


Figure **3.2.2**: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

Example 3.2.1

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

a. Se
b. Mg
c. Ge

Solution

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

Exercise 3.2.1

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

Answer

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure **3.2.3**. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

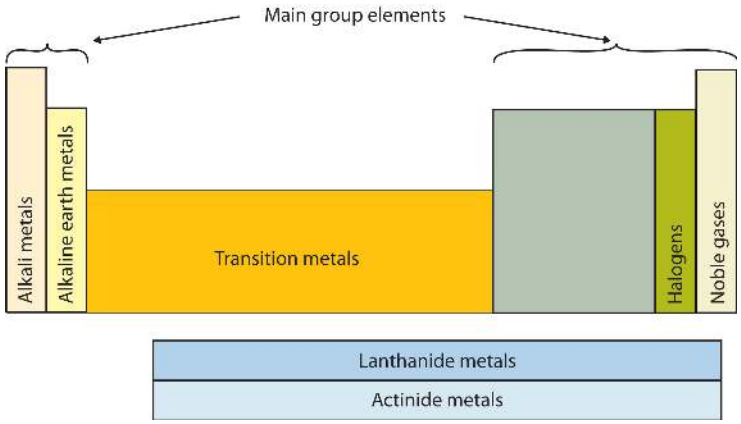


Figure **3.2.3**: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Sections of period table: Alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanide metals, actinide metals.

Descriptive Names

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

Group 1: The Alkali Metals

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

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



Video: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via https://youtu.be/6ZY6d6jrq-0)

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 of the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth's ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

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

Example 3.2.2: Groups

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

Solution

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

Exercise 3.2.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a:

Fluorine is a halogen (Group 17).

Answer b:

Calcium is a alkaline earth metal (Group 2).

Answer c:

Krypton is a noble gas (Group 18).

Example 3.2.3: Classification of Elements

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

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

Solution

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

Exercise 3.2.3: Classification of Elements

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

a. F
b. U
c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

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

Answer c:

Copper is a metal (and a transition metal too).

Summary

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

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

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
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3.3: Indivisible - The Atomic Theory

Learning Objectives

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

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

Democritus and the Greek Philosophers

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



Figure **3.3.1**: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

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

Greeks: "All Thought and No Action"

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

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British

scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

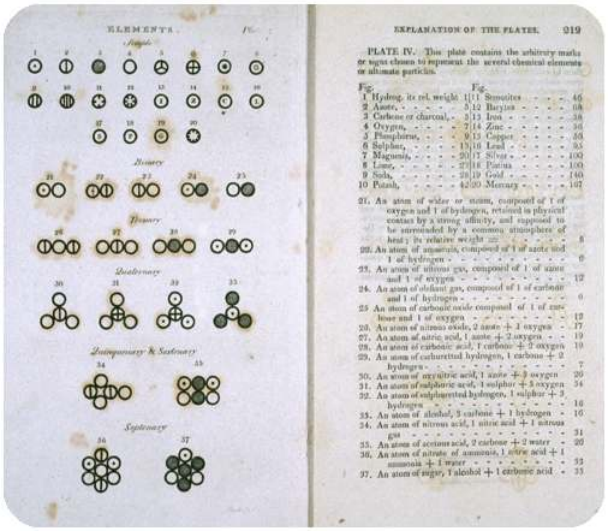
The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure **3.3.1**), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual 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).

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

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

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

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


ELEMENTS.

EXPLANATION OF THE PLATES. 219

PLATE IV. This plate contains the arbitrary marks or signs chosen to represent the several chemical elements or ultimate particles.

Fig.		Fig.	
1	Hydrog. its rel. weight	11	Strontites
2	Azote,	12	Barytes
3	Carbone or charcoal,	13	Iron
4	Oxygen,	14	Zinc
5	Phosphorus,	15	Copper
6	Sulphur,	16	Lead
7	Magnesia,	17	Silver
8	Lime,	18	Platina
9	Soda,	19	Gold
10	Potash,	20	Mercury

21. An atom of water or steam, composed of 1 of oxygen and 1 of hydrogen, retained in physical contact by a strong affinity, and supposed to be surrounded by a common atmosphere of heat; its relative weight
22. An atom of ammonia, composed of 1 of azote and 1 of hydrogen
23. An atom of nitrous gas, composed of 1 of azote and 1 of oxygen
24. An atom of olefiant gas, composed of 1 of carbone and 1 of hydrogen
25. An atom of carbonic oxide composed of 1 of carbone and 1 of oxygen
26. An atom of nitrous oxide, 2 azote + 1 oxygen
27. An atom of nitric acid, 1 azote + 2 oxygen
28. An atom of carbonic acid, 1 carbone + 2 oxygen
29. An atom of carburetted hydrogen, 1 carbone + 2 hydrogen
30. An atom of oxynitric acid, 1 azote + 2 oxygen
31. An atom of sulphuric acid, 1 sulphur + 3 oxygen
32. An atom of sulphurretted hydrogen, 1 sulphur + 3 hydrogen
33. An atom of alcohol, 3 carbone + 1 hydrogen
34. An atom of nitrous acid, 1 nitric acid + 1 nitrous gas
35. An atom of acetous acid, 2 carbone + 2 water
36. An atom of nitrate of ammonia, 1 nitric acid + 1 ammonia + 1 water
37. An atom of sugar, 1 alcohol + 1 carbonic acid

Figure **3.3.2**: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video **3.3.1**, control of individual atoms can be use used create animations.



Video **3.3.1**: A Boy And His Atom: The World's Smallest Movie.

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

Summary

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

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3.4: The Properties of Protons, Neutrons, and Electrons

Learning Objectives

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

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

Electrons

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

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



Figure **3.4.1**: Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

Protons

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

Neutrons

Atoms of all elements—except for most atoms of hydrogen—have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge—they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic

mass unit (amu). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

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

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

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

Table 3.4.1: Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p^+	1	1	+1	inside the nucleus
electron	e^-	5.45×10^{-4}	0.00055	−1	outside the nucleus
neutron	n^0	1	1	0	inside the nucleus

Table 3.4.1 gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth of the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

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

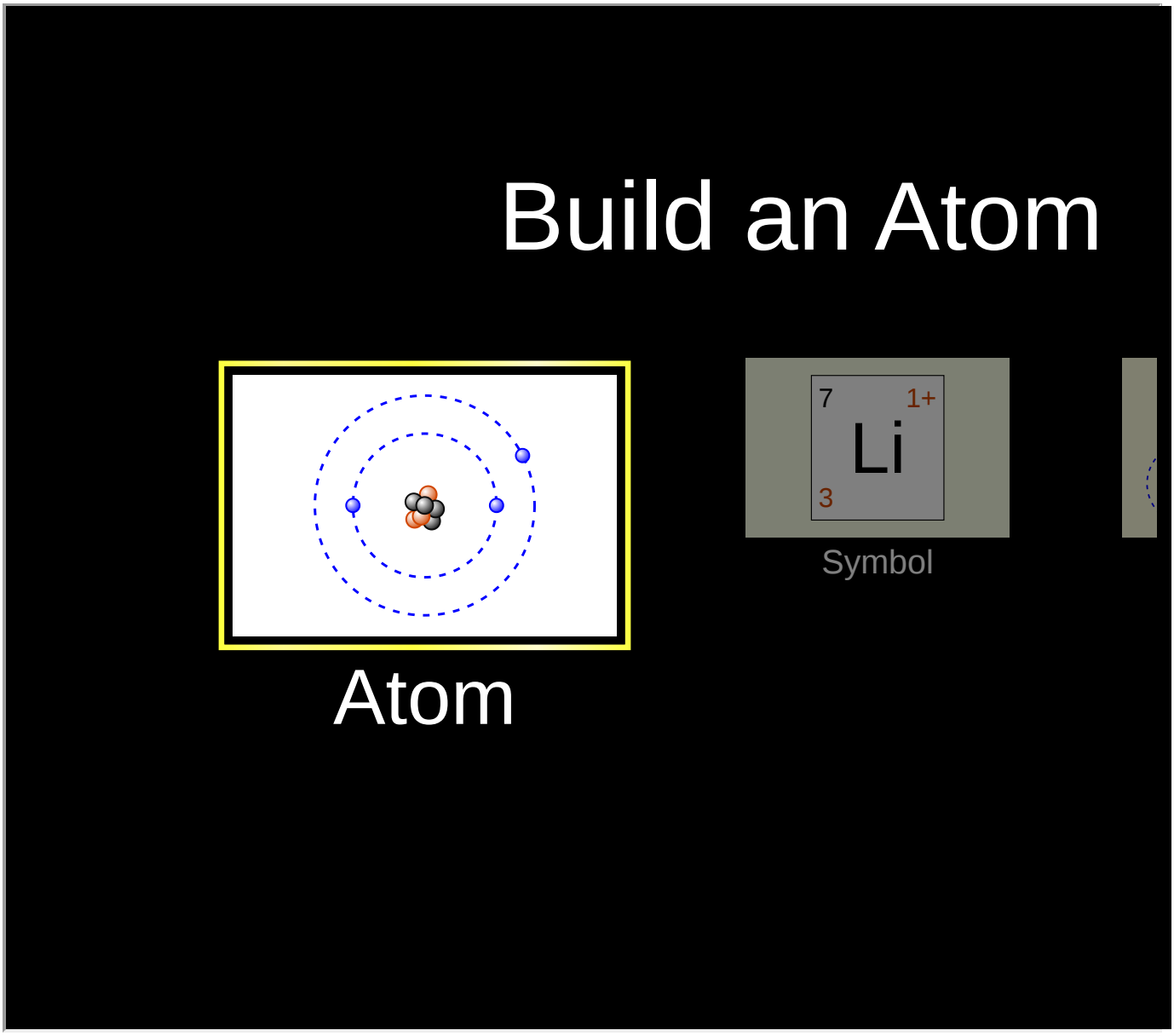
Summary

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

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3.5: Activity: Build an Atom



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3.6: Atomic Number



Figure **3.6.1** (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

What is unique about each one of us?

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

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

Organizing the Elements

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

Atomic Number

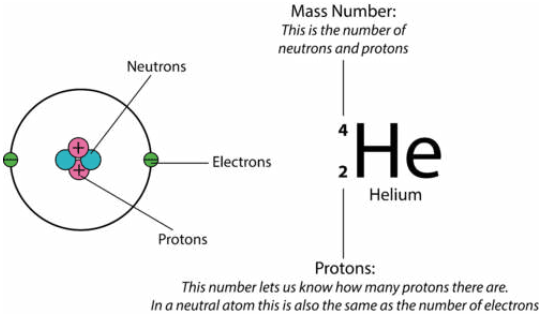


Figure **3.6.2**: How can you determine the atomic number of an element. (Public Domain; User:Materialscientist/Wikimedia Commons via Wikipedia)

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

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

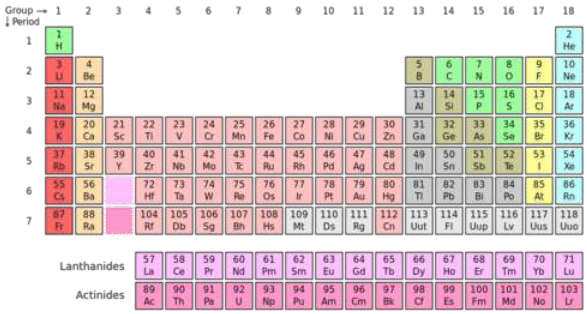


Figure **3.6.3**: The periodic table classifies elements by atomic number. (Public Domain; User:Cepheus/Wikimedia Commons via Wikipedia)

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



Summary

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

Review

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

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3.7: Mass Number

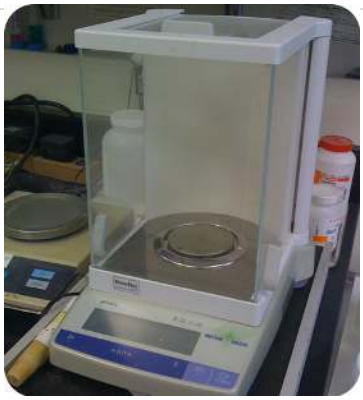


Figure **3.7.1** (Public Domain; User:Ebultoof/Wikipedia via Wikipedia)

How can you determine the mass of a chemical?

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

History of Atomic Weight Determinations

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

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

Mass Number

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

Mass number = atomic number + number of neutrons

Consider Table **3.7.1** below that shows data from the first six elements of the periodic table.

Table **3.7.1**: Atoms of the First Six Elements

Name	Symbol	Protons	Neutrons	Electrons	Atomic Number	Mass Number
Hydrogen	**H**	1	0	1	1	1
Helium	**He**	2	2	2	2	4
Lithium	**Li**	3	4	3	3	7
Beryllium	**Be**	4	5	4	4	9
Boron	**B**	5	6	5	5	11
Carbon	**C**	6	6	6	6	12

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

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

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

Example 3.7.1

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

Solution

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

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

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

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

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



Summary

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

Review

1. Who first determined atomic weights for elements?
2. What were the original atomic weights based on?

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

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3.8: Isotopes



Figure **3.8.1** (Public Domain; Chief Photographer's Mate Johnny Bivera, U.S. Navy via Wikipedia)

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

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

What are Isotopes?

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

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

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

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



Summary

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

Review

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

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3.9: Atomic Models of the Twentieth Century

Learning Objectives

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

Plum Pudding and Bohr Models

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

Figure **3.9.1**: Plum pudding **Figure 3.9.2:** Thomson's depiction of a plum pudding atom.

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

Figure **3.9.3**: The Bohr model of the atom illustrating levels of electrons.

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



Figure **3.9.4**: Scientists who attended the 1927 Fifth Solvay International Conference on Electrons and Photons. Niels Bohr is seated in the second row on the far right. Image taken from: https://www.flickr.com/photos/iharsten/8710094055

Energy Levels

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

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized**, meaning that only specific amounts are possible. To imagine this phenomenon, note the ladder in the image below. Ladders contain rungs that are spaced at specified distances from the ground. On the ladder, an individual must remain on one of the rungs. He or she cannot step between the rungs and remain on the ladder successfully. The lowest rung required the least amount of energy for a person to exert. Moving to higher levels on the ladder would involve inputting more energy or effort.

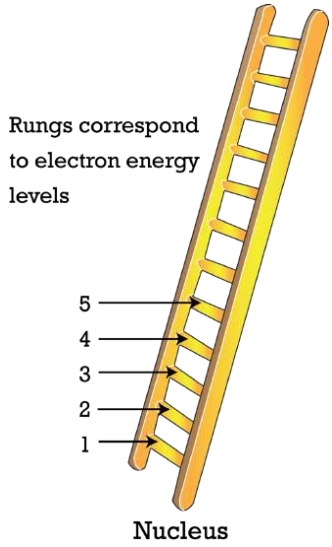


Figure **3.9.5**: The concept of energy levels can be view within the context of a ladder. Each allowed electronic state has an energy. For an atom, the rungs are not equally spaced.

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

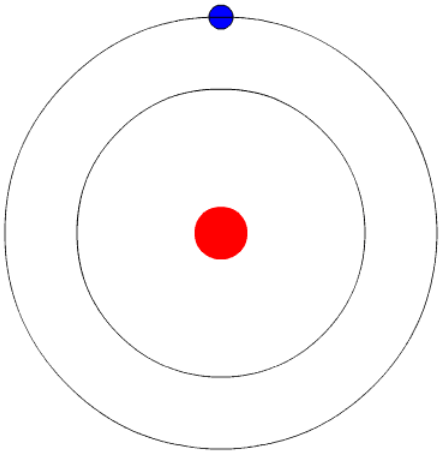


Figure **3.9.6**: In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. (CC BY-SA 3.0; Kurzon).

Drawing Bohr Models

Although not completely correct, the Bohr model can give us a visual representation of electronic levels in atoms. Bohr used the formula $2n^2$ to calculate maximum number of electrons that could be housed in an energy level or shell. In this formula, "n" represents the particular energy level. When substituting an integer above the number zero, one can obtain maximum electrons for each shell. For example, the first energy level would be n =1 and this level could hold $2(1)^2 = 2$ maximum electrons. The second energy level could accommodate $2(2)^2 = 8$ electrons. In drawing Bohr diagrams, different amounts of shells are required to store a specific atom's electrons. Looking at a Bohr diagram for the carbon atom, one can see that two energy levels/shells are required for its six electrons.

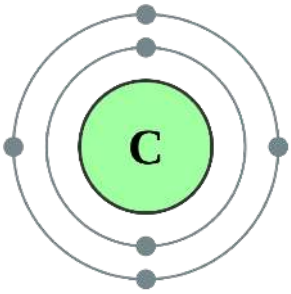


Figure **3.9.7**: Electron shell diagram for carbon, the 6th element in the periodic table of elements. (CC BY-SA 2.0(UK); Pumbaa)

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

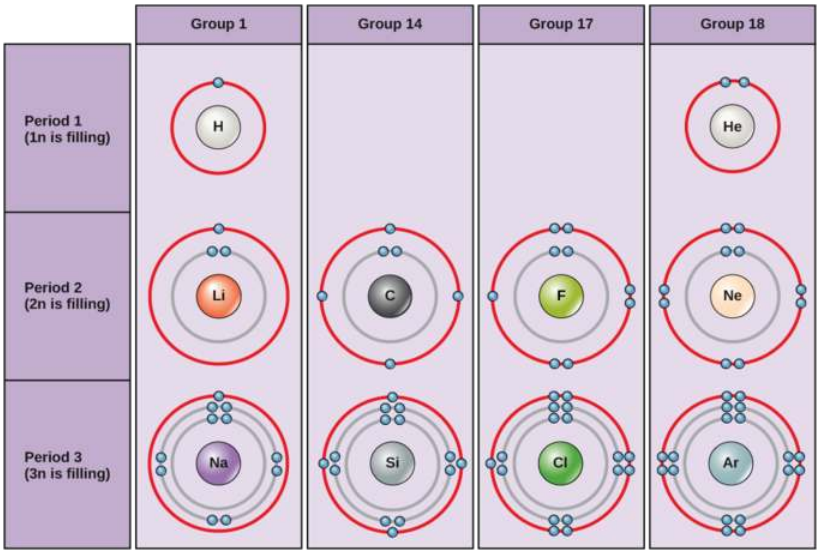


Figure **3.9.8**: Bohr models of different groups/families of elements from the periodic table.

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

✔ Example **3.9.1**: Drawing BOhr Diagrams

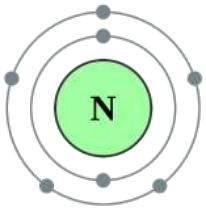
Draw modified Bohr diagrams for

a. nitrogen
b. calcium
c. sulfur

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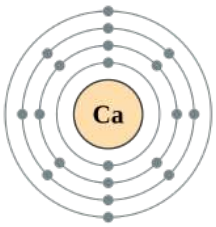
commons.wikimedia.org/wiki/E...-_no_label.svg

Answer a



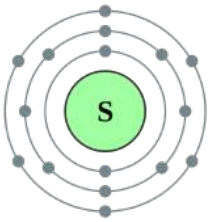
Bohr model of nitrogen: (CC BY-SA 2.0 uk; *Greg Robson):*

Answer b



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

Answer b



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

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

Example **3.9.2**: Identifying Valence Electrons

From the Bohr Diagrams constructed in Example **3.9.1**, how many valence electrons are on each atom.

Answer a

Five.

Answer b

Two

Answer c

Six

Quantum Theory

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

Summary

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

Vocabulary

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

Contributors and Attributions

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-
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3.10: Bohr's Atomic Model



Figure **3.10.1** (Public Domain; the US Department of Agriculture via Wikipedia)

How does this worker's energy change as he climbs up and down the ladder?

Climbing a ladder takes energy. At every step, you push yourself up against gravity, and accumulate potential energy. Coming back down releases that potential energy as you descend step by step. If you are not careful, you can release that potential energy all at once by falling off the ladder (never a good thing). In addition, you take the climb or descent in steps. There is no "in-between" position on the ladder—your foot either hits a rung, or it hits empty space, and you are in trouble until you find a rung to stand on.

Bohr's Atomic Model

Following the discoveries of hydrogen emission spectra and the photoelectric effect, the Danish physicist Niels Bohr (1885-1962) proposed a new model of the atom in 1915. Bohr proposed that electrons do not radiate energy as they orbit the nucleus, but exist in states of constant energy that he called **stationary states**. This means that the electrons orbit at fixed distances from the nucleus (see below). Bohr's work was primarily based on the emission spectra of hydrogen. This is also referred to as the planetary model of the atom. It explained the inner workings of the hydrogen atom. Bohr was awarded the Nobel Prize in physics in 1922 for his work.

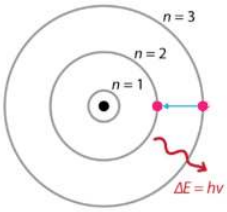


Figure **3.10.2**: Bohr's atomic model hydrogen emission spectra. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Bohr explained that electrons can be moved into different orbits with the addition of energy. When the energy is removed, the electrons return back to their ground state, emitting a corresponding amount of energy—a quantum of light, or photon. This was the basis for what later became known as **quantum theory**. This is a theory based on the principle that matter and energy have the properties of both particles and waves. It accounts for a wide range of physical phenomena, including the existence of discrete packets of energy and matter, the uncertainty principle, and the exclusion principle.

According to the Bohr model, often referred to as a **planetary model**, the electrons encircle the nucleus of the atom in specific allowable paths called orbits. When the electron is in one of these orbits, its energy is fixed. The ground state of the hydrogen atom, where its energy is lowest, is when the electron is in the orbit that is closest to the nucleus. The orbits that are further from the nucleus are all of successively greater energy. The electron is not allowed to occupy any of the spaces in between the orbits. An everyday analogy to the Bohr model is the rungs of a ladder. As you move up or down a ladder, you can only occupy specific rungs and cannot be in the spaces in between rungs. Moving up the ladder increases your potential energy, while moving down the ladder decreases your energy.

Bohr's work had a strong influence on our modern understanding of the inner workings of the atom. However, his model worked well as an explanation for the emissions of the hydrogen atom, but was seriously limited when applied to other atoms. Shortly after

Bohr published his planetary model of the atom, several new discoveries were made, which resulted in, yet again, a revised view of the atom.



Summary

- The Bohr model postulates that electrons orbit the nucleus at fixed energy levels.
- Orbits further from the nucleus exist at higher energy levels.
- When electrons return to a lower energy level, they emit energy in the form of light.

Review

1. When did Bohr propose his model of the atom?
2. What is a stationary state?
3. What is the ground state?
4. Can the electron occupy any space between the orbits?

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3.11: Understanding Atomic Spectra

Learning Objectives

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

Continuous Spectrum

A rainbow is an example of a continuous spectrum. Here, the colors displayed are within the visible spectrum (between 380-760 nm). Light in this wavelength range is visible to the naked eye. Unlike the visible spectrum, light that is of different wavelengths (see the electromagnetic spectrum below) is not visible. Looking at Figure **3.11.1**, note the different areas of the light (or electromagnetic spectrum). Areas of light that possess short wavelengths are located on the left of the spectrum. All wavelengths from UV (ultraviolet) to γ (gamma) range have the potential to ionize tissues and/or DNA. As a result, individuals who have been exposed to large amounts of these types of radiation in acute time periods could develop cancer. In contrast, visible light (see the rainbow area in Figure **3.11.1**) and radiation on the right side of the spectrum has longer wavelengths and does not have the potential to ionize tissues and/or DNA. Visible, infrared (labeled IR), microwave, and radio waves are classified as being nonionizing radiation and have not been linked to cancer.

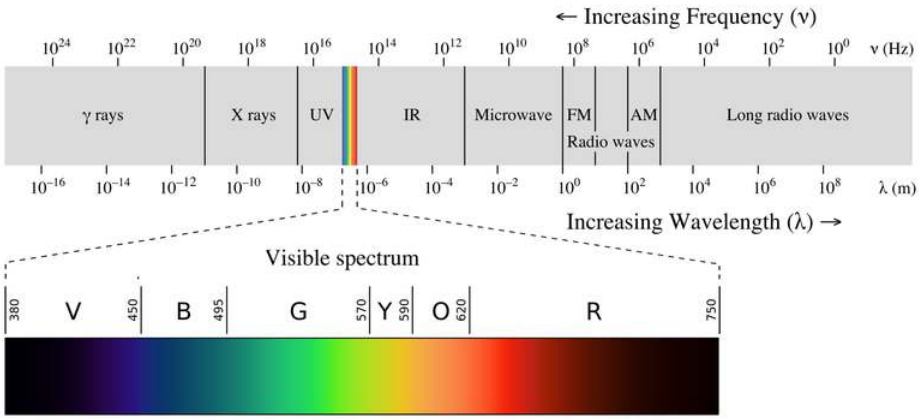


Figure **3.11.1**: The electromagnetic spectrum is an example of a continous spectrum. It contains various frequencies (cycle/time) and wavelengths. (CC BY-SA 3.0; Philip Ronan, Gringer).

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

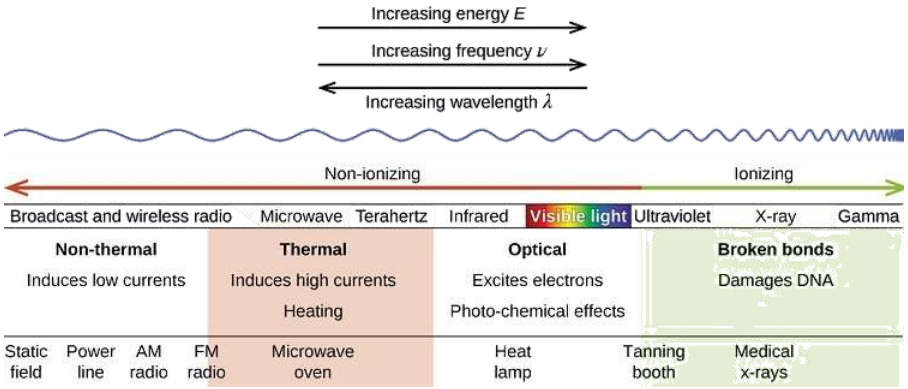


Figure **3.11.2**: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing. (CC BY-SA, OpenStax).

Examples of Radiation

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

Atomic Emission Spectra

The electrons in an atom tend to be arranged in such a way that the energy of the atom is as low as possible. The **ground state** of an atom is the lowest energy state of the atom. When those atoms are given energy, the electrons absorb the energy and move to a higher energy level. These energy levels of the electrons in atoms are quantized, meaning again that the electron must move from one energy level to another in discrete steps rather than continuously. An **excited state** of an atom is a state where its potential energy is higher than the ground state. An atom in the excited state is not stable. When it returns back to the ground state (Figure **3.11.3**), it releases the energy that it had previously gained often in the form of electromagnetic radiation (although it can be released via heat).

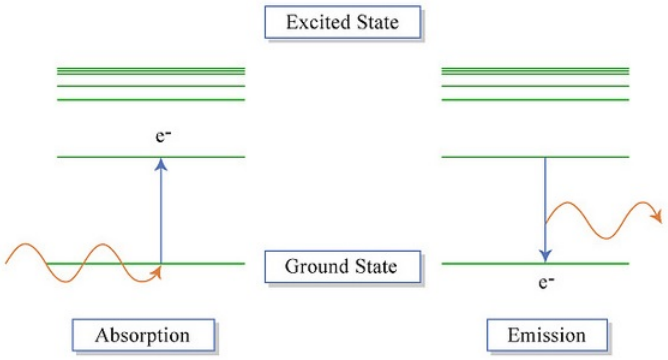


Figure **3.11.3**: Atomic absorption and emission in excited and ground states. (CC BY-SA-NC 4.0; MIT OpenCourseWare).

Atoms can gain energy to induce these transitions from various sources. The gases in the image below have been excited with the use of electrical current. The atoms in each of these noble gases produce distinctive colors that can be used to identify the elements (Figure **3.11.4**). Each of these species contains a different number of electrons that can undergo different types of excitations. In turn, each gas produces a signature color.

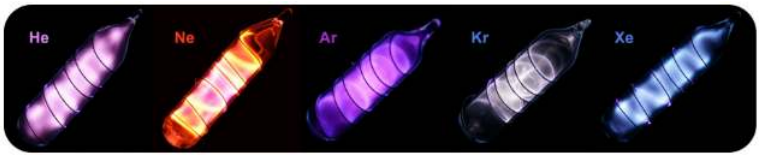


Figure **3.11.4**: Gas discharge tubes are enclosed glass tubes filled with a gas at low pressure through which an electric current is passed. Electrons in the gaseous atoms first become excited, and then fall back to lower energy levels, emitting light of a distinctive color in the process. Shown ar gas discharge tubes of helium, neon, argon, krypton, and xenon.

Flames can be utilized to excite atoms as well. During a flame test experiment, metal chlorides are directly placed into a flame. The intense heat will promote the metal's electrons to an excited state. Upon emission, this extra energy is released in the form of visible light. If a reference panel is provided, the flame color can be used to identify a metal atom (Table **3.11.1**).



Video **3.11.1**: Burning Salts/Flame Test demonstration: Sully Science

Table **3.11.1**: Flame colors corresponding to elements

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

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



Video 3.11.2: Guinness World Record For World's Largest Aerial Firework Shell

Discontinous Spectra

If the light emitted from the excited atoms is viewed through a prism, then individual patterns of lines will be produced. These lines are called spectra and correspond to fingerprint wavelengths (symbol for wavelength is λ) for a specific element. Figure 3.11.5 illustrates how the light from excited electrons can be diffracted to produce line spectra for the elements of hydrogen, helium, and iron. The specific elements produce wavelengths within the visible spectrum (between 400-700 nm) and can be seen by the naked eye. In order to obtain the numerical wavelengths (in nanometers), one would need to employ some type of detector.

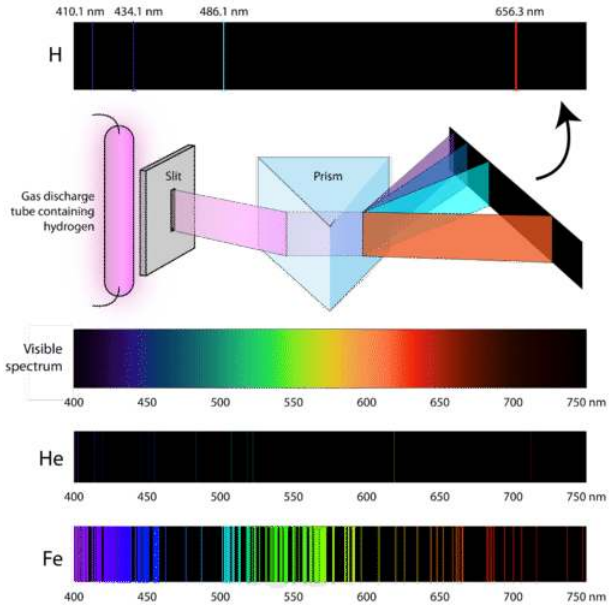


Figure 3.11.5: When light from a hydrogen gas discharge tube is passed through a prism, the light is split into four visible lines. Each of these spectral lines corresponds to a different electron transition from a higher energy state to a lower energy state. Every element has a unique atomic emission spectrum, as shown by the examples of helium (He) and iron (Fe). (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

Atomic Absorption Spectra

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

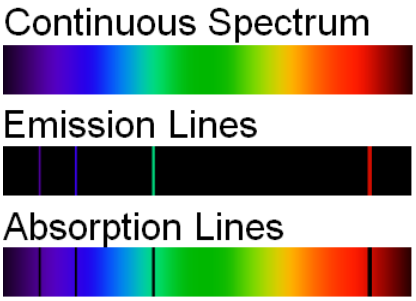


Figure **3.11.6**: Spectral lines. (Public Domain; Jhausauer via Wikipedia).

Example 3.11.1:

Which statements is/are true?

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

Answer a

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

Answer b

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

Answer c

True, scientists need those fingerprint wavelengths to identify atoms.

Answer d

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

Answer e

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

Answer f

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

Answer g

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

Summary

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

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

- The existence of the atomic spectra is support for Bohr's model of the atom.
- Bohr's model was only successful in calculating energy levels for the hydrogen atom.
- The emitted light corresponds to energies of the specific electrons.

Vocabulary

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

Contributors and Attributions

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3.12: Predicting Ion Charges

Learning Objectives

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

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

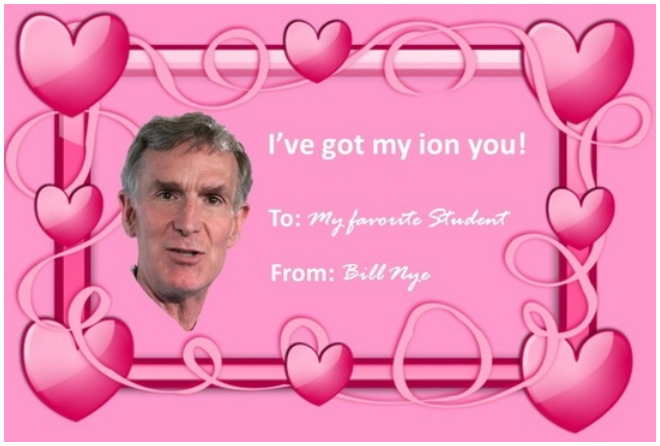


Figure **3.12.1**: Valentine image constructed from Pink Hearts Frame from 365PSD.com and photo of Bill Nye by Gage Skidmore (CC-BY=SA 2.0).

Cations

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

$$\mathbf{Na} \rightarrow \mathbf{Na^+ + e^-}$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure **3.12.1** is a graphical depiction of this process.

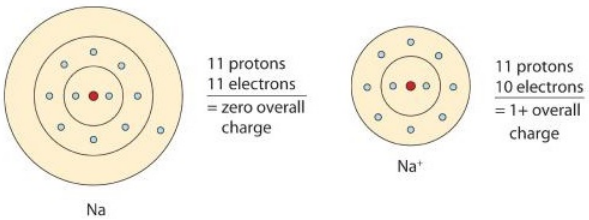


Figure **3.12.2**: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. (CC BY-SA-NC 3.0; anonymous)

Atoms vs. Cations

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



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

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



Follow-up questions

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

Anions

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

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

$$e^- + Cl \longrightarrow Cl^-$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure **3.12.2** is a graphical depiction of this process.

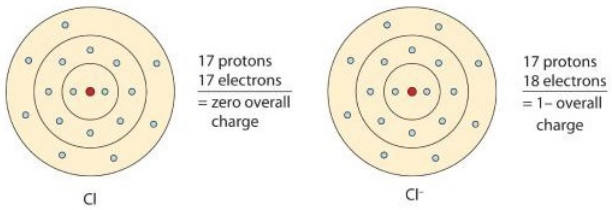
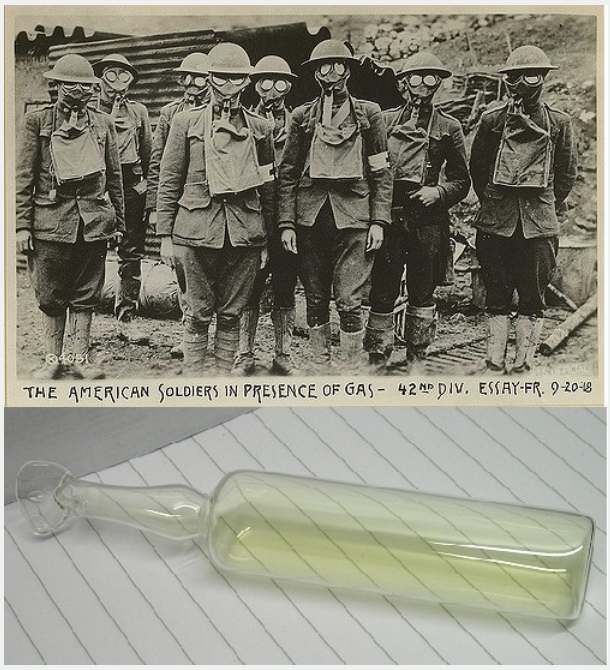


Figure **3.12.3**: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge. (CC BY-SA-NC 3.0; anonymous)

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

Anions Vs. Atoms

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



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

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-

last column, the halogens, form ions having a 1− charge. Figure **3.12.3** shows how the charge on many ions can be predicted by the location of an element on the periodic table.

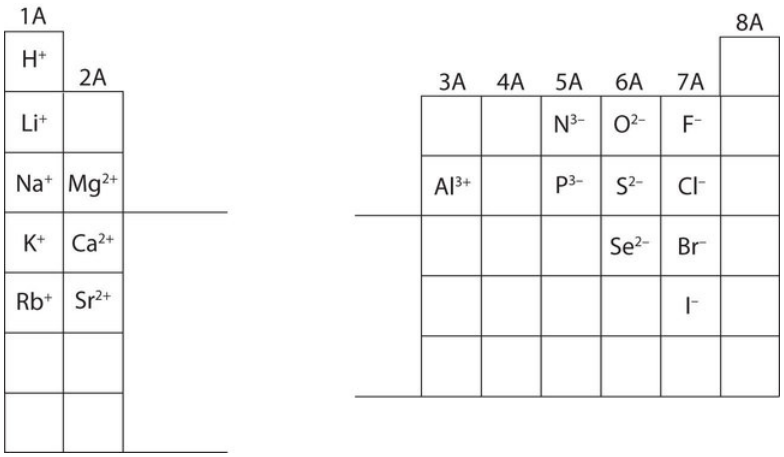


Figure **3.12.3**: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge. Be sure to memorize silver (Ag^{+}) and zinc (Zn^{2+}) charges in addition to this chart.

Example 3.12.1:

Will the species below lose or gain electrons?

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

Answer a

Aluminum is a metal that will always lose three electrons.

Answer b

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

Answer c

Metals will always lose electrons to form cations.

Answer d

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

Answer e

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

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CHAPTER OVERVIEW

4: Light and Waves

4.1: Prelude to Electromagnetic Waves

4.2: The Electromagnetic Spectrum

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CHAPTER OVERVIEW

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4: Light and Waves

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4.1: Prelude to Electromagnetic Waves

The beauty of a coral reef, the warm radiance of sunshine, the sting of sunburn, the X-ray revealing a broken bone, even microwave popcorn—all are brought to us by **electromagnetic waves**. The list of the various types of electromagnetic waves, ranging from radio transmission waves to nuclear gamma-ray (γ-ray) emissions, is interesting in itself.

Even more intriguing is that all of these widely varied phenomena are different manifestations of the same thing—electromagnetic waves. (See Figure 2.) What are electromagnetic waves? How are they created, and how do they travel? How can we understand and organize their widely varying properties? What is their relationship to electric and magnetic effects? These and other questions will be explored.



Figure **4.1.1**: Human eyes detect these orange “sea goldie” fish swimming over a coral reef in the blue waters of the Gulf of Eilat (Red Sea) using visible light. (credit: Daviddarom, Wikimedia Commons)

MISCONCEPTION ALERT: SOUND WAVES VS. RADIO WAVES:

Many people confuse sound waves with **radio waves**, one type of electromagnetic (EM) wave. However, sound and radio waves are completely different phenomena. Sound creates pressure variations (waves) in matter, such as air or water, or your eardrum. Conversely, radio waves are *electromagnetic waves*, like visible light, infrared, ultraviolet, X-rays, and gamma rays. EM waves don’t need a medium in which to propagate; they can travel through a vacuum, such as outer space.

A radio works because sound waves played by the D.J. at the radio station are converted into electromagnetic waves, then encoded and transmitted in the radio-frequency range. The radio in your car receives the radio waves, decodes the information, and uses a speaker to change it back into a sound wave, bringing sweet music to your ears.

Discovering a New Phenomenon

It is worth noting at the outset that the general phenomenon of electromagnetic waves was predicted by theory before it was realized that light is a form of electromagnetic wave. The prediction was made by James Clerk Maxwell in the mid-19th century when he formulated a single theory combining all the electric and magnetic effects known by scientists at that time. “Electromagnetic waves” was the name he gave to the phenomena his theory predicted.

 Figure **4.1.2**: The electromagnetic waves sent and received by this 50-foot radar dish antenna at Kennedy Space Center in Florida are not visible, but help track expendable launch vehicles with high-definition imagery. The first use of this C-band radar dish was for the launch of the Atlas V rocket sending the New Horizons probe toward Pluto. (credit: NASA)

Such a theoretical prediction followed by experimental verification is an indication of the power of science in general, and physics in particular. The underlying connections and unity of physics allow certain great minds to solve puzzles without having all the pieces. The prediction of electromagnetic waves is one of the most spectacular examples of this power. Certain others, such as the prediction of antimatter, will be discussed in later modules.

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4.2: The Electromagnetic Spectrum

Learning Objectives

By the end of this section, you will be able to:

- List three “rules of thumb” that apply to the different frequencies along the electromagnetic spectrum.
- Explain why the higher the frequency, the shorter the wavelength of an electromagnetic wave.
- Draw a simplified electromagnetic spectrum, indicating the relative positions, frequencies, and spacing of the different types of radiation bands.
- List and explain the different methods by which electromagnetic waves are produced across the spectrum.

In this module we examine how electromagnetic waves are classified into categories such as radio, infrared, ultraviolet, and so on, so that we can understand some of their similarities as well as some of their differences. We will also find that there are many connections with previously discussed topics, such as wavelength and resonance. A brief overview of the production and utilization of electromagnetic waves is found in Table **4.2.1**.

Table **4.2.1**: Electromagnetic Wave

Type of EM wave	Production	Applications	Life sciences aspect	Issues
Radio & TV	Acceleration charges	Communications Remote controls	MRI	Requires controls for band use
Microwaves	Accelerating charges & thermal agitation	Communication Ovens Radar	Deep heating	Cell phone use
Infrared	Thermal agitations & electronic transitions	Thermal imaging Heating	Absorbed by atmosphere	Greenhouse effect
Visible light	Thermal agitations & electronic transitions	All pervasive	Photosynthesis Human vision	
Ultraviolet	Thermal agitations & electronic transitions	Sterilization Cancer control	Vitamin D production	Ozone depletion Cancer causing
X-rays	Inner electronic transitions and fast collisions	Medical Security	Medical diagnosis Cancer therapy	Cancer causing
Gamma rays	Nuclear decay	Nuclear medicineSecurity	Medical diagnosis Cancer therapy	Cancer causing Radiation damage

Wave

There are many types of waves, such as water waves and even earthquakes. Among the many shared attributes of waves are propagation speed, frequency, and wavelength. These are always related by the expression $vw = f\lambda$. This module concentrates on EM waves, but other modules contain examples of all of these characteristics for sound waves and submicroscopic particles.

As noted before, an electromagnetic wave has a frequency and a wavelength associated with it and travels at the speed of light, or c. The relationship among these wave characteristics can be described by $vw = f\lambda$, where vw is the propagation speed of the wave, f is the frequency, and λ is the wavelength. Here $vw = c$, so that for all electromagnetic waves,

$$c = f\lambda. \tag{4.2.1}$$

Thus, for all electromagnetic waves, the greater the frequency, the smaller the wavelength.

Figure 4.2.1 shows how the various types of electromagnetic waves are categorized according to their wavelengths and frequencies -- that is, it shows the electromagnetic spectrum. Many of the characteristics of the various types of electromagnetic waves are related to their frequencies and wavelengths, as we shall see.

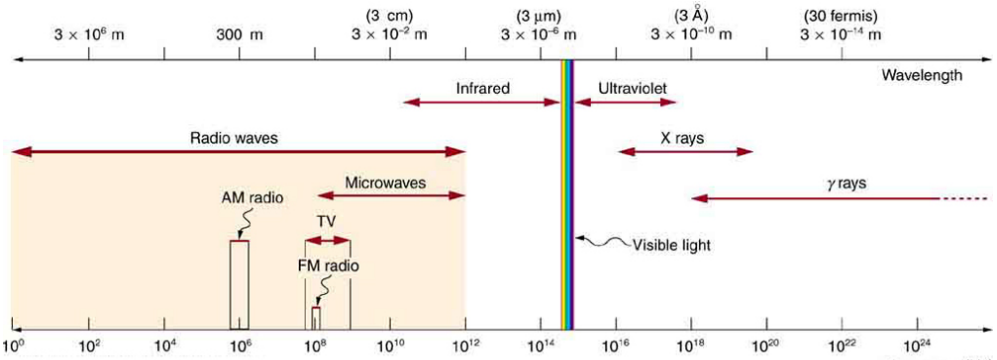


Figure 4.2.1: The electromagnetic spectrum, showing the major categories of electromagnetic waves. The range of frequencies and wavelengths is remarkable. The dividing line between some categories is distinct, whereas other categories overlap.

ELECTROMAGNETIC SPECTRUM: RULES OF THUMB

Three rules that apply to electromagnetic waves in general are as follows:

- High-frequency electromagnetic waves are more energetic and are more able to penetrate than low-frequency waves.
- High-frequency electromagnetic waves can carry more information per unit time than low-frequency waves.
- The shorter the wavelength of any electromagnetic wave probing a material, the smaller the detail it is possible to resolve.

Note that there are exceptions to these rules of thumb.

Transmission, Reflection, and Absorption

What happens when an electromagnetic wave impinges on a material? If the material is transparent to the particular frequency, then the wave can largely be transmitted. If the material is opaque to the frequency, then the wave can be totally reflected. The wave can also be absorbed by the material, indicating that there is some interaction between the wave and the material, such as the thermal agitation of molecules.

Of course it is possible to have partial transmission, reflection, and absorption. We normally associate these properties with visible light, but they do apply to all electromagnetic waves. What is not obvious is that something that is transparent to light may be opaque at other frequencies. For example, ordinary glass is transparent to visible light but largely opaque to ultraviolet radiation. Human skin is opaque to visible light -- we cannot see through people -- but transparent to X-rays.

Radio and TV Waves

The broad category of **radio waves** is defined to contain any electromagnetic wave produced by currents in wires and circuits. Its name derives from their most common use as a carrier of audio information (i.e., radio). The name is applied to electromagnetic waves of similar frequencies regardless of source. Radio waves from outer space, for example, do not come from alien radio stations. They are created by many astronomical phenomena, and their study has revealed much about nature on the largest scales.

There are many uses for radio waves, and so the category is divided into many subcategories, including microwaves and those electromagnetic waves used for AM and FM radio, cellular telephones, and TV.

The lowest commonly encountered radio frequencies are produced by high-voltage AC power transmission lines at frequencies of 50 or 60 Hz. (Figure 4.2.2). These extremely long wavelength electromagnetic waves (about 6000 km!) are one means of energy loss in long-distance power transmission.



Figure 4.2.2: This high-voltage traction power line running to Eutingen Railway Substation in Germany radiates electromagnetic waves with very long wavelengths. (credit: Zonk43, Wikimedia Commons)

There is an ongoing controversy regarding potential health hazards associated with exposure to these electromagnetic fields (***E***-fields). Some people suspect that living near such transmission lines may cause a variety of illnesses, including cancer. But demographic data are either inconclusive or simply do not support the hazard theory. Recent reports that have looked at many European and American epidemiological studies have found no increase in risk for cancer due to exposure to ***E***-fields.

Extremely low frequency (ELF) radio waves of about 1 kHz are used to communicate with submerged submarines. The ability of radio waves to penetrate salt water is related to their wavelength (much like ultrasound penetrating tissue) -- the longer the wavelength, the farther they penetrate. Since salt water is a good conductor, radio waves are strongly absorbed by it, and very long wavelengths are needed to reach a submarine under the surface. (Figure 4.2.3).

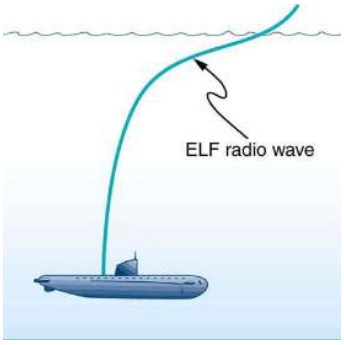


Figure 4.2.3: Very long wavelength radio waves are needed to reach this submarine, requiring extremely low frequency signals (ELF). Shorter wavelengths do not penetrate to any significant depth.

AM radio waves are used to carry commercial radio signals in the frequency range from 540 to 1600 kHz. The abbreviation AM stands for **amplitude modulation**, which is the method for placing information on these waves (Figure 4.2.4). A **carrier wave** having the basic frequency of the radio station, say 1530 kHz, is varied or modulated in amplitude by an audio signal. The resulting wave has a constant frequency, but a varying amplitude.

A radio receiver tuned to have the same resonant frequency as the carrier wave can pick up the signal, while rejecting the many other frequencies impinging on its antenna. The receiver's circuitry is designed to respond to variations in amplitude of the carrier wave to replicate the original audio signal. That audio signal is amplified to drive a speaker or perhaps to be recorded.

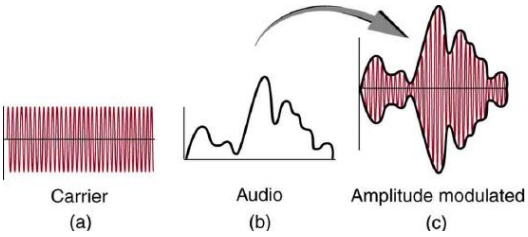


Figure 4.2.4: Amplitude modulation for AM radio. (a) A carrier wave at the station's basic frequency. (b) An audio signal at much lower audible frequencies. (c) The amplitude of the carrier is modulated by the audio signal without changing its basic frequency.

FM Radio Waves

FM radio waves are also used for commercial radio transmission, but in the frequency range of 88 to 108 MHz. FM stands for **frequency modulation**, another method of carrying information (Figure 4.2.5). Here a carrier wave having the basic frequency of

the radio station, perhaps 105.1 MHz, is modulated in frequency by the audio signal, producing a wave of constant amplitude but varying frequency.

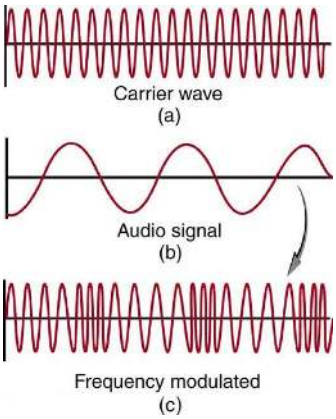


Figure 4.2.5: Frequency modulation for FM radio. (a) A carrier wave at the station's basic frequency. (b) An audio signal at much lower audible frequencies. (c) The frequency of the carrier is modulated by the audio signal without changing its amplitude.

Since audible frequencies range up to 20 kHz (or 0.020 MHz) at most, the frequency of the FM radio wave can vary from the carrier by as much as 0.020 MHz. Thus the carrier frequencies of two different radio stations cannot be closer than 0.020 MHz. An FM receiver is tuned to resonate at the carrier frequency and has circuitry that responds to variations in frequency, reproducing the audio information.

FM radio is inherently less subject to noise from stray radio sources than AM radio. The reason is that amplitudes of waves add. So an AM receiver would interpret noise added onto the amplitude of its carrier wave as part of the information. An FM receiver can be made to reject amplitudes other than that of the basic carrier wave and only look for variations in frequency. It is thus easier to reject noise from FM, since noise produces a variation in amplitude.

Television is also broadcast on electromagnetic waves. Since the waves must carry a great deal of visual as well as audio information, each channel requires a larger range of frequencies than simple radio transmission. TV channels utilize frequencies in the range of 54 to 88 MHz and 174 to 222 MHz. (The entire FM radio band lies between channels 88 MHz and 174 MHz.) These TV channels are called VHF (for **very high frequency**). Other channels called UHF (for **ultra high frequency**) utilize an even higher frequency range of 470 to 1000 MHz.

The TV video signal is AM, while the TV audio is FM. Note that these frequencies are those of free transmission with the user utilizing an old-fashioned roof antenna. Satellite dishes and cable transmission of TV occurs at significantly higher frequencies and is rapidly evolving with the use of the high-definition or HD format.

Example 4.2.1: Calculating Wavelengths of Radio Waves:

Calculate the wavelengths of a 1530-kHz AM radio signal, a 105.1-MHz FM radio signal, and a 1.90-GHz cell phone signal.

Strategy

The relationship between wavelength and frequency is $c = f\lambda$, where $c = 3.00 \times 10^8 m/s$ is the speed of light (the speed of light is only very slightly smaller in air than it is in a vacuum). We can rearrange this equation to find the wavelength for all three frequencies.

Solution

Rearranging gives

$$\lambda = \frac{c}{f}. \tag{4.2.2}$$

a. For the $f = 1530 kHz$ AM radio signal, then,

$$\lambda = \frac{3.00 \times 10^8 m/s}{1530 \times 10^3 cycles/s} \tag{4.2.3}$$

$$= 196m. \tag{4.2.4}$$

b. For the $f = 105.1MHz$ AM radio signal, then,

$$\lambda = \frac{3.00 \times 10^8 m/s}{105.1 \times 10^6 cycles/s} \tag{4.2.5}$$

$$= 2.85m. \tag{4.2.6}$$

c. For the $f = 1.90GHz$ AM radio signal, then,

$$\lambda = \frac{3.00 \times 10^8 m/s}{1.90 \times 10^9 cycles/s} \tag{4.2.7}$$

$$= 0.158m. \tag{4.2.8}$$

Discussion

These wavelengths are consistent with the spectrum in Figure **4.2.1**. The wavelengths are also related to other properties of these electromagnetic waves, as we shall see.

The wavelengths found in the preceding example are representative of AM, FM, and cell phones, and account for some of the differences in how they are broadcast and how well they travel. The most efficient length for a linear antenna, such as discussed in 24.3, is $\lambda/2$, half the wavelength of the electromagnetic wave. Thus a very large antenna is needed to efficiently broadcast typical AM radio with its carrier wavelengths on the order of hundreds of meters.

One benefit to these long AM wavelengths is that they can go over and around rather large obstacles (like buildings and hills), just as ocean waves can go around large rocks. FM and TV are best received when there is a line of sight between the broadcast antenna and receiver, and they are often sent from very tall structures. FM, TV, and mobile phone antennas themselves are much smaller than those used for AM, but they are elevated to achieve an unobstructed line of sight (Figure **4.2.6**).



Figure **4.2.6**: (a) A large tower is used to broadcast TV signals. The actual antennas are small structures on top of the tower -- they are placed at great heights to have a clear line of sight over a large broadcast area. (credit: Ozizo, Wikimedia Commons) (b) The NTT Dokomo mobile phone tower at Tokorozawa City, Japan. (credit: tokoroten, Wikimedia Commons)

Radio Wave Interference

Astronomers and astrophysicists collect signals from outer space using electromagnetic waves. A common problem for astrophysicists is the “pollution” from electromagnetic radiation pervading our surroundings from communication systems in general. Even everyday gadgets like our car keys having the facility to lock car doors remotely and being able to turn TVs on and off using remotes involve radio-wave frequencies. In order to prevent interference between all these electromagnetic signals, strict regulations are drawn up for different organizations to utilize different radio frequency bands.

One reason why we are sometimes asked to switch off our mobile phones (operating in the range of 1.9 GHz) on airplanes and in hospitals is that important communications or medical equipment often uses similar radio frequencies and their operation can be affected by frequencies used in the communication devices.

For example, radio waves used in magnetic resonance imaging (MRI) have frequencies on the order of 100 MHz, although this varies significantly depending on the strength of the magnetic field used and the nuclear type being scanned. MRI is an important

medical imaging and research tool, producing highly detailed two- and three-dimensional images. Radio waves are broadcast, absorbed, and reemitted in a resonance process that is sensitive to the density of nuclei (usually protons or hydrogen nuclei).

The wavelength of 100-MHz radio waves is 3 m, yet using the sensitivity of the resonant frequency to the magnetic field strength, details smaller than a millimeter can be imaged. This is a good example of an exception to a rule of thumb (in this case, the rubric that details much smaller than the probe's wavelength cannot be detected). The intensity of the radio waves used in MRI presents little or no hazard to human health.

Microwaves

Microwaves are the highest-frequency electromagnetic waves that can be produced by currents in macroscopic circuits and devices. Microwave frequencies range from about $10^9 Hz$ to the highest practical LC resonance at nearly $10^{12} Hz$. Since they have high frequencies, their wavelengths are short compared with those of other radio waves -- hence the name "microwave."

Microwaves can also be produced by atoms and molecules. They are, for example, a component of electromagnetic radiation generated by **thermal agitation.** The thermal motion of atoms and molecules in any object at a temperature above absolute zero causes them to emit and absorb radiation.

Since it is possible to carry more information per unit time on high frequencies, microwaves are quite suitable for communications. Most satellite-transmitted information is carried on microwaves, as are land-based long-distance transmissions. A clear line of sight between transmitter and receiver is needed because of the short wavelengths involved.

Radar is a common application of microwaves that was first developed in World War II. By detecting and timing microwave echoes, radar systems can determine the distance to objects as diverse as clouds and aircraft. A Doppler shift in the radar echo can be used to determine the speed of a car or the intensity of a rainstorm. Sophisticated radar systems are used to map the Earth and other planets, with a resolution limited by wavelength. (Figure 4.2.7). The shorter the wavelength of any probe, the smaller the detail it is possible to observe.

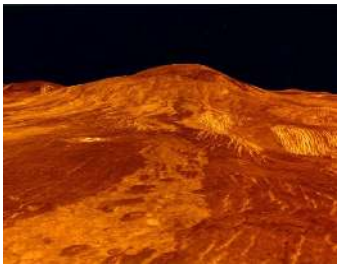


Figure 4.2.7: An image of Sif Mons with lava flows on Venus, based on Magellan synthetic aperture radar data combined with radar altimetry to produce a three-dimensional map of the surface. The Venusian atmosphere is opaque to visible light, but not to the microwaves that were used to create this image. (credit: NSSDC, NASA/JPL)

Heating with Microwaves

How does the ubiquitous microwave oven produce microwaves electronically, and why does food absorb them preferentially? Microwaves at a frequency of 2.45 GHz are produced by accelerating electrons. The microwaves are then used to induce an alternating electric field in the oven.

Water and some other constituents of food have a slightly negative charge at one end and a slightly positive charge at one end (called polar molecules). The range of microwave frequencies is specially selected so that the polar molecules, in trying to keep orienting themselves with the electric field, absorb these energies and increase their temperatures -- called dielectric heating.

The energy thereby absorbed results in thermal agitation heating food and not the plate, which does not contain water. Hot spots in the food are related to constructive and destructive interference patterns. Rotating antennas and food turntables help spread out the hot spots.

Another use of microwaves for heating is within the human body. Microwaves will penetrate more than shorter wavelengths into tissue and so can accomplish "deep heating" (called microwave diathermy). This is used for treating muscular pains, spasms, tendonitis, and rheumatoid arthritis.

MAKING CONNECTIONS: TAKE-HOME EXPERIMENT - MICROWAVE OVENS

1. Look at the door of a microwave oven. Describe the structure of the door. Why is there a metal grid on the door? How does the size of the holes in the grid compare with the wavelengths of microwaves used in microwave ovens? What is this wavelength?
2. Place a glass of water (about 250 ml) in the microwave and heat it for 30 seconds. Measure the temperature gain (the ΔT). Assuming that the power output of the oven is 1000 W, calculate the efficiency of the heat-transfer process.
3. Remove the rotating turntable or moving plate and place a cup of water in several places along a line parallel with the opening. Heat for 30 seconds and measure the ΔT for each position. Do you see cases of destructive interference?

Microwaves generated by atoms and molecules far away in time and space can be received and detected by electronic circuits. Deep space acts like a blackbody with a 2.7 K temperature, radiating most of its energy in the microwave frequency range. In 1964, Penzias and Wilson detected this radiation and eventually recognized that it was the radiation of the Big Bang's cooled remnants.

Infrared Radiation

The microwave and infrared regions of the electromagnetic spectrum overlap (Figure 4.2.1). **Infrared radiation** is generally produced by thermal motion and the vibration and rotation of atoms and molecules. Electronic transitions in atoms and molecules can also produce infrared radiation.

The range of infrared frequencies extends up to the lower limit of visible light, just below red. In fact, infrared means "below red." Frequencies at its upper limit are too high to be produced by accelerating electrons in circuits, but small systems, such as atoms and molecules, can vibrate fast enough to produce these waves.

Water molecules rotate and vibrate particularly well at infrared frequencies, emitting and absorbing them so efficiently that the emissivity for skin is $e = 0.97$ in the infrared. Night-vision scopes can detect the infrared emitted by various warm objects, including humans, and convert it to visible light.

We can examine radiant heat transfer from a house by using a camera capable of detecting infrared radiation. Reconnaissance satellites can detect buildings, vehicles, and even individual humans by their infrared emissions, whose power radiation is proportional to the fourth power of the absolute temperature. More mundanely, we use infrared lamps, some of which are called quartz heaters, to preferentially warm us because we absorb infrared better than our surroundings.

The Sun radiates like a nearly perfect blackbody (that is, it has $e = 1$), with a 6000 K surface temperature. About half of the solar energy arriving at the Earth is in the infrared region, with most of the rest in the visible part of the spectrum, and a relatively small amount in the ultraviolet. On average, 50 percent of the incident solar energy is absorbed by the Earth.

The relatively constant temperature of the Earth is a result of the energy balance between the incoming solar radiation and the energy radiated from the Earth. Most of the infrared radiation emitted from the Earth is absorbed by CO_2 and H_2O in the atmosphere and then radiated back to Earth or into outer space. This radiation back to Earth is known as the greenhouse effect, and it maintains the surface temperature of the Earth about $40°C$ higher than it would be if there is no absorption. Some scientists think that the increased concentration of CO_2 and other greenhouse gases in the atmosphere, resulting from increases in fossil fuel burning, has increased global average temperatures.

Visible Light

Visible light is the narrow segment of the electromagnetic spectrum to which the normal human eye responds. Visible light is produced by vibrations and rotations of atoms and molecules, as well as by electronic transitions within atoms and molecules. The receivers or detectors of light largely utilize electronic transitions. We say the atoms and molecules are excited when they absorb and relax when they emit through electronic transitions.

Figure 4.2.8 shows this part of the spectrum, together with the colors associated with particular pure wavelengths. We usually refer to visible light as having wavelengths of between 400 nm and 750 nm. (The retina of the eye actually responds to the lowest ultraviolet frequencies, but these do not normally reach the retina because they are absorbed by the cornea and lens of the eye.)

Red light has the lowest frequencies and longest wavelengths, while violet has the highest frequencies and shortest wavelengths. Blackbody radiation from the Sun peaks in the visible part of the spectrum but is more intense in the red than in the violet, making

the Sun yellowish in appearance.

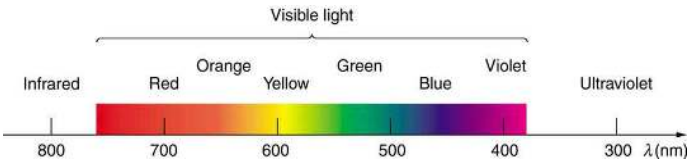


Figure 4.2.8: A small part of the electromagnetic spectrum that includes its visible components. The divisions between infrared, visible, and ultraviolet are not perfectly distinct, nor are those between the seven rainbow colors.

Living things--plants and animals-- have evolved to utilize and respond to parts of the electromagnetic spectrum they are embedded in. Visible light is the most predominant and we enjoy the beauty of nature through visible light. Plants are more selective. Photosynthesis makes use of parts of the visible spectrum to make sugars.

Example 4.2.2: Integrated Concept Problem: Correcting Vision with Lasers

During laser vision correction, a brief burst of 193-nm ultraviolet light is projected onto the cornea of a patient. It makes a spot 0.80 mm in diameter and evaporates a layer of cornea $0.30\mu m$ thick. Calculate the energy absorbed, assuming the corneal tissue has the same properties as water; it is initially at $34°C$. Assume the evaporated tissue leaves at a temperature of $100°C$.

Strategy

The energy from the laser light goes toward raising the temperature of the tissue and also toward evaporating it. Thus we have two amounts of heat to add together. Also, we need to find the mass of corneal tissue involved.

Solution

To figure out the heat required to raise the temperature of the tissue to $100°C$, we can apply concepts of thermal energy. We know that

$$Q = mc\Delta T, \tag{4.2.9}$$

where Q is the heat required to raise the temperature, ΔT is the desired change in temperature, m is the mass of tissue to be heated, and c is the specific heat of water equal to 4186 J/kg/K.

Without knowing the mass m at this point, we have

$$\begin{aligned} Q &= m\,(4186 J/kg/K)\,(100°C - 34°C) \\ &= m\,(276,276 J/kg) \\ &= m\,(276 J/kg)\,. \end{aligned}$$

The latent heat of vaporization of water is 2256 kJ/kg, so that the energy needed to evaporate mass m is

$$Q_v = mL_v = m\,(2256 kJ/kg)\,.$$

To find the mass m, we use the equation $\rho = m/V$, where ρ is the density of the tissue and V is its volume. For this case,

$$\begin{aligned} m &= \rho V \\ &= (1000 kg/m^3)\,(area \times thickness\,(m^3)) \\ &= (1000 kg/m^3)\left(\pi\,(0.80\times10^{-3} m)^2/4\right)(0.30 \times 10^{-6} m) \\ &= 0.151 \times 10^{-9} kg. \end{aligned}$$

Therefore, the total energy absorbed by the tissue in the eye is the sum of Q and Q_v:

$$\begin{aligned} Q_{tot} &= m\,(c\Delta T + L_v) \\ &= (0.151 \times 10^{-9} kg)\,(276 kJ/kg + 2256 kJ/kg) \\ &= 382 \times 10^{-9} kJ. \end{aligned}$$

Discussion

The lasers used for this eye surgery are excimer lasers, whose light is well absorbed by biological tissue. They evaporate rather than burn the tissue, and can be used for precision work. Most lasers used for this type of eye surgery have an average power rating of about one watt. For our example, if we assume that each laser burst from this pulsed laser lasts for 10 ns, and there are 400 bursts per second, then the average power is

$$Q_{tot} \times 400 = 150\,mW$$

Optics is the study of the behavior of visible light and other forms of electromagnetic waves. Optics falls into two distinct categories. When electromagnetic radiation, such as visible light, interacts with objects that are large compared with its wavelength, its motion can be represented by straight lines like rays. Ray optics is the study of such situations and includes lenses and mirrors.

When electromagnetic radiation interacts with objects about the same size as the wavelength or smaller, its wave nature becomes apparent. For example, observable detail is limited by the wavelength, and so visible light can never detect individual atoms, because they are so much smaller than its wavelength. Physical or wave optics is the study of such situations and includes all wave characteristics.

TAKE-HOME EXPERIMENT: COLORS THAT MATCH

When you light a match you see largely orange light; when you light a gas stove you see blue light. Why are the colors different? What other colors are present in these?

Ultraviolet Radiation

Ultraviolet means “above violet.” The electromagnetic frequencies of **ultraviolet radiation (UV)** extend upward from violet, the highest-frequency visible light. Ultraviolet is also produced by atomic and molecular motions and electronic transitions. The wavelengths of ultraviolet extend from 400 nm down to about 10 nm at its highest frequencies, which overlap with the lowest X-ray frequencies. It was recognized as early as 1801 by Johann Ritter that the solar spectrum had an invisible component beyond the violet range.

Solar UV radiation is broadly subdivided into three regions: UV-A (320–400 nm), UV-B (290–320 nm), and UV-C (220–290 nm), ranked from long to shorter wavelengths (from smaller to larger energies). Most UV-B and all UV-C is absorbed by ozone (O_3) molecules in the upper atmosphere. Consequently, 99% of the solar UV radiation reaching the Earth’s surface is UV-A.

Human Exposure to UV Radiation

It is largely exposure to UV-B that causes skin cancer. It is estimated that as many as 20% of adults will develop skin cancer over the course of their lifetime. Again, treatment is often successful if caught early. Despite very little UV-B reaching the Earth’s surface, there are substantial increases in skin-cancer rates in countries such as Australia, indicating how important it is that UV-B and UV-C continue to be absorbed by the upper atmosphere.

All UV radiation can damage collagen fibers, resulting in an acceleration of the aging process of skin and the formation of wrinkles. Because there is so little UV-B and UV-C reaching the Earth’s surface, sunburn is caused by large exposures, and skin cancer from repeated exposure. Some studies indicate a link between overexposure to the Sun when young and melanoma later in life.

The tanning response is a defense mechanism in which the body produces pigments to absorb future exposures in inert skin layers above living cells. Basically UV-B radiation excites DNA molecules, distorting the DNA helix, leading to mutations and the possible formation of cancerous cells.

Repeated exposure to UV-B may also lead to the formation of cataracts in the eyes -- a cause of blindness among people living in the equatorial belt where medical treatment is limited. Cataracts, clouding in the eye’s lens and a loss of vision, are age related; 60% of those between the ages of 65 and 74 will develop cataracts. However, treatment is easy and successful, as one replaces the lens of the eye with a plastic lens. Prevention is important. Eye protection from UV is more effective with plastic sunglasses than those made of glass.

A major acute effect of extreme UV exposure is the suppression of the immune system, both locally and throughout the body.

Low-intensity ultraviolet is used to sterilize haircutting implements, implying that the energy associated with ultraviolet is deposited in a manner different from lower-frequency electromagnetic waves. (Actually this is true for all electromagnetic waves with frequencies greater than visible light.)

Flash photography is generally not allowed of precious artworks and colored prints because the UV radiation from the flash can cause photo-degradation in the artworks. Often artworks will have an extra-thick layer of glass in front of them, which is especially designed to absorb UV radiation.

UV Light and the Ozone Layer

If all of the Sun's ultraviolet radiation reached the Earth's surface, there would be extremely grave effects on the biosphere from the severe cell damage it causes. However, the layer of ozone (O_3) in our upper atmosphere (10 to 50 km above the Earth) protects life by absorbing most of the dangerous UV radiation.

Unfortunately, today we are observing a depletion in ozone concentrations in the upper atmosphere. This depletion has led to the formation of an "ozone hole" in the upper atmosphere. The hole is more centered over the southern hemisphere, and changes with the seasons, being largest in the spring. This depletion is attributed to the breakdown of ozone molecules by refrigerant gases called chlorofluorocarbons (CFCs).

The UV radiation helps dissociate the CFC's, releasing highly reactive chlorine (Cl) atoms, which catalyze the destruction of the ozone layer. For example, the reaction of $CFCl_3$ with a photon of light ($h\nu$) can be written as:

$$\mathrm{CFCl_3 + h\nu \rightarrow CFCl_2 + Cl.} \tag{4.2.10}$$

The Cl atom then catalyzes the breakdown of ozone as follows:

$$\mathrm{Cl + O_3 \rightarrow ClO + O_2} \tag{4.2.11}$$

and

$$\mathrm{ClO + O_3 \rightarrow Cl + 2\,O_2.} \tag{4.2.12}$$

A single chlorine atom could destroy ozone molecules for up to two years before being transported down to the surface. The CFCs are relatively stable and will contribute to ozone depletion for years to come. CFCs are found in refrigerants, air conditioning systems, foams, and aerosols.

International concern over this problem led to the establishment of the "Montreal Protocol" agreement (1987) to phase out CFC production in most countries. However, developing-country participation is needed if worldwide production and elimination of CFCs is to be achieved. Probably the largest contributor to CFC emissions today is India. But the protocol seems to be working, as there are signs of an ozone recovery (Figure 4.2.9).

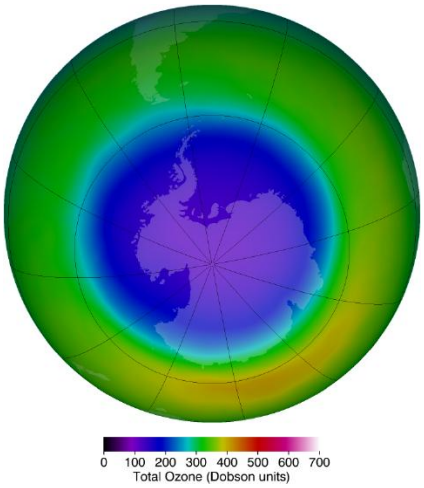


Figure 4.2.9: This map of ozone concentration over Antarctica in October 2011 shows severe depletion suspected to be caused by CFCs. Less dramatic but more general depletion has been observed over northern latitudes, suggesting the effect is global. With less ozone, more ultraviolet radiation from the Sun reaches the surface, causing more damage. (credit: NASA Ozone Watch)

Benefits of UV Light

Besides the adverse effects of ultraviolet radiation, there are also benefits of exposure in nature and uses in technology. Vitamin D production in the skin (epidermis) results from exposure to UVB radiation, generally from sunlight. A number of studies indicate lack of vitamin D can result in the development of a range of cancers (prostate, breast, colon), so a certain amount of UV exposure is helpful. Lack of vitamin D is also linked to osteoporosis. Exposures (with no sunscreen) of 10 minutes a day to arms, face, and legs might be sufficient to provide the accepted dietary level. However, in the winter time north of about **37°** latitude, most UVB gets blocked by the atmosphere.

UV radiation is used in the treatment of infantile jaundice and in some skin conditions. It is also used in sterilizing workspaces and tools, and killing germs in a wide range of applications. It is also used as an analytical tool to identify substances.

When exposed to ultraviolet, some substances, such as minerals, glow in characteristic visible wavelengths, a process called fluorescence. So-called black lights emit ultraviolet to cause posters and clothing to fluoresce in the visible. Ultraviolet is also used in special microscopes to detect details smaller than those observable with longer-wavelength visible-light microscopes.

THINGS GREAT AND SMALL: A SUBMICROSCOPIC VIEW OF X-RAY PRODUCTION

X-rays can be created in a high-voltage discharge. They are emitted in the material struck by electrons in the discharge current. There are two mechanisms by which the electrons create X-rays.

The first method is illustrated in the Figure **4.2.10**. An electron is accelerated in an evacuated tube by a high positive voltage. The electron strikes a metal plate (e.g., copper) and produces X-rays. Since this is a high-voltage discharge, the electron gains sufficient energy to ionize the atom.

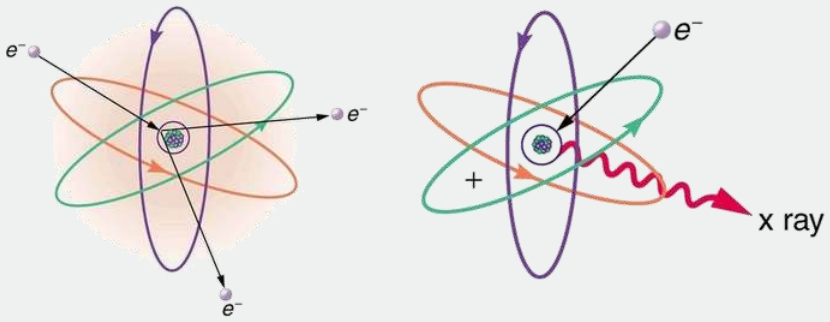


Figure **4.2.10**: Artist's conception of an electron ionizing an atom followed by the recapture of an electron and emission of an X-ray. An energetic electron strikes an atom and knocks an electron out of one of the orbits closest to the nucleus. Later, the atom captures another electron, and the energy released by its fall into a low orbit generates a high-energy EM wave called an X-ray.

In the case shown, an inner-shell electron (one in an orbit relatively close to and tightly bound to the nucleus) is ejected. A short time later, another electron is captured and falls into the orbit in a single great plunge. The energy released by this fall is given to an EM wave known as an X-ray. Since the orbits of the atom are unique to the type of atom, the energy of the X-ray is characteristic of the atom, hence the name characteristic X-ray.

The second method by which an energetic electron creates an X-ray when it strikes a material is illustrated in the figure below. The electron interacts with charges in the material as it penetrates. These collisions transfer kinetic energy from the electron to the electrons and atoms in the material.

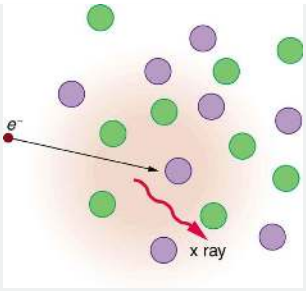


Figure 4.2.11: Artist's conception of an electron being slowed by collisions in a material and emitting X-ray radiation. This energetic electron makes numerous collisions with electrons and atoms in a material it penetrates. An accelerated charge radiates EM waves, a second method by which X-rays are created.

A loss of kinetic energy implies an acceleration, in this case decreasing the electron's velocity. Whenever a charge is accelerated, it radiates EM waves. Given the high energy of the electron, these EM waves can have high energy. We call them X-rays. Since the process is random, a broad spectrum of X-ray energy is emitted that is more characteristic of the electron energy than the type of material the electron encounters. Such EM radiation is called "bremsstrahlung" (German for "braking radiation").

X-Rays

In the 1850s, scientists (such as Faraday) began experimenting with high-voltage electrical discharges in tubes filled with rarefied gases. It was later found that these discharges created an invisible, penetrating form of very high frequency electromagnetic radiation. This radiation was called an **X-ray**, because its identity and nature were unknown.

As described in "Things Great and Small," there are two methods by which X-rays are created --both are submicroscopic processes and can be caused by high-voltage discharges. While the low-frequency end of the X-ray range overlaps with the ultraviolet, X-rays extend to much higher frequencies (and energies).

X-rays have adverse effects on living cells similar to those of ultraviolet radiation, and they have the additional liability of being more penetrating, affecting more than the surface layers of cells. Cancer and genetic defects can be induced by exposure to X-rays. Because of their effect on rapidly dividing cells, X-rays can also be used to treat and even cure cancer.

The widest use of X-rays is for imaging objects that are opaque to visible light, such as the human body or aircraft parts. In humans, the risk of cell damage is weighed carefully against the benefit of the diagnostic information obtained. However, questions have risen in recent years as to accidental overexposure of some people during CT scans --a mistake at least in part due to poor monitoring of radiation dose.

The ability of X-rays to penetrate matter depends on density, and so an X-ray image can reveal very detailed density information. Figure 4.2.12 shows an example of the simplest type of X-ray image, an X-ray shadow on film. The amount of information in a simple X-ray image is impressive, but more sophisticated techniques, such as CT scans, can reveal three-dimensional information with details smaller than a millimeter.


Figure 4.2.12: This shadow X-ray image shows many interesting features, such as artificial heart valves, a pacemaker, and the wires used to close the sternum. (credit: P. P. Urone)

The use of X-ray technology in medicine is called radiology -- an established and relatively cheap tool in comparison to more sophisticated technologies. Consequently, X-rays are widely available and used extensively in medical diagnostics. During World War I, mobile X-ray units, advocated by Madame Marie Curie, were used to diagnose soldiers.

Because they can have wavelengths less than 0.01 nm, X-rays can be scattered (a process called X-ray diffraction) to detect the shape of molecules and the structure of crystals. X-ray diffraction was crucial to Crick, Watson, and Wilkins in the determination of the shape of the double-helix DNA molecule.

X-rays are also used as a precise tool for trace-metal analysis in X-ray induced fluorescence, in which the energy of the X-ray emissions are related to the specific types of elements and amounts of materials present.

Gamma Rays

Soon after nuclear radioactivity was first detected in 1896, it was found that at least three distinct types of radiation were being emitted. The most penetrating nuclear radiation was called **a gamma ray** (γ ray) (again a name given because its identity and character were unknown), and it was later found to be an extremely high frequency electromagnetic wave.

In fact, γ rays are any electromagnetic radiation emitted by a nucleus. This can be from natural nuclear decay or induced nuclear processes in nuclear reactors and weapons. The lower end of the **γ-ray** frequency range overlaps the upper end of the X-ray range, but γ rays can have the highest frequency of any electromagnetic radiation.

Gamma rays have characteristics identical to X-rays of the same frequency -- they differ only in source. At higher frequencies, γ rays are more penetrating and more damaging to living tissue. They have many of the same uses as X-rays, including cancer therapy. Gamma radiation from radioactive materials is used in nuclear medicine.

Figure 13 shows a medical image based on γ rays. Food spoilage can be greatly inhibited by exposing it to large doses of γ radiation, thereby obliterating responsible microorganisms. Damage to food cells through irradiation occurs as well, and the long-term hazards of consuming radiation-preserved food are unknown and controversial for some groups. Both X-ray and **γ-ray** technologies are also used in scanning luggage at airports.

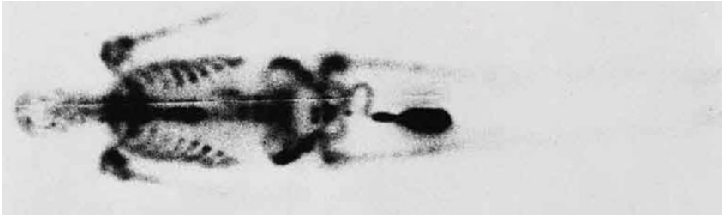


Figure **4.2.13**: This is an image of the γ rays emitted by nuclei in a compound that is concentrated in the bones and eliminated through the kidneys. Bone cancer is evidenced by nonuniform concentration in similar structures. For example, some ribs are darker than others. (credit: P. P. Urone)

Detecting Electromagnetic Waves from Space

A final note on star gazing. The entire electromagnetic spectrum is used by researchers for investigating stars, space, and time. As noted earlier, Penzias and Wilson detected microwaves to identify the background radiation originating from the Big Bang. Radio telescopes such as the Arecibo Radio Telescope in Puerto Rico and Parkes Observatory in Australia were designed to detect radio waves.

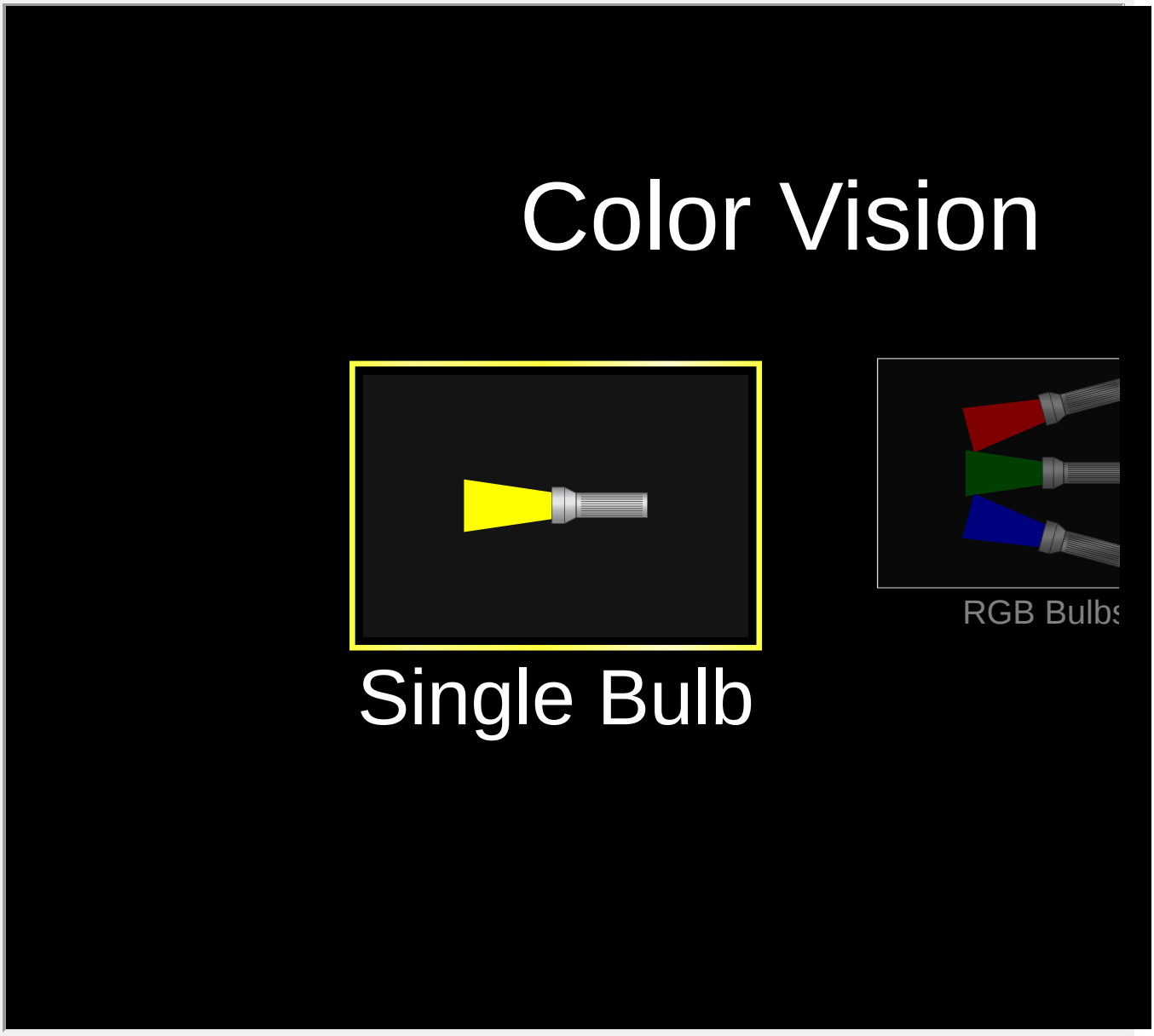
Infrared telescopes need to have their detectors cooled by liquid nitrogen to be able to gather useful signals. Since infrared radiation is predominantly from thermal agitation, if the detectors were not cooled, the vibrations of the molecules in the antenna would be stronger than the signal being collected.

The most famous of these infrared sensitive telescopes is the James Clerk Maxwell Telescope in Hawaii. The earliest telescopes, developed in the seventeenth century, were optical telescopes, collecting visible light. Telescopes in the ultraviolet, X-ray, and γ-ray regions are placed outside the atmosphere on satellites orbiting the Earth.

The Hubble Space Telescope (launched in 1990) gathers ultraviolet radiation as well as visible light. In the X-ray region, there is the Chandra X-ray Observatory (launched in 1999), and in the γ-ray region, there is the new Fermi Gamma-ray Space Telescope (launched in 2008—taking the place of the Compton Gamma Ray Observatory, 1991–2000.).

PHET EXPLORATIONS: COLOR VISION

Make a whole rainbow by mixing red, green, and blue light. Change the wavelength of a monochromatic beam or filter white light. View the light as a solid beam, or see the individual photons.



Summary

- The relationship among the speed of propagation, wavelength, and frequency for any wave is given by $v_w = f\lambda$, so that for electromagnetic waves,

$$c = f\lambda,$$

  where f is the frequency, λ is the wavelength, and c is the speed of light.
- The electromagnetic spectrum is separated into many categories and subcategories, based on the frequency and wavelength, source, and uses of the electromagnetic waves.
- Any electromagnetic wave produced by currents in wires is classified as a radio wave, the lowest frequency electromagnetic waves. Radio waves are divided into many types, depending on their applications, ranging up to microwaves at their highest frequencies.

- Infrared radiation lies below visible light in frequency and is produced by thermal motion and the vibration and rotation of atoms and molecules. Infrared's lower frequencies overlap with the highest-frequency microwaves.
- Visible light is largely produced by electronic transitions in atoms and molecules, and is defined as being detectable by the human eye. Its colors vary with frequency, from red at the lowest to violet at the highest.
- Ultraviolet radiation starts with frequencies just above violet in the visible range and is produced primarily by electronic transitions in atoms and molecules.
- X-rays are created in high-voltage discharges and by electron bombardment of metal targets. Their lowest frequencies overlap the ultraviolet range but extend to much higher values, overlapping at the high end with gamma rays.
- Gamma rays are nuclear in origin and are defined to include the highest-frequency electromagnetic radiation of any type.

Glossary

electromagnetic spectrum
the full range of wavelengths or frequencies of electromagnetic radiation

radio waves
electromagnetic waves with wavelengths in the range from 1 mm to 100 km; they are produced by currents in wires and circuits and by astronomical phenomena

microwaves
electromagnetic waves with wavelengths in the range from 1 mm to 1 m; they can be produced by currents in macroscopic circuits and devices

thermal agitation
the thermal motion of atoms and molecules in any object at a temperature above absolute zero, which causes them to emit and absorb radiation

radar
a common application of microwaves. Radar can determine the distance to objects as diverse as clouds and aircraft, as well as determine the speed of a car or the intensity of a rainstorm

infrared radiation (IR)
a region of the electromagnetic spectrum with a frequency range that extends from just below the red region of the visible light spectrum up to the microwave region, or from $0.74\mu m$ to $300\mu m$

ultraviolet radiation (UV)
electromagnetic radiation in the range extending upward in frequency from violet light and overlapping with the lowest X-ray frequencies, with wavelengths from 400 nm down to about 10 nm

visible light
the narrow segment of the electromagnetic spectrum to which the normal human eye responds

amplitude modulation (AM)
a method for placing information on electromagnetic waves by modulating the amplitude of a carrier wave with an audio signal, resulting in a wave with constant frequency but varying amplitude

extremely low frequency (ELF)
electromagnetic radiation with wavelengths usually in the range of 0 to 300 Hz, but also about 1kHz

carrier wave
an electromagnetic wave that carries a signal by modulation of its amplitude or frequency

frequency modulation (FM)

a method of placing information on electromagnetic waves by modulating the frequency of a carrier wave with an audio signal, producing a wave of constant amplitude but varying frequency

TV

video and audio signals broadcast on electromagnetic waves

very high frequency (VHF)

TV channels utilizing frequencies in the two ranges of 54 to 88 MHz and 174 to 222 MHz

ultra-high frequency (UHF)

TV channels in an even higher frequency range than VHF, of 470 to 1000 MHz

X-ray

invisible, penetrating form of very high frequency electromagnetic radiation, overlapping both the ultraviolet range and the γ-ray range

gamma ray

(γ ray); extremely high frequency electromagnetic radiation emitted by the nucleus of an atom, either from natural nuclear decay or induced nuclear processes in nuclear reactors and weapons. The lower end of the γ-ray frequency range overlaps the upper end of the X-ray range, but γ rays can have the highest frequency of any electromagnetic radiation

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5.1: Ions - Losing and Gaining Electrons

Learning Objectives

- Define the two types of ions.

Atoms tend to lose, gain, or share some valance electrons, making bonds to acquire the electron configuration of the nearest noble gas, i.e., usually eight valence electrons for the main group elements, called an octet. Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while others lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called **cations**. Most metals become cations when they make ionic compounds.

Cations

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

$$Na \rightarrow Na^+ + e^-$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure 5.1.1 is a graphical depiction of this process.

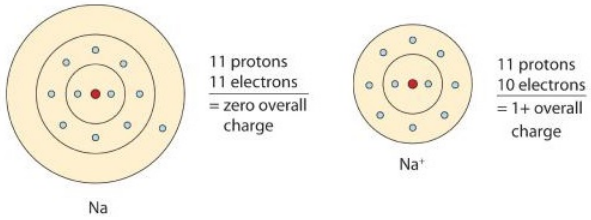


Figure 5.1.1: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

Anions

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

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

$$e^- + Cl \longrightarrow Cl^-$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure 5.1.2 is a graphical depiction of this process.

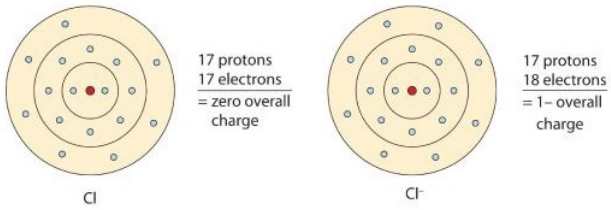


Figure **5.1.2**: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge. Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a -1 overall charge.

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

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1− charge. Figure **5.1.3** shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a ion with multiple charges. The barium cation is written Ba^{2+}, not Ba^{+2}.

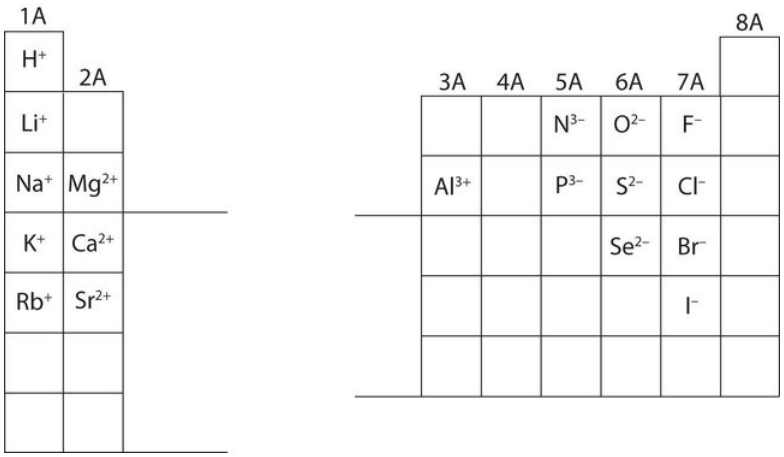


Figure **5.1.3**: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

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5.2: Sugar and Salt

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

$$\mathrm{2\,Na(s) + Cl_2(g) \rightarrow 2\,NaCl(s)} \tag{5.2.1}$$

The element sodium (Figure 5.2.1*a*) is a very reactive metal; given the opportunity, it will react with the sweat on your hands and form sodium hydroxide, which is a very corrosive substance. The element chlorine (Figure 5.2.1*b*) is a pale yellow, corrosive gas that should not be inhaled due to its poisonous nature. Bring these two hazardous substances together, however, and they react to make the ionic compound sodium chloride (Figure 5.2.1*c*), known simply as salt.

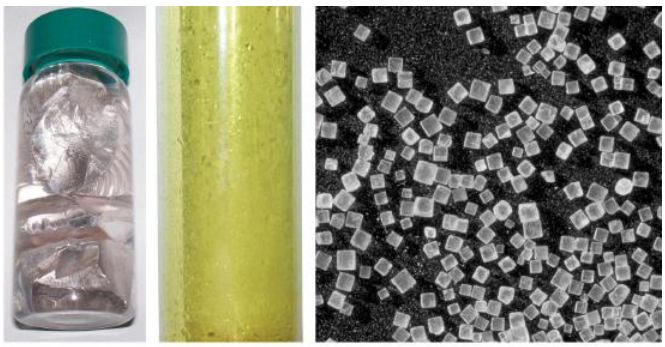


Figure 5.2.1: Sodium + Chlorine = Sodium Chloride (a) Sodium is a very reactive metal. (b) Chlorine is a pale yellow, noxious gas. (c) Together, sodium and chlorine make sodium chloride—salt—which is necessary for our survival. Source: Photo on the left courtesy of reenhorn1 and photo in the center courtesy of Benjah-bmm27. Photo on the right © Thinkstock.

Salt is necessary for life. $\mathrm{Na^+}$ ions are one of the main ions in the human body and are necessary to regulate the fluid balance in the body. $\mathrm{Cl^-}$ ions are necessary for proper nerve function and respiration. Both of these ions are supplied by salt. The taste of salt is one of the fundamental tastes; salt is probably the most ancient flavoring known, and one of the few rocks we eat. Clearly when the elemental sodium and chlorine combine (Equation 5.2.1), the resulting salt product has radically different properties (both physical and chemical). This reaction is spectacular to observe (Video 5.2.1).



Video* 5.2.1*: *Making Table Salt using Sodium Metal and Chlorine gas*

Another compound is sugar, which is the generic name for sweet, soluble carbohydrates, many of which are used in food. Sugar has the chemical formulate $\mathrm{C_{12}H_{22}O_{11}}$ and is constructed from different elements than salt: carbon, hydrogen and oxygen. While sugar qualitatively resembles table salt (often confused in the kitchen), the two have distinctly different physical and chemical properties. There are various types of sugar derived from different sources. While sugar is made with carbon, hydrogen, and oxygen, it is considerably harder to synthesize from its constituent elements than table salt is (Equation 5.2.1). However, the thermal

decomposition is considerably easier and can be represented as a dehydration of sucrose to pure carbon and water vapor in Equation 5.2.2, and demonstrated in Video 5.2.2.

$$\mathrm{C_{12}H_{22}O_{11}\,(s) + heat \longrightarrow 12\,C\,(s) + 11\,H_2O\,(g)} \tag{5.2.2}$$



***Video* 5.2.2:** *A science experiment in the kitchen shows what happens to sugar molecules when they are heated. The experiment did not disappoint!*

As with salt, sugar has radically different properties (both physical and chemical) than its constituent elements. This difference in properties, of constituent elements and compounds, is a central feature of chemical reactions.

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5.3: Writing Formulas for Ionic Compounds

Learning Objectives

- Write the correct formula for an ionic compound.
- Recognize polyatomic ions in chemical formulas.

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattice containing many ions each of the cation and anion. An ionic formula, like $NaCl$, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

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

Example 5.3.1: Aluminum Nitride and Lithium Oxide

Write the formulas for aluminum nitride and lithium oxide.

Solution

Solution to Example 5.5.1

	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Al^{3+} N^{3-}	Li^{+} O^{2-}
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

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

Example 5.3.2: The Crisscross Method for Lead (IV) Oxide

Write the formula for lead (IV) oxide.

Solution

Solution to Example 5.5.2

Crisscross Method	Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Pb^{4+} O^{2-}
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all subscripts that are 1.	PbO_2

? Exercise 5.3.2

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

> *Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.*

✓ Example 5.3.3: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Solution to Example 5.5.3

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Na^{+} S^{2-}
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all subscripts that are 1.	Na_2S

? Exercise 5.3.3

Write the formula for each ionic compound.

a. sodium bromide

b. lithium chloride
c. magnesium oxide

Answer a:
NaBr
Answer b:
LiCl
Answer c:
MgO

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Table **5.3.1** lists the most common polyatomic ions.

Table **5.3.1**: Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 5.3.4: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Solution to Example 5.5.4

Crisscross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Ca^{2+} NO_3^-
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 5.3.5

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion.

Solution

Solution to Example 5.5.5

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 5.3.5

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a:

$MgCO_3$

Answer b:

$Al(CH_3COO)_3$

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds.

Method 1

Compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2 [or 2A] metal) and a nonmetallic element (bromine, a group 17 [or 7A] nonmetal). Therefore, it is most likely an ionic compound (in fact, it *is* ionic). In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 15 [or 5A], and oxygen, from group 16 [or 6A]. It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Method 2

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, NO_3^-. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2− charge from the two nitrate ions. Thus, this compound is also ionic.

Example 5.3.6

Identify each compound as ionic or not ionic.

a. Na_2O
b. PCl_3
c. NH_4Cl
d. OF_2

Solution

Solution to Example 5.5.6

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, Na_2O is expected to be ionic via method 1.	Na_2O, ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic via method 1	PCl_3, not ionic
c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic via method 2	NH_4Cl, ionic
d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic via method 1	OF_2, not ionic

Exercise 5.3.6

Identify each compound as ionic or not ionic.

a. N_2O
b. $FeCl_3$
c. $(NH_4)_3PO_4$
d. $SOCl_2$

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Summary

Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

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5.4: Naming Inorganic Compounds

Learning Objectives

- To describe the composition of a chemical compound.
- To name covalent compounds that contain up to three elements.

As with ionic compounds, the system for naming covalent compounds enables chemists to write the molecular formula from the name and vice versa. This and the following section describe the rules for naming simple covalent compounds, beginning with inorganic compounds and then turning to simple organic compounds that contain only carbon and hydrogen.

When chemists synthesize a new compound, they may not yet know its molecular or structural formula. In such cases, they usually begin by determining its empirical formula, the relative numbers of atoms of the elements in a compound, reduced to the smallest whole numbers. Because the empirical formula is based on experimental measurements of the numbers of atoms in a sample of the compound, it shows only the ratios of the numbers of the elements present. The difference between empirical and molecular formulas can be illustrated with butane, a covalent compound used as the fuel in disposable lighters. The molecular formula for butane is $\mathbf{C_4H_{10}}$. The ratio of carbon atoms to hydrogen atoms in butane is 4:10, which can be reduced to 2:5. The empirical formula for butane is therefore $\mathbf{C_2H_5}$. The formula unit is the absolute grouping of atoms or ions represented by the empirical formula of a compound, either ionic or covalent. Butane has the empirical formula $\mathbf{C_2H_5}$, but it contains two $\mathbf{C_2H_5}$ formula units, giving a molecular formula of $\mathbf{C_4H_{10}}$.

Because ionic compounds **do not** contain discrete molecules, empirical formulas are used to indicate their compositions. All compounds, whether ionic or covalent, must be electrically neutral. Consequently, the positive and negative charges in a formula unit must exactly cancel each other. If the cation and the anion have charges of equal magnitude, such as $\mathbf{Na^+}$ and $\mathbf{Cl^-}$, then the compound must have a 1:1 ratio of cations to anions, and the empirical formula must be $\mathbf{NaCl}$. If the charges are not the same magnitude, then a cation:anion ratio other than 1:1 is needed to produce a neutral compound. In the case of $\mathbf{Mg^{2+}}$ and $\mathbf{Cl^-}$, for example, two Cl^- ions are needed to balance the two positive charges on each Mg^{2+} ion, giving an empirical formula of $\mathbf{MgCl_2}$. Similarly, the formula for the ionic compound that contains Na^+ and O^{2-} ions is Na_2O.

Ionic compounds do not contain discrete molecules, so empirical formulas are used to indicate their compositions.

Binary Ionic Compounds

An ionic compound that contains only two elements, one present as a cation and one as an anion, is called a binary ionic compound. One example is $\mathbf{MgCl_2}$, a coagulant used in the preparation of tofu from soybeans. For binary ionic compounds, the subscripts in the empirical formula can also be obtained by crossing charges: use the absolute value of the charge on one ion as the subscript for the other ion. This method is shown schematically as follows:



Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges.

When crossing charges, it is sometimes necessary to reduce the subscripts to their simplest ratio to write the empirical formula. Consider, for example, the compound formed by Mg^{2+} and O^{2-}. Using the absolute values of the charges on the ions as subscripts gives the formula $\mathbf{Mg_2O_2}$:



This simplifies to its correct empirical formula MgO. The empirical formula has one Mg^{2+} ion and one O^{2-} ion.

Example 5.4.1: Binary Ionic Compounds

Write the empirical formula for the simplest binary ionic compound formed from each ion or element pair.

a. Ga^{3+} and As^{3-}
b. Eu^{3+} and O^{2-}
c. calcium and chlorine

Given: ions or elements

Asked for: empirical formula for binary ionic compound

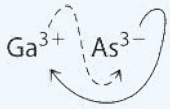
Strategy:

A. If not given, determine the ionic charges based on the location of the elements in the periodic table.
B. Use the absolute value of the charge on each ion as the subscript for the other ion. Reduce the subscripts to the lowest numbers

to write the empirical formula. Check to make sure the empirical formula is electrically neutral.

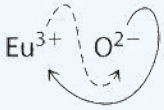
Solution

a. **B** Using the absolute values of the charges on the ions as the subscripts gives $\mathbf{Ga_3As_3}$:



Reducing the subscripts to the smallest whole numbers gives the empirical formula GaAs, which is electrically neutral [+3 + (−3) = 0]. Alternatively, we could recognize that Ga^{3+} and As^{3-} have charges of equal magnitude but opposite signs. One Ga^{3+} ion balances the charge on one As^{3-} ion, and a 1:1 compound will have no net charge. Because we write subscripts only if the number is greater than 1, the empirical formula is GaAs. GaAs is gallium arsenide, which is widely used in the electronics industry in transistors and other devices.

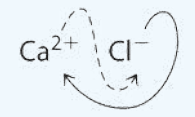
b. **B** Because Eu^{3+} has a charge of +3 and O^{2-} has a charge of −2, a 1:1 compound would have a net charge of +1. We must therefore find multiples of the charges that cancel. We cross charges, using the absolute value of the charge on one ion as the subscript for the other ion:



The subscript for Eu^{3+} is 2 (from O^{2-}), and the subscript for O^{2-} is 3 (from Eu^{3+}), giving Eu_2O_3; the subscripts cannot be reduced further. The empirical formula contains a positive charge of 2(+3) = +6 and a negative charge of 3(−2) = −6, for a net charge of 0. The compound Eu_2O_3 is neutral. Europium oxide is responsible for the red color in television and computer screens.

c. **A** Because the charges on the ions are not given, we must first determine the charges expected for the most common ions derived from calcium and chlorine. Calcium lies in group 2, so it should lose two electrons to form Ca^{2+}. Chlorine lies in group 17, so it should gain one electron to form Cl^-.

B Two Cl^- ions are needed to balance the charge on one Ca^{2+} ion, which leads to the empirical formula $CaCl_2$. We could also cross charges, using the absolute value of the charge on Ca^{2+} as the subscript for Cl and the absolute value of the charge on Cl^- as the subscript for Ca:



The subscripts in $CaCl_2$ cannot be reduced further. The empirical formula is electrically neutral [+2 + 2(−1) = 0]. This compound is calcium chloride, one of the substances used as "salt" to melt ice on roads and sidewalks in winter.

? Exercise 5.4.1

Write the empirical formula for the simplest binary ionic compound formed from each ion or element pair.

a. Li^+ and N^{3-}
b. Al^{3+} and O^{2-}
c. lithium and oxygen

Answer a

Li_3N

Answer b

Al_2O_3

Answer c

Li_2O



Nomenclature of Metals: Nomenclature of Metals(opens in new window) [youtu.be]

Polyatomic Ions

Polyatomic ions are groups of atoms that bear net electrical charges, although the atoms in a polyatomic ion are held together by the same covalent bonds that hold atoms together in molecules. Just as there are many more kinds of molecules than simple elements, there are many more kinds of polyatomic ions than monatomic ions. Two examples of polyatomic cations are the ammonium (NH_4^+) and the methylammonium ($CH_3NH_3^+$) ions. Polyatomic anions are much more numerous than polyatomic cations; some common examples are in Table **5.4.1**.

Table **5.4.1**: Common Polyatomic Ions and Their Names

Formula	Name of Ion	Formula	Name of Ion
NH_4^+	ammonium	HPO_4^{2-}	hydrogen phosphate
$CH_3NH_3^+$	methylammonium	$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide	ClO^-	hypochlorite
O_2^{2-}	peroxide	ClO_2^-	chlorite
CN^-	cyanide	ClO_3^-	chlorate
SCN^-	thiocyanate	ClO_4^-	perchlorate
NO_2^-	nitrite	MnO_4^-	permanganate
NO_3^-	nitrate	$CrO4_2^-$	chromate
CO_3^{2-}	carbonate	$Cr_2O_7^{2-}$	dichromate
HCO_3^-	hydrogen carbonate, or bicarbonate	$C_2O_4^{2-}$	oxalate
SO_3^{2-}	sulfite	HCO_2^-	formate
SO_4^{2-}	sulfate	$CH_3CO_2^-$	acetate
HSO_4^-	hydrogen sulfate, or bisulfate	$C_6H_5CO_2^-$	benzoate
PO_4^{3-}	phosphate		

The method used to predict the empirical formulas for ionic compounds that contain monatomic ions can also be used for compounds that contain polyatomic ions. The overall charge on the cations must balance the overall charge on the anions in the formula unit. Thus, K^+ and NO_3^- ions combine in a 1:1 ratio to form KNO_3 (potassium nitrate or saltpeter), a major ingredient in black gunpowder. Similarly, Ca^{2+} and SO_4^{2-} form $CaSO_4$ (calcium sulfate), which combines with varying amounts of water to form gypsum and plaster of Paris. The polyatomic ions NH_4^+ and NO_3^- form NH_4NO_3 (ammonium nitrate), a widely used fertilizer and, in the wrong hands, an explosive. One example of a compound in which the ions have charges of different magnitudes is calcium phosphate, which is composed of Ca^{2+} and PO_4^{3-} ions; it is a major component of bones. The compound is electrically neutral because the ions combine in a ratio of three Ca^{2+} ions [3(+2) = +6] for every two ions [2(−3) = −6], giving an empirical formula of $Ca_3(PO_4)_2$; the parentheses around PO_4 in the empirical formula indicate that it is a polyatomic ion. Writing the formula for calcium phosphate as $Ca_3P_2O_8$ gives the correct number of each atom in the formula unit, but it obscures the fact that the compound contains readily identifiable PO_4^{3-} ions.

Example 5.4.2

Write the empirical formula for the compound formed from each ion pair.

a. Na^+ and HPO_4^{2-}
b. potassium cation and cyanide anion
c. calcium cation and hypochlorite anion

Given: ions

Asked for: empirical formula for ionic compound

Strategy:

A. If it is not given, determine the charge on a monatomic ion from its location in the periodic table. Use Table 5.4.1 to find the charge on a polyatomic ion.
B. Use the absolute value of the charge on each ion as the subscript for the other ion. Reduce the subscripts to the smallest whole numbers when writing the empirical formula.

Solution:

a. **B** Because HPO_4^{2-} has a charge of −2 and Na^+ has a charge of +1, the empirical formula requires two Na^+ ions to balance the charge of the polyatomic ion, giving Na_2HPO_4. The subscripts are reduced to the lowest numbers, so the empirical formula is Na_2HPO_4. This compound is sodium hydrogen phosphate, which is used to provide texture in processed cheese, puddings, and instant breakfasts.
b. **A** The potassium cation is K^+, and the cyanide anion is CN^-. **B** Because the magnitude of the charge on each ion is the same, the empirical formula is KCN. Potassium cyanide is highly toxic, and at one time it was used as rat poison. This use has been discontinued, however, because too many people were being poisoned accidentally.
c. **A** The calcium cation is Ca^{2+}, and the hypochlorite anion is ClO^-. **B** Two ClO^- ions are needed to balance the charge on one Ca^{2+} ion, giving $Ca(ClO)_2$. The subscripts cannot be reduced further, so the empirical formula is $Ca(ClO)_2$. This is calcium hypochlorite, the “chlorine” used to purify water in swimming pools.

Exercise 5.4.2

Write the empirical formula for the compound formed from each ion pair.

a. Ca^{2+} and $H_2PO_4^-$
b. sodium cation and bicarbonate anion
c. ammonium cation and sulfate anion

Answer a

$Ca(H_2PO_4)_2$: calcium dihydrogen phosphate is one of the ingredients in baking powder.

Answer b

$NaHCO_3$: sodium bicarbonate is found in antacids and baking powder; in pure form, it is sold as baking soda.

Answer c

$(NH_4)_2SO_4$: ammonium sulfate is a common source of nitrogen in fertilizers.



Polyatomics: Polyatomics, YouTube(opens in new window) [youtu.be] (opens in new window)

Hydrates

Many ionic compounds occur as hydrates, compounds that contain specific ratios of loosely bound water molecules, called waters of hydration. Waters of hydration can often be removed simply by heating. For example, calcium dihydrogen phosphate can form a solid that contains one molecule of water per $Ca(H_2PO_4)_2$ unit and is used as a leavening agent in the food industry to cause baked goods to rise. The empirical formula for the solid is $Ca(H_2PO_4)_2 \cdot H_2O$. In contrast, copper sulfate usually forms a blue solid that contains five waters of hydration per formula unit, with the empirical formula $CuSO_4 \cdot 5H_2O$. When heated, all five water molecules are lost, giving a white solid with the empirical formula $CuSO_4$.



Figure 5.4.1: Loss of Water from a Hydrate with Heating. (left) When blue $CuSO_4 \cdot 5H_2O$ is heated, two molecules of water are lost at 30°C, two more at 110°C, and the last at 250°C to give white "anhydrous" $CuSO_4$ (right). Images used with permission from Wikipedia.

Compounds that differ only in the numbers of waters of hydration can have very different properties. For example, $CaSO_4 \cdot \frac{1}{2}H_2O$ is *plaster of Paris*, which was often used to make sturdy casts for broken arms or legs, whereas $CaSO_4 \cdot 2H_2O$ is the less dense, flakier gypsum, a mineral used in drywall panels for home construction. When a cast would set, a mixture of plaster of Paris and water crystallized to give solid $CaSO_4 \cdot 2H_2O$. Similar processes are used in the setting of cement and concrete.

Binary Acids

Some compounds containing hydrogen are members of an important class of substances known as acids. The chemistry of these compounds is explored in more detail in later chapters of this text, but for now, it will suffice to note that many acids release hydrogen ions, H^+, when dissolved in water. To denote this distinct chemical property, a mixture of water with an acid is given a name derived from the compound's name. If the compound is a binary acid (comprised of hydrogen and one other nonmetallic element):

1. The word "hydrogen" is changed to the prefix *hydro-*
2. The other nonmetallic element name is modified by adding the suffix *-ic*
3. The word "acid" is added as a second word

For example, when the gas HCl (hydrogen chloride) is dissolved in water, the solution is called *hydrochloric acid*. Several other examples of this nomenclature are shown in Table 5.4.2.

Table 5.4.2: Names of Some Simple Acids

Name of Gas	Name of Acid
HF(*g*), hydrogen fluoride	HF(*aq*), hydrofluoric acid
HCl(*g*), hydrogen chloride	HCl(*aq*), hydrochloric acid
HBr(*g*), hydrogen bromide	HBr(*aq*), hydrobromic acid
HI(*g*), hydrogen iodide	HI(*aq*), hydroiodic acid
H_2S(*g*), hydrogen sulfide	H_2S(*aq*), hydrosulfuric acid

Oxyacids

Many compounds containing three or more elements (such as organic compounds or coordination compounds) are subject to specialized nomenclature rules that you will learn later. However, we will briefly discuss the important compounds known as oxyacids, compounds that contain hydrogen, oxygen, and at least one other element, and are bonded in such a way as to impart acidic properties to the compound (you will learn the details of this in a later chapter). Typical oxyacids consist of hydrogen combined with a polyatomic, oxygen-containing ion. To name oxyacids:

1. Omit "hydrogen"
2. Start with the root name of the anion
3. Replace *–ate* with *–ic*, or *–ite* with *–ous*
4. Add "acid"

For example, consider H_2CO_3 (which you might be tempted to call "hydrogen carbonate"). To name this correctly, "hydrogen" is omitted; the *–ate* of carbonate is replace with *–ic*; and acid is added—so its name is carbonic acid. Other examples are given in Table **5.4.3**. There are some exceptions to the general naming method (e.g., H_2SO_4 is called sulfuric acid, not sulfic acid, and H_2SO_3 is sulfurous, not sulfous, acid).

Table **5.4.3**: Names of Common Oxyacids

Formula	Anion Name	Acid Name
$HC_2H_3O_2$	acetate	acetic acid
HNO_3	nitrate	nitric acid
HNO_2	nitrite	nitrous acid
$HClO_4$	perchlorate	perchloric acid
H_2CO_3	carbonate	carbonic acid
H_2SO_4	sulfate	sulfuric acid
H_2SO_3	sulfite	sulfurous acid
H_3PO_4	phosphate	phosphoric acid



Nomenclature of Acids: Nomenclature of Acids, YouTube(opens in new window) [youtu.be]

Bases

We will present more comprehensive definitions of bases in later chapters, but virtually every base you encounter in the meantime will be an ionic compound, such as sodium hydroxide (NaOH) and barium hydroxide [$Ba(OH)_2$], that contain the hydroxide ion and a metal cation. These have the general formula $M(OH)_n$. It is important to recognize that alcohols, with the general formula ROH, are covalent compounds, not ionic compounds; consequently, they do *not* dissociate in water to form a basic solution (containing OH^- ions). When a base reacts with any of the acids we have discussed, it accepts a proton (H^+). For example, the hydroxide ion (OH^-) accepts a proton to form H_2O. Thus bases are also referred to as *proton acceptors*.

Concentrated aqueous solutions of ammonia (NH_3) contain significant amounts of the hydroxide ion, even though the dissolved substance is *not* primarily ammonium hydroxide (NH_4OH) as is often stated on the label. Thus aqueous ammonia solution is also a common base. Replacing a hydrogen atom of NH_3 with an alkyl group results in an amine (RNH_2), which is also a base. Amines have pungent odors—for example, methylamine (CH_3NH_2) is one of the compounds responsible for the foul odor associated with spoiled fish. The physiological importance of amines is suggested in the word *vitamin*, which is derived from the phrase *vital amines*. The word was coined to describe dietary substances that were effective at preventing scurvy, rickets, and other diseases because these substances were assumed to be amines. Subsequently, some vitamins have indeed been confirmed to be amines.

Binary Inorganic Compounds

Binary covalent compounds—covalent compounds that contain only two elements—are named using a procedure similar to that used for simple ionic compounds, but prefixes are added as needed to indicate the number of atoms of each kind. The procedure, diagrammed in Figure **5.4.2** consists of the following steps:

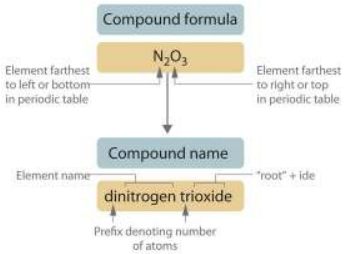


Figure **5.4.2**: Naming a Covalent Inorganic Compound

Breakdown of formula names and compound formula. The atom farthest to the left on the periodic table is named first.

1. Place the elements in their proper order.
   - The element farthest to the left in the periodic table is usually named first. If both elements are in the same group, the element closer to the bottom of the column is named first.
   - The second element is named as if it were a monatomic anion in an ionic compound (even though it is not), with the suffix *-ide* attached to the root of the element name.
2. Identify the number of each type of atom present.
   1. Prefixes derived from Greek stems are used to indicate the number of each type of atom in the formula unit (Table **5.4.3**). The prefix *mono-* ("one") is used only when absolutely necessary to avoid confusion, just as the subscript 1 is omitted when writing molecular formulas.

To demonstrate steps 1 and 2a, HCl is named hydrogen chloride (because hydrogen is to the left of chlorine in the periodic table), and PCl_5 is phosphorus pentachloride. The order of the elements in the name of BrF_3, bromine trifluoride, is determined by the fact that bromine lies below fluorine in Group 17.

Table **5.4.3**: Prefixes for Indicating the Number of Atoms in Chemical Names

Prefix	Number
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10
undeca-	11
dodeca-	12

2. If a molecule contains more than one atom of both elements, then prefixes are used for both. Thus N_2O_3 is *di*nitrogen *tri*oxide, as shown in Figure 2.13.
3. In some names, the final *a* or *o* of the prefix is dropped to avoid awkward pronunciation. Thus OsO_4 is osmium tetroxide rather than osmium tetraoxide.

3. Write the name of the compound.
   1. Binary compounds of the elements with oxygen are generally named as "element oxide," with prefixes that indicate the number of atoms of each element per formula unit. For example, CO is carbon monoxide. The only exception is binary compounds of oxygen with fluorine, which are named as oxygen fluorides.
   2. Certain compounds are *always* called by the common names that were assigned before formulas were used. For example, H_2O is water (not dihydrogen oxide); NH_3 is ammonia; PH_3 is phosphine; SiH_4 is silane; and B_2H_6, a *dimer* of BH_3, is diborane. For many compounds, the systematic name and the common name are both used frequently, requiring familiarity with both. For example, the systematic name for NO is nitrogen monoxide, but it is much more commonly called nitric oxide. Similarly, N_2O is usually called nitrous oxide rather than dinitrogen monoxide. Notice that the suffixes *-ic* and *-ous* are the same ones used for ionic compounds.

Start with the element at the far left in the periodic table and work to the right. If two or more elements are in the same group, start with the bottom element and work up.

Example 5.4.3: Binary Covalent Compounds

Write the name of each binary covalent compound.

a. SF_6
b. N_2O_4
c. ClO_2

Given: molecular formula

Asked for: name of compound

Strategy:

A. List the elements in order according to their positions in the periodic table. Identify the number of each type of atom in the chemical formula and then use Table **5.4.2** to determine the prefixes needed.

B. If the compound contains oxygen, follow step 3a. If not, decide whether to use the common name or the systematic name.

Solution:

a. **A** Because sulfur is to the left of fluorine in the periodic table, sulfur is named first. Because there is only one sulfur atom in the formula, no prefix is needed. **B** There are, however, six fluorine atoms, so we use the prefix for six: *hexa-* (Table **5.4.2**). The compound is sulfur hexafluoride.

b. **A** Because nitrogen is to the left of oxygen in the periodic table, nitrogen is named first. Because more than one atom of each element is present, prefixes are needed to indicate the number of atoms of each. According to Table **5.4.2** "Prefixes for Indicating the Number of Atoms in Chemical Names", the prefix for two is *di-*, and the prefix for four is *tetra-*. **B** The compound is dinitrogen tetroxide (omitting the *a* in *tetra-* according to step 2c) and is used as a component of some rocket fuels.

c. **A** Although oxygen lies to the left of chlorine in the periodic table, it is not named first because ClO_2 is an oxide of an element other than fluorine (step 3a). Consequently, chlorine is named first, but a prefix is not necessary because each molecule has only one atom of chlorine. **B** Because there are two oxygen atoms, the compound is a dioxide. Thus the compound is chlorine dioxide. It is widely used as a substitute for chlorine in municipal water treatment plants because, unlike chlorine, it does not react with organic compounds in water to produce potentially toxic chlorinated compounds.

Example 5.4.3

Write the name of each binary covalent compound.

a. IF_7
b. N_2O_5
c. OF_2

Answer a

iodine heptafluoride

Answer b

dinitrogen pentoxide

Answer c

oxygen difluoride

Example 5.4.4

Write the formula for each binary covalent compound.

a. sulfur trioxide
b. diiodine pentoxide

Given: name of compound

Asked for: formula

Strategy:

List the elements in the same order as in the formula, use Table **5.4.2** to identify the number of each type of atom present, and then indicate this quantity as a subscript to the right of that element when writing the formula.

Solution:

a. Sulfur has no prefix, which means that each molecule has only one sulfur atom. The prefix *tri-* indicates that there are three oxygen atoms. The formula is therefore SO_3. Sulfur trioxide is produced industrially in huge amounts as an intermediate in the synthesis of sulfuric acid.
b. The prefix *di-* tells you that each molecule has two iodine atoms, and the prefix *penta-* indicates that there are five oxygen atoms. The formula is thus I_2O_5, a compound used to remove carbon monoxide from air in respirators.

Exercise 5.4.4

Write the formula for each binary covalent compound.

a. silicon tetrachloride
b. disulfur decafluoride

Answer a

$SiCl_4$

Answer b

S_2F_{10}

The structures of some of the compounds in Examples **5.4.3** and **5.4.4** are shown in Figure **5.4.2** along with the location of the "central atom" of each compound in the periodic table. It may seem that the compositions and structures of such compounds are entirely random, but this is not true. After mastering the material discussed later on this course, one is able to predict the compositions and structures of compounds of this type with a high degree of accuracy.

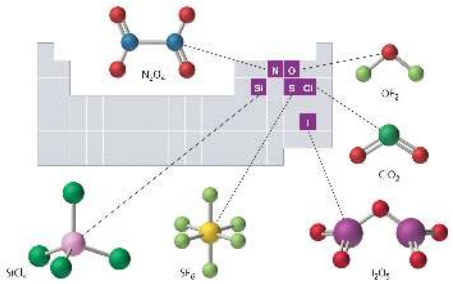


Figure **5.4.2**: The Structures of Some Covalent Inorganic Compounds and the Locations of the "Central Atoms" in the Periodic Table. The compositions and structures of covalent inorganic compounds are not random and can be predicted from the locations of the component atoms in the periodic table.

Central atoms of compounds N 2 O 4, O F 2, C l O 2, I 2 O 5, S F 6, and S i C l 4 shown on the periodic table.



Nomenclature of Nonmetals: Nomenclature of Nonmetals, YouTube(opens in new window) [youtu.be] (opens in new window)

Summary

The composition of a compound is represented by an empirical or molecular formula, each consisting of at least one formula unit. Covalent inorganic compounds are named using a procedure similar to that used for ionic compounds, whereas hydrocarbons use a system based on the number of bonds between carbon atoms. Covalent inorganic compounds are named by a procedure similar to that used for ionic compounds, using prefixes to indicate the numbers of atoms in the molecular formula. An empirical formula gives the relative numbers of atoms of the elements in a compound, reduced to the lowest whole numbers. The formula unit is the absolute grouping represented by the empirical formula of a compound, either ionic or covalent. Empirical formulas are particularly useful for

describing the composition of ionic compounds, which do not contain readily identifiable molecules. Some ionic compounds occur as hydrates, which contain specific ratios of loosely bound water molecules called waters of hydration.

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5.5: Naming Ionic Compounds

Learning Objectives

- To use the rules for naming ionic compounds.

After learning a few more details about the names of individual ions, you will be one step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 3.3). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system**, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word *ion*. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table **5.5.1**: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table **5.5.1** lists the elements that use the common system, along with their respective cation names.

Table **5.5.2**: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 5.5.2 lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

Example 5.5.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. SO_3^{2-}
d. NH_4^+
e. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the sulfite ion
d. the ammonium ion
e. the copper(I) ion or the cuprous ion

Exercise 5.5.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. SO_4^{2-}
d. Ba^{2+}
e. HCO_3^-

Answer a:
iron(II) ion

Answer b:
iron(III) ion

Answer c:
sulfate ion

Answer d:

barium ion

Answer e:

hydrogen carbonate ion or bicarbonate ion

Example 5.5.2

Write the formula for each ion.

a. the bromide ion
b. the phosphate ion
c. the cupric ion
d. the magnesium ion

Solution

a. Br^-
b. PO_4^{3-}
c. Cu^{2+}
d. Mg^{2+}

Exercise 5.5.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the stannous ion
d. the potassium ion

Answer a:

F^-

Answer b:

CO_3^{2-}

Answer c:

Sn^{2+}

Answer d:

K^+

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure **5.5.1** for the compound $BaCl_2$. The word *ion* is dropped from both parts.

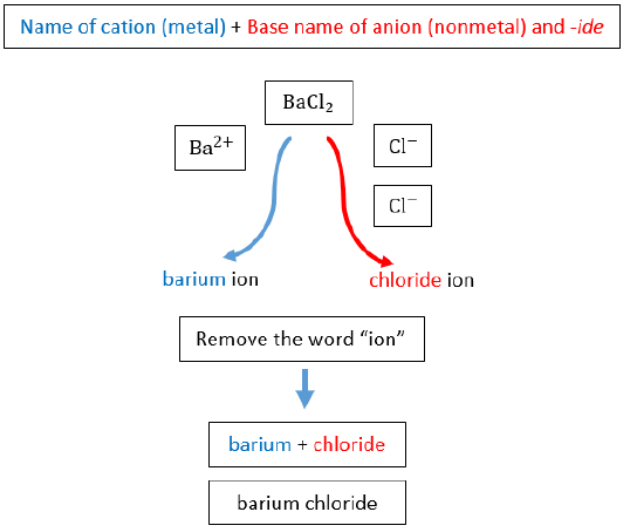


Figure 5.5.1: Naming $BaCl_2$
Naming formula: Name of metal cation + base
name of nonmetal anion + suffix -ide. BaCl2 is named as barium chloride.

Subscripts in the formula do not affect the name.

Example 5.5.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride.
b. The name of this ionic compound is aluminum fluoride.
c. The name of this ionic compound is potassium chloride

Exercise 5.5.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

Answer a:

silver iodide

Answer b:

magnesium oxide

Answer c:

calcium phosphide

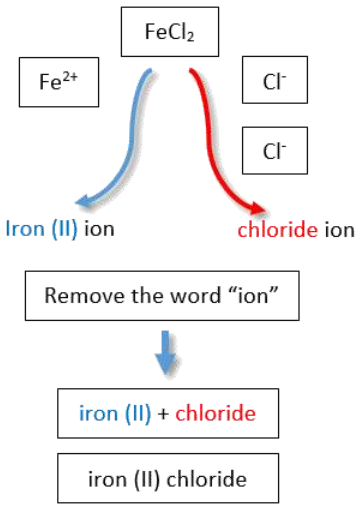
Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

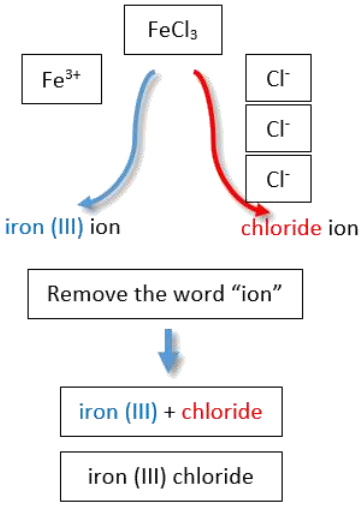
If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron

ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure **5.5.2**).

Table **5.5.3**: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Modern/Stock System.

Name of cation (metal) + (Roman Numeral in parenthesis) + Base name of anion (nonmetal) and *-ide*

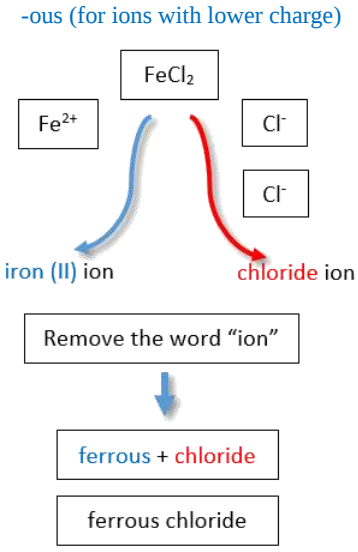


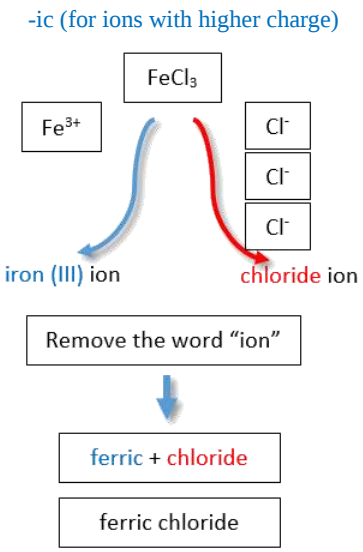


If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure **5.5.3**) .

Table **5.5.4**: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Old/Common System.

"Old" base name of cation (metal) and *-ic or -ous* + Base name of anion (nonmetal) and *-ide*





Example 5.5.4:

Name each ionic compound.

a. Co_2O_3
b. FeCl2

Solution

Solutions to Example 5.7.4

	Explanation	Answer
a	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2− charge, so with three oxide ions, we have a total negative charge of 6−. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
b	Iron can also have more than one possible charge. • Chloride always has a 1− charge, so with two chloride ions, we have a total negative charge of 2−. • This means that the one iron ion must have a 2+ charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

Exercise 5.5.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium sulfate compound in Figure **5.5.6**.

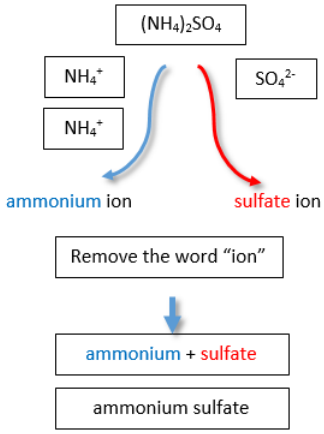


Figure **5.5.2**: Naming Ionic Compounds with Polyatomic Ions

Example 5.5.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Solutions to Example 5.7.5

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound’s name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound’s name is iron(II) phosphate.	iron(II) phosphate

Exercise 5.5.5A

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Figure 5.5.1 is a synopsis of how to name simple ionic compounds.

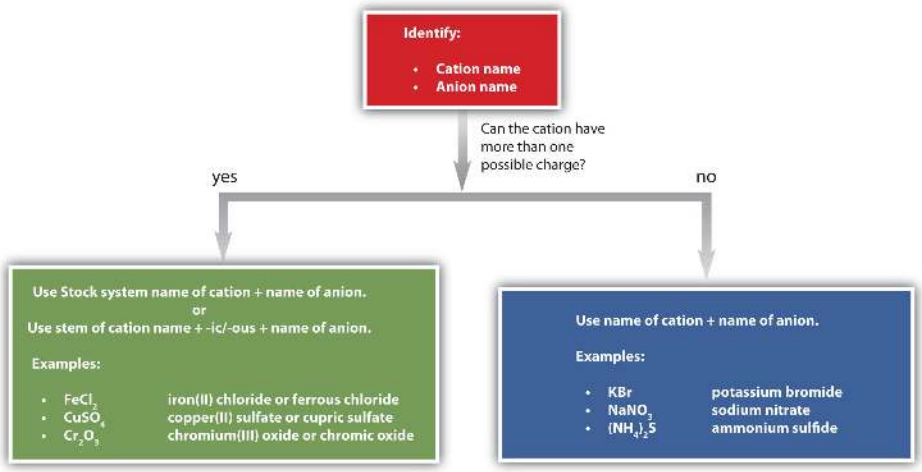


Figure 5.5.3: A Guide to Naming Simple Ionic Compounds.

? Exercise 5.5.5*B*

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a:

zinc bromide

Answer b:

aluminum oxide

Answer c:

ammonium phosphate

Answer d:

gold(III) fluoride or auric fluoride

Answer e:

silver fluoride

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parentheses.
- Ternary compounds are composed of three or more elements.

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5.6: Naming Acids

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

Acids

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



Figure 5.6.1: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

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

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

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed here.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. These anions can either be monatomic or polyatomic.

Naming Binary acids (in aqueous form)

A **binary acid** is an acid that consists of hydrogen and one other element. The most common binary acids contain a halogen. The acid name begins with the prefix *hydro-*. followed by the base name of the anion, followed by the suffix *-ic*.

Hydro- and Base name of Non metal and *-ic + acid*

Example: $HCl_{(aq)}$

Cl^-, chloride ion

hydrochloric acid

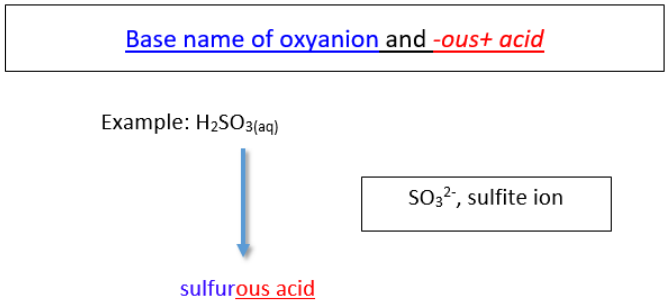
Formula for naming acids: Hydro- and Base name of nonmetal and -ic + acid. Example: HCl is hydrochloric acid.

Naming Oxyacids

An **oxyacid** is an acid that consists of hydrogen, oxygen, and a third element. The third element is usually a nonmetal.

a. Oxyanions with ***-ite*** ending.

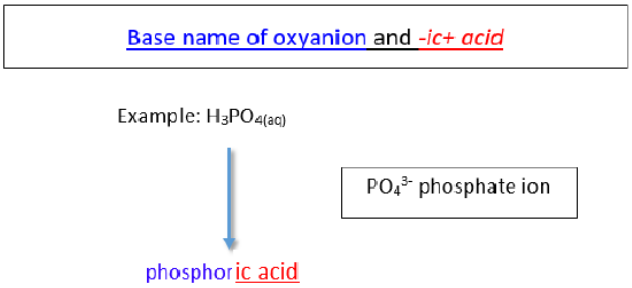
The name of the acid is the root of the anion followed by the suffix ***-ous***. There is no prefix.



Formula for naming oxyanions with -ite ending: Base name of oxyanion and -ous + acid. Example: H2SO3 is sulfurous acid.

b. Oxyanions with ***-ate*** ending.

The name of the acid is the root of the anion followed by the suffix ***-ic***. There is no prefix.

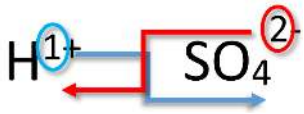


Formula for naming oxyanions with -ate ending: Base name of oxyanion and -ic + acid. Example: H3PO4 is phosphoric acid.

> **Note**
>
> The base name for sulfur containing oxyacid is ***sulfur-*** instead of just ***sulf-***. The same is true for a phosphorus containing oxyacid. The base name is ***phosphor-*** instead of simply ***phosph-***.

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H^+ ions. Another way to think about writing the correct formula is to utilize the crisscross method, shown below for sulfuric acid.



H2SO4: H has +1 charge and SO4 has -2 charge, so there must be 2 H+ and 1 SO4 for the charges to balance out.

Formula: H_2SO_4

Figure 5.6.2: *Crisscross approach to writing formula for sulfuric acid.*

Summary

- Acids are molecular compounds that release hydrogen ions.
- A binary acid consists of hydrogen and one other element.
- Oxyacids contain hydrogen, oxygen, and one other element.

- The name of the acid is based on the anion attached to the hydrogen.

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CHAPTER OVERVIEW

6: Covalent Compounds

6.1: Molecules and Molecular Compounds
6.2: Lewis Dot and Bonding
6.3: Covalent Lewis Structures: Electrons Shared
6.4: Writing Lewis Structures for Covalent Compounds
6.5: Predicting the Shapes of Molecules
6.6: Activity: Molecule Shapes
6.7: Naming Molecular Compounds
6.8: Nomenclature Summary

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6.1: Molecules and Molecular Compounds

Learning Objectives

- to understand the differences between covalent and ionic bonding.

The atoms in all substances that contain multiple atoms are held together by electrostatic interactions—interactions between electrically charged particles such as protons and electrons. Electrostatic attraction between oppositely charged species (positive and negative) results in a force that causes them to move toward each other, like the attraction between opposite poles of two magnets. In contrast, electrostatic repulsion between two species with the same charge (either both positive or both negative) results in a force that causes them to repel each other, as do the same poles of two magnets. Atoms form chemical compounds when the attractive electrostatic interactions between them are stronger than the repulsive interactions. Collectively, the attractive interactions between atoms are called chemical bonds.

Chemical bonds are generally divided into two fundamentally different types: ionic and covalent. In reality, however, the bonds in most substances are neither purely ionic nor purely covalent, but lie on a spectrum between these extremes. Although purely ionic and purely covalent bonds represent extreme cases that are seldom encountered in any but very simple substances, a brief discussion of these two extremes helps explain why substances with different kinds of chemical bonds have very different properties. Ionic compounds consist of positively and negatively charged ions held together by strong electrostatic forces, whereas covalent compounds generally consist of molecules, which are groups of atoms in which one or more pairs of electrons are shared between bonded atoms. In a covalent bond, atoms are held together by the electrostatic attraction between the positively charged nuclei of the bonded atoms and the negatively charged electrons they share. This discussion of structures and formulas begins by describing covalent compounds. The energetic factors involved in bond formation are described in more quantitative detail in later.

> Ionic compounds consist of ions of opposite charges held together by strong electrostatic forces, whereas pairs of electrons are shared between bonded atoms in covalent compounds.

Covalent Molecules and Compounds

Just as an atom is the simplest unit that has the fundamental chemical properties of an element, a molecule is the simplest unit that has the fundamental chemical properties of a covalent compound. Some pure elements exist as covalent molecules. Hydrogen, nitrogen, oxygen, and the halogens occur naturally as the diatomic (“two atoms”) molecules H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2 (part (a) in Figure **6.1.1**). Similarly, a few pure elements exist as polyatomic (“many atoms”) molecules, such as elemental phosphorus and sulfur, which occur as P_4 and S_8 (part (b) in Figure **6.1.1**).

Each covalent compound is represented by a molecular formula, which gives the atomic symbol for each component element, in a prescribed order, accompanied by a subscript indicating the number of atoms of that element in the molecule. The subscript is written only if the number of atoms is greater than 1. For example, water, with two hydrogen atoms and one oxygen atom per molecule, is written as H_2O. Similarly, carbon dioxide, which contains one carbon atom and two oxygen atoms in each molecule, is written as CO_2.

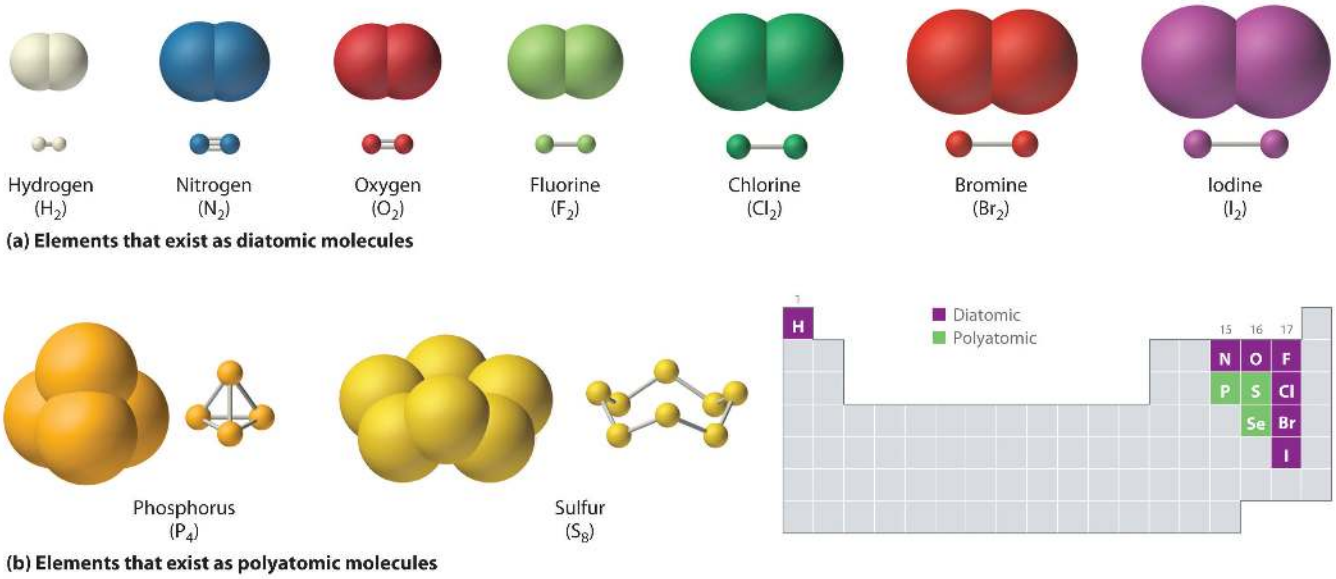


Figure **6.1.1**: Elements That Exist as Covalent Molecules. (a) Several elements naturally exist as diatomic molecules, in which two atoms (E) are joined by one or more covalent bonds to form a molecule with the general formula E2. (b) A few elements naturally exist as polyatomic molecules, which contain more than two atoms. For example, phosphorus exists as P4 tetrahedra—regular polyhedra with four triangular sides—with a phosphorus atom at each vertex. Elemental sulfur consists of a puckered ring of eight sulfur atoms connected by single bonds. Selenium is not shown due to the complexity of its structure.

A: Elements that exist as diatomic molecules: H2, N2, O2, F2, Cl2, Br2, and I2. B: Elements that exist as polyatomic molecules: P4 and S8.

Covalent compounds that predominantly contain carbon and hydrogen are called organic compounds. The convention for representing the formulas of organic compounds is to write carbon first, followed by hydrogen and then any other elements in alphabetical order (e.g., CH_4O is methyl alcohol, a fuel). Compounds that consist primarily of elements other than carbon and hydrogen are called inorganic compounds; they include both covalent and ionic compounds. In inorganic compounds, the component elements are listed beginning with the one farthest to the left in the periodic table, as in CO_2 or SF_6. Those in the same group are listed beginning with the lower element and working up, as in ClF. By convention, however, when an inorganic compound contains both hydrogen and an element from groups 13–15, hydrogen is usually listed last in the formula. Examples are ammonia (NH_3) and silane (SiH_4). Compounds such as water, whose compositions were established long before this convention was adopted, are always written with hydrogen first: Water is always written as H_2O, not OH_2. The conventions for inorganic acids, such as hydrochloric acid (HCl) and sulfuric acid (H_2SO_4), are described in Section 2.8.

For organic compounds: write C first, then H, and then the other elements in alphabetical order. For molecular inorganic compounds: start with the element at far left in the periodic table; list elements in same group beginning with the lower element and working up.

✔ Example 6.1.1

Write the molecular formula of each compound.

a. The phosphorus-sulfur compound that is responsible for the ignition of so-called strike anywhere matches has 4 phosphorus atoms and 3 sulfur atoms per molecule.
b. Ethyl alcohol, the alcohol of alcoholic beverages, has 1 oxygen atom, 2 carbon atoms, and 6 hydrogen atoms per molecule.
c. Freon-11, once widely used in automobile air conditioners and implicated in damage to the ozone layer, has 1 carbon atom, 3 chlorine atoms, and 1 fluorine atom per molecule.

Given: identity of elements present and number of atoms of each

Asked for: molecular formula

Strategy:

A. Identify the symbol for each element in the molecule. Then identify the substance as either an organic compound or an inorganic compound.
B. If the substance is an organic compound, arrange the elements in order beginning with carbon and hydrogen and then list the other elements alphabetically. If it is an inorganic compound, list the elements beginning with the one farthest left in the periodic table. List elements in the same group starting with the lower element and working up.
C. From the information given, add a subscript for each kind of atom to write the molecular formula.

Solution

a

A. The molecule has 4 phosphorus atoms and 3 sulfur atoms. Because the compound does not contain mostly carbon and hydrogen, it is inorganic.
B. Phosphorus is in group 15, and sulfur is in group 16. Because phosphorus is to the left of sulfur, it is written first.
C. Writing the number of each kind of atom as a right-hand subscript gives P_4S_3 as the molecular formula.

b.

A. Ethyl alcohol contains predominantly carbon and hydrogen, so it is an organic compound.
B. The formula for an organic compound is written with the number of carbon atoms first, the number of hydrogen atoms next, and the other atoms in alphabetical order: CHO.
C. Adding subscripts gives the molecular formula]C_2H_6O.

c

A. Freon-11 contains carbon, chlorine, and fluorine. It can be viewed as either an inorganic compound or an organic compound (in which fluorine has replaced hydrogen). The formula for Freon-11 can therefore be written using either of the two conventions.
B. According to the convention for inorganic compounds, carbon is written first because it is farther left in the periodic table. Fluorine and chlorine are in the same group, so they are listed beginning with the lower element and working up: CClF. Adding subscripts gives the molecular formula CCl_3F.
C. We obtain the same formula for Freon-11 using the convention for organic compounds. The number of carbon atoms is written first, followed by the number of hydrogen atoms (zero) and then the other elements in alphabetical order, also giving CCl_3F.

? Exercise 6.1.1

Write the molecular formula for each compound.

a. Nitrous oxide, also called “laughing gas,” has 2 nitrogen atoms and 1 oxygen atom per molecule. Nitrous oxide is used as a mild anesthetic for minor surgery and as the propellant in cans of whipped cream.
b. Sucrose, also known as cane sugar, has 12 carbon atoms, 11 oxygen atoms, and 22 hydrogen atoms.
c. Sulfur hexafluoride, a gas used to pressurize “unpressurized” tennis balls and as a coolant in nuclear reactors, has 6 fluorine atoms and 1 sulfur atom per molecule.

Answer a

N_2O

Answer b

$C_{12}H_{22}O_{11}$

Answer c

SF_6

Representations of Molecular Structures

Molecular formulas give only the elemental composition of molecules. In contrast, structural formulas show which atoms are bonded to one another and, in some cases, the approximate arrangement of the atoms in space. Knowing the structural formula of a compound enables chemists to create a three-dimensional model, which provides information about how that compound will behave physically and chemically.

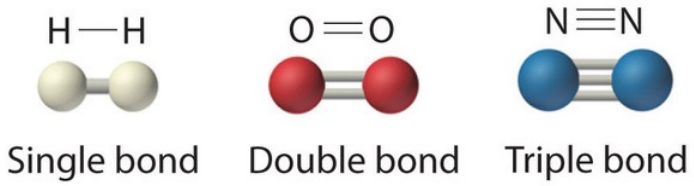


Figure **6.1.2**: Molecules That Contain Single, Double, and Triple Bonds. Hydrogen (H_2) has a single bond between atoms. Oxygen (O_2) has a double bond between atoms, indicated by two lines (=). Nitrogen (N_2) has a triple bond between atoms, indicated by three lines (≡). Each bond represents an electron pair.

The structural formula for H_2 can be drawn as H–H and that for I_2 as I–I, where the line indicates a single pair of shared electrons, a single bond. Two pairs of electrons are shared in a double bond, which is indicated by two lines—for example, O_2 is O=O. Three electron pairs are shared in a triple bond, which is indicated by three lines—for example, N_2 is N≡N (see Figure **6.1.2**). Carbon is unique in the extent to which it forms single, double, and triple bonds to itself and other elements. The number of bonds formed by an atom in its covalent compounds is not arbitrary. Hydrogen, oxygen, nitrogen, and carbon have very strong tendencies to form substances in which they have one, two, three, and four bonds to other atoms, respectively (Table **6.1.1**).

Table **6.1.1**: The Number of Bonds That Selected Atoms Commonly Form to Other Atoms

Atom	Number of Bonds
H (group 1)	1
O (group 16)	2
N (group 15)	3
C (group 14)	4

The structural formula for water can be drawn as follows:



Because the latter approximates the experimentally determined shape of the water molecule, it is more informative. Similarly, ammonia (NH_3) and methane (CH_4) are often written as planar molecules:

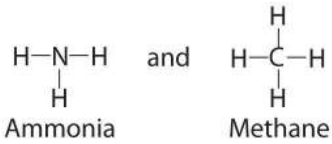


Figure **6.1.3**, is tetrahedral: the hydrogen atoms are positioned at every other vertex of a cube. Many compounds—carbon compounds, in particular—have four bonded atoms arranged around a central atom to form a tetrahedron.

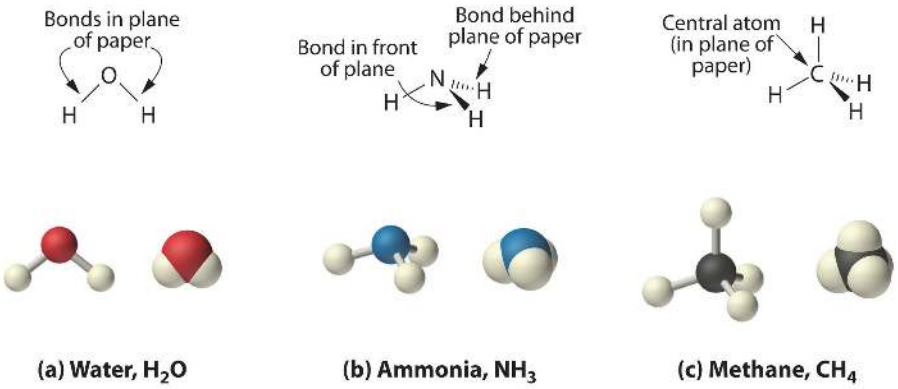
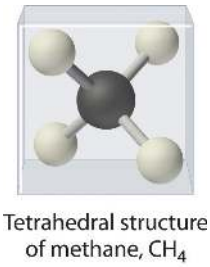


Figure 6.1.3: The Three-Dimensional Structures of Water, Ammonia, and Methane. (a) Water is a V-shaped molecule, in which all three atoms lie in a plane. (b) In contrast, ammonia has a pyramidal structure, in which the three hydrogen atoms form the base of the pyramid and the nitrogen atom is at the vertex. (c) The four hydrogen atoms of methane form a tetrahedron; the carbon atom lies in the center.

A: Bent configuration of H 2 O. B: Trigonal pyramidal configuration of ammonia. C: Tetrahedral configuration of methane.



Methane CH_4 has a three-dimensional, tetrahedral structure.

Figures 6.1.1, 6.1.2, and 6.1.3 illustrate different ways to represent the structures of molecules. It should be clear that there is no single "best" way to draw the structure of a molecule; the method used depends on which aspect of the structure should be emphasized and how much time and effort is required. Figure 6.1.4 shows some of the different ways to portray the structure of a slightly more complex molecule: methanol. These representations differ greatly in their information content. For example, the molecular formula for methanol (Figure 6.1.4*a*) gives only the number of each kind of atom; writing methanol as CH_4O tells nothing about its structure. In contrast, the structural formula (Figure 6.1.4*b*) indicates how the atoms are connected, but it makes methanol look as if it is planar (which it is not). Both the ball-and-stick model (part (c) in Figure 6.1.4) and the perspective drawing (Figure 6.1.4*d*) show the three-dimensional structure of the molecule. The latter (also called a wedge-and-dash representation) is the easiest way to sketch the structure of a molecule in three dimensions. It shows which atoms are above and below the plane of the paper by using wedges and dashes, respectively; the central atom is always assumed to be in the plane of the paper. The space-filling model (part (e) in Figure 6.1.4) illustrates the approximate relative sizes of the atoms in the molecule, but it does not show the bonds between the atoms. In addition, in a space-filling model, atoms at the "front" of the molecule may obscure atoms at the "back."

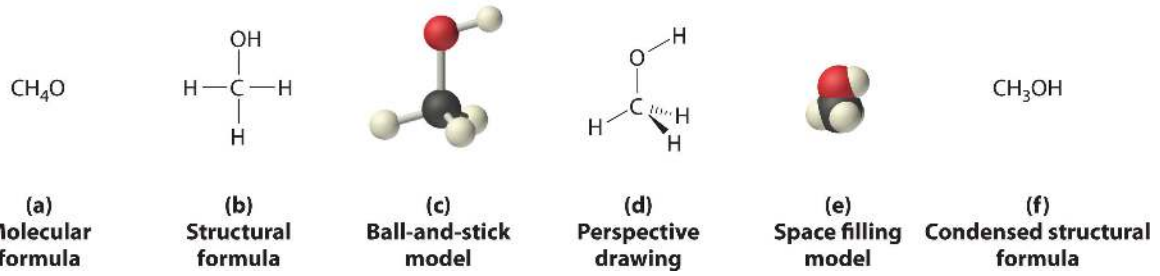


Figure **6.1.4**: Different Ways of Representing the Structure of a Molecule. (a) The molecular formula for methanol gives only the number of each kind of atom present. (b) The structural formula shows which atoms are connected. (c) The ball-and-stick model shows the atoms as spheres and the bonds as sticks. (d) A perspective drawing (also called a wedge-and-dash representation) attempts to show the three-dimensional structure of the molecule. (e) The space-filling model shows the atoms in the molecule but not the bonds. (f) The condensed structural formula is by far the easiest and most common way to represent a molecule.

Although a structural formula, a ball-and-stick model, a perspective drawing, and a space-filling model provide a significant amount of information about the structure of a molecule, each requires time and effort. Consequently, chemists often use a condensed structural formula (part (f) in Figure **6.1.4**), which omits the lines representing bonds between atoms and simply lists the atoms bonded to a given atom next to it. Multiple groups attached to the same atom are shown in parentheses, followed by a subscript that indicates the number of such groups. For example, the condensed structural formula for methanol is CH_3OH, which indicates that the molecule contains a CH_3 unit that looks like a fragment of methane (CH_4). Methanol can therefore be viewed either as a methane molecule in which one hydrogen atom has been replaced by an –OH group or as a water molecule in which one hydrogen atom has been replaced by a $–CH_3$ fragment. Because of their ease of use and information content, we use condensed structural formulas for molecules throughout this text. Ball-and-stick models are used when needed to illustrate the three-dimensional structure of molecules, and space-filling models are used only when it is necessary to visualize the relative sizes of atoms or molecules to understand an important point.

Example 6.1.2: Molecular Formulas

Write the molecular formula for each compound. The condensed structural formula is given.

a. Sulfur monochloride (also called disulfur dichloride) is a vile-smelling, corrosive yellow liquid used in the production of synthetic rubber. Its condensed structural formula is ClSSCl.
b. Ethylene glycol is the major ingredient in antifreeze. Its condensed structural formula is $HOCH_2CH_2OH$.
c. Trimethylamine is one of the substances responsible for the smell of spoiled fish. Its condensed structural formula is $(CH_3)_3N$.

Given: condensed structural formula

Asked for: molecular formula

Strategy:

A. Identify every element in the condensed structural formula and then determine whether the compound is organic or inorganic.
B. As appropriate, use either organic or inorganic convention to list the elements. Then add appropriate subscripts to indicate the number of atoms of each element present in the molecular formula.

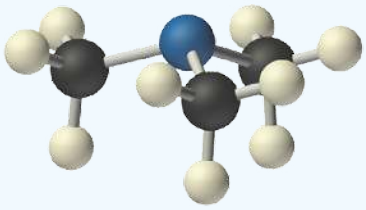
Solution:

The molecular formula lists the elements in the molecule and the number of atoms of each.

a. **A** Each molecule of sulfur monochloride has two sulfur atoms and two chlorine atoms. Because it does not contain mostly carbon and hydrogen, it is an inorganic compound. **B** Sulfur lies to the left of chlorine in the periodic table, so it is written first in the formula. Adding subscripts gives the molecular formula S_2Cl_2.

b. **A** Counting the atoms in ethylene glycol, we get six hydrogen atoms, two carbon atoms, and two oxygen atoms per molecule. The compound consists mostly of carbon and hydrogen atoms, so it is organic. **B** As with all organic compounds, C and H are written first in the molecular formula. Adding appropriate subscripts gives the molecular formula $C_2H_6O_2$.

c. **A** The condensed structural formula shows that trimethylamine contains three CH_3 units, so we have one nitrogen atom, three carbon atoms, and nine hydrogen atoms per molecule. Because trimethylamine contains mostly carbon and hydrogen, it is an organic compound. **B** According to the convention for organic compounds, C and H are written first, giving the molecular formula C_3H_9N.



Trimethylamine

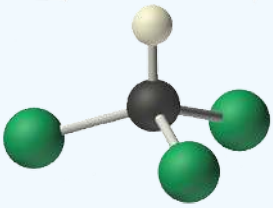
? Exercise 6.1.2: Molecular Formulas

Write the molecular formula for each molecule.

a. Chloroform, which was one of the first anesthetics and was used in many cough syrups until recently, contains one carbon atom, one hydrogen atom, and three chlorine atoms. Its condensed structural formula is $\mathbf{CHCl_3}$.

b. Hydrazine is used as a propellant in the attitude jets of the space shuttle. Its condensed structural formula is $\mathbf{H_2NNH_2}$.

c. Putrescine is a pungent-smelling compound first isolated from extracts of rotting meat. Its condensed structural formula is H2NCH2CH2CH2CH2NH2. This is often written as $\mathbf{H_2N(CH_2)_4NH_2}$ to indicate that there are four CH2 fragments linked together.



Chloroform

Answer a

$CHCl_3$

Answer b

N_2H_4

Answer c

$C_4H_{12}N_2$

Summary

There are two fundamentally different kinds of chemical bonds (covalent and ionic) that cause substances to have very different properties. The atoms in chemical compounds are held together by attractive electrostatic interactions known as chemical bonds. The molecular formula of a covalent compound gives the types and numbers of atoms present. Compounds that contain predominantly carbon and hydrogen are called organic compounds, whereas compounds that consist primarily of elements other than carbon and hydrogen are inorganic compounds. Diatomic molecules contain two atoms, and polyatomic molecules contain more than two. A structural formula indicates the composition and approximate structure and shape of a molecule. Covalent

molecular compounds, in contrast, consist of discrete molecules held together by weak intermolecular forces and can be gases, liquids, or solids at room temperature and pressure.

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6.2: Lewis Dot and Bonding

Learning Objectives

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

Bonding (Ionic and Covalent Basics)

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



Figure **6.2.1**: NaCl crystals. (Public Domain; NASA).

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



Figure **6.2.2**: Covalent compounds NO_2 and N_2O_4 contained in glass bottles. Image taken from: commons.wikimedia.org/wiki/File:NO2-N2O4.jpg

Table **6.2.1**: Comparing ionic and covalent compounds

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

Lewis Symbols

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

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

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

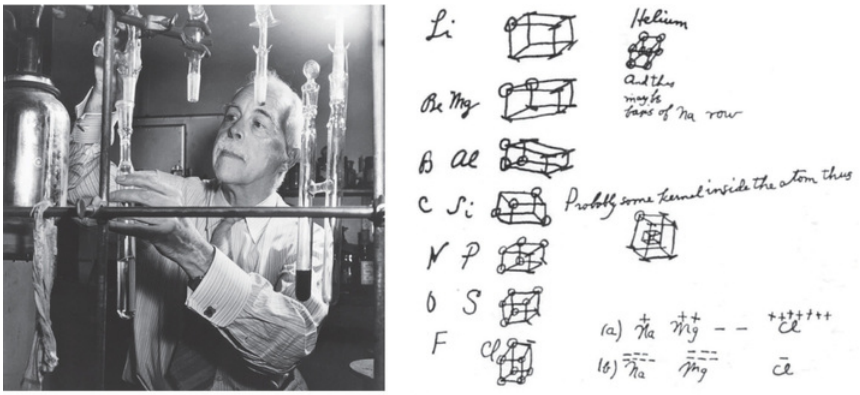


Figure 6.2.3: G. N. Lewis and the Octet Rule. (a) Lewis is working in the laboratory. (b) In Lewis's original sketch for the octet rule, he initially placed the electrons at the corners of a cube rather than placing them as we do now.

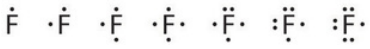
Lewis Dot symbols

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

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



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



The number of dots in the Lewis dot symbol is the same as the number of valence electrons, which is the same as the last digit of the element's group number in the periodic table. Lewis dot symbols and electron configurations (not a topic for this course) for the elements in period 2 are given in Table **6.2.2**.

***Table* 6.2.2:** *Lewis Dot Symbols for the Elements in Period 2*

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

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

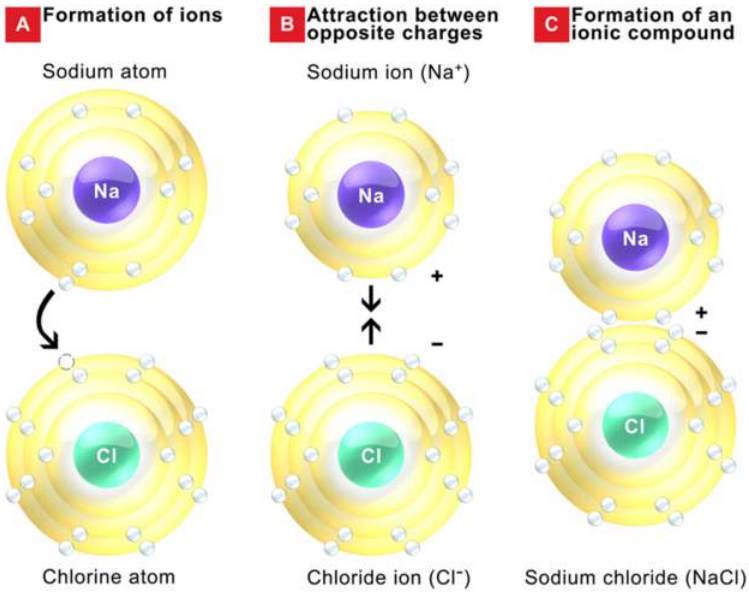


Figure **6.2.4**: Formation of an ionic bond between sodium and chlorine. Image courtesy of By BruceBlaus - Own work, CC BY-SA 4.0, commons.wikimedia.org/w/inde...curid=44968540

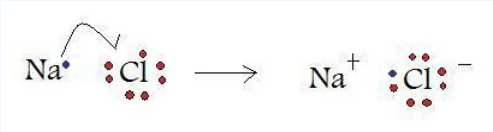
Rules for Drawing IOnic Lewis structures

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

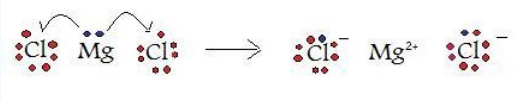
a. Locate the elements of interest and recall how many valence electrons each species has.
b. Draw dots around each element individually.
c. To obtain successful octets, the metal must transfer all of its valence electrons to the nonmetal.

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

Example 6.2.1: Chloride Salts



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



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

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

Contributors and Attributions

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6.3: Covalent Lewis Structures: Electrons Shared

Learning Objectives

- Define *covalent bond*.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won't give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

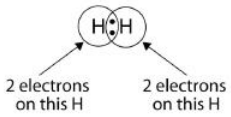
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \quad \cdot\mathbf{H}$$

The two H atoms can share their electrons:

$$\mathbf{H:H}$$

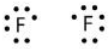
We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity's sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:

H–H

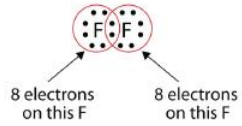
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:



These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.



Note that each F atom has a complete octet around it now:



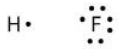
We can also write this using a dash to represent the shared electron pair:



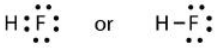
There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F

atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

Example 6.3.1:

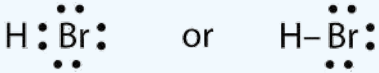
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:



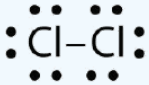
The two atoms can share their unpaired electron:



Exercise 6.3.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

Answer



When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far, and there are still some electrons that remain unattached. You can't just leave them there. So where do you put them?

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of $\mathbf{C_2H_4}$. The carbon atoms are bonded together, with each carbon also bonded to two hydrogen atoms.

$$\text{two } \mathbf{C} \text{ atoms } = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four } \mathbf{H} \text{ atoms } = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of } 12 \text{ valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

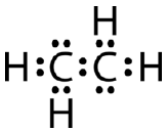


Figure **6.3.1**: Incorrect dot structure of ethene. (CK12 License)

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

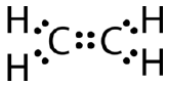


Figure **6.3.2**: Correct dot structure for ethene. (CK12 License)

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as shown in the figure below.

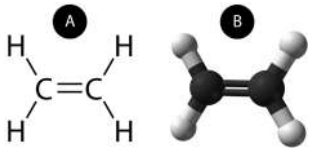


Figure **6.3.3**: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

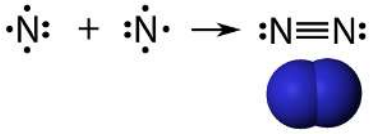


Figure **6.3.4**: Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons, and six electrons that are shared between the atoms.

Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

Contributions & Attributions

-
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6.4: Writing Lewis Structures for Covalent Compounds

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

  For CO_3^{2-}, for example, we add two electrons to the total because of the -2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 6.4.1: Water

Write the Lewis Structure for H_2O.

Solution

Solutions to Example 10.4.1

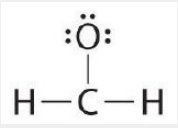
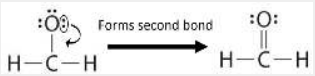
Steps for Writing Lewis Structures	Example **6.4.1**
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives H -O- H with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

Example 6.4.2

Write the Lewis structure for the CH_2O molecule

Solution

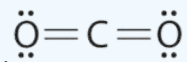
Solutions to Example 10.4.2

Steps for Writing Lewis Structures	Example **6.4.2**
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  6 electrons are used, and 6 are left over.
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.
5. If any electrons are left over, place them on the central atom.	Not necessary. There are no electrons left to place on the central atom.
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

? Exercise 6.4.1

Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

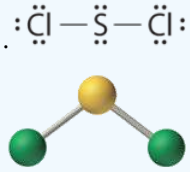
Answer CO_2





Carbon dioxide

Answer SCl_2



Sulfur dichloride

The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions (CK-12)

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, $\mathbf{NH_4^+}$, is formed when a hydrogen ion ($\mathbf{H^+}$) attaches to the lone pair of an ammonia ($\mathbf{NH_3}$) molecule in a coordinate covalent bond.

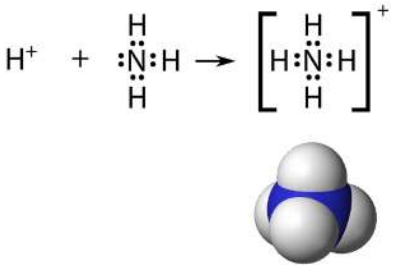


Figure **6.4.3**: The ammonium ion. (CK12 License)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$$\mathbf{1\ N}\text{ atom } = \mathbf{5}\text{ valence electrons}$$

$$\mathbf{4\ H}\text{ atoms } = \mathbf{4 \times 1 = 4}\text{ valence electrons}$$

subtract 1 electron for the $\mathbf{1+}$ charge of the ion

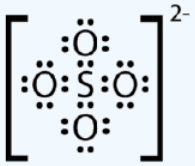
total of 8 valence electrons in the ion

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

Exercise 6.4.2

Draw the Lewis electron dot structure for the sulfate ion.

Answer (CK12 License)



Exceptions to the Octet Rule (BC Campus)

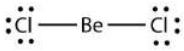
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

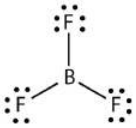


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

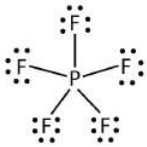
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

Example 6.4.3: Octet Violations

Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

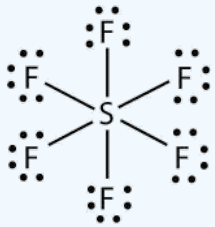
a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:



b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



? Exercise 6.4.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer



The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.

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6.5: Predicting the Shapes of Molecules

Learning Objective

- Determine the shape of simple molecules.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is called **valence shell electron pair repulsion (VSEPR).** VSEPR says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from one another as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2:



Figure **6.5.1**: Beryllium hydride and carbon dioxide bonding.

The two molecules, shown in the figure below in a "ball and stick" model.

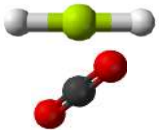


Figure **6.5.2**: Beryllium hydride and carbon dioxide models. (CK12 Licence)

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3:

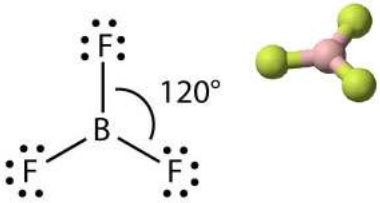


Figure **6.5.3**: Boron trifluoride bonding. (CK12 Licence)

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2:

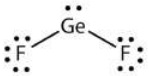


Figure **6.5.4**: Germanium difluoride bonding.

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure **6.5.1** Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

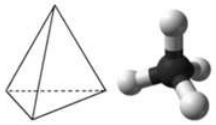


Figure **6.5.5**: Tetrahedral structure of methane. (CK12 Licence)

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

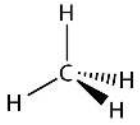


Figure **6.5.6**: Methane bonding. (CK12 Licence)

NH_3 is an example of a molecule whose central atom has four electron groups, but only three of them are bonded to surrounding atoms.

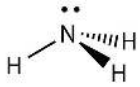


Figure **6.5.7**: Ammonia bonding. (CK12 Licence)

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups, but only two of them are bonded to surrounding atoms.



Figure **6.5.8**: Water bonding.

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom, but only one electron group bonded to another atom, is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure **6.5.9**.



Figure **6.5.9**: Lewis Electron Dot Diagram of Formaldehyde.

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

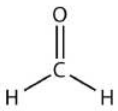


Figure **6.5.10**: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

Table **6.5.1** summarizes the shapes of molecules based on the number of electron groups and surrounding atoms.

Table **6.5.1**: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

Example 6.5.1

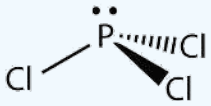
What is the approximate shape of each molecule?

a. PCl_3
b. NOF

Solution

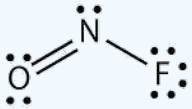
The first step is to draw the Lewis structure of the molecule.

For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

Exercise 6.5.1

What is the approximate molecular shape of CH_2Cl_2?

Answer

Tetrahedral

? Exercise 6.5.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. (Hint: hydrogen is a terminal atom.)

Answer

Trigonal planar about both central C atoms.

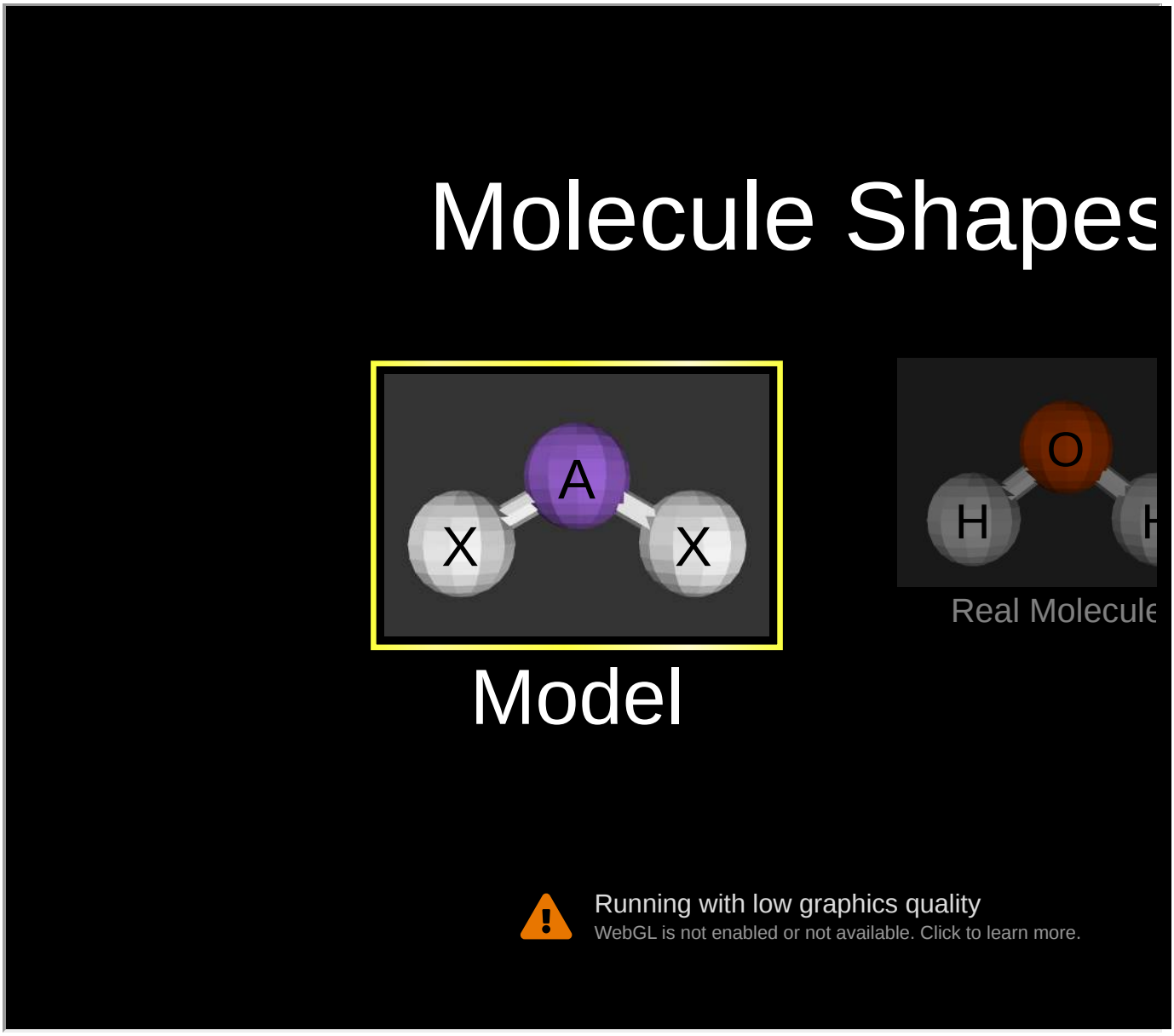
Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

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6.6: Activity: Molecule Shapes



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6.7: Naming Molecular Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

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

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

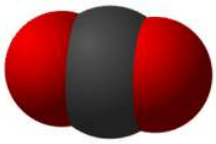


Figure 6.7.1: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

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

Naming Binary Molecular Compounds

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



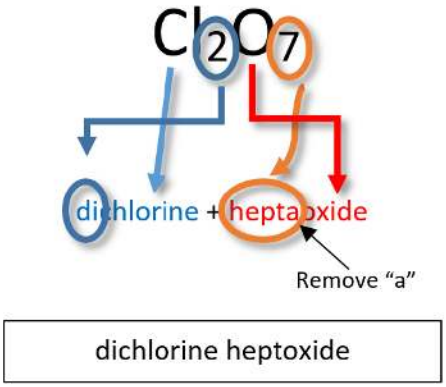
Figure 6.7.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

Naming *binary* (two-element) molecular compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 6.7.1 lists these numerical prefixes.

Table 6.7.1: Numerical Prefixes for Naming Binary Covalent Compounds

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

*This prefix is not used for the first element's name.



Note

- Generally, the less electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3. The order of common nonmetals in binary compound formulas is C, P, N, H, S, I, Br, Cl, O, F.
- The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms, is tetr**o**xide instead of tetr**ao**xide.
- The prefix is "mono" is not added to the first element's name if there is only one atom of the first element in a molecule.

Some examples of molecular compounds are listed in Table **6.7.2**.

Table **6.7.2**

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

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

? Exercise 6.7.1

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

Simple molecular compounds with common names

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane
- H_2O_2: hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.

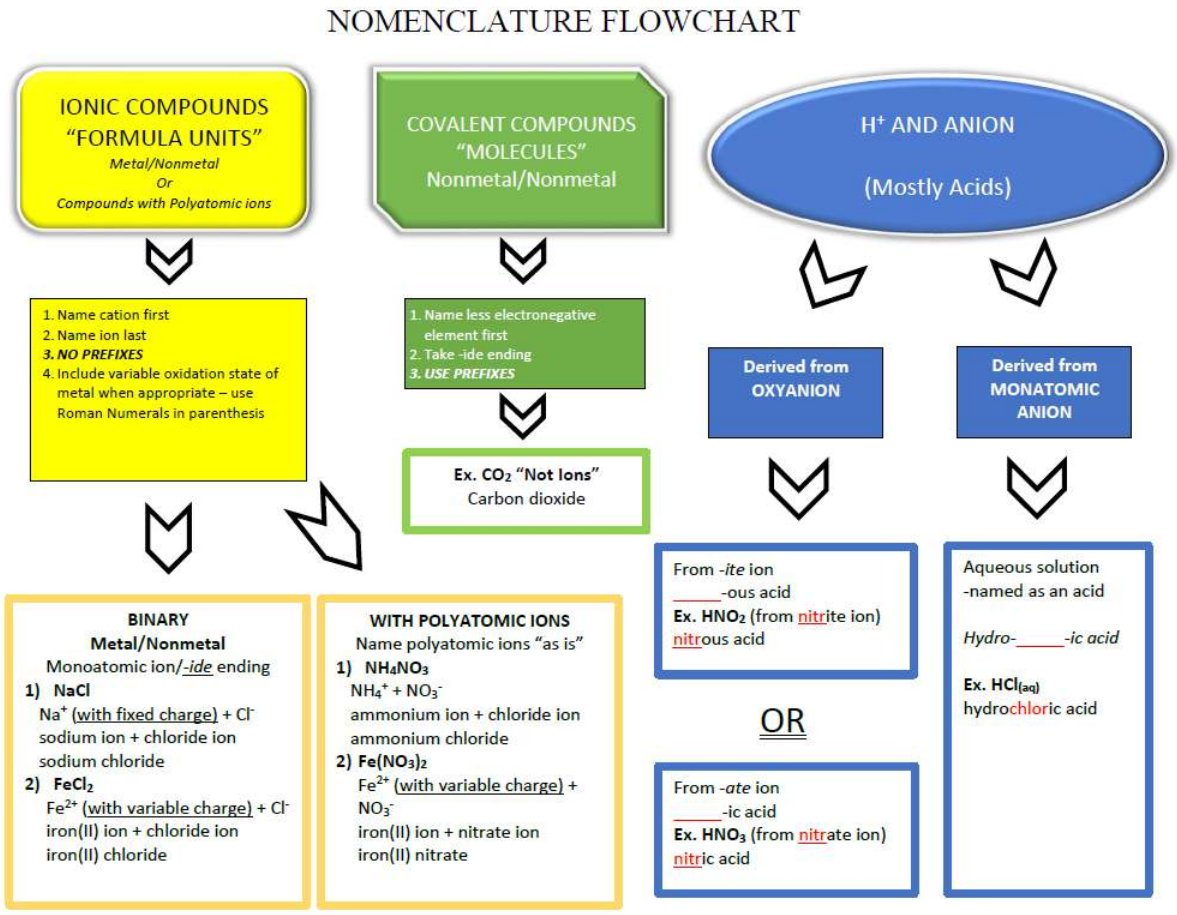
Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix -ide. Numerical prefixes are used to specify the number of atoms in a molecule.

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6.8: Nomenclature Summary



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7.1: Climate and Global Change – Chapter Introduction

The Earth's climate is changing. The scientific consensus is that by altering the composition of the atmosphere humans are increasing the average temperature of the Earth's surface. This process has already begun – the planet is measurably warmer than it was at the start of the last century – but scientists predict the change that will occur over the 21st century will be even greater. This increase will have unpredictable impacts on weather patterns around the globe. We are all experiencing climate change. Our descendants will likely experience far more.

This chapter focuses on the science of climate change. We recognize that climate change can be a controversial subject, and that prescriptions for solutions quickly take on a political character, which can raise suspicions of bias. Some argue that the climate is too complicated to predict, and others suggest that natural variations can explain the observed changes in the climate.

These objections have some merit. It should be no surprise that the Earth's climate is a complicated subject. First, the atmosphere is vast: it extends over 600 km (370 miles) above the ground, and it weighs over five quadrillion tons (that's a five followed by 15 zeros). Second, the atmosphere is a dynamic system, creating blizzards, hurricanes, thunderstorms, and all the other weather we experience. And it is true that this dynamic system is largely controlled by natural processes – the Earth's climate has been changing continually since the atmosphere was produced.

And yet scientists can still confidently say that humans are responsible for the current warming. We can do this because this complicated system obeys underlying principles. In the modules 5.2 and 5.3, we will describe how these principles work, and how they have been observed by scientists. We can then use these principles to understand how, by producing greenhouse gases, humans are altering the physical properties of the atmosphere in such a way as to increase its ability to retain heat.

In the module 5.4 we show how this theoretical prediction of a warming world is borne out by ever stronger evidence. Temperatures have been measured, and are shown to be increasing. These increases in temperatures are significant and have observable effects on the world: glaciers are shrinking, sea ice is retreating, sea levels are rising – even cherry blossoms are blooming earlier in the year.

In the module 5.5, we describe how we can attempt to predict the climate of the future. This is a doubly difficult problem, as it involves not only physics, but, harder yet, people. What will today's societies do with the foreknowledge of the consequences of our actions? The climate has become yet another natural system whose immediate fate is connected to our own. The reader may find it either reassuring or frightening that we hold the climate's future in our hands.

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7.2: Climate Processes; External and Internal Controls

Learning Objectives

After reading this module, students should be able to

- define both "climate" and "weather" and explain how the two are related
- use the Celsius temperature scale to describe climate and weather
- discuss the role and mechanisms of the major controls on Earth's climate using the concepts of insolation, albedo and greenhouse gases
- identify and describe the mechanisms by which major external and internal changes to the climate (including solar output variation, volcanoes, biological processes, changes in glacial coverage, and meteorite impacts) operate
- know that the Earth's climate has changed greatly over its history as a result of changes in insolation, albedo, and atmospheric composition
- describe the processes that can lead to a "Snowball Earth" using the "positive feedback" concept, and be able to contrast the climate factors that influenced this period of Earth's history with others, including the dominant factors that operated during the Cretaceous
- state the major ways in which carbon dioxide is both added to and removed from the atmosphere, and be able to describe why levels of carbon dioxide and other greenhouse gases can be kept in balance

Introduction

The Earth's climate is continually changing. If we are to understand the current climate and predict the climate of the future, we need to be able to account for the processes that control the climate. One hundred million years ago, much of North America was arid and hot, with giant sand dunes common across the continent's interior. Six hundred and fifty million years ago it appears that the same land mass—along with the rest of the globe—was covered in a layer of snow and ice. What drives these enormous changes through Earth's history? If we understand these fundamental processes we can explain why the climate of today may also change.

In discussing climate in this chapter, we will be using degrees Celsius (^{o}C) as the unit of temperature measurement.

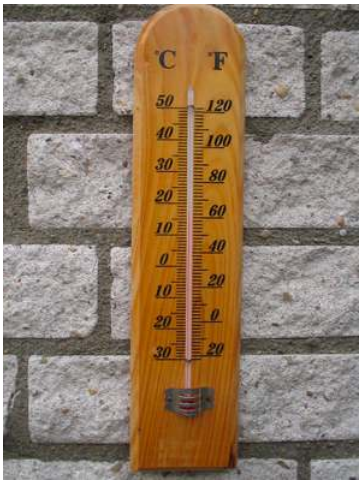


Figure **7.2.1** This thermometer shows how the two scales compare for typical atmospheric temperatures. A change of one degree Celsius (1 ^{o}C) is equivalent to a change of one and four fifths degrees Fahrenheit (1.8 ^{o}F). *Source: Michiel1972 at nl.Wikipedia.*

The Celsius scale is the standard international unit for temperature that scientists use when discussing the climate. In the Celsius scale, water freezes at 0 ^{o}C and boils at 100 ^{o}C. A comfortable room might be heated to 20 ^{o}C (which is equivalent to 68 ^{o}F). Temperatures can be converted from the Celsius scale to the Fahrenheit scale with the following equation:

$$°F = \frac{9}{5}°C + 32 \tag{7.2.1}$$

Weather describes the short term state of the atmosphere. This includes such conditions as wind, air pressure, precipitation, humidity and temperature. Climate describes the typical, or average, atmospheric conditions. Weather and climate are different as the short term state is always changing but the long-term average is not. On The 1st of January, 2011, Chicago recorded a high temperature of 6 ^{o}C; this is a measure of the weather. Measurements of climate include the averages of the daily, monthly, and yearly weather patterns, the seasons, and even a description of how often extraordinary events, such as hurricanes, occur. So if we consider the average Chicago high temperature for the 1st of January (a colder 0.5 ^{o}C) or the average high temperature for the entire year (a warmer 14.5 ^{o}C) we are comparing the city's weather with its climate. The climate is the average of the weather.

Insolation, Albedo and Greenhouse Gases

What controls the climate? The average temperature of the Earth is about 15 ^{o}C (which is the yearly average temperature for the city of San Francisco), so most of the Earth's water is in a liquid state. The average temperature of Mars is about -55 ^{o}C (about the same as the average winter temperature of the South Pole), so all of the water on the Martian surface is frozen. This is a big difference! One reason Earth is so much hotter than Mars is that Earth is closer to the Sun. Mars receives less than half as much energy from the Sun per unit area as Earth does. This difference in insolation, which is the measure of the amount of solar radiation falling on a surface, is a very important factor in determining the climate of the Earth.

On Earth, we notice the effects of varying insolation on our climate. Sunlight falls most directly on the equator, and only obliquely (at an angle) on the poles. This means that the sunlight is more concentrated at the equator. As shown in Figure **7.2.2**, the same amount of sunlight covers twice as much area when it strikes a surface at an angle of 30° compared to when it strikes a surface directly: the same energy is spread more thinly, weakening its ability to warm the Earth.

Firgure 7.2.2 Insolation angle. Insolation is the effect of incidence angle on sunlight intensity. Note that the same amount of sunlight is spread out over twice the area when it strikes the surface at a 30-degree angle. *Source: Wikipedia*

As a consequence, the tropics receive about twice the insolation as the area inside the Arctic Circle – see Figure **7.2.3**. This difference in energy explains why the equator has a hot climate and the poles have a cold climate. Differences in insolation also explain the existence of seasons. The Earth's axis is tilted at 23° compared to its orbit, and so over the course of the year each hemisphere alternates between directly facing the Sun and obliquely facing the Sun. When the Northern hemisphere is most directly facing the Sun (the months of May, June and July) insolation is thus higher, and the climate is warmer. This variation in insolation explains why summer and winter occur (we get less energy from the Sun in winter then we do in summer), and why the timing of the seasons is opposite in the Southern and Northern hemispheres.

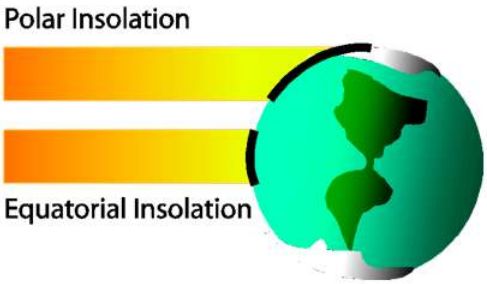


Figure **7.2.3** Insolation effect. A cartoon of how latitude is important in determining the amount of insolation. The same amount of sunlight (yellow bars) is spread out over twice the planet's surface area when the rays strike the Earth at an angle (compare the length of the dark lines at the equator and at the poles). *Source: Jonathan H. Tomkin.*

Figure **7.2.4** shows both the equatorial and seasonal impacts of insolation. High levels of insolation are shown in warm colors (red and pink) and low levels of insolation are shown in cold colors (blue). Notice that in January (top map) the maximum levels of insolation are in the Southern Hemisphere, as this is when the Southern Hemisphere is most directly facing the sun. The Arctic receives very little insolation at this time of year, as it experiences its long polar night. The reverse is true in April (bottom map).

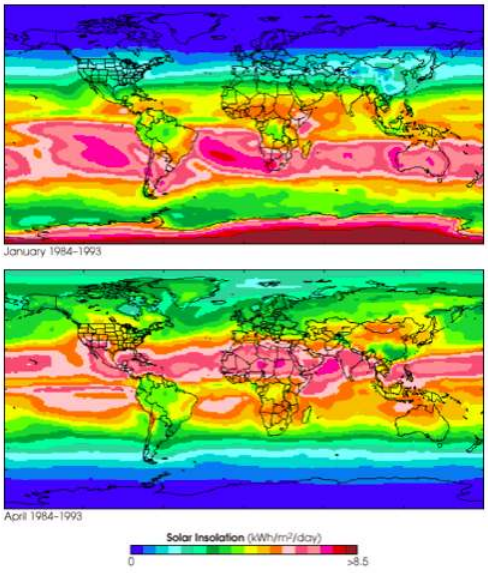


Figure **7.2.4** Insolation. Average insolation over ten years for the months of January (top) and April (bottom). *Source: Roberta DiPasquale, Surface Meteorology and Solar Energy Project, NASA Langley Research Center, and the ISCCP Project. Courtesy of NASA's Earth Observatory.*

The equator always receives plenty of sunlight, however, and has a much higher average temperature as a consequence; compare the average temperature of the equator with that of the poles in Figure **7.2.5**.

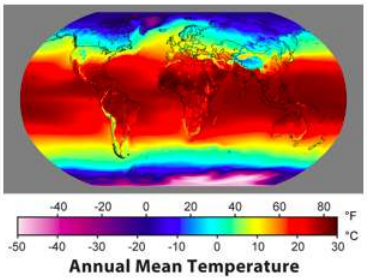


Figure **7.2.5** The Earth's average annual temperature. *Source: Robert A. Rohde for Global Warming Art.*

The level of insolation affecting Earth depends on the amount of light (or solar radiation) emitted by the Sun. Over the current geologic period, this is very slowly changing—solar radiation is increasing at a rate of around 10% every billion years. This change is much too slow to be noticeable to humans. The sun also goes through an 11-year solar cycle, in which the amount of solar radiation increases and decreases. At the solar cycle peak, the total solar radiation is about 0.1% higher than it is at the trough.

The Earth's orbit is not perfectly circular, so sometimes the Earth is closer to or further from the Sun than it is on average. This also changes the amount of insolation, as the closer the Earth is to the Sun the more concentrated the solar radiation. As we shall see in the next section, these orbital variations have made a big difference in conditions on the Earth during the period in which humans have inhabited it.

In addition to considering how much energy enters the Earth system via insolation, we also need to consider how much energy leaves. The climate of the Earth is controlled by the Earth's energy balance, which is the movement of energy into and out of the Earth system. Energy flows into the Earth from the Sun and flows out when it is radiated into space. The Earth's energy balance is determined by the amount of sunlight that shines on the Earth (the insolation) and the characteristics of the Earth's surface and atmosphere that act to reflect, circulate and re-radiate this energy. The more energy in the system the higher the temperature, so either increasing the amount of energy arriving or decreasing the rate at which it leaves would make the climate hotter.

One way to change how quickly energy exits the Earth system is to change the reflectivity of the surface. Compare the difference in dark surface of tilled soil (Figure **7.2.6**a) with the blinding brightness of snow-covered ice (Figure **7.2.6**b).



6a: Tilled soil. *Source:* Tim Hallam.



6b: The snow surface at Dome C Station, Antarctica *Source:* Stephen Hudson

Figure **7.2.6** The reflectivity of the surface of the earth.

The dark soil is absorbing the sun's rays and in so doing is heating the Earth surface, while the brilliant snow is reflecting the sunlight back into space. Albedo is a measure of how reflective a surface is. The higher the albedo the more reflective the material: a perfectly black surface has zero albedo, while a perfectly white surface has an albedo of 1 - it reflects 100% of the incident light. If a planet has a high albedo, much of the radiation from the Sun is reflected back into space, lowering the average temperature. Today, Earth has an average albedo of just over 30%, but this value depends on how much cloud cover there is and what covers the surface. Covering soil with grass increases the amount of light reflected from 17% to 25%, while adding a layer of fresh snow can increase the amount reflected to over 80%. Figure **7.2.7** is a composite photograph of the Earth with the cloud cover removed. As you can see, forests and oceans are dark (low albedo) while snow and deserts are bright (high albedo).

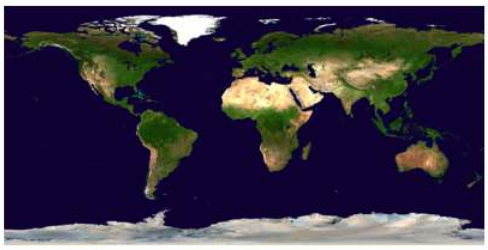


Figure **7.2.7** The surface of the Earth with cloud cover removed. The poles and deserts are much brighter than the oceans and forests. *Source:* *NASA Goddard Space Flight Center Image by Reto Stöckli. Courtesy of NASA's Earth Observatory.*

Changes in albedo can create a positive feedback that reinforces a change in the climate. A positive feedback is a process which amplifies the effect of an initial change. If the climate cools, (the initial change), snow covers more of the surface of the land, and sea-ice covers more of the oceans. Because snow has a higher albedo than bare ground, and ice has a higher albedo than water, this initial cooling increases the amount of sunlight that is reflected back into space, cooling the Earth further (the amplification, or positive feedback). Compare the brightness of Figure **7.2.7** with a similar photo montage from February (Figure **7.2.7**): the extra snow has increased the Earth's albedo. Imagine what would happen if the Earth produced even more snow and ice as a result of this further cooling. The Earth would then reflect more sunlight into space, cooling the planet further and producing yet more snow. If such a loop continued for long enough, this process could result in the entire Earth being covered in ice! Such a feedback loop is

known as the Snowball Earth hypothesis, and scientists have found much supporting geological evidence. The most recent period in Earth's history when this could have occurred was around 650 Million years ago. Positive feedbacks are often described as "runaway" processes; once they are begun they continue without stopping.



Figure 7.2.8 This image shows the surface of the Earth in February (the Northern Hemisphere winter) with cloud cover removed. The seasonal snow cover is brighter (and so has a higher albedo) than the land surface it covers. *Source: NASA Goddard Space Flight Center Image by Reto Stöckli. Courtesy of NASA's Earth Observatory*

Albedo does not explain everything, however. The Earth and the Moon both receive the same amount of insolation. Although the Moon is only slightly more reflective than the Earth, it is much colder. The average temperature on Earth is 15 ^{o}C, while the Moon's average temperature is -23 ^{o}C. Why the difference? A planet's energy balance is also regulated by its atmosphere. A thick atmosphere can act to trap the energy from sunlight, preventing it from escaping directly into space. Earth has an atmosphere while the Moon does not. If the Earth did not have an atmosphere, it would have an average temperature of -18 ^{o}C; slightly warmer than the Moon since it has a lower albedo.

How does the atmosphere trap the energy from the Sun? Shouldn't the Earth's atmosphere reflect as much incoming radiation as it traps? It is true the atmosphere reflects incoming solar radiation—in fact, only around half the insolation that strikes the top of the atmosphere reaches the Earth's surface. The reason an atmosphere generally acts to warm a planet is that the nature of light radiation changes as it reaches the planet's surface. Atmospheres trap more light than they reflect.

Humans see the Earth's atmosphere as largely transparent; that is, we can see a long way in air. This is because we see light in the visible spectrum, which is the light radiation in the range of wavelengths the human eye is able to perceive, and visible light is able to travel a long way through the Earth's atmosphere before it is absorbed. Light is also transmitted in wavelengths we can't see, such as in the infrared spectrum, which is sometimes referred to as infrared light, heat, or thermal radiation. Compared to visible light, infrared light cannot travel very far in the Earth's atmosphere before it is absorbed. Solar radiation striking the Earth is largely in the visible part of the spectrum. The surface of the Earth absorbs this energy and re-radiates it largely in the infrared part of the spectrum. This means that solar radiation enters the Earth in the form of visible light, unhindered, but tries to leave in the form of infrared light, which is trapped. Thicker atmospheres keep this infrared radiation trapped for longer, and so warm the Earth—just like an extra blanket makes you warmer in bed.

This effect is shown in Figure 7.2.9. The visible light radiation enters the atmosphere, and quickly exits as infrared radiation if there is no atmosphere (top Earth). With our atmosphere (the middle Earth), visible light enters unhindered but the infrared light is partially reflected back to the surface, increasing the amount of energy and thus the temperature at the Earth's surface. If the atmosphere is made thicker (bottom Earth) the infrared radiation is trapped for longer, further warming the planet's surface.

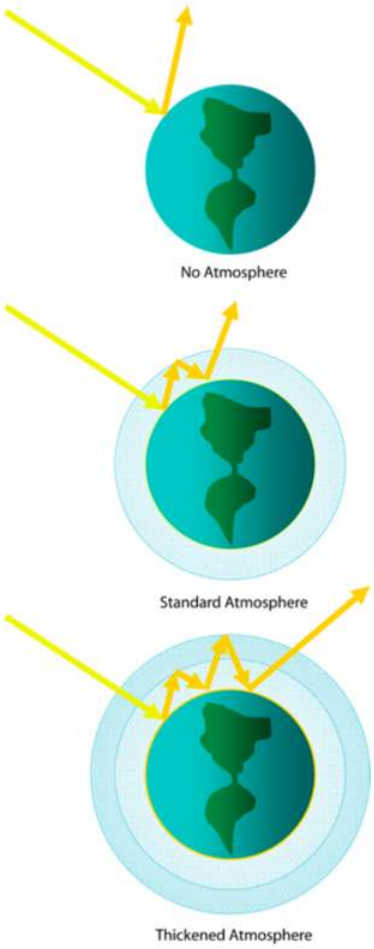


Figure **7.2.9** A cartoon of the greenhouse effect. (Top) Visible light radiation emitted by the sun (yellow arrows) strikes the Earth and reflects as infrared radiation (orange arrow); (middle) an atmosphere reflects some of the infrared radiation back toward the planet; (bottom) a thickened atmosphere reflects greater amounts of infrared radiation. *Source:* *Jonathan H. Tomkin.*

Earth's Changing Atmosphere

The composition of Earth's atmosphere has changed over geologic time. The atmosphere has largely come from volcanic venting of gas from Earth's interior (see Figure **7.2.10**), but biology has also made important changes by producing oxygen and removing carbon dioxide. Greenhouse gases currently make up only a small fraction of the Earth's atmosphere—99% of air consists of nitrogen and oxygen molecules.



Figure **7.2.10** The Mt. Bromo volcano in Indonesia emitting gas into the atmosphere. *Source:* *Jan-Pieter Nap, taken on July 11, 2004.*

While volcanoes can warm the Earth by adding carbon dioxide to the atmosphere, which produces a greenhouse effect, they can also cool the Earth by injecting ash and sulfur into the atmosphere. These additions raise the albedo of the atmosphere, allowing less sunlight to reach the surface of the Earth. The effect lasts until the particles settle out of the atmosphere, typically within a few years. Volcanic eruptions have impacted human societies throughout history; the Mt. Tambora eruption in 1815 cooled the Earth so much that snow fell during June in New England, and the more recent Mt. Pinatubo eruption in 1991 (see Figure **7.2.11**) ejected so much sulfuric acid into the atmosphere that global temperatures were lowered by about 0.5 ^{o}C in the following year.



Figure **7.2.11** The 1991 eruption of Mt. Pinatubo. *Source: U.S. Geological Survey photograph, by Richard P. Hoblitt.*

Evidence from the geologic past indicates that similar events have caused mass extinctions wherein a significant fraction of all species on Earth were wiped out in a relatively short amount of time. Sustained outgassing from continuous volcanic eruptions is thought to have produced so much ash and aerosols that light sufficient to support photosynthesis in plants was unable to penetrate the atmosphere, causing the food chain to collapse. The ash particles produced by extended eruptions would also have increased the Earth's albedo, making conditions inhospitably cool for plants and animals adapted to a warmer environment.

Asteroid impacts can also cause the climate to suddenly cool. When large asteroids strike the Earth, ash is ejected into the atmosphere, which increases albedo in the same way as volcanic eruptions. Everyday clouds (made up of water droplets) both cool and warm the Earth. They can cool the Earth by increasing the planet's albedo, reflecting sunlight into space before it reaches the surface. Clouds can also warm the Earth, by reflecting infrared radiation emitted by the surface back towards the planet. Different types of clouds, and different conditions, determine which effect predominates. On a hot summer's day, for example, clouds cool us by shielding us from the sun's rays, but on a winter's night a layer of cloud can act as a warming blanket.

The composition of the Earth's atmosphere is not fixed; greenhouse gases can be added to and removed from the atmosphere over time. For example, carbon dioxide is added by volcanoes and the decay or burning of organic matter. It is removed by photosynthesis in plants, when it is dissolved in the oceans and when carbonate sediments (a type of rock) are produced. Over geologic time, these processes have significantly reduced the proportion of carbon dioxide in the atmosphere. Atmospheric carbon dioxide levels just prior to the industrial revolution are thought to have been only one twentieth of those of 500 million years ago. Natural processes also remove carbon dioxide added by human activity, but only very slowly. It is estimated that it would take the Earth around a thousand years to naturally remove most of the carbon dioxide released by the industrial consumption of fossil fuels up to the present.

Greenhouse gases other than carbon dioxide are shorter-lived: methane is removed from the atmosphere in around a decade, and chlorofluorocarbons break down within a century. Individual water molecules spend only a few days at a time in the atmosphere, but unlike the other greenhouse gases, the total amount of water vapor in the atmosphere remains constant. Water evaporated from the oceans replaces water lost by condensation and precipitation.

Changing the composition of the Earth's atmosphere also changes the climate. Do you remember the snowball earth? — how increasing ice cover also increased the Earth's albedo, eventually covering the entire planet in ice and snow? Today's climate is temperate—so we must have escaped this frozen trap. But how? The leading hypothesis is that the composition of the Earth's

atmosphere changed, with volcanoes slowly adding more and more carbon dioxide to it. Without access to the oceans, plants, or surface rocks, this carbon dioxide was not removed from the atmosphere and so continued to build up over millions of years. Eventually, the additional warming caused by the increase in greenhouse gases overcame the cooling caused by the snow's high albedo, and temperatures rose enough to melt the ice, freeing the Earth.

For most of Earth's history, carbon dioxide concentrations have been higher than they are today. As a consequence, past climates have often been very warm. During the late stage of the dinosaur era (the cretaceous era, a period that lasted between 65 and 145 million years ago), carbon dioxide levels were about 5 times higher than they are today, and the average global temperatures were more than 10 ^{o}C higher than today's. There were no large ice sheets, and dinosaur fossils from this period have been found as far north as Alaska. These animals would not survive the cold conditions found in the arctic today. Further south, fossil crocodiles from 60 million years ago have been found in North Dakota. The modern average winter temperature in North Dakota is around -10 ^{o}C –but being cold-blooded, crocodiles are most at home when the air temperature is around 30 ^{o}C! The climate was warmer in the past when the amount of carbon dioxide was higher.

Review Questions

1. The text describes how the high albedo of snow acts as a positive feedback—if the Earth is made cooler, the highly reflective snow can act to further cool the Earth. Today, part of the Earth is covered with snow and ice. Can you describe a mechanism by which warmer temperatures would also produce a positive feedback—this time heating the Earth further—through a similar albedo mechanism?
2. Mars is colder than the Earth. Venus, on the other hand, is much hotter, with average surface temperatures of around 450 ^{o}C. Venus is closer to the Sun than the Earth is, and so receives about twice as much solar radiation. Venus's atmosphere is also different than Earth's, as it is much thicker and mainly consists of carbon dioxide. Using the terms insolation and greenhouse gasses, can you suggest reasons why Venus is so hot?
3. Oxygen makes up over 20% of Earth's atmosphere, while carbon dioxide makes up less than 0.04%. Oxygen is largely transparent to both visible and infrared light. Explain why carbon dioxide is a more important greenhouse gas in the Earth's atmosphere than oxygen, even though there is much more oxygen than carbon dioxide.
4. Figure **7.2.4** shows the insolation at the surface of the Earth. The Earth is spherical, so we would expect the values to be the same for places of the same latitude. But notice that this is not true – compare, for example, central Africa with the Atlantic Ocean at the same latitude. What feature of the atmosphere might explain this variation, and why?

Resources

The National Aeronautical and Space Administration (NASA) Earth Observatory website has an array of climate resources. For a more in-depth discussion of Earth's energy budget, go to http://earthobservatory.nasa.gov/Features/EnergyBalance/

Are you interested in finding more about the controversial Snowball Earth hypothesis? The National Science foundation and Harvard University have set up a website with more about the hypothesis and the evidence. Go to http://www.snowballearth.org/

Glossary

albedo

A measure of how reflective a surface is. A perfectly black surface has an albedo of 0, while a perfectly white surface has an albedo of 1.

climate

The average of the weather.

cretaceous period

The period between 65 and 145 million years ago, which was the final period of Earth's history that included dinosaurs.

greenhouse effect

The process by which the atmosphere acts to trap heat, warming the climate.

greenhouse gases

Those gases in the atmosphere that warm the climate, most importantly, water vapor, carbon dioxide, methane, and ozone.

infrared spectrum

The light radiation just below the range of wavelengths visible to the human eye. Also referred to as thermal radiation.

insolation

The measure of the amount of solar radiation falling on a surface.

positive feedback

A runaway process which amplifies the effect of an initial change.

snowball earth

A condition in which the entire planet is covered in ice, last thought to have happened 650 million years ago.

solar radiation

The energy emitted by the sun in the form of light.

visible spectrum

The light radiation that is in the range of wavelengths that is visible to the human eye.

weather

A description of the short term state of the atmosphere.

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7.3: The Carbon Cycle

Carbon, just like all other elements, cycles through the environment and is constantly in the process of changing forms and locations. In this section, as in many other pieces of scientific literature, we will periodically refer to carbon by its chemical symbol, C. There is no new carbon in the world, rather, all carbon is continuously recycled from one form to another. All plants, animals (including humans!), fungi, bacteria, and archaea are made of mostly carbon-based molecules such as lipids, carbohydrates, proteins, and nucleic acids. Carbon is also prevalent in soils, rocks and sediments, water bodies (dissolved), and the atmosphere. These locations where carbon resides are known as pools or reservoirs, and the processes that move carbon from one location to another are called fluxes. Figure 7.1.1 shows a simplified version of the global carbon cycle.

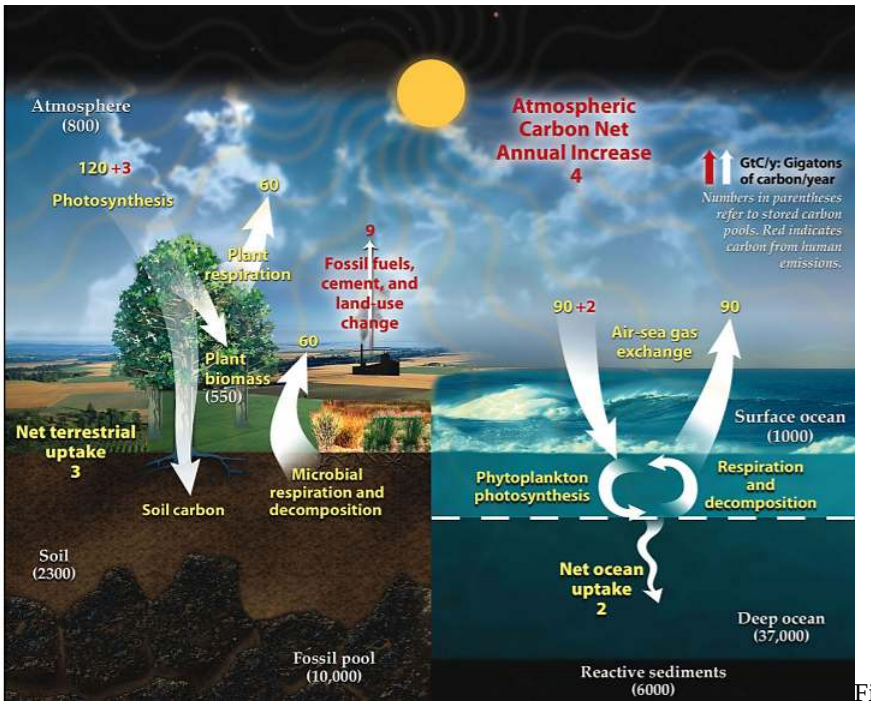


Figure **7.3.1**: A simplified carbon cycle. Diagram adapted from U.S. DOE, Biological and Environmental Research Information System.

Some reservoirs hold on to carbon for only a short time. **Aerobic** (oxygen-using) organisms convert carbohydrates created by other organisms into **carbon dioxide** (CO_2) almost instantaneously, which they exhale into the atmosphere. When considering the flux of respiration, living organisms are the **source** of carbon, and the atmosphere is the **sink**. The carbon stays in the reservoir of living organisms for a relatively short time, depending on their life span, from hours and days to years and decades. In contrast, the **residence time** of carbon in the fossil pool is dramatically different. Fossil fuels form over a course of 300-400 million years, forming from ancient plants and animals that decomposed slowly under very specific, **anaerobic** (without oxygen) conditions in wetland environments. Their bodies were gradually transformed by the heat and pressure of the Earth's crust into the fossil fuels that we mine today to provide petroleum oil, natural gas, and coal (see more on this in chapter 4).

Reservoirs and fluxes of importance

The two largest reservoirs of carbon on Earth are the oceans, which cover the majority of Earth's surface, and the **lithosphere** (the mineral fraction of Earth: soils, rocks, and sediments). Each of these reservoirs holds more carbon than all of the other reservoirs combined. Much of the carbon stored in these reservoirs, especially deep in the lithosphere or in deep ocean environments, has an extremely long residence time, and does not actively participate in rapid fluxes. The notable exceptions here, of course, are fossil fuels, which are mined by humans and converted into gaseous forms of carbon through combustion.

Biomass, which is biological material derived from living, or recently living organisms, is a much smaller reservoir of carbon. The amount of carbon stored in all of the terrestrial vegetation (550 Gt C) (Gt = gigatonne = 10^9 metric tons = 10^{15} g) is just a fraction of that stored in the oceans (38,000 Gt C) and lithosphere (18,000 Gt C). All of the carbon that is currently stored in all of the vegetation on Earth got there through the process of **photosynthesis**. Plants and other photosynthetic organisms are called **primary**

producers, because they "fix" atmospheric CO_2 into organic carbon, such as sugar, a form that is usable by animals and other organisms that need to consume their carbon molecules.

Photosynthetic organisms, such as plants, algae, and cyanobacteria, bring in CO_2 from the atmosphere and, using energy from the sun, convert CO_2 and water into glucose molecules (organic carbon). The products of photosynthesis are oxygen and glucose (Equation **7.3.1**). These glucose molecules are simple sugars that **autotrophs** ("self-feeders") can "burn" for energy, or transform into other usable carbon molecules through the process of cellular respiration (described in the next paragraph), or to build plant biomass. Photosynthesis takes place in organelles called **chloroplasts**, shown in Figure **7.3.2**. Photosynthesis accounts for 123 Gt of C per year that is removed from the atmosphere and stored in plant biomass. Such a massive amount of photosynthesis occurs on Earth that no other single flux moves as much carbon in the same timeframe.

$$6CO_2 + 6H_2O + \text{solar energy} \rightarrow C_6H_{12}O_6 + 6O_2 \tag{7.3.1}$$

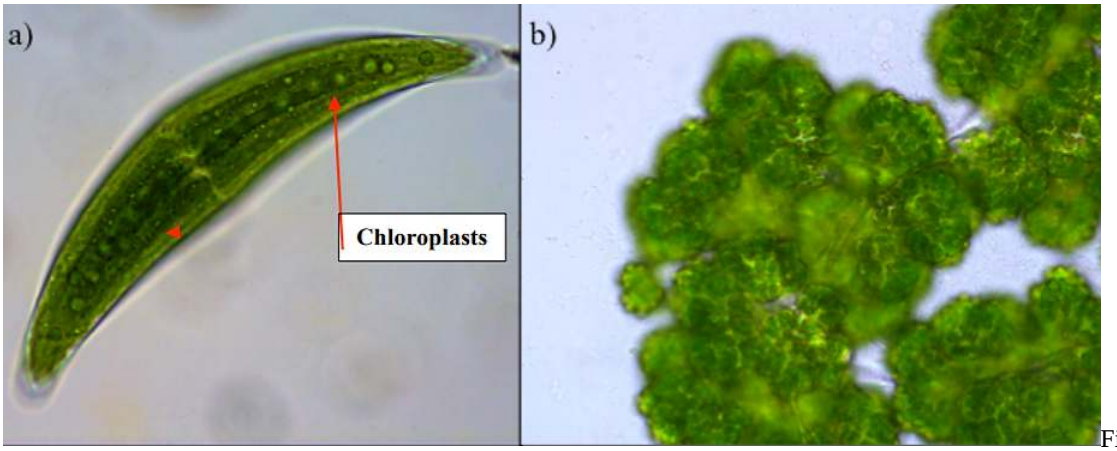


Figure **7.3.2**: Chloroplasts visible in freshwater algae. Chloroplasts are green in color due to the chlorophyll a they contain, and are the site of photosynthesis. Chlorophyll a is the green pigment that allows plants, algae, and cyanobacteria to absorb the energy they need for photosynthesis from sunlight. a) Closterium moniliferum Ralfs, (Chlorophyta) green coccoid algae; b) Botryococcus braunii Kützing, (Chlorophyta) green coccoid algae with discoid chloroplasts. Image credit: K. Manoylov, Lake Sinclair, GA

Biomass in the carbon cycle, including plants and animals, is the reservoir of carbon that we are most likely most familiar with, and also the reservoir that is most readily available to us. We all participate in the flux of **consumption** of carbon when we eat food. All of our food is simply plant and/or animal biomass. Our body takes the carbon molecules contained in this biomass, and uses them, along with the oxygen we breathe in, for **cellular respiration** to create the adenosine triphosphate (ATP) we need for energy. The products of cellular respiration include the CO_2 we exhale, water, and energy that is stored in ATP (Equation **7.3.2**). Our bodies also builds additional biomass out of the carbon molecules in this food, allowing us to create new cells for growth or replenishment. This is the only way we, and all other **heterotrophs** ("other-eaters"), can bring in the carbon we need to build and maintain our bodies. Remember, you are what you eat!

$$C_6H_{12}O_6 + CO_2 \rightarrow 6CO_2 + 6H_2O + \text{energy} \tag{7.3.2}$$

Cellular respiration is an important flux in the carbon cycle, and one that contributes carbon to the atmosphere. Remember that animals and other heterotrophs complete cellular respiration using the carbon molecules that they bring in through their food. Plants and other photosynthetic autotrophs complete cellular respiration using the carbon molecules they formed from CO_2 through photosynthesis. Any carbon molecules that are left over after the organism has acquired sufficient energy through cellular respiration make up the biomass of the plant. As plants and animals die and decompose, their bodies are consumed by decomposer organisms such as fungi and bacteria. Through the flux of **decomposition**, some decaying biomass is converted into atmospheric carbon by the decomposers, while most of the biomass is buried into the soil, contributing to soil carbon. In oxygen-rich environments, decomposers rapidly consume dead and decaying biomass using the same process of aerobic cellular respiration described above. In oxygen-deficient environments, decomposers complete other metabolic pathways, and very slowly consume the organic matter. Some of the gases produced from anaerobic decomposition include **methane** (CH_4), **nitrous oxide** (N_2O), and the foul-smelling hydrogen sulfide (H_2S).

The biomass reservoir of the carbon cycle is also important to us as a source of energy. Through the flux of **combustion**, we convert the **potential energy** held in biomass into heat energy that we can use, and release carbon dioxide in the process. If you have ever burned logs on a campfire, or even burned food on the stove, you have completed this flux of biomass combustion. Of course, this happens naturally as well, the best example being natural forest fires caused by lightning strikes. The chemical reaction for combustion is identical to the chemical reaction for cellular respiration. The difference is that in cellular respiration, energy is released in a controlled fashion, and captured in ATP molecules. In combustion, all of this energy is released rapidly in the form of light and heat.

As all of the fluxes we've discussed so far involve the atmosphere, we have not yet discussed the flux that connects the atmosphere to the oceans. Carbon can enter the oceans through two primary fluxes: first through photosynthesis by algae or cyanobacteria (also called phytoplankton in Figure **7.3.1**), and second through the chemical reaction of **ocean-atmosphere exchange**. The ocean, as with all surface water bodies, always contains some dissolved CO_2. This CO_2 is in **equilibrium** with the CO_2 in the air. Some atmospheric CO_2 is constantly dissolving into the ocean, while some dissolved CO_2 is constantly diffusing into the atmosphere. Under normal conditions, these two fluxes will be happening at equal rates. As you can see in Figure **7.3.1**, however, this is no longer the case. In the section Human impacts on the carbon cycle, we will discuss why this is the case.

Activity: Better understanding the carbon cycle

To further review the carbon cycle, and better understand the human impacts on it, use this interactive graphic from Woods Hole laboratories: http://www.whoi.edu/feature/carboncycle/. As you will see, the information described in this text is only a small portion of the total carbon cycle on Earth. Finally, complete Table **7.3.1** as a way to review the sink/source relationship within this cycle. See if you can correctly identify the source and sink of carbon for each of these important fluxes in the carbon cycle.

Table **7.3.1**: Practice understanding the sink/source relationship with cycles

Carbon flux	Carbon source	Carbon sink
Cellular respiration	Carbohydrates in living organisms	CO_2 in the atmosphere
Photosynthesis		
Consumption		
Combustion		
Decomposition		
Ocean/atmosphere exchange		
Fossil fuel formation		

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7.4: Human Impacts on the Carbon Cycle

Humans, just like all other living organisms, have impacted the global carbon cycle since the dawn of our species. However, the magnitude of our impacts has changed dramatically throughout history. The **Industrial Revolution**, which occurred around the turn of the 19th century, began to make major changes in the use of resources around the world. Beginning in Britain, industrialization eventually affected the whole world. The development of coal-fueled steam power, and later transportation following the discovery of large oil deposits, had enormous influence on the economic and social structure of the world. As the world accelerated in the production and transportation of manufactured goods, the production and consumption of fossil fuels grew. As economic growth continued to increase, so did the production of carbon dioxide through fossil fuel combustion. See Figure **7.4.2** later in this text.

Some of the human impacts on the carbon cycle have been quantified for you in Figure 7.1.1. Changes to fluxes in the carbon cycle that humans are responsible for include: increased contribution of CO_2 and other **greenhouse gases** to the atmosphere through the combustion of fossil fuels and biomass; increased contribution of CO_2 to the atmosphere due to land-use changes; increased CO_2 dissolving into the ocean through ocean-atmosphere exchange; and increased terrestrial photosynthesis. The first two impacts, both contributing excess CO_2 to the atmosphere at a rate of 4 Gt of carbon per year have, by far, the largest impact on our planet. For this reason, this is the change that we will most often focus on throughout this section. The excess CO_2 in the atmosphere is responsible for the increased CO_2 dissolving into the ocean, which we will discuss later in this section. This is also, in part, responsible for the increased terrestrial photosynthesis that can be observed, as additional CO_2 is available to plants for photosynthesis. However, intensive agricultural and forestry practices also contribute to the change in this flux.

One characteristic example of a human impact on the carbon cycle is illustrated in Figure **7.4.1**. Throughout most of our recent human history, people have been physically altering the landscape around them in order to have more control over their surroundings and increase their odds of survival. One way that people have done this is through agriculture. In order for most forms of agriculture to be successful, native vegetation is eliminated or minimized. Resources from this native vegetation, such as wood, may be used for combustion to provide heat, sanitation, or fuel for cooking. Combustion may also be used as an efficient way to clear the land and make way for crops or grazing lands for livestock. Often, settlements are formed around these newly fashioned agricultural fields, and the land is used in a similar fashion for many years in the future.

Let's identify the ways in which humans are impacting the carbon cycle in this scenario of agricultural establishment. You should be able to identify from the above paragraph that the flux of combustion will release CO_2 previously held in vegetation into the atmosphere. In addition, remember that the land that used to house native vegetation is now home to agricultural lands. In most controlled agricultural environments, there is less total vegetative biomass than there would be under natural conditions. This decreased biomass leads to lower total photosynthesis rates, thereby decreasing the amount of CO_2 that is removed from the atmosphere and turned into plant biomass. Also, open soil on the fields between crops, during the winter months, or as a result of overgrazing allows for the air to penetrate deep into the soil structure. This provides the environment necessary for enhanced aerobic respiration by soil microorganisms. This decreases soil carbon, which can lead to erosion and soil degradation, and also releases additional CO_2 to the atmosphere.

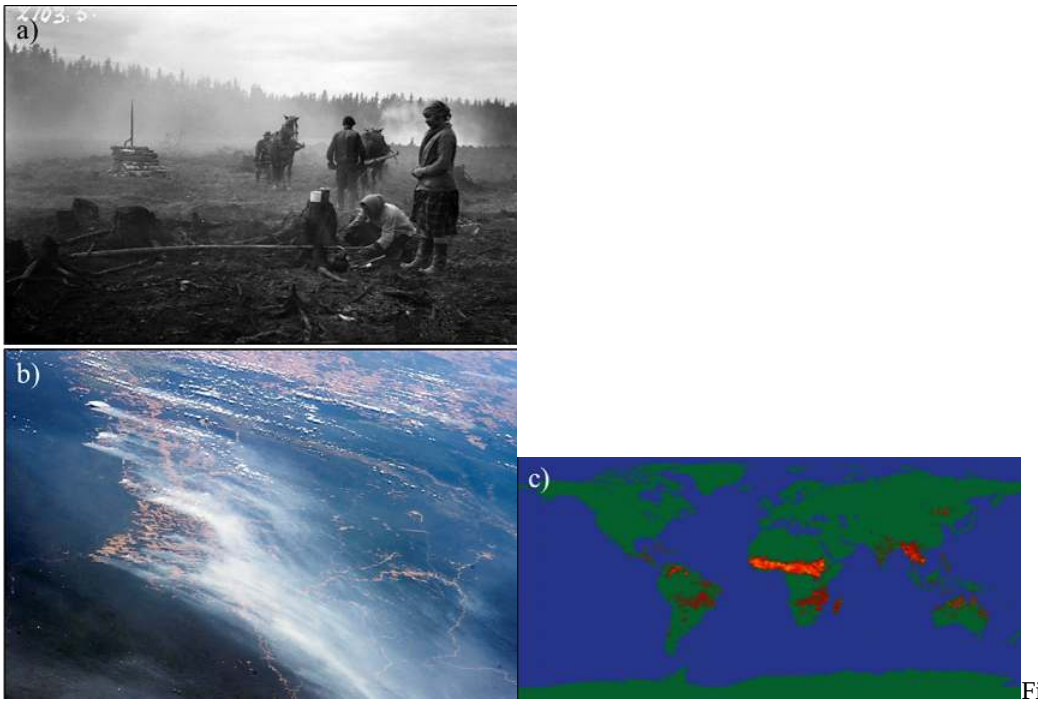


Figure 7.4.1: Impacts of slash and burn agriculture and biomass burning. a) Slash and burn agriculture in Maaninka, Finland. This kind of agriculture was still in use in Finland at the end of the 1920s. b) Slash and burn agriculture at the margins of the Amazon Rainforest in South America captured by astronauts on the International Space Station in August, 2014. Credit: Astronaut photograph ISS040-E-103496. c) Global distribution of biomass fires, represented by red, orange, and yellow dots (lighter colors indicates more fires), based on nighttime measurements obtained by the DMSP Operational Linescan System. Credit: Julia Cole, NASA Earth Observatory

As you learned in Chapter 5, biomass is an important form of energy to human civilization. Prior to the Industrial Revolution, this was essentially the only form of fuel to which most people on Earth had access. In many **less-industrialized countries**, combustion of biomass such as wood or animal dung is still the primary energy source that many citizens, particularly in rural areas, depend on for domestic use (heating, sanitation, and cooking) as it is inexpensive, relatively efficient, and readily available. Figure 7.4.1c shows the global distribution of biomass fires in the world. While the burning of biomass for domestic use contributes to some of these fires, it is the so-called **slash-and-burn agriculture** that makes up a larger contribution. Take a minute to compare the areas highlighted in Figure 7.4.1c to the countries of the world that are currently experiencing rapid population growth (Chapter 3). If you need a refresher, use the CIA World Factbook website to view current global population growth values by country: www.cia.gov/library/publicat.../2002rank.html.

While biomass burning still has a significant impact on the global carbon cycle, human impacts on fluxes such as fossil fuel extraction and combustion continue to grow. For a review of the impacts of non-renewable energy sources such as fossil fuels, see Chapter 4. Burning of any fossil fuel (coal, natural gas, crude oil) moves carbon from a previously-sequestered state deep within the Earth's crust into carbon dioxide in the atmosphere. As countries become more industrialized, their reliance on and combustion of fossil fuels tends to increase. Look at the graph in Figure 7.4.2, which compares CO_2 emissions from fossil fuels of regions across the globe.

How has the use and distribution of fossil fuels changed throughout the past 250 years?

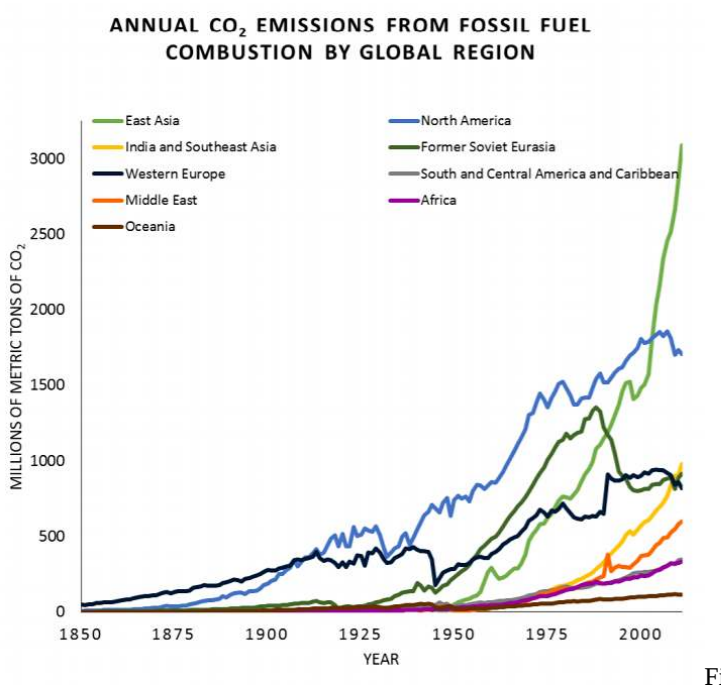


Figure **7.4.2**: Total annual CO_2 emissions from fossil fuel combustion 1850-2011 by global region. Data from the Carbon Dioxide Information Analysis Center (CDIAC) cdiac.esd.ornl.gov/

The data shown in Figure **7.4.2** reveals much about the regions of the world it depicts. The effects of historic events such as the Great Depression of 1929-1939, World Wars, the fall of the Soviet Union in 1991, and the Kuwait oil fires of 1991 can be seen. Furthermore, between 1850 and 2011, different regions have gone in and out of the lead position as top producer of CO_2 from fossil fuel emissions. Population is one reason why fossil fuel use has changed throughout time. This is particularly apparent when comparing the data for Western Europe to that of India and Southeast Asia.

As countries industrialize, their relationship with agriculture also changes. **More-industrialized countries** rely very little on slash-and-burn agriculture. Their agricultural practices, however, are no less impactful on the environment. The growing population (Chapter 3) in many countries has required agriculture to become industrialized in order to meet demand. As a person living in the United States, **industrialized agriculture** probably produces the vast majority of the food you eat, including grains, fruits and vegetables, dairy and eggs, meats, and even fish. Industrialized agriculture can refer to a variety of practices, but has several main components: the use of motorized machinery; the use of chemicals such as fertilizers, pesticides, hormones, and/or antibiotics; and the intense and efficient production of one product across a large area of land.

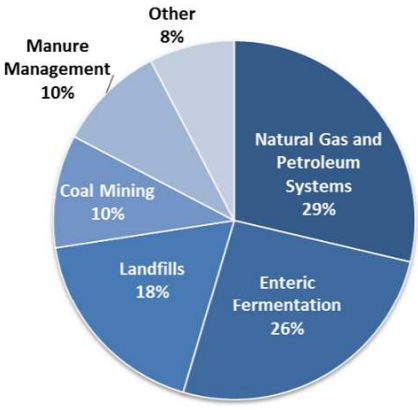


Figure **7.4.3**: US methane emissions by source. 36% of the US methane production comes from agriculture: enteric fermentation (production of methane by anaerobic bacteria within the ruminant gut) and manure management. All emissions estimates from the Inventory of US Greenhouse Gas Emissions and Sinks: 1990-2013. US EPA

One example of the impacts of industrialized agriculture is the production of methane (CH_4), a potent greenhouse gas. You will learn more methaneabout methane later in this section. As you saw earlier, methane is a common product of anaerobic metabolisms. The gut of **ruminant animals** (such as sheep, cattle, and goats) has evolved to allow the animals to digest the very

tough carbon molecules, such as cellulose, in grass. They do this through symbiosis, or cooperation, with anaerobic bacteria who live in the gut tract. These anaerobic bacteria produce methane and other gases as a result of their metabolism when they break down molecules like cellulose. This is sometimes called enteric fermentation. The methane gas is excreted from the animal, and this contributes significantly to total methane emissions (Figure **7.4.3**). A similar type of bacteria live in the fecal matter, or manure, of livestock. As the manure is handled or stored for future use, methane is also released to the environment.

The methane excretions of one cow or a few sheep would be miniscule and insignificant. If you were a small farmer with only enough livestock to feed your family, your contribution to total methane emissions would be close to zero. However, the demand for animal protein from meat, dairy, and eggs is very large in the United States. As of January 2015, the United States had a total cattle inventory of 89.9 million animals, and in 2014, 25.5 billion pounds of beef was consumed in the United States (statistics: National Cattlemen's Beef Association). The impacts of enteric fermentation and manure management for almost 90 million animals are very significant, as seen in Figure **7.4.3**. In both cases, carbon that was previously stored in biomass (cattle feed) is moved into the atmosphere, this time in the form of CH_4. This is another example of how humans have impacted the carbon cycle.

Previously in this chapter, you identified other ways the carbon cycle is impacted by human agriculture. Through industrialized agriculture, we must also account for the fossil fuels that are used. In order to deliver agricultural products to consumers, fossil fuels are used numerous times: deliveries of fertilizer, feed, and/or seed to farms; farm machinery; delivery of products to processers; food processing; delivery of foods to supermarkets; etc.

As animal products, especially meat, are expensive, the demand is typically greater in more industrialized countries than it is in less-industrialized countries. This makes industrialized agriculture, and especially industrialized animal agriculture, one of the major contributors to greenhouse gas emissions in more-industrialized countries.

Knowledge check – answer these questions on your own to further explore the impacts of biomass and fossil fuel burning on the global carbon cycle.

1. Why is there a correlation between population growth rate and global distribution of biomass fires?

2. Do you think this correlation is more likely due to personal biomass fires for activities such as cooking, or due to slash-and-burn agriculture? Why?

3. Given any other knowledge you might have about the areas highlighted in Figure **7.4.1**c, what other environmental impacts may be occurring here besides carbon cycle alterations?

4. Compare the production of CO_2 emissions from fossil fuel combustion across world regions in 1900, 1950, and 2011 in Figure **7.4.2**. What has accounted for these differences?

5. Has the total worldwide production of CO_2 from fossil fuels increased evenly relative to human population growth during the time period displayed in Figure **7.4.2**? Why or why not?

6. What are the differences in contributions of greenhouse gas emissions from more industrialized countries and less-industrialized countries? What are the similarities?

Resources

Carbon Dioxide Information Analysis Center cdiac.esd.ornl.gov/

NOAA Earth System Research Laboratory: Carbon Cycle Science www.esrl.noaa.gov/research/themes/carbon/

Sass, Ronald. Q2: What are the Causes of Global Climate Change? OpenStax CNX. Sep 22, 2009 http://cnx.org/contents/5d263a29-7bd...c233619bca33@3

USDA Climate Change and Agriculture in the United States: Effects and Adaptation www.usda.gov/oce/climate_chan...griculture.htm

US EPA Overview of Greenhouse Gases: Methane http://epa.gov/climatechange/ghgemis...gases/ch4.html

Woods Hole Oceanographic Institution: Carbon Around the Earth http://www.whoi.edu/feature/carboncycle/

Terms list

Aerobic

Less-industrialized country

Anaerobic

Lithosphere

Autotroph

Methane

Biomass

More-industrialized country

Carbon

Nitrous oxide

Carbon dioxide

Ocean-atmosphere

Cellular respiration exchange

Chloroplast

Photosynthesis

Combustion

Potential energy

Consumption

Primary producer

Decomposition

Reservoir

Equilibrium

Residence time

Flux

Ruminant animal

Greenhouse gas

Sink

Heterotroph

Slash-and-burn agriculture

Industrial Revolution

Source

Industrialized agriculture

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article=1003&context=biology-textbooks.

7.5: The Science of Climate Change

What is causing global climate change?

Scientists have identified the source of our current global climate change as being the increased human-caused emissions of greenhouse gases such as carbon dioxide (CO_2), methane (CH_4), and nitrous oxide (N_2O), since the industrial revolution. Greenhouse gases are defined as large (at least three atoms) gas molecules that participate in the **greenhouse effect**. While you already know about the "big three" greenhouse gases (CO_2, CH_4, and N_2O), 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, is it still an essential component of the natural greenhouse effect that occurs in our atmosphere.

The Earth receives energy from the sun and in turn radiates energy back into space. When these two energies are equal, a stable temperature of the Earth is achieved. This temperature can be calculated from basic physics and is equal to about -18°C (0°F). This thermal equilibrium temperature is obviously much colder than that of the surface of the Earth. The actual average value of the Earth's surface temperature is about 15°C (59°F). The difference between these temperatures is due primarily to the natural greenhouse gas concentrations in the atmosphere, causing the greenhouse effect. If the Earth had no naturally occurring atmospheric greenhouse gases, the temperature at the surface of the Earth would equal the thermal equilibrium temperature. The influence of these greenhouse gases, mainly water and some CO_2, moderates the Earth's climate and makes life possible (Figure 7.5.1).

As **solar radiation** reaches the Earth's atmosphere, there are a variety of possibilities for its fate. Some solar radiation is reflected by the Earth and its atmosphere, and does not contribute to warming. Some passes through the atmosphere and reaches the surface of the Earth. When this solar radiation is absorbed by objects on Earth's surface, it is re-emitted as infrared radiation (heat) that escapes to space. However, some of this heat is intercepted in the atmosphere by greenhouse gases. These gases absorb and re-emit the radiation in all directions. This creates a warming impact on the Earth's surface Radiation can be bounced around from one greenhouse gas molecule to another, becoming trapped, and increasing its warming potential. For this reason, an increased greenhouse gas concentration causes an increase in the overall warming potential of the Earth's atmosphere.

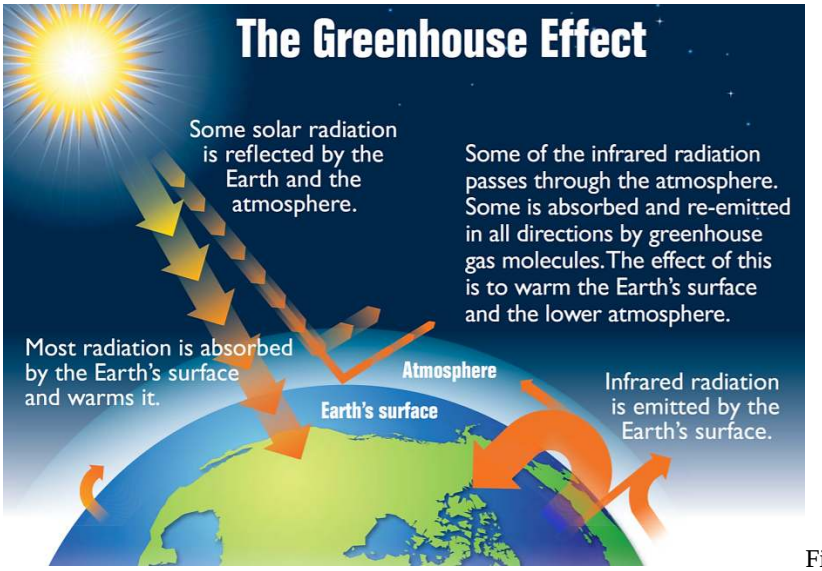


Figure 7.5.1: This diagram shows the Earth's greenhouse effect. The Earth absorbs some of the energy it receives from the sun and radiates the rest back toward space. However, certain gases in the atmosphere, called greenhouse gases, absorb some of the energy radiated from the Earth and trap it in the atmosphere. These gases essentially act as a blanket, making the Earth's surface warmer than it otherwise would be. While this greenhouse effect occurs naturally, making life as we know it possible, human activities in the past century have substantially increased the amount of greenhouse gases in the atmosphere, causing the atmosphere to trap more heat and leading to changes in the Earth's temperature. Credit: US EPA

On a geological time scale, the climate has changed many times in the past, even before the presence of humans. These changes occurred naturally because man had not yet evolved. A well-known example of past climate change is the occurrence of ice ages.

Ice ages have occurred repeatedly throughout Earth's history, the most severe ice age of which scientists have reliable data occurred around 650,000 years ago. During this time, glacial ice covered much of Canada, the northern United States, and northern Europe; the level of the ocean decreased 120 m, and the temperature decreased by 5°C.

A geologic history of ice events is preserved in the ice sheets covering Antarctica and Greenland. This history has been uncovered over the past decades by scientists who have cored deeply into the ice and deciphered the temperature and atmospheric composition records stored in the ice. This process of obtaining **ice cores** is shown in Figure 7.5.2. The temperature at which the ice originally formed can be obtained from an interpretation of the measured ratio of the **stable isotopes** (see Chapter 1 supplement for a description of isotopes) of oxygen in the molecules of water forming the ice. The atmospheric gas composition is taken from air bubbles trapped in the ice at the time of formation. From these data, scientists have gathered a set of reliable data that track atmospheric temperature and gas concentrations that dates back 800,000 years. These data helped scientists come to the conclusion that the Earth's temperature and greenhouse gas concentrations are directly correlated to one another (Figure 7.5.3). During the ice age 650,000 years ago, the Earth was experiencing depressed temperature and atmospheric CO_2 concentrations below 200 **parts per million** (ppm). We can also see from these data, that CO_2 concentrations can be naturally elevated to as high as 300 ppm, correlating with increased temperatures.

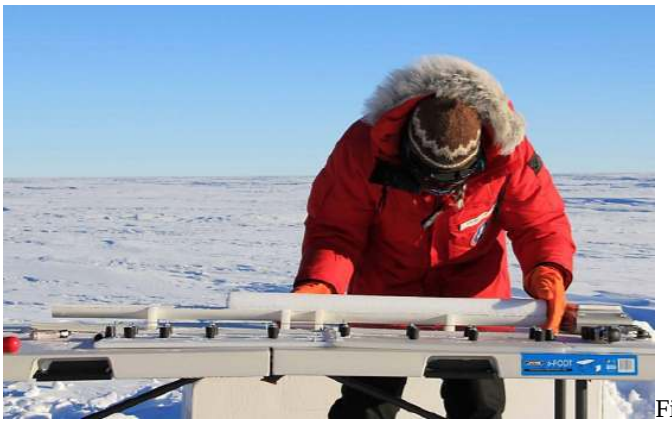


Figure 7.5.2: On Dec. 8, 2010, Michelle Koutnik, of the University of Copenhagen's Center for Ice and Climate, prepared ore of Antarctic ice to be wrapped and put into core tubes for transport back to labs at Brigham Young University in Utah. But first, Koutnik measured the core's length, diameter and weight. The traverse was the first of two field campaigns to study snow accumulation on the West Antarctic Ice Sheet and tie the information back to larger-scale data collected from satellites. 120 m, and the global average Credit: NASA/Lora Koenig.

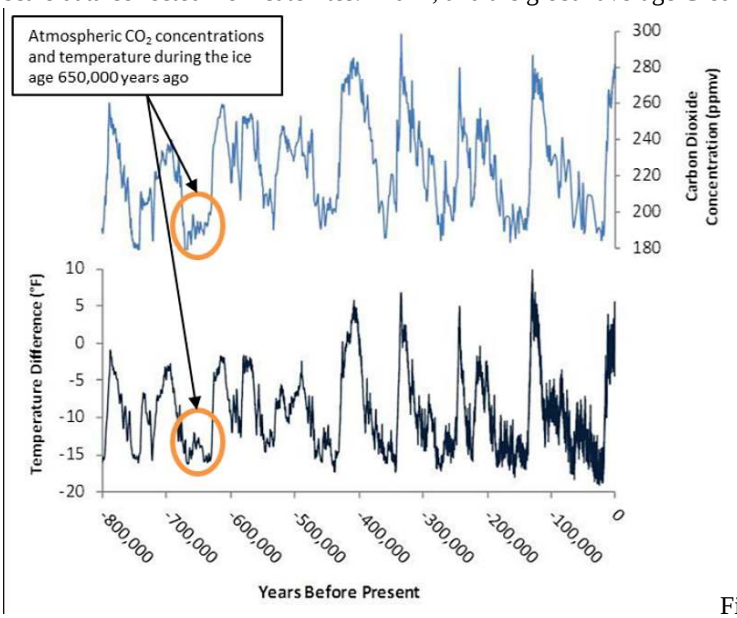


Figure 7.5.3: Estimates of the Earth's changing CO_2 concentration (top) and Antarctic temperature (bottom), based on analysis of ice core data extending back 800,000 years. Until the past century, natural factors caused atmospheric CO_2 concentrations to vary within a range of about 180 to 300 ppm. Warmer

periods coincide with periods of relatively high CO_2 concentrations. NOTE: The past century's temperature changes and rapid CO_2 rise (to 400 ppm in 2015) are not shown here. Source: Based on data appearing in NRC (2010).

The 100,000 year major cycle of the ice ages and some variations within the cycles agree very well with predicted periodic relationships between the Earth's orbit around the sun, generally referred to as the **Milankovitch cycles**. Milankovitch cycles describe the very slight "wobbles" that occur in the Earth's tilt and path as it moves around the sun. The Earth is always slightly tilted on its axis with respect to the sun. The angle of this tilt, however, changes periodically, varying from about 22° to about 25°. A less severe tilt will cause milder summers and winters close to the poles, preventing full summer ice melt in the northern- and southernmost regions, and allowing for a buildup of ice from year to year.

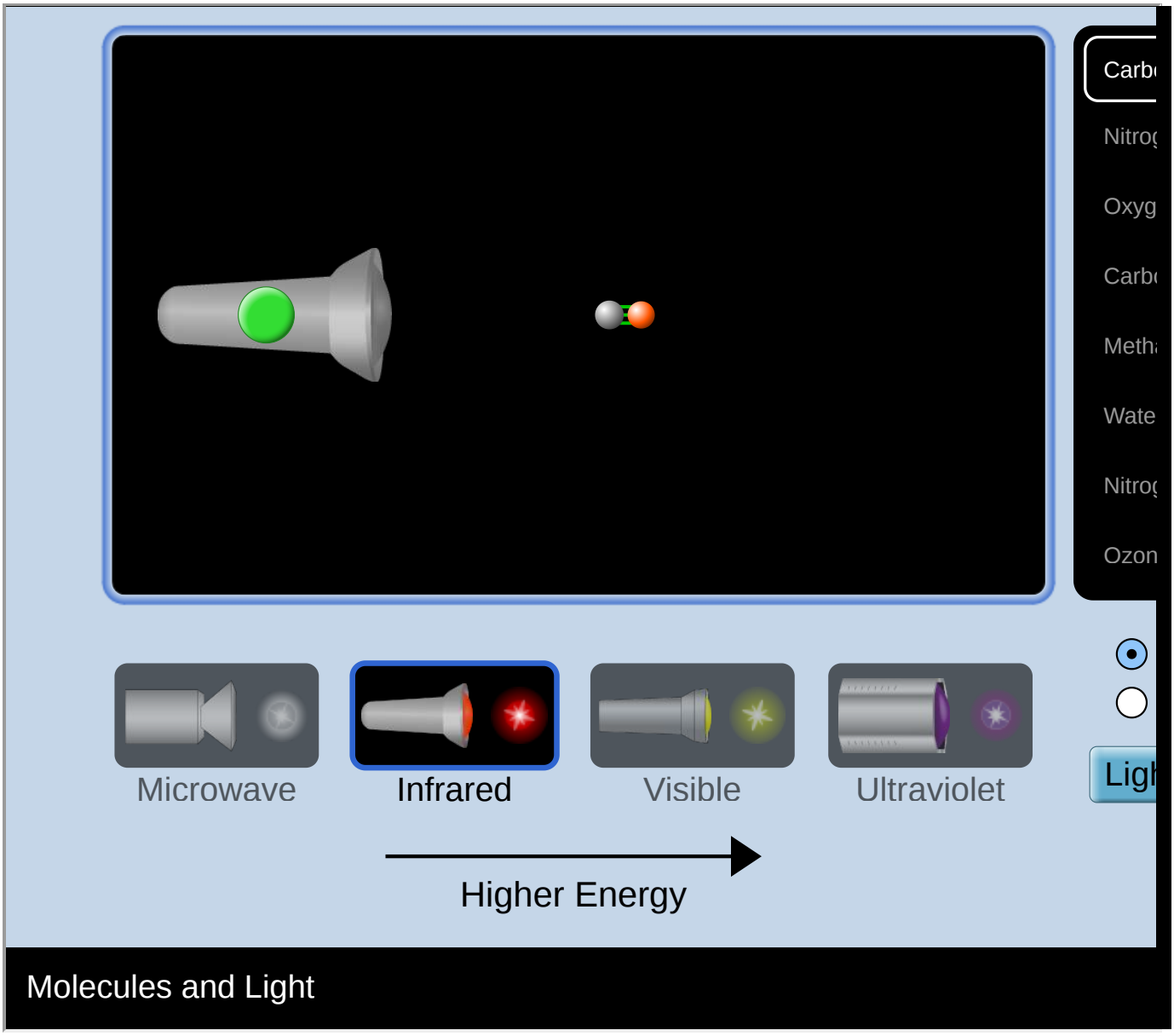
The path through which the Earth travels on its journey around the sun also changes from a more circular to a more elongated shape. Again, a round orbit will cause milder summers and winters close to the poles. These are very long term changes, and the results of the Milankovitch cycles can be observed in the changes in temperature and atmospheric CO_2 concentration shown in Figure **7.5.3**. The climate change event that scientists are currently documenting is occurring much more rapidly than could be explained by Milankovitch cycles. Therefore, scientists agree that the cause of our currently changing climate is due to human impacts and not natural forces.



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7.6: Activity: Molecules and Light



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7.7: Additional Reading - Milankovitch Cycles

Learning Objectives

After reading this module, students should be able to

- describe the changing climate of the Quaternary
- explain why Milankovitch cycles explain the variations of climate over the Quaternary, in terms of the similar periods of orbital variations and glacial cycles
- explain how the glacier/climate system is linked via albedo feedbacks
- describe how sediment and ice cores provide information about past climates
- use the mechanisms that cause stable isotope fractionation to predict the impact of changing climate on stable isotope records

Introduction

In Module 5.2 we saw the major drivers of the climate—the energy that comes from the Sun (insolation) and the properties of the planet that determine how long that energy stays in the Earth system (albedo, greenhouse gases). In this section, we will look at the recent natural changes in Earth's climate, and we will use these drivers to understand why the climate has changed.

The most recent period of Earth's geologic history—spanning the last 2.6 million years—is known as the quaternary period. This is an important period for us because it encompasses the entire period over which humans have existed—our species evolved about 200,000 years ago. We will examine how the climate has changed over this period in detail. By understanding recent natural processes of climate change, we will be able to better understand why scientists attribute the currently observed changes in global climate as being the result of human activity.

Quaternary Climate — Information From Ice Cores

How do we know about the Quaternary climate? After all, most of the period predates human existence, and we have only been recording the conditions of climate for a few centuries. Scientists are able to make informed judgments about the climates of the deep past by using proxy data. Proxy data is information about the climate that accumulates through natural phenomena. In the previous module, for example, we discussed how 60-million-year-old crocodile fossils have been found in North Dakota. This gives us indirect information about the climate of the period—that the climate of the region was warmer than it is today. Although not as precise as climate data recorded by instruments (such as thermometers), proxy data has been recovered from a diverse array of natural sources, and provides a surprisingly precise picture of climate change through deep time.

One highly detailed record of past climate conditions has been recovered from the great ice sheets of Greenland and Antarctica. These ice sheets are built by snow falling on the ice surface and being covered by subsequent snowfalls. The compressed snow is transformed into ice. It is so cold in these polar locations that the ice doesn't melt even in the summers, so the ice is able to build up over hundreds of thousands of years. Because the ice at lower depths was produced by progressively earlier snowfalls, the age of the ice increases with depth, and the youngest ice is at the surface. The Antarctic ice sheet is up to three miles thick. It takes a long time to build up this much ice, and the oldest ice found at the bottom of the Antarctica ice sheet is around 800,000 years old.

Scientists drill into these ice sheets to extract *ice cores*, which record information about past climates. Figure **7.7.1** shows what these cores look like when they are cut open. Like tree rings, ice cores indicate years of growth. Note how the middle core (which required over a mile of drilling to extract!) has distinct layers—this is because the seasons leave an imprint in the layers of snow. Scientists can use this imprint to help calculate the age of the ice at different depths, although the task becomes more difficult the deeper the core sample, since the ice layers become more compressed. The ice records several different types of climate information: the temperature of the core, the properties of the water that make up the ice, trapped dust, and tiny entombed bubbles of ancient atmosphere.

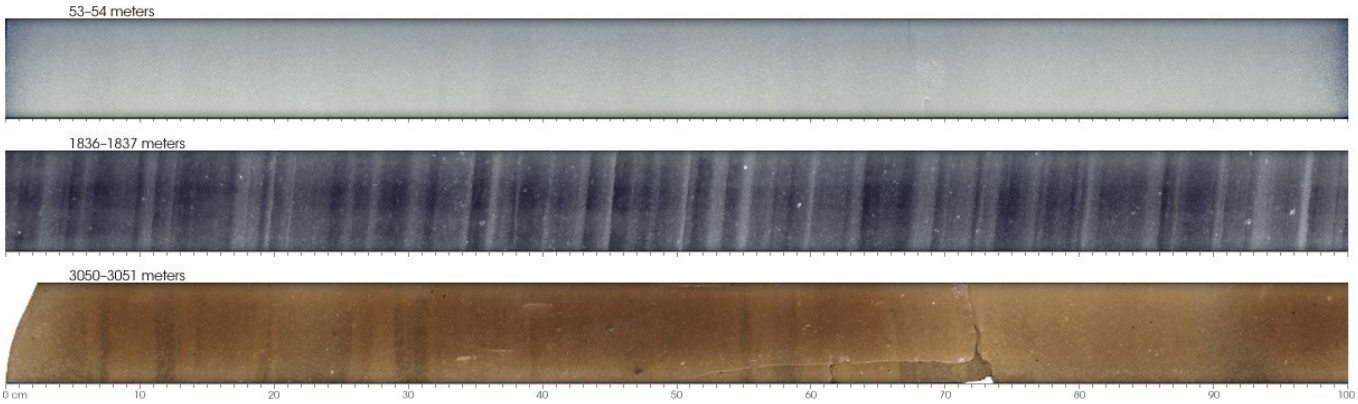


Figure 7.7.1 Ice cores. Three different sections of an ice core. The seasonal layers are most clear in the middle section (note the dark and light bands). The deepest section (bottom core) is taken from almost two miles down and is colored brown by rocky debris from the ground under the ice. *Source: National Ice Core Laboratory*

The water molecules that make up the ice record information about the temperature of the atmosphere. Each water molecule is made up of two hydrogen atoms and one oxygen atom (and so has the chemical name H_2O). Not all oxygen atoms are the same however; some are "light" and some are "heavy". These different types of oxygen are called isotopes, which are atoms that have same number of protons but different numbers of neutrons. The heavy isotope of oxygen (oxygen-18, or ^{18}O) is more than 10% heavier than the light isotope (oxygen-16 or ^{16}O). This means that some water molecules weigh more than others. This is important because lighter water molecules are more easily evaporated from the ocean, and once in the atmosphere, heavier water molecules are more likely to condense and fall as precipitation. As we can see from Figure **7.7.2**, the water in the ice sheets is lighter (has a higher proportion of ^{16}O relative to ^{18}O) than the water in the oceans.

The process of differentiation between heavy and light water molecules is temperature dependent. If the atmosphere is warm, there is more energy available to evaporate and hold the heavier ^{18}O water in the atmosphere, so the snow that falls on the polar ice sheets is relatively higher in ^{18}O. When the atmosphere is cold, the amount of energy is less, and so less ^{18}O makes it to the poles to be turned into glacial ice. We can compare the amount of ^{18}O in different parts of the ice core to see how the atmosphere's temperature—the climate—has changed.

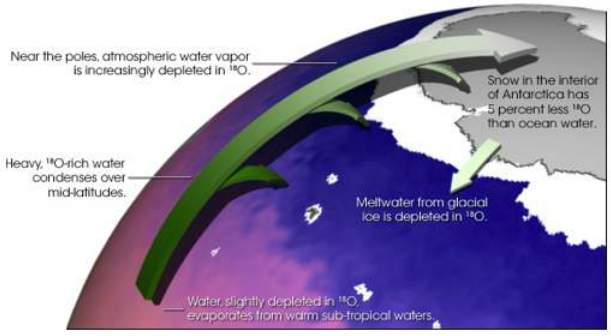


Figure **7.7.2** Water becomes lighter as it travels toward the poles. The heavy (18O) water drops out of the atmosphere (as rain or snow) before reaching the ice sheet. This means that the snow that forms the glacial ice is lighter than the ocean water (has more 16O than 18O, compared to ocean water). *Source: Robert Simmon, NASA GSFC, NASA Earth Observatory*

Figure **7.7.3** shows what this record looks like over the last 400,000 years. The blue and green lines depict two different Antarctic ice cores (taken from ice about 350 miles apart) and the variations in oxygen isotopes are converted into temperature changes. The y-axis shows temperature change; today's climate is at zero—the dashed line. Notice that the Earth's climate has not been stable! Sometimes the temperature is higher than it is today—the blue and green lines are higher than the dashed about 120,000 years ago, for example. Most of the time the climate is much colder than today's, however: the most common value is around -6 ^{o}C (-13 ^{o}F). On average, the earth's temperature between 25,000 and 100,000 years ago was about 6 ^{o}C lower than it is today. These changes

can be double-checked by measuring the temperature of the ice in the cores directly. Ice that is 30,000 years old is indeed colder than the ice made today, just as the isotope data predicts.

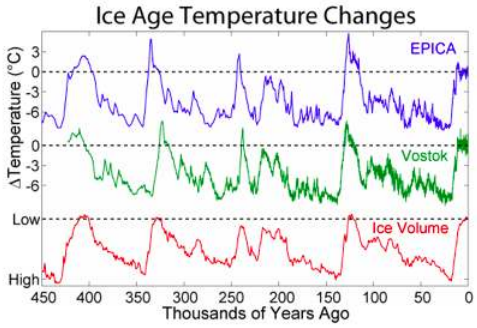


Figure **7.7.3** Ice age temperature. The blue and green lines depict two different Antarctic ice cores (taken from ice about 350 miles apart) and the variations in oxygen isotopes are converted into temperature changes. The red line depicts global ice volume. The y-axis shows temperature change; today's climate is at zero – the dashed line. *Source:* *Robert A. Rohde*

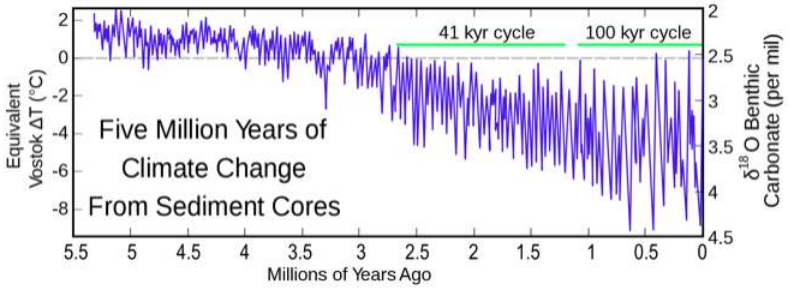


Figure **7.7.4** A comparison of the age of sediment (x-axis) and the change in temperature over time (left y-axis) as derived from oxygen isotope ratios (right y-axis). The dashed line shows today's climate. Note that the climate is cooling over the last few million years, but it is highly variable. In the last one million years the climate alternates between warm and cool conditions on a 100,000-year time scale ("100 kyr cycle"), before this it alternated on a 41,000 year cycle. Both these period lengths are the same as Milankovitch cycles. These cores suggest that today's temperature is higher than almost all of that of the Quaternary (the last 2.6 Million years). *Source:* *Jo Weber*

The changes in climate recorded in ice sheets are thought to be worldwide. The same climate changes observed in Antarctica are also found in cores taken from Greenland, which is on the other side of the Earth. Isotope data can also be taken from sediment cored from the ocean floor—all over the planet—and these cores also show the same changes in climate, alternating between cold and warm. Because ocean sediment is deposited over millions of years, the sediment can give an indication of the climate across the whole of the Quaternary and beyond. Figure **7.7.4** shows how temperature has changed over time (blue line), compared with today (dashed line). The temperature has, on average, gotten colder over the Quaternary, but it also appears to oscillate between warm and cold periods. We'll investigate these periodic changes in the next section of this chapter.

As falling snow accumulates on the ground, tiny bubbles of air become trapped in it. These bubbles are retained as the snow transforms to ice, and constitute tiny samples of the ancient atmosphere that can be analyzed to find out if the changes in temperature (as recorded in the oxygen isotopes) are related to changes in the atmosphere. The temperature recorded by the isotopes in the ice is directly related to the amount of carbon dioxide in the trapped air (Figure **7.7.5**): the times with higher carbon dioxide are also times of high temperature.

Falling snow also captures and entombs atmospheric dust, which is topsoil born aloft by the wind, and which is especially prevalent during droughts. The fact that more dust occurs in the ice accumulated during cold periods suggests that the glacial climate was dry, as well as cold.

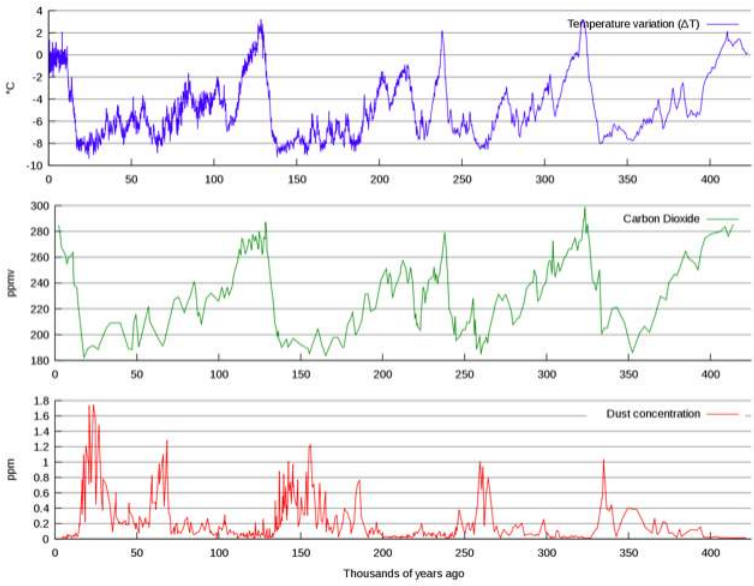


Figure **7.7.5** Vostok Petit Data. These graphs depict how changes in temperature—inferred from changes in isotope ratios (blue line)--correspond to changes in atmospheric carbon dioxide (green line) and dust (red line) over the last 400,000 years as recorded in an ice core extracted from Antarctica. Carbon dioxide varies directly with temperature – the warmer the climate the higher the carbon dioxide level. Atmospheric dust is highest during the coolest periods (such as 25,000 and 150,000 years ago). *Source: William M. Connolley produced figure using data from the National Oceanic and Atmospheric Administration, U.S. Department of Commerce, Paleoclimatology branch, Vostok Ice Core Data.*



Figure **7.7.6** Ice Age Earth- an artist's impression of the Earth during an ice age. Note that the Northern parts of North America and Europe (including Canada and Scandinavia) are entirely covered by ice-sheets. *Source: Itti*

During the Quaternary, the Earth has cycled between glacial periods (sometimes referred to as "ice ages") and interglacial periods. The ice was at its most recent extreme around 20,000 years ago in a period known as the last glacial maximum, or LGM. As we can see from the ice core record, the Quaternary climate is usually cold (see Figure **7.7.3**), with long periods of cold punctuated with shorter (10,000 year long, or so) periods of warmer conditions, like those we experience today. In many ways, our current climate is exceptional—for most of human existence, the Earth has been a much colder place.

What was the Earth like during these glacial periods? Almost all the world was cold; average temperatures were around 6 ^{o}C (-13 ^{o}F) colder than today. Such conditions allow ice sheets to grow—much of North America, Asia and Europe were covered under mile-thick ice (see Figure **7.7.3**). Because this ice was made of water that was once in the oceans, sea levels were much lower. At the LGM, sea level was about 120 meters (or about 400 feet) lower than it is today. As the seas retreated, the continents grew larger, creating land bridges that joined Asia with North America, Britain with Europe, and Australia with Papua New Guinea.

During glacial periods the climate was also much drier, as evidenced by the increase in atmospheric dust (Figure **7.7.5**). The lands at and near the poles were covered with ice, and dry grasslands occupied areas where temperate forests occur today. Deserts were much larger than they are now, and tropical rainforests, having less water and less warmth, were small. The animals and plants of glacial periods were different in their distribution than they are today, as they were adapted to these different conditions. Fossils of

Mastodons (Figure 7.7.7) have been found from all across what is now the United States, including from Florida, which currently enjoys a subtropical climate.



Figure 7.7.7 An artist's impression of a Knight Mastodon, an elephant-like mammal with a thick wooly coat. Mastodon fossils dating from past glacial periods have been found across North America—from Florida to Alaska. *Source:* Charles R. Knight

During glacial periods humans would have been unable to occupy the globe as they do today because all landmasses experienced different climactic conditions. Some countries of the present could not exist, as they would be almost completely covered by ice. As examples, look for Canada, Iceland and The United Kingdom in Figure 7.7.9.

Milankovitch Cycles

Why has the Earth cycled through hot and cold climates throughout the Quaternary? As we learned in the previous module, the Earth's climate is controlled by several different factors—insolation, greenhouse gases, and albedo are all important. Scientists believe that changes in insolation are responsible for these climate swings, and the insolation varies as a result of wobbles in the Earth's orbit.

The Earth's orbit is not fixed – it changes regularly over time. These periodic changes in Earth's orbit named are referred to as Figure 7.7.8, and are illustrated in Figure 7.7.8. Changes in the Earth's orbit alter the pattern of insolation that the Earth receives. There are three principle ways in which the Earth's orbit varies:

1. Eccentricity (or *Orbital shape*): The Earth's orbit is not perfectly circular, but instead follows an ellipse. This means that the Earth is, through the course of the year, sometimes closer and sometimes further away from the Sun. Currently, the Earth is closest to the Sun in early January, and furthest from the Sun in Early July. This changes the amount of insolation by a few percent, so Northern Hemisphere seasons are slightly milder than they would be if the orbital was circular (cooler summers and warmer winters). The orbital shape changes over time: the Earth moves between being nearly circular and being mildly elliptical. There are two main periods over which this change occurs, one takes around 100,000 years (this is the time over which the orbit goes from being circular, to elliptic, and back to circular), another takes around 400,000 years.
2. Axial Tilt (or Obliquity): The Earth axis spins at an angle to its orbit around the Sun – currently this angle is 23.5 degrees (this angle is known as the *axial tilt*). This difference in orbit creates the seasons (as each hemisphere takes turns being tilted towards and away from the Sun over the course of the year). If the axis of spin lined up with the direction of the Earth's orbit (so that the tilt angle was zero) there would be no seasons! This axial tilt also changes over time, varying between 22.1 and 24.5 degrees. The larger the angle, the larger the temperature difference between summer and winter. It takes about 41,000 year for the axial tilt to change from one extreme to the other, and back again. Currently, the axial tilt is midway between the two extremes and is decreasing—which will make the seasons weaker (cooler summers and warmer winters) over the next 20,000 years.
3. Axial Precession: The direction of Earth's axis of rotation also changes over time relative to the stars. Currently, the North Pole points towards the star Polaris, but the axis of rotation cycles between pointing to that star and the star Vega. This impacts the Earth's climate as it determines when the seasons occur in Earth's orbit. When the axis is pointing at Vega, the Northern Hemisphere's peak summer is in January, not July. If this were true today, it would mean that the Northern Hemisphere would experience more extreme seasons, because January is when the Earth is closest to the Sun (as discussed above in eccentricity). This cycle takes around 20,000 years to complete.

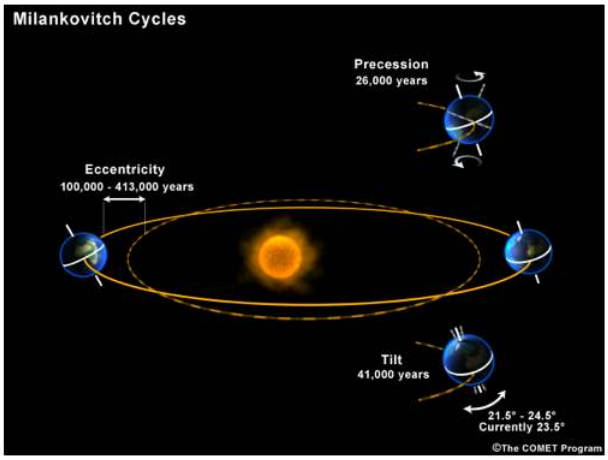


Figure 8. Milankovitch Cycles Illustration of the three variables in Earth's orbit, with periods of variation marked. *Source:* *COMET® at the University Corporation for Atmospheric Research (UCAR) pursuant to a Cooperative Agreements with the National Oceanic and Atmospheric Administration, U.S. Department of Commerce. ©1997-2009 University Corporation for Atmospheric Research. All Rights Reserved.*

The three cycles described above have different periods, all of which are long by human standards: 20,000, 40,000, 100,000 and 400,000 years. If we look at the temperature data from ice and sediment cores, we see that these periods are reflected in Earth's climate. In the last million or so years, the 100,000-year eccentricity in the orbit has determined the timing of glaciations, and before that the 40,000-year axial tilt was dominant (Figure **7.7.8**). These cycles have been important for a long time; geologists have even found evidence of these periods in rocks that are hundreds of millions of years old.

But how do the Milankovitch Cycles change our climate? These orbital cycles do not have much impact on the *total* insolation the Earth receives: they change only the *timing* of that insolation. Since the total insolation does not change, these orbital variations have the power to make the Earth's seasons stronger or weaker, but the average annual temperature should stay the same. The best explanation for long term changes in average annual temperature is that the Milankovitch cycles initiate a positive feedback that amplifies the small change in insolation.

Insolation and the Albedo Feedback

Today, the Earth's orbit is not very eccentric (it is almost circular), but at the beginning of each of the recent ice age periods, the orbit was much more elliptical. This meant that the Earth was further away from the sun during the northern hemisphere summers, reducing the total insolation. Lower insolation meant that the summer months were milder than they would otherwise be, with cooler temperatures. Summer temperatures were also lower when the Earth's axial tilt was smaller, so the two different orbital parameters could reinforce one another's effects, in this case producing especially mild summers.

It is thought that these mild northern summers produced an albedo feedback that made the whole planet slip into an ice age. The northern hemisphere has continents near the poles—Europe, Asia, and North America. Today, these continents have largely temperate climates. During the winter, snow falls across much of the land (see Figure **7.7.9** in the previous module) only to melt during the summer months. If the summers are not hot enough to melt all the snow and ice, glaciers can advance, covering more of the land. Because ice has a high albedo, more sunlight is reflected than before, and the Earth is made cooler. This creates a positive feedback, as the cooler conditions allow the ice to advance further—which, in turn, increases the albedo and cools the Earth! Eventually, a large proportion of the northern continents became covered in ice (Figure **7.7.9**).

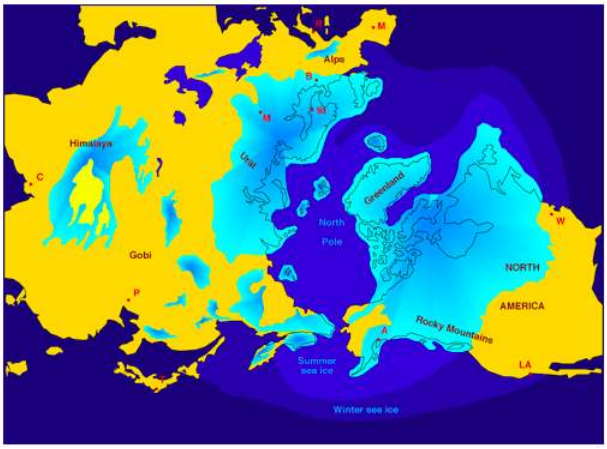


Figure **7.7.9** Glacial coverage (light blue) of the northern hemisphere during the ice ages. *Source: Hannes Grobe*

This positive feedback process works in the other direction, as well. The interglacial periods are ushered in when the orbital parameters create summers that are unusually warm, which melts some of the ice. When the ice sheets shrink, the Earth's albedo decreases, which further warms the system. The giant northern ice sheets shriveled up in a few thousand years as warm summers and decreasing albedo worked together.

These cycles of alternating cooling and warming are also related to changes in the amount of greenhouse gases in the atmosphere. As we observed in Figure **7.7.5**, the climate contains higher levels of carbon dioxide during interglacial periods. Although this appears to make sense—carbon dioxide is a greenhouse gas, and so should produce warmer climates—it is also a puzzle, because it is not clear how changes in Milankovitch cycles lead to higher levels of carbon dioxide in the atmosphere. It is clear that these changes in carbon dioxide are important in making the change in temperature between interglacial and glacial periods so extreme. Several different hypotheses have been proposed to explain why glacial periods produce lower levels of carbon dioxide (it may be related to how the physical changes influence the Earth's ecosystems ability to absorb carbon dioxide: perhaps lower sea levels increase the nutrient supply in the ocean, or the drop in sea level destroys coral reefs, or iron-rich dust from new deserts fertilizes the oceans) but further work on this question remains to be done.

It is a concern for all of us that there are gaps in our understanding of how the feedbacks between insolation, albedo and greenhouse gases operate, as it makes it hard to predict what the consequences of any changes in the climate system might lead to. The current level of atmospheric carbon dioxide is unprecedented in human experience; it is at the highest level ever recorded in the Quaternary. Will the current increase in greenhouse gases lead to a positive feedback, warming the Earth even more?

Review Questions

1. In the text, we discuss how polar ice has a smaller ^{18}O to ^{16}O ratio (that is, it has proportionally less heavy isotope water) than ocean water does. Hydrogen also has isotopes, the two most common being hydrogen-1 (^{1}H) and hydrogen-2 (^{2}H, also known as deuterium). Water is made up of both hydrogen and oxygen, and scientists analyze both elements when examining ice cores. Do you predict that polar ice sheets would have a higher ratio or a lower ratio of ^{1}H to ^{2}H than ocean water? Will colder global temperatures increase or decrease the amount of ^{2}H in polar ice?
2. In the text, we discuss how polar ice has a smaller ^{18}O to ^{16}O ratio (that is, it has proportionally less heavy-isotope water) when the climate is cooler. We also discuss how changes in the ratio of ^{18}O to ^{16}O ratio in sediment cores can also be used to determine the climate's average temperature. In ocean sediments, the ratio of ^{18}O to ^{16}O increases when the climate is cooler (that is, it has proportionally more heavy isotope water). Explain why isotope ratios in ocean sediment have the opposite reaction to those in polar ice.
3. There are three different ways in which the Earth's orbit changes through time. What combination of orbital parameters would be most likely to start an ice age? (Hint: Ice ages require cool northern summers.)

Resources

Do you want to know more about how ice cores are extracted and analyzed? NASA's Earth Observatory has details about the practical issues of drilling ice cores (deep ice needs to "relax" for as long as a year at the surface before being cut open – or it can

shatter!) and how chemical data is interpreted. Go to http://earthobservatory.nasa.gov/Features/Paleoclimatology_IceCores/ for an in-depth article with great links.

Glossary

axial precession

The movement in the axis of rotation, which change in the direction of Earth's axis of rotation relative to the stars.

axial tilt

The angle between a planet's axis of rotation and the line perpendicular to the plane in which it orbits. The Earth's current axial tilt is 23.5 degrees.

eccentricity

A measure of how much an ellipse departs from circularity.

glacial period

A long period of time in which ice -sheets and glaciers are advanced in their extent.

ice sheets

Glaciers big enough to cover a continent. Currently, ice sheets are found in Antarctica and Greenland, but during glacial periods, ice sheets have covered other land masses, including North America.

interglacial period

The warm periods of the Quaternary in which glaciers and ice-sheets retreat. These occur between the longer glacial periods.

isotopes

Atoms that have same number of protons but different numbers of neutrons. This means that they are the same element (e.g. oxygen), have the same chemical properties, but different masses.

last glacial maximum

The time at which ice sheets were at their greatest extent during the latest glacial period.

milankovitch cycles

Periodic variations in the Earth's orbit that influence its climate. These cycles are named after Milutin Milankovitch, a mathematician who quantified the theory.

quaternary period

The most recent geological period, spanning the time from 2.6 million years ago to today.

obliquity

See Axial Tilt.

proxy data

Information about the climate that accumulates through natural phenomena.

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7.8: Greenhouse Gases

We will be covering the four major categories of greenhouse gases that have been impacted by humans the most. See Table **7.8.1** 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, as you saw in section 7.1. 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. For a review of how fossil fuels are mined, see Chapter 4. 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 shows the increases in all three gases following the industrial revolution. 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. Comparing Figure **7.8.1** to Figure 7.3.3, above, what is likely to happen to global temperature following this unprecedented rise in greenhouse gas levels?

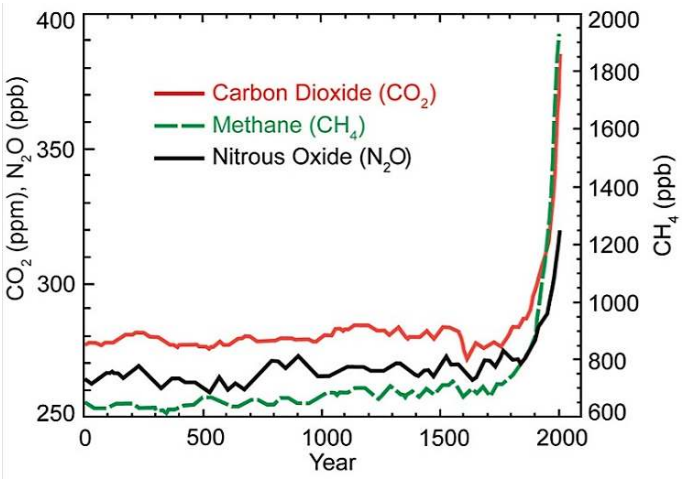


Figure 7.8.1: Increase in greenhouse gas concentrations in the atmosphere over the last 2,000 years. Increases in concentrations of these gases since 1750 are due to human activities in the industrial era. Concentration units are parts per million (ppm) or parts per billion (ppb), indicating the number 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 7.8.1. Some of these chemicals, as well as the older **chlorofluorocarbons** (CFCs), have been phased out by international environmental legislation under the Montreal Protocol (see Chapter 6). 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 7.8.1. 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. How does this impact the comparison of the environmental effects of agricultural practices in lessindustrialized and more-industrialized countries that we completed in section 7.1?

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 **7.8.2**.



Figure **7.8.2**: 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

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7.9: Fossil Fuels and Greenhouse Gases

Fossil fuels are made up mainly of hydrogen and carbon. When burned, the carbon combines with oxygen to create carbon dioxide (CO_2). The amount of CO_2 produced depends on the carbon content of the fuel. For example, for the same amount of energy produced, natural gas produces about half and petroleum produces about three-fourths of the amount of CO_2 produced by coal. Energy-related CO_2 emissions, resulting from the combustion of coal, petroleum, and natural gas, account for about 80% of total U.S. human-caused (anthropogenic) greenhouse gas (GHG) emissions. There are many sources of non-energy CO_2 emissions, but those emissions account for a relatively small share of total GHG emissions. See chapter 7 for a discussion of the results of GHG emissions.

Energy use is largely driven by economic growth and by weather patterns that affect heating and cooling needs. The fuels used in electricity generation also have an impact on the amount of GHG emissions. In the United States, most of the electricity generated comes from coal power plants and consequently, the majority of the carbon dioxide emission resulting from electricity generation is from coal combustion (Figure **7.9.1**). Although the industrial sector is the largest consumer of energy (including direct fuel use and purchased electricity), the transportation sector emits more carbon dioxide because of its near complete dependence on petroleum fuels. The residential and commercial sectors have lower emission levels (most of which comes from fossil energy combustion to produce electricity) than the transportation and industry sectors.

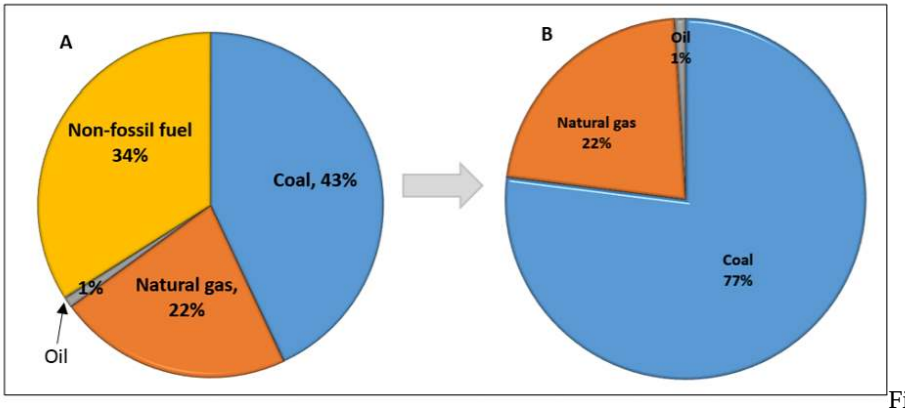


Figure **7.9.1**: *A) Major fuel/energy sources for U.S. electricity generation, 2013. B) Resulting carbon dioxide emissions from electricity generation by fuel type, 2013. Based on data from U.S. Energy Information Administration*

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7.10: Consequences of Climate Change

As part of your assigned reading for this section, read the article “The Coming Storm” by Don Belt published in National Geographic: http://ngm.nationalgeographic.com/pr...desh/belt-text In this section, we will discuss the effects of climate change, both those that have already been observed, as well as future predictions based on scientific climate models (see section 7.5 for a discussion of scientific models). Here, the differences between the terms **global warming** and **climate change** become apparent. 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. We have already observed this increase occurring, as you saw in Figure 7.5.2 from section 7.5. We have also seen, and expect to continue to see, other changes occurring in the **climate** of the Earth. Furthermore, changes have been observed, and we expect to continue to observe, changes in other chemical, physical, and biological aspects of the Earth’s environment. 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 (see Chapter 6). 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 (about 0.6 °C, according to Figure 7.5.2 from section 7.5). 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 **7.10.1**, 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?

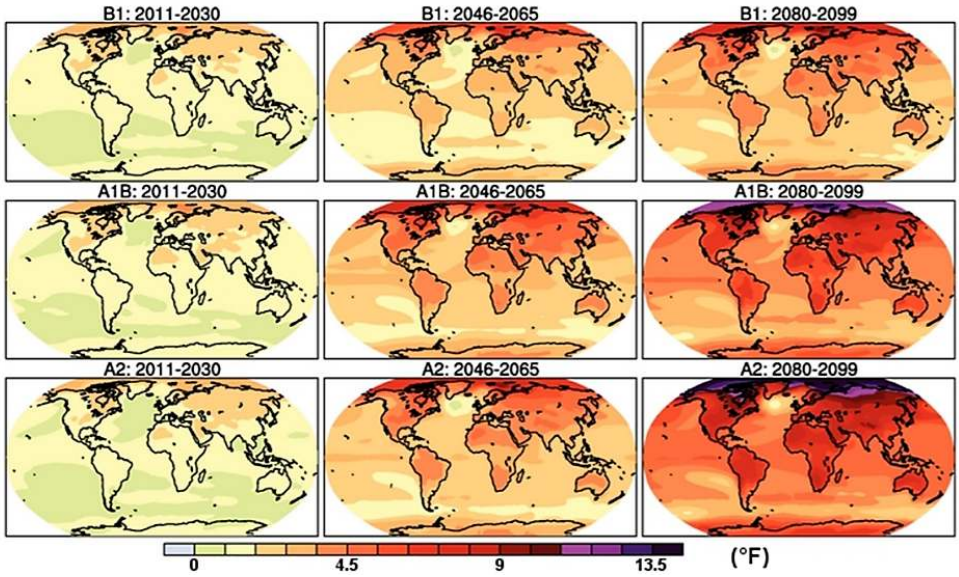


Figure **7.10.1**: Projected changes in global average temperatures under three emissions scenarios (rows) for three different time periods (columns). Changes in temperatures are relative to 1961- 1990 averages. The scenarios come from the IPCC Special Report on Emissions Scenarios: B1 is a low emissions 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 **7.10.2**, 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?

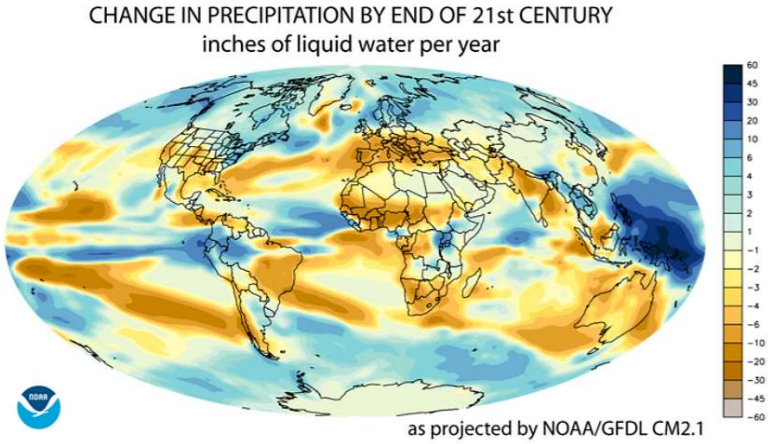


Figure 7.6.1). The plotted precipitation differences were computed as the difference between the 2081 to 2100 20-year averages minus the 1951 to 2000 50-year average. Blue areas project increases in precipitation; brown areas project decreases.

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 (see Chapter 8 for information on the water cycle), 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 **7.10.3**).

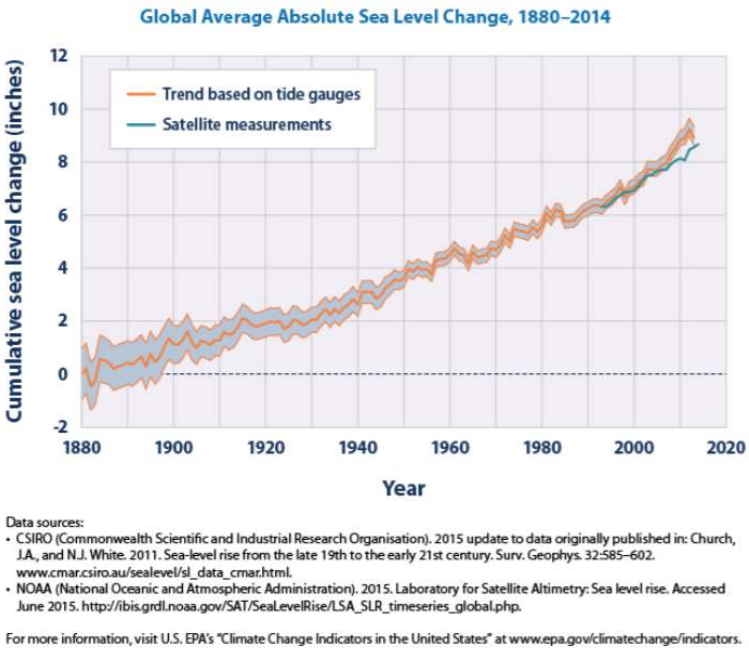


Figure **7.10.3**: *This graph shows average absolute cumulative changes in sea level for the world's oceans since 1880, based on a combination of tide gauge measurements and recent satellite measurements. The shaded band shows the likely range of values, based on the number of measurements collected and the precision of the methods used.*

From the data in Figure **7.10.3**, 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 **7.10.4**) 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/sci ence/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 7.10.4: 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 7.10.5 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.

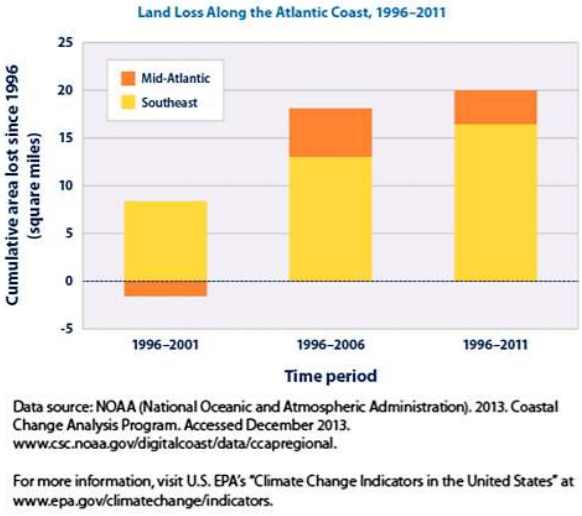


Figure 7.6.5: *This graph shows the net amount of land converted to open water along the Atlantic coast during three time periods: 1996–2001, 1996–2006, and 1996–2011. The results are divided into two regions: the Southeast and the Mid-Atlantic. Negative numbers show where land loss is outpaced by the 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 **7.10.1** and **7.10.2**).

$$CO_2 + H_2O \leftrightarrow H^+ + HCO_3^- \tag{7.10.1}$$

$$HCO_3^- \leftrightarrow H^+ + CO_3^{2-} \tag{7.10.2}$$

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 **7.10.1** and **7.10.2** 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 **7.10.6** 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.

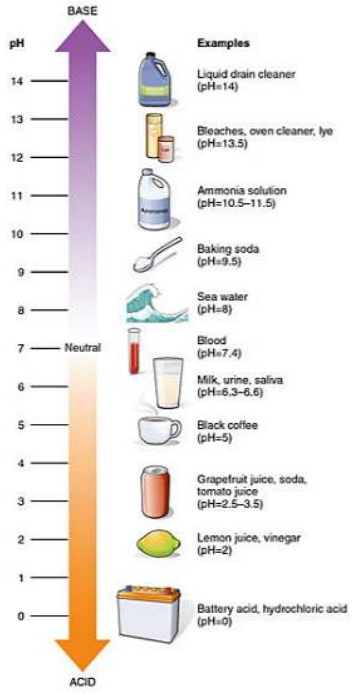


Figure 7.6.6: *The pH scale and relative acidity. Illustration from Anatomy & Physiology, Connexions Web site. http://cnx.org/content/col11496/1.6/, Jun 19, 2013.*

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 **7.10.7**.

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/.

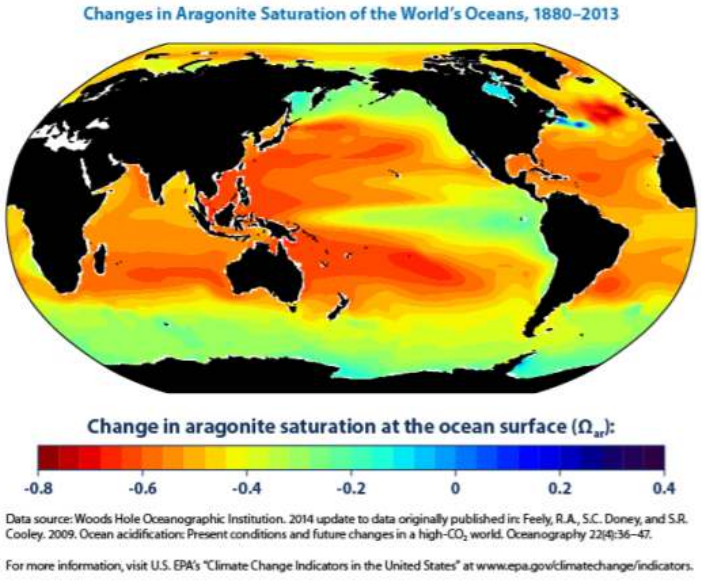


Figure **7.10.7**: This map shows changes in the aragonite saturation level of ocean surface waters between the 1880s and the most recent decade (2004–2013). Aragonite is a form of calcium carbonate that many marine animals use to build their skeletons and shells. A negative change represents a decrease in saturation.

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7.11: Looking Forward: 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 8).

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 7.7.1).

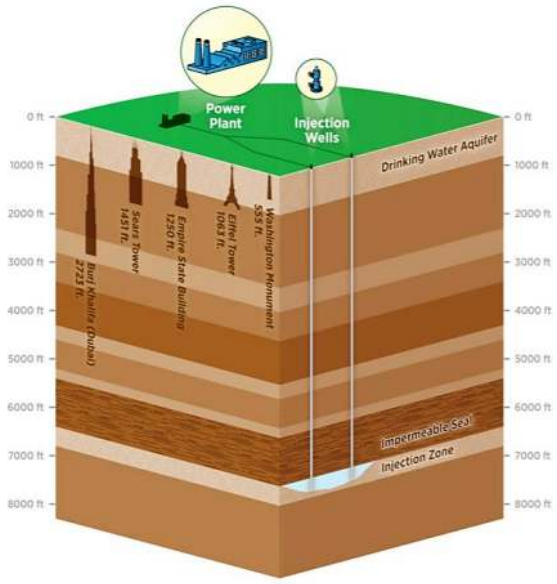


Figure 7.7.1: *Carbon capture and sequestration schematic with landmarks shown to scale for depth reference. Source: US EPA.*

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 5) 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.

Resources

Belt, Don. "The Coming Storm." National Geographic Magazine, May 2011. http://ngm.nationalgeographic.com/pr...desh/belt-text

IPCC Fourth Assessment Report: Climate Change 2007 Synthesis Report. 4. Adaptation and mitigation options www.ipcc.ch/publications_and.../en/spms4.html

NOAA Coral Reef Conservation Program: Climate Change coralreef.noaa.gov/threats/climate/

NOAA Geophysical Fluid Dynamics Laboratory: Will the wet get wetter and the dry drier? www.gfdl.noaa.gov/will-the-we...-the-dry-drier

Ontl, Todd and Lisa Schulte (2012) Soil Carbon Storage. Nature Education Knowledge 3(10):35 http://www.nature.com/scitable/knowl...orage-84223790

US EPA Carbon Dioxide Capture and Sequestration http://www.epa.gov/climatechange/ccs/

US EPA Carbon Footprint Calculator http://www3.epa.gov/carbon-footprint-calculator/

US EPA Climate Change Indicators in the United States http://www.epa.gov/climatechange/sci...ors/index.html

Terms list

Acidity

Climate

pH

Adaptation

Climate change

Precipitation

Aragonite

Coral bleaching

Runoff

Bicarbonate

Desalination

Sea ice

Calcium carbonate

Evaporation

Thermal expansion

Carbon capture and sequestration

Glacier

Wetland

Global warming

Carbon footprint

Hydrogen ions

Carbon tax

Land ice

Carbonate

Mitigation

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7.12: Water Cycle and Fresh Water Supply

Learning Objectives

After reading this module, students should be able to

- understand how the water cycle operates
- understand the principles controlling groundwater resources and how they also can affect surface water resources
- know the causes and effects of depletion in different water reservoirs
- understand how we can work toward solving the water supply crisis

Water Reservoirs and Water Cycle

Water is the only substance that occurs naturally on earth in three forms: solid, liquid and gas. It is distributed in various locations, called water resevoirs. The oceans are by far the largest of the reservoirs with about 97% of all water but that water is too saline for most human uses (see Figure **7.12.1**). Ice caps and glaciers are the largest reservoirs of fresh water but this water is inconveniently located, mostly in Antarctica and Greenland. Shallow groundwater is the largest reservoir of usable fresh water. Although rivers and lakes are the most heavily used water resources, they represent only a tiny amount of the world's water. If all of world's water was shrunk to the size of 1 gallon, then the total amount of fresh water would be about 1/3 cup, and the amount of readily usable fresh water would be 2 tablespoons.

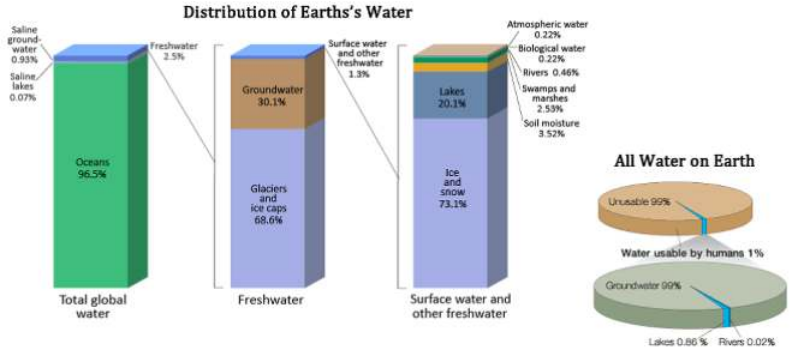


Figure **7.12.1** Earth's Water Resevoir. Bar chart Distribution of Earth's water including total global water, fresh water, and surface water and other fresh water and Pie chart Water usable by humans and sources of usable water. *Source: United States Geographical Survey Igor Skiklomanov's chapter "World fresh water resources" in Peter H. Gleick (editor), 1993, Water in Crisis: A Guide to the World's Fresh Water Resources*

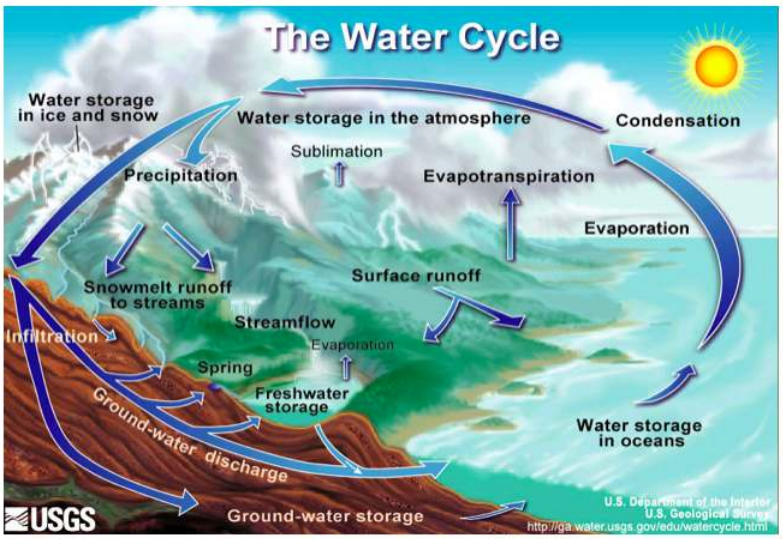


Figure **7.12.2** The Water Cycle. Arrows depict movement of water to different reservoirs located above, at, and below Earth's surface. *Source: United States Geological Survey*

An important part of the water cycle is how water varies in salinity, which is the abundance of dissolved ions in water. Ocean water is called salt water because it is highly saline, with about 35,000 mg of dissolved ions per liter of seawater. Evaporation (where water changes from liquid to gas at ambient temperatures) is a distillation process that produces nearly pure water with almost no dissolved ions. As water vaporizes, it leaves the dissolved ions in the original liquid phase. Eventually, condensation (where water changes from gas to liquid) forms clouds and sometimes precipitation (rain and snow). After rainwater falls onto land, it dissolves minerals, which increases its salinity. Most lakes, rivers, and near-surface groundwater have a relatively low salinity and are called fresh water. The next several sections discuss important parts of the water cycle relative to fresh water resources.

Primary Fresh Water Resources: Precipitation

Precipitation is a major control of fresh water availability, and it is unevenly distributed around the globe (see Figure **7.12.3**). More precipitation falls near the equator, and landmasses there are characterized by a tropical rainforest climate. Less precipitation tends to fall near 20–30° north and south latitude, where the world's largest deserts are located. These rainfall and climate patterns are related to global wind circulation cells. The intense sunlight at the equator heats air, causing it to rise and cool, which decreases the ability of the air mass to hold water vapor and results in frequent rainstorms. Around 30° north and south latitude, descending air conditions produce warmer air, which increases its ability to hold water vapor and results in dry conditions. Both the dry air conditions and the warm temperatures of these latitude belts favor evaporation. Global precipitation and climate patterns are also affected by the size of continents, major ocean currents, and mountains.

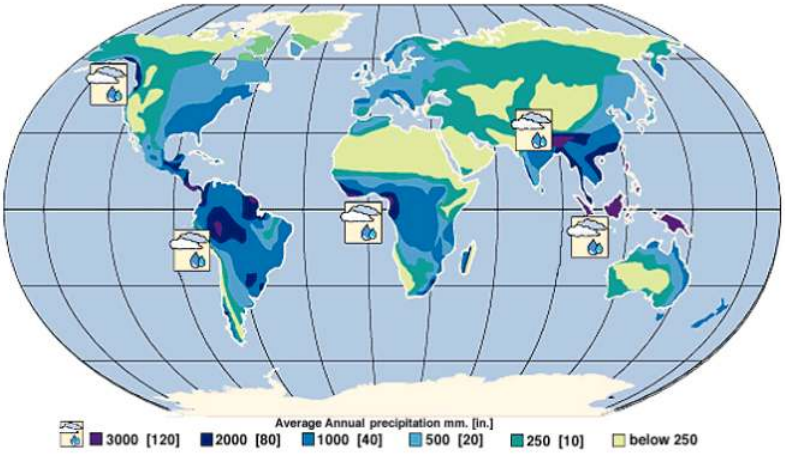


Figure **7.12.3** World Rainfall Map. The false-color map above shows the amount of rain that falls around the world. Areas of high rainfall include Central and South America, western Africa, and Southeast Asia. Since these areas receive so much rainfall, they are where most of the world's rainforests grow. Areas with very little rainfall usually turn into deserts. The desert areas include North Africa, the Middle East, western North America, and Central Asia. *Source: United States Geological Survey Earth Forum, Houston Museum Natural Science*

Surface Water Resources: Rivers, Lakes, Glaciers

Flowing water from rain and melted snow on land enters river channels by surface runoff (see Figure **7.12.4**) and groundwater seepage (see Figure **7.12.5**). River discharge describes the volume of water moving through a river channel over time (see Figure **7.12.6**). The relative contributions of surface runoff vs. groundwater seepage to river discharge depend on precipitation patterns, vegetation, topography, land use, and soil characteristics. Soon after a heavy rainstorm, river discharge increases due to surface runoff. The steady normal flow of river water is mainly from groundwater that discharges into the river. Gravity pulls river water downhill toward the ocean. Along the way the moving water of a river can erode soil particles and dissolve minerals, creating the river's load of moving sediment grains and dissolved ions. Groundwater also contributes a large amount of the dissolved ions in river water. The geographic area drained by a river and its tributaries is called a drainage basin. The Mississippi River drainage basin includes approximately 40% of the U.S., a measure that includes the smaller drainage basins (also called watersheds), such as the Ohio River and Missouri River that help to comprise it. Rivers are an important water resource for irrigation and many cities around the world. Some of the world's rivers that have had international disputes over water supply include the Colorado (Mexico, southwest U.S.), Nile (Egypt, Ethiopia, Sudan), Euphrates (Iraq, Syria, Turkey), Ganges (Bangladesh, India), and Jordan (Israel, Jordan, Syria).



Figure **7.12.4** Surface runoff, part of overland flow in the water cycle *Source: James M. Pease at Wikimedia Commons*



Figure **7.12.5** Groundwater seepage can be seen in Box Canyon in Idaho, where approximately 10 cubic meters per second of seepage emanates from its vertical headwall. *Source: NASA*



Figure **7.12.6** Colorado River, U.S.. Rivers are part of overland flow in the water cycle and an important surface water resource. *Source: Gonzo fan2007 at Wikimedia Commons*

Lakes can also be an excellent source of fresh water for human use. They usually receive water from surface runoff and groundwater discharge. They tend to be short-lived on a geological time-scale because they are constantly filling in with sediment supplied by rivers. Lakes form in a variety of ways including glaciation (Great Lakes, North America, See Figure **7.12.7**), recent tectonic uplift (Lake Tanganyika, Africa), and volcanic eruptions (Crater Lake, Oregon). People also create artificial lakes (resevoirs) by damming rivers. Large changes in climate can result in major changes in a lake's size. As Earth was coming out of the last Ice Age about fifteen thousand years ago, the climate in the western U.S. changed from cool and moist to warm and arid, which caused more than100 large lakes to disappear. The Great Salt Lake in Utah is a remnant of a much larger lake called Lake Bonneville.



Figure **7.12.7** Great Lakes from Space. The Great Lakes hold 21% of the world's surface fresh water. Lakes are an important surface water resource. *Source: SeaWiFS Project, NASA/Goddard Space Flight Center, and ORBIMAGE*

Although glaciers represent the largest reservoir of fresh water, they generally are not used as a water source because they are located too far from most people (see Figure **7.12.8**). Melting glaciers do provide a natural source of river water and groundwater. During the last Ice Age there was as much as 50% more water in glaciers than there is today, which caused sea level to be about 100 m lower. Over the past century, sea level has been rising in part due to melting glaciers. If Earth's climate continues to warm, the melting glaciers will cause an additional rise in sea level.



Figure **7.12.8** Mountain Glacier in Argentina. Glaciers are the largest reservoir of fresh water but they are not used much as a water resource directly by society because of their distance from most people. *Source: Luca Galuzzi - www.galuzzi.it*

Groundwater Resources

Although most people in the U.S. and the world use surface water, groundwater is a much larger reservoir of usable fresh water, containing more than 30 times more water than rivers and lakes combined. Groundwater is a particularly important resource in arid climates, where surface water may be scarce. In addition, groundwater is the primary water source for rural homeowners, providing 98% of that water demand in the U.S.. Groundwater is water located in small spaces, called pore space, between mineral grains and fractures in subsurface earth materials (rock or sediment, i.e., loose grains). Groundwater is not located in underground rivers or lakes except where there are caves, which are relatively rare. Between the land surface and the depth where there is groundwater is the unsaturated zone, where pore spaces contain only air and water films on mineral grains (see Figure **7.12.9**)[1]. Below the unsaturated zone is the saturated zone, where groundwater completely fills pore spaces in earth materials. The interface between the unsaturated zone and saturated zone is the water table. Most groundwater originates from rain or snowmelt, which infiltrates the

ground and moves downward until it reaches the saturated zone. Other sources of groundwater include seepage from surface water (lakes, rivers, reservoirs, and swamps), surface water deliberately pumped into the ground, irrigation, and underground wastewater treatment systems, i.e., septic tanks. Recharge areas are locations where surface water infiltrates the ground rather than running off into rivers or evaporating. Wetlands and flat vegetated areas in general are excellent recharge areas.

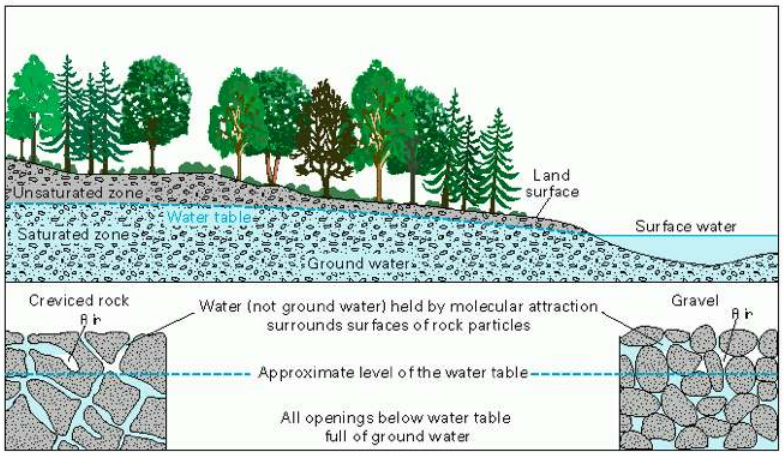


Figure **7.12.9** Subsurface Water Terminology. Groundwater in pore spaces and fractures of earth materials, saturated zone, unsaturated zone, and water table, which follows land surface but in a more subdued way. *Source:* *United States Geological Survey*

Groundwater is in constant motion due to interconnection between pore spaces. Porosity is the percentage of pore space in an earth material and it gives a measure of how much groundwater an earth material can hold. Permeability is a measure of the speed that groundwater can flow through an earth material, and it depends on the size and degree of interconnection among pores. An earth material that is capable of supplying groundwater from a well at a useful rate—i.e., it has relatively high permeability and medium to high porosity—is called an aquifier. Examples of aquifers are earth materials with abundant, large, well-connected pore spaces such as sand, gravel, uncemented sandstone, and any highly fractured rock. An earth material with low hydraulic conductivity is an aquitard. Examples of aquitards include clay, shale (sedimentary rock with abundant clay), and igneous and metamorphic rock, if they contain few fractures.

As discussed above, groundwater flows because most earth materials near the surface have finite (nonzero) porosity and permeability values. Another reason for groundwater movement is that the surface of the water table commonly is not completely flat but mimics the topography of the land surface, especially in humid climates. There is "topography" to the water table because groundwater moves slowly through rock and soil, so it builds up in higher elevation areas. In fact, when groundwater flows slowly through aquitards and deep underground, it can take many thousands of years to move relatively short distances. An unconfined aquifier has no aquitard above it and, therefore, it is exposed to the atmosphere and surface waters through interconnected pores (See Figure **7.12.10**). In an unconfined aquifer, groundwater flows because of gravity to lower water table levels, where it eventually may discharge or leave the groundwater flow system. Discharge areas include rivers, lakes, swamps, reservoirs, water wells, and springs (see Figure **7.12.11**). Springs are rivers that emerge from underground due to an abrupt intersection of the land surface and the water table caused by joints, caves, or faults that bring permeable earth materials to the surface. A confined aquifier is bounded by aquitards below and above, which prevents recharge from the surface immediately above. Instead, the major recharge occurs where the confined aquifer intercepts the land surface, which may be a long distance from water wells and discharge areas (see Figure **7.12.12**). Confined aquifers are commonly inclined away from recharge areas, so groundwater in a confined aquifer is under greater-than-atmospheric pressure due to the weight of water in the upslope direction. Similar to river discharge, groundwater discharge describes the volume of water moving through an aquifer over time. Total groundwater discharge depends on the permeability of the earth material, the pressure that drives groundwater flow, and the size of the aquifer. It is important to determine groundwater discharge to evaluate whether an aquifer can meet the water needs of an area.

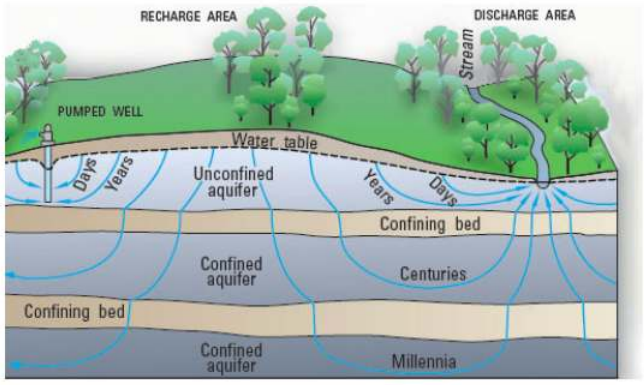


Figure **7.12.10** Flowing Groundwater. Blue lines show the direction of groundwater in unconfined aquifers, confined aquifers, and confining beds. Deep groundwater moves very slowly especially through low permeability layers.*Source: United States Geological Survey*



Figure **7.12.11** Fatzael Springs in Jordan Valley. A spring is a river that emerges from underground due to an abrupt intersection of the water table with the land surface such as alongside a hill. *Source: Hanay at Mediawiki Commons*

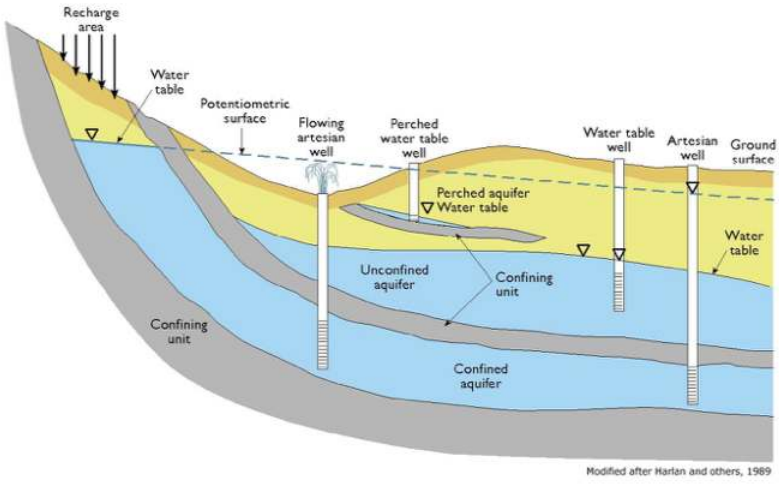


Figure **7.12.12** Schematic Cross Section of Aquifer Types. This figure shows different types of aquifers and water wells, including unconfined aquifer, confined aquifer, water table well, artesian well, and flowing artesian well. Point of triangle is water level in each well and water table in other parts of figure.

Water level in artesian well is at potentiometric surface and above local water table (dashed blue line) due to extra pressure on groundwater in confined aquifer. Water in flowing artesian well moves above land surface. *Source: Colorado Geological Survey*

Most shallow water wells are drilled into unconfined aquifers. These are called water table wells because the water level in the well coincides with the water table (See Figure **7.12.13**). 90% of all aquifers for water supply are unconfined aquifers composed of sand or gravel. To produce water from a well, you simply need to drill a hole that reaches the saturated zone and then pump water to the surface. Attempting to pump water from the unsaturated zone is like drinking root beer with a straw immersed only in the foam at the top.

To find a large aquifer for a city, hydrogeologists (geologists who specialize in groundwater) use a variety of information including knowledge of earth materials at the surface and sub-surface as well as test wells. Some people search for water by dowsing, where someone holds a forked stick or wire (called a divining rod) while walking over an area. The stick supposedly rotates or deflects downward when the dowser passes over water. Controlled tests show that a dowser's success is equal to or less than random chance. Nevertheless, in many areas water wells are still drilled on dowser's advice sometimes for considerable money. There is no scientific basis to dowsing.

Wells into confined aquifers typically are deeper than those into unconfined aquifers because they must penetrate a confining layer. The water level in a well drilled into a confined aquifer, which is an artesian well, (see Figure **7.12.12**), moves above the local water table to a level called the potentiometric surface because of the greater pressure on the groundwater. Water in a flowing well (see Figure **7.12.14**) moves all of the way to the land surface without pumping.



Figure **7.12.13** A Flowing Well. Flowing artesian well where water moves above the land surface due to extra pressure on the groundwater in a confined aquifer. *Source: Environment Canada*

A confined aquifer tends to be depleted from groundwater pumping more quickly than an unconfined aquifer, assuming similar aquifer properties and precipitation levels. This is because confined aquifers have smaller recharge areas, which may be far from the pumping well. Conversely, an unconfined aquifer tends to be more susceptible to pollution because it is hydrologically connected to the surface, which is the source of most pollution.

Groundwater and surface water (rivers, lakes, swamps, and reservoirs) are strongly interrelated because both are part of the same overall resource. Major groundwater removal (from pumping or drought) can lower the levels of surface water and vice versa. We can define two types of streams: gaining (effluent) streams and losing (influent) streams (see Figure **7.12.14**). Gaining streams tend to be perennial (flow year round), are characteristic of humid climates, have the water table sloping towards the river, and therefore gain water from groundwater discharge. Losing streams tend to be ephemeral (flow only after significant rain), are characteristic of arid climates, are located above the water table (which slopes away from the river), and therefore lose water to groundwater

recharge. Pollution that is dumped into a losing stream will tend to move into the ground and could also contaminate local groundwater.

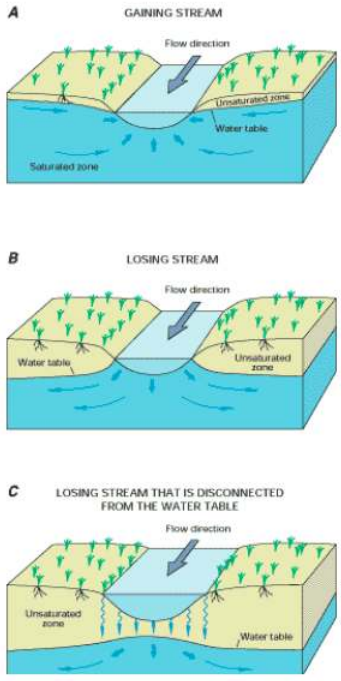


Figure **7.12.14** Interaction of Streams and Ground Water. A) Gaining stream where water table slopes toward river and groundwater discharges into river, B) Losing stream where water table slopes away from river and river water discharges into groundwater, C) Losing stream where water table is separated from and below river. *Source: United States Geological Survey*

Water Use in the U.S. and World

People need water to produce the food, energy, and mineral resources they use—commonly large amounts of it. Consider, for example, these approximate water requirements for some things people in the developed world use every day: one tomato = 3 gallons; one kilowatt-hour of electricity (from a thermoelectric power plant) = 21 gallons; one loaf of bread = 150 gallons; one pound of beef = 1,600 gallons; and one ton of steel = 63,000 gallons. Human beings require only about 1 gallon per day to survive, but a typical person in a U.S. household uses approximately 100 gallons per day, which includes cooking, washing dishes and clothes, flushing the toilet, and bathing.

The water demand of an area is a function of the population and other uses of water. There are several general categories of water use, including offstream use, which removes water from its source, e.g., irrigation, thermoelectric power generation (cooling electricity-producing equipment in fossil fuel, nuclear, and geothermal power plants), industry, and public supply; consumptive use, which is a type of offstream use where water does not return to the surface water or groundwater system immediately after use, e.g., irrigation water that evaporates or goes to plant growth; and instream use, which is water used but not removed from a river, mostly for hydroelectric power generation. The relative size of these three categories are instream use >> offstream use > consumptive use. In 2005, the U.S. used approximately 3,300 billion gallons per day for instream use, 410 billion gallons per day for offstream use, and 100 billion gallons per day for consumptive use. The major offstream uses of that water were thermoelectric (49%), irrigation (31%), public supply (11%), and industry (4%, see Figure **7.12.15**). About 15% of the total water withdrawals in the U.S. in 2005 were saline water, which was used almost entirely for thermoelectric power generation. Almost all of the water used for thermoelectric power generation is returned to the river, lake, or ocean from where it came but about half of irrigation water does not return to the original source due to evaporation, plant transpiration, and loss during transport, e.g., leaking pipes. Total withdrawals of water in the U.S. actually decreased slightly from 1980 to 2005, despite a steadily increasing population. This is because the two largest categories of water use (thermoelectric and irrigation) stabilized or decreased over that time period due to better water management and conservation. In contrast, public supply water demand increased steadily from 1950 (when estimates

began) through 2005. Approximately 77% of the water for offstream use in the U.S. in 2005 came from surface water and the rest was from groundwater (see Figure 7.12.15).

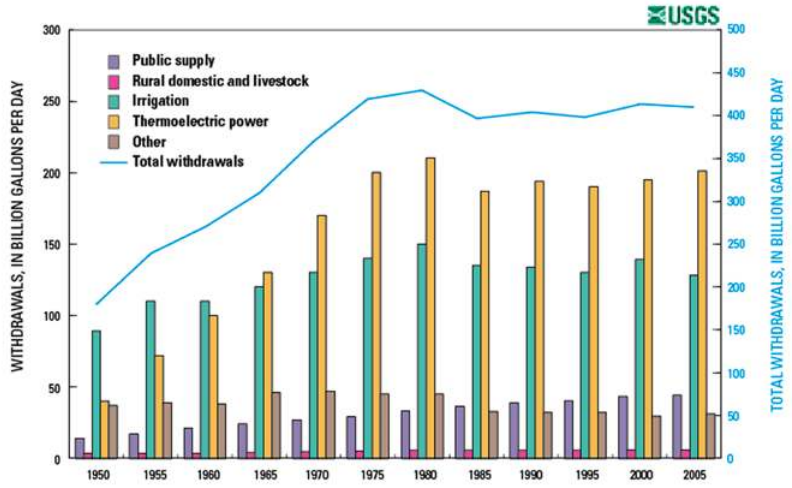


Figure 7.12.15 Trends in Total Water Withdrawals by Water-use Category, 1950-2005. Trends in total water withdrawals in the U.S. from 1950 to 2005 by water use category, including bars for thermoelectric power, irrigation, public water supply, and rural domestic and livestock. Thin blue line represents total water withdrawals using vertical scale on right. *Source: United States Geological Survey*

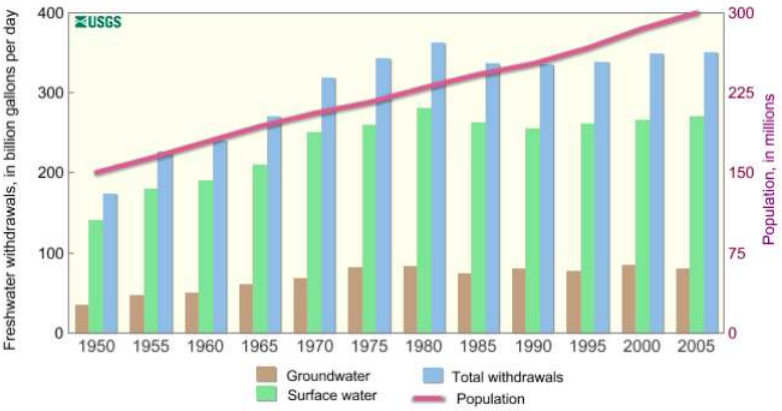


Figure 7.12.16 Trends in Source of Fresh Water Withdrawals in the U.S. from 1950 to 2005. Trends in source of fresh water withdrawals in the U.S. from 1950 to 2005, including bars for surface water, groundwater, and total water. Red line gives U.S. population using vertical scale on right. *Source: United States Geological Survey*

In contrast to trends in the U.S., global total water use is steadily increasing at a rate greater than world population growth (see Figure 7.12.17). During the twentieth century global population tripled and water demand grew by a factor of six. The increase in global water demand beyond the rate of population growth is due to improved standard of living without an offset by water conservation. Increased production of goods and energy entails a large increase in water demand. The major global offstream water uses are irrigation (68%), public supply (21%), and industry (11%).

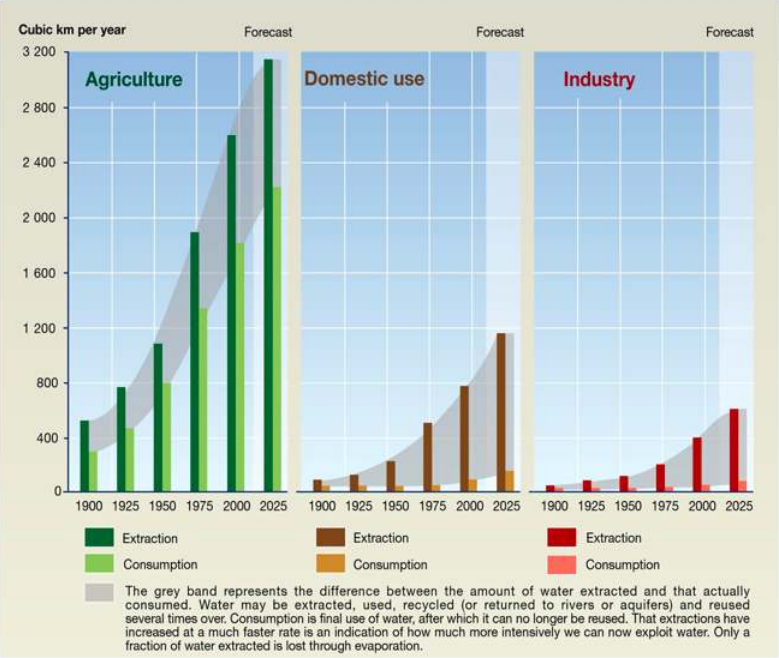


Figure **7.12.17** Trends in World Water Use from 1900 to 2000 and Projected to 2025. For each water major use category, including trends for agriculture, domestic use, and industry. Darker colored bar represents total water extracted for that use category and lighter colored bar represents water consumed (i.e., water that is not quickly returned to surface water or groundwater system) for that use category. *Source: Igor A. Shiklomanow, State Hydrological Institute (SHI, St. Petersburg) and United Nations Educational, Scientific and Cultural Organisation (UNESCO, Paris), 1999*

Water Supply Problems: Resource Depletion

As groundwater is pumped from water wells, there usually is a localized drop in the water table around the well called a cone of depression (see Figure **7.12.18**). When there are a large number of wells that have been pumping water for a long time, the regional water table can drop significantly. This is called groundwater mining, which can force the drilling of deeper, more expensive wells that commonly encounter more saline groundwater. The occurrence of mining does not mean that groundwater will *never* be recharged, but in many cases the recharge rate is negligible on a human time-scale. Confined aquifers are more susceptible to groundwater mining due to their limited recharge areas. Urban development usually worsens groundwater mining because natural recharge rates drop with the proliferation of impermeable pavement, buildings, and roads. Extensive groundwater pumping around Chicago has created a gigantic cone of depression there. Because the water table dropped up to 250 m (800 ft) in the area (see Figure **7.12.19**), many local public water suppliers have switched to Lake Michigan water. Chicago is fortunate to have a large alternate supply of fresh water; many arid locations don't have that luxury. Other places where groundwater mining is a serious problem include the High Plains (Ogallala Aquifer) and the Desert Southwest of the U.S., Mexico, the Middle East, India, and China. Rivers, lakes, and artificial lakes (reservoirs) can also be depleted due to overuse. Some large rivers, such as the Colorado in the U.S. and Yellow in China, run dry in some years. The case history of the Aral Sea discussed below involves depletion of a lake. Finally, glaciers are being depleted due to accelerated melting associated with global warming over the past century.

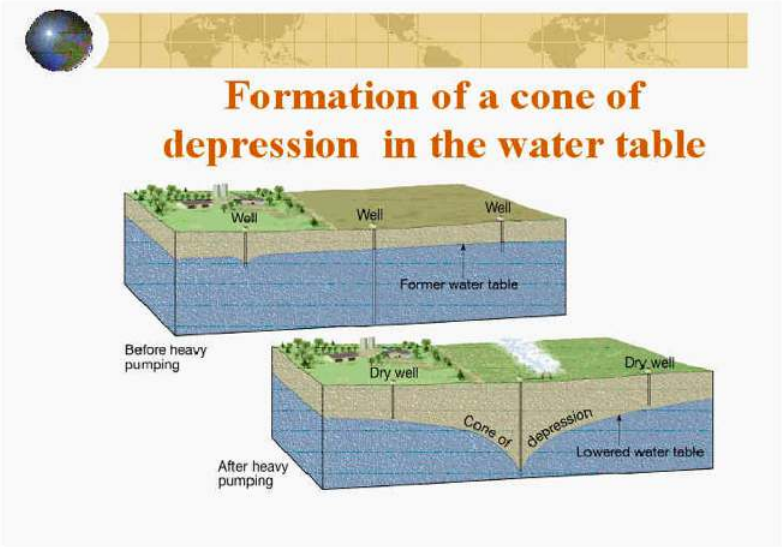


Figure 7.12.18 Formation of a Cone of Depression around a Pumping Water Well. *Source:* *Fayette County Groundwater Conservation District, TX*



Figure 7.12.19 Drop in Water Table in a Confined Aquifer in the Area of Chicago, Illinois and Milwaukee, Wisconsin, U.S. from 1864 - 1980. *Source:* *United States Geological Survey*

Another water resource problem associated with groundwater mining is saltwater intrusion, where overpumping of fresh water aquifers near ocean coastlines causes saltwater to enter fresh water zones. Saltwater intrusion is a significant problem in many coastal areas of the U.S. including Long Island, New York; Cape Cod, Massachusetts; and southeastern and Gulf Coastal states. The drop of the water table around a cone of depression in an unconfined aquifer can change the regional groundwater flow direction, which could send nearby pollution toward the pumping well instead of away from it. Finally, problems of subsidence (gradual sinking of the land surface over a large area) and sinkholes (rapid sinking of the land surface over a small area) can develop due to a drop in the water table.

The Water Supply Crisis

The water crisis refers to a global situation where people in many areas lack access to sufficient water or clean water or both. This section describes the global situation involving water shortages, also called water stress. The next section covers the water crisis involving water pollution. Figure 7.12.20 shows areas of the world experiencing water stress as defined by a high percentage of water withdrawal compared to total available water. Due to population growth the 2025 projection for global water stress is significantly worse than water stress levels in 1995. In general, water stress is greatest in areas with very low precipitation (major

deserts) or large population density (e.g., India) or both. Future global warming could worsen the water crisis by shifting precipitation patterns away from humid areas and by melting mountain glaciers that recharge rivers downstream. Melting glaciers will also contribute to rising sea level, which will worsen saltwater intrusion in aquifers near ocean coastlines. Compounding the water crisis is the issue of social injustice; poor people generally get less access to clean water and commonly pay more for water than wealthy people.

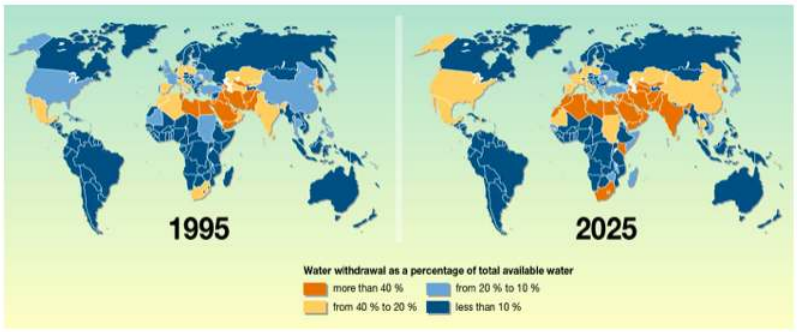


Figure **7.12.20** Countries Facing Water Stress in 1995 and Projected in 2025. Water stress is defined as having a high percentage of water withdrawal compared to total available water in the area. *Source: Philippe Rekacewicz (Le Monde diplomatique), February 2006*

According to a 2006 report by the United Nations Development Programme, in 2005, 700 million people (11% of the world's population) lived under water stress with a per capita water supply below 1,700 $m^3/year^2$ (Watkins, 2006). Most of them live in the Middle East and North Africa. By 2025, the report projects that more than 3 billion people (about 40% of the world's population) will live in water-stressed areas with the large increase coming mainly from China and India. The water crisis will also impact food production and our ability to feed the ever-growing population. We can expect future global tension and even conflict associated with water shortages and pollution. Historic and future areas of water conflict include the Middle East (Euphrates and Tigris River conflict among Turkey, Syria, and Iraq; Jordan River conflict among Israel, Lebanon, Jordan, and the Palestinian territories), Africa (Nile River conflict among Egypt, Ethiopia, and Sudan), Central Asia (Aral Sea conflict among Kazakhstan, Uzbekistan, Turkmenistan, Tajikistan, and Kyrgyzstan), and south Asia (Ganges River conflict between India and Pakistan).

Sustainable Solutions to the Water Supply Crisis?

The current and future water crisis described above requires multiple approaches to extending our fresh water supply and moving towards sustainability. Some of the longstanding traditional approaches include dams and aqueducts. Reservoirs that form behind dams in rivers can collect water during wet times and store it for use during dry spells (see Figure **7.12.21**). They also can be used for urban water supplies. New York City has a large number of reservoirs and controlled lakes up to 200 km away to meet the water demands of its large population. Other benefits of dams and reservoirs are hydroelectricity, flood control, and recreation. Some of the drawbacks are evaporative loss of reservoir water in arid climates, downstream river channel erosion, and impact on the ecosystem including a change from a river to lake habitat and interference with fish migration and spawning. Aqueducts can move water from where it is plentiful to where it is needed (see Figure **7.12.22**). Southern California has a large and controversial network of aqueducts that brings in water from the Sierra Nevada Mountains in the north, the valleys in northern and central California, and the Colorado River to the east (see Figure **7.12.23**). Aqueducts can be controversial and politically difficult especially if the water transfer distances are large. One drawback is the water diversion can cause drought in the area from where the water is drawn. For example, Owens Lake and Mono Lake in central California began to disappear after their river inflow was diverted to the Los Angeles aqueduct. Owens Lake remains almost completely dry, but Mono Lake has recovered more significantly due to legal intervention.

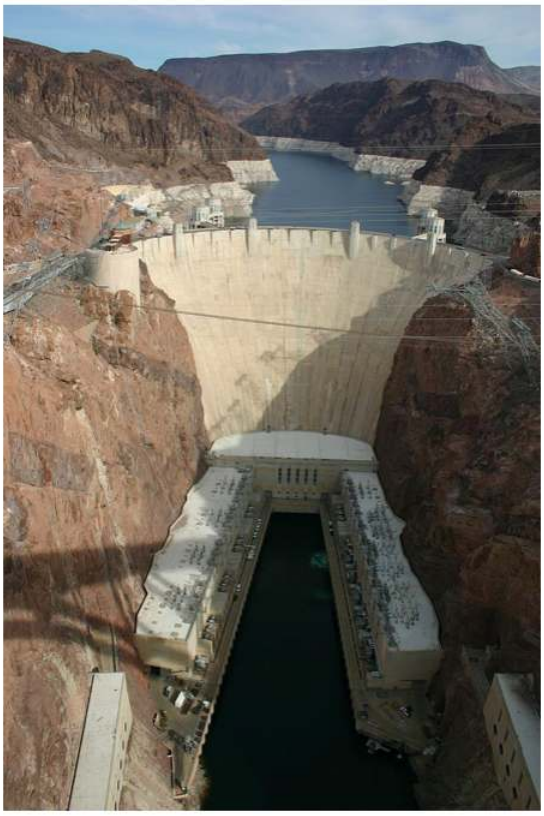


Figure **7.12.21** Hoover Dam, Nevada, U.S.- Hoover Dam, Nevada, U.S.. Behind the dam is Lake Mead, the largest reservoir in U.S.. White band reflects the lowered water levels in the reservoir due to drought conditions from 2000 - 2010. *Source: Cygnusloop99 at Wikimedia Commons*



Figure **7.12.22** The California Aqueduct. California



Aqueduct in southern California, U.S. *Source: David Jordan at en.Wikipedia*

Figure **7.12.23** Map of California Aqueducts. Map of California aqueducts that bring water to southern California from central and northern California and from the Colorado River to the east. *Source: Central Basin Municipal Water District*

The Colorado River, probably the most exploited river in the U.S., has many dams, some huge reservoirs, and several large aqueducts so that it can provide large amounts of fresh water to 7 states in the arid southwestern U.S. and Mexico. The primary use for the water is for a few large cities (Las Vegas, Phoenix, and Tuscon) and irrigation. Allocation of Colorado River water is strictly regulated. Fortunately, not all states use all of their water allocation because the total amount of allocated water is more than the typical Colorado River discharge. Colorado River water gets so saline due to evaporation along its course that the U.S. was forced to build a desalination plant near the border with Mexico so that it could be used for drinking and irrigation. The wetlands of the Colorado River delta and its associated ecosystem have been sadly degraded by the water overuse; some years, no river flow even reaches the ocean.

One method that actually can increase the amount of fresh water on Earth is desalination, which involves removing dissolved salt from seawater or saline groundwater. There are several ways to desalinate seawater including boiling, filtration, electrodialysis, and freezing. All of these procedures are moderately to very expensive and require considerable energy input, making the produced water much more expensive than fresh water from conventional sources. In addition, the processes create highly saline wastewater, which must be disposed of. Desalination is most common in the Middle East, where energy from oil is abundant but water is scarce.

Conservation means using less water and using it more efficiently. Around the home, conservation can involve both engineered features, such as high-efficiency clothes washers and low-flow showers and toilets, as well as behavioral decisions, such as growing native vegetation that require little irrigation in desert climates, turning off the water while you brush your teeth, and fixing leaky faucets. Rainwater harvesting involves catching and storing rainwater for reuse before it reaches the ground. Efficient irrigation is extremely important because irrigation accounts for a much larger water demand than public water supply. Water conservation strategies in agriculture include growing crops in areas where the natural rainfall can support them, more efficient irrigation systems such as drip systems that minimize losses due to evaporation, no-till farming that reduces evaporative losses by covering the soil, and reusing treated wastewater from sewage treatment plants. Recycled wastewater has also been used to recharge aquifers. There are a great many other specific water conservation strategies. Sustainable solutions to the water crisis must use a variety of approaches but they should have water conservation as a high priority.

Review Questions

1. What is the water cycle and why is it important to fresh water resources?
2. What are the relative merits of using surface water vs. groundwater as a water resource?
3. What should society learn from the case history of the Aral Sea?
4. Why is society facing a crisis involving water supply and how can we solve it?

References

Watkins, K. (2006). Beyond scarcity: Power, poverty and the global water crisis. *Human Development Report 2006, United Nations Development Programme*. Retrieved from http://hdr.undp.org/en/reports/global/hdr2006/

Footnotes

1. Groundwater is the name for water in the saturated zone and soil moisture describes water in the unsaturated zone. Therefore, groundwater is the underground water resource used by society but soil moisture is the principal water supply for most plants and is an important factor in agricultural productivity.
2. Although 1,700 m^3/year sounds like a lot of water for every person, it is the minimum amount that hydrologists consider is needed to grow food, support industry, and maintain the environment in general.

Glossary

aqueduct

An aqueduct is a water supply or navigable channel constructed to convey water. In modern engineering, the term is used for any system of pipes, ditches, canals, tunnels, and other structures used for this purpose.

aquifer

Rock or sediment that is capable of supplying groundwater from a well at a useful rate.

aquitard

Earth material with low hydraulic conductivity.

artesian well

Water well drilled into a confined aquifer where the water level in the well moves above the local water table.

condensation

Change in the physical state of water where it goes from gas to liquid.

cone of depression

A localized drop in the water table around a pumping well.

confined aquifer

An aquifer that is bounded by aquitards below and above.

consumptive water use

A societal use of water that is a type of offstream use where water does not return to the river or groundwater system immediately after use.

dam

A barrier built across a river to obstruct the flow of water.

desalination

Removing dissolved salt from seawater or saline groundwater.

discharge area

Location on Earth where groundwater leaves the groundwater flow system.

drainage basin

Geographic area drained by a river and its tributaries.

evaporation

Where water changes from liquid to gas at ambient temperatures.

groundwater

Water located in small spaces between mineral grains and fractures in subsurface rock or sediment.

groundwater mining

A depletion in groundwater resources caused by a large number of water wells that pumped water for a long time.

instream water use

A societal use of water that does not remove it from its source.

offstream water use

A societal use of water that removes it from its source.

permeability

Measure of the speed that groundwater can flow through rock or sediment.

pore space

Small spaces between mineral grains in subsurface rock or sediment.

porosity

Percentage of pore space in rock or sediment.

rainwater harvesting

Catching and storing rainwater for reuse before it reaches the ground.

recharge area

Location on Earth where surface water infiltrates into the ground rather than runs off into rivers or evaporates.

reservoir

Large artificial lake used as a source of water.

river discharge

Volume of water moving through a river channel over time.

saltwater intrusion

Saltwater that enters an aquifer due to overpumping of freshwater aquifers near ocean coastlines.

saturated zone

Subsurface area where groundwater completely fills pore spaces in rock or sediment.

soil moisture

Water in the unsaturated zone.

spring

River that emerges from underground due to an abrupt intersection of the water table with the land surface.

surface runoff

Unchannelized overland flow of water.

transpiration

Loss of water by plants to the atmosphere.

unconfined aquifer

Aquifer with no aquitard above it.

unsaturated zone

Subsurface area where pore spaces contain only air and water films on mineral grains.

water conservation

Using less water and using it more efficiently

water crisis

A global situation where people in many areas lack access to sufficient water or clean water or both.

water cycle

The continuous movement of water through water reservoirs located on, above, and below Earth's surface.

water reservoir (in water cycle)

General location on Earth where water is located including oceans, atmosphere, glaciers, groundwater, lakes, rivers, and biosphere.

water table

Interface between the unsaturated zone and saturated zone.

water table well

Water well drilled into an unconfined aquifer where the water level in the well coincides with the water table.

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7.13: Case Study: The Aral Sea - Going, Going, Gone

The Aral Sea is a lake located east of the Caspian Sea between Uzbekistan and Kazakhstan in central Asia (see Figure **7.13.1**). This area is part of the Turkestan desert, which is the fourth largest desert in the world; it is produced from a rain shadow effect by Afghanistan's high mountains to the south. Due to the arid and seasonally hot climate there is extensive evaporation and limited surface waters in general. Summer temperatures can reach 60° C (140° F)! The water supply to the Aral Sea is mainly from two rivers, the Amu Darya and Syr Darya, which carry snowmelt from mountainous areas. In the early 1960s the then-Soviet Union diverted the Amu Darya and Syr Darya Rivers for irrigation of one of the driest parts of Asia to produce rice, melons, cereals, and especially cotton. The Soviets wanted cotton or “white gold” to become a major export. They were successful and today Uzbekistan is one of the world's largest exporters of cotton. Unfortunately this action essentially eliminated any river inflow to the Aral Sea and caused it to disappear almost completely.



Figure **7.13.1** Map of Aral Sea Area. Map shows lake size in 1960 and political boundaries of 2011. Countries in yellow are at least partially in Aral Sea drainage basin. *Source:* *Wikimedia Commons*

In 1960 Aral Sea was the fourth largest inland water body; only the Caspian Sea, Lake Superior, and Lake Victoria were larger. Since then, it has progressively shrunk due to evaporation and lack of recharge by rivers (see Figure **7.13.2**). Before 1965 the Aral Sea received 20–60 km^3 of fresh water per year from rivers and by the early 1980s it received none. By 2007 the Aral Sea shrank to about 10% of its original size and its salinity increased from about 1% dissolved salt to about 10% dissolved salt, which is 3 times more saline than seawater. These changes caused an enormous environmental impact. A once thriving fishing industry is dead as are the 24 species of fish that used to live there; the fish could not adapt to the more saline waters. The current shoreline is tens of kilometers from former fishing towns and commercial ports. Large fishing boats lie in the dried up lakebed of dust and salt (see Figure **7.13.2**). A frustrating part of the river diversion project is that many of the irrigation canals were poorly built, allowing abundant water to leak or evaporate. An increasing number of dust storms blow salt, pesticides, and herbicides into nearby towns causing a variety of respiratory illnesses including tuberculosis.



Figure **7.13.2** Shrinking Aral Sea. Blue area gives size of Aral Sea in 1960, 1970, 1980, 1990, 2000, 2004, 2008, and 2009 *Source: NordNordWest at Wikimedia Commons*



Figure **7.13.3** An Abandoned Ship. This abandoned ship lies in a dried up lake bed that was the Aral Sea near Aral, Kazakhstan *Source: Staecker at Wikimedia Commons*

The wetlands of the two river deltas and their associated ecosystems have disappeared. The regional climate is drier and has greater temperature extremes due to the absence of moisture and moderating influence from the lake. In 2003 some lake restoration work began on the northern part of the Aral Sea and it provided some relief by raising water levels and reducing salinity somewhat. The southern part of the Aral Sea has seen no relief and remains nearly completely dry. The destruction of the Aral Sea is one of the planet's biggest environmental disasters and it is caused entirely by humans. Lake Chad in Africa is another example of a massive lake that has nearly disappeared for the same reasons as the Aral Sea. Aral Sea and Lake Chad are the most extreme examples of large lakes destroyed by unsustainable diversions of river water. Other lakes that have shrunk significantly due to human diversions of water include the Dead Sea in the Middle East, Lake Manchar in Pakistan, and Owens Lake and Mono Lake, both in California.

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7.14: Water Pollution

Learning Objectives

After reading this module, students should be able to

- understand the major kinds of water pollutants and how they degrade water quality
- understand how and why the lack of safe drinking water in some parts of the world is a major problem
- know what sewage treatment does and why it is important
- know why it is more difficult to remediate groundwater pollution than surface water pollution
- understand how we can work toward solving the crisis involving water pollution

The Water Pollution Crisis

The Module 7.3 described one aspect of the global water crisis, the water shortages that afflict many arid and densely populated areas. The global water crisis also involves water pollution, because to be useful for drinking and irrigation, water 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 improved (safe) drinking water (World Health Statistics, 2010) (See Figure **7.14.1**). At the same time, about 2.6 billion people (or 40% of world population) lived without improved sanitation (see Figure **7.14.2**), 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, e.g., hand washing with soap. Almost all of these deaths are in developing countries, and around 90% of them occur among children under the age of 5 (see Figure **7.14.3**). 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, 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. Some chemical pollutants have serious and well-known health effects; however, many others have poorly known long-term health effects. In the U.S. currently more than 40,000 water bodies fit the definition of “impaired” set by EPA (See Figure **7.14.4**), which means they could neither support a healthy ecosystem nor meet water quality standards. 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.

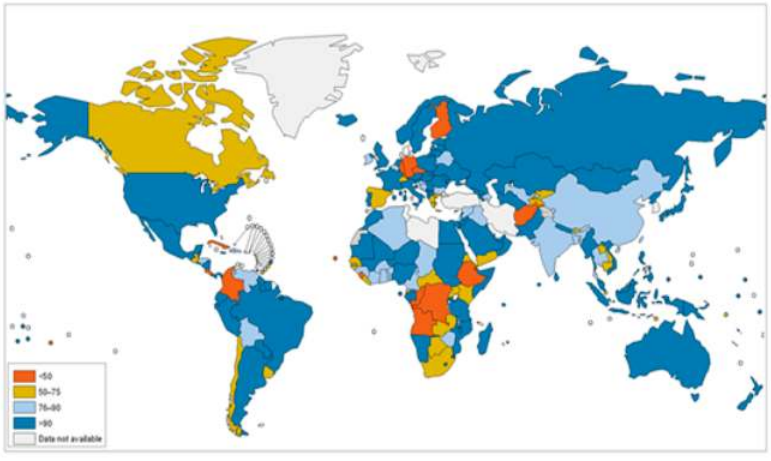


Figure **7.14.1** Proportion of Population by Country Using Improved Drinking Water Sources in 2008. Improved drinking water sources, e.g., household connections, public standpipes, boreholes, protected dug wells and springs, and rainwater collections, are defined as those more likely to provide safe water than unimproved water sources, e.g., unprotected wells and springs, vendor-provided water, bottled water (unless water for other uses is available from an improved source), and tanker truck-provided water. *Source:* *World Health Organization*

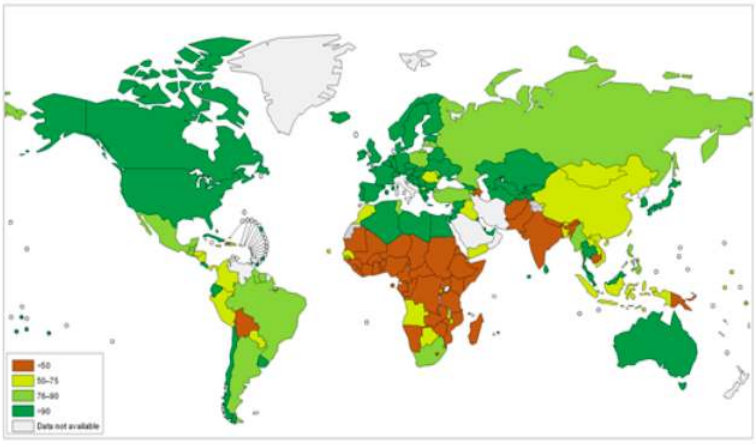


Figure **7.14.2** Proportion of Population by Country Using Improved Sanitation Facilities in 2008. Improved sanitation facilities, e.g., connection to public sewers or septic systems, pour-flush latrines, pit latrines, and ventilated improved pit latrines, are defined as those more likely to be sanitary than unimproved facilities, e.g., bucket latrines, public latrines, and open pit latrines. *Source:* [World Health Organization]

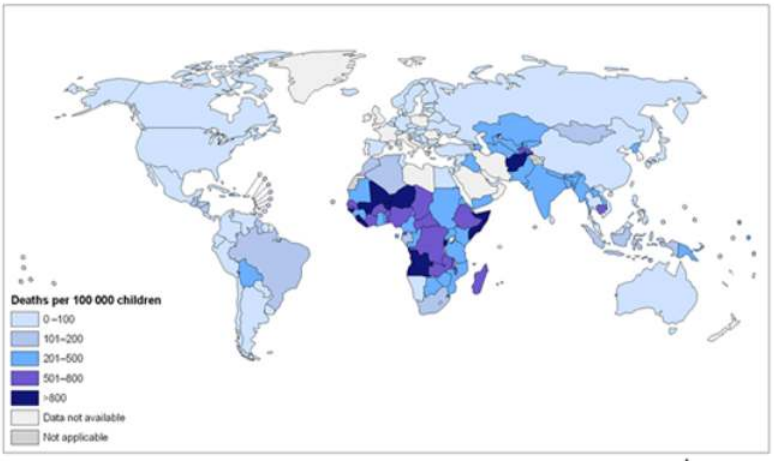


Figure **7.14.3** Deaths by Country from Diarrhea Caused by Unsafe Water, Unimproved Sanitation, and Poor Hygiene in Children Less than 5 Years Old, 2004. *Source:* [World Health Organization]

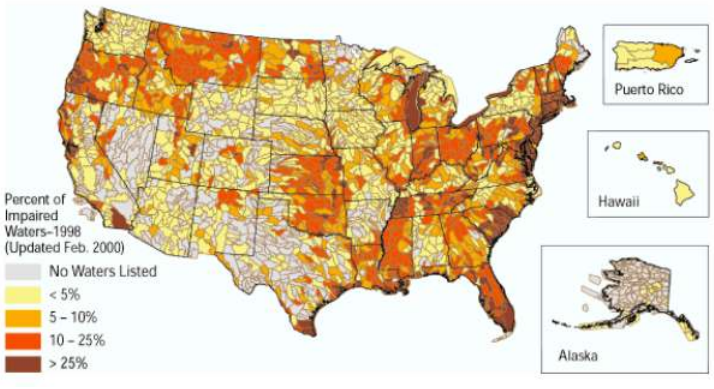


Figure **7.14.4** Percentage of Impaired Water Bodies in a Watershed by State in USA Based on US EPA Data in 2000. Map of watersheds containing impaired water bodies from the U.S. Environmental Protection Agency's 1998 list of impaired waters *Source:* [U.S. Geological Survey]

Water Chemistry Overview

Compared to other molecules of similar molecular weight, water (H_2O) has unique physical properties including high values for melting and boiling point, surface tension (water's cohesion, or "stickiness"), and capacity to dissolve soluble minerals, i.e., act as a solvent. These properties are related to its asymmetrical structure and *polar nature*, which means it is electrically neutral overall but it has a net positive charge on the side with the two hydrogen atoms and a net negative charge on the oxygen side (see Figure

7.14.5). This separation of the electrical charge within a water molecule results in *hydrogen bonds*with other water molecules, mineral surfaces (hydrogen bonding produces the water films on minerals in the unsaturated zone of the subsurface), and *dissolved ions* (atoms with a negative or positive charge). Many minerals and pollutants dissolve readily in water because water forms *hydration shells* (spheres of loosely coordinated, oriented water molecules) around ions.

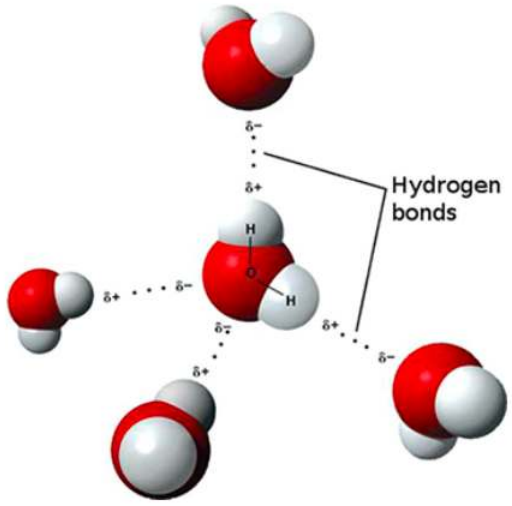


Figure **7.14.5** Structure of Water, Polar Charge of Water, and Hydrogen Bonds between Water Molecules. *Source: Michal Maňas at Wikimedia Commons*

Any natural water contains dissolved chemicals; some of these are important human nutrients, while others can be harmful to human health. The abundance of a water pollutant is commonly given in very small concentration units such as parts per million (ppm) or even parts per billion (ppb). An arsenic concentration of 1 ppm means 1 part of arsenic per million parts of water. This is equivalent to one drop of arsenic in 50 liters of water. To give you a different perspective on appreciating small concentration units, converting 1 ppm to length units is 1 cm (0.4 in) in 10 km (6 miles) and converting 1 ppm to time units is 30 seconds in a year. Total dissolved solids (TDS) represent the total amount of dissolved material in water. Average TDS (salinity) values for rainwater, river water, and seawater are about 4 ppm, 120 ppm, and 35,000 ppm. As discussed in Module 5.2, the most important processes that affect the salinity of natural waters are evaporation, which distills nearly pure water and leaves the dissolved ions in the original water, and chemical weathering, which involves mineral dissolution that adds dissolved ions to water. Fresh water is commonly defined as containing less than either 1,000 or 500 ppm TDS, but the US Environmental Protection Agency (EPA) recommends that drinking water not exceed 500 ppm TDS or else it will have an unpleasant salty taste.

Water Pollution Overview

Water pollution is the contamination of water by an excess amount of a substance that can cause harm to human beings and 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 (see Figure 7.14.6). Water pollutants can move through different water reservoirs, as the water carrying them progresses through stages of the water cycle (see Figure 7.14.7). Water residence time (the average time that a water molecule spends in a water reservoir) is very important to pollution problems because it affects pollution potential. Water in rivers has a relatively short residence time, so pollution usually is there only briefly. Of course, pollution in rivers may simply move to another reservoir, such as the ocean, where it can cause further problems. Groundwater is typically characterized by slow flow and longer residence time, which can make groundwater pollution particularly problematic. Finally, pollution residence time can be much greater than the water residence time because a pollutant may be taken up for a long time within the ecosystem or absorbed onto sediment.



Figure 7.14.6 Water Pollution. Obvious water pollution in the form of floating debris; invisible water pollutants sometimes can be much more harmful than visible ones. *Source: Stephen Codrington at Wikimedia Commons*

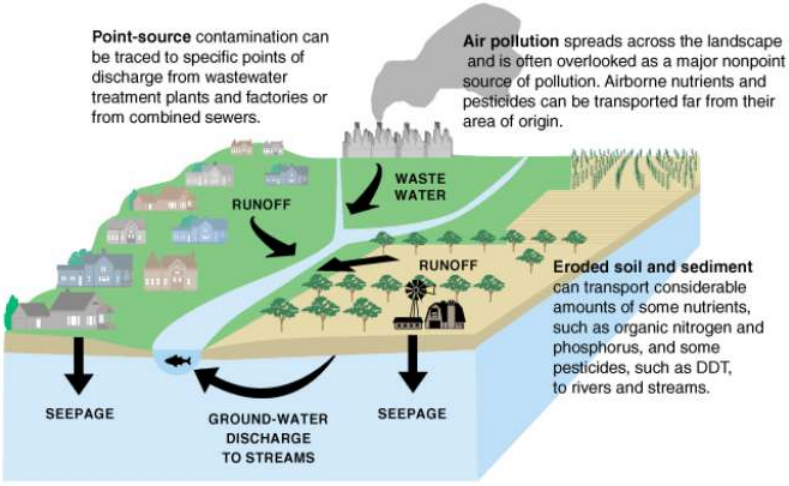


Figure 7.14.7 Sources of Water Contamination. Sources of some water pollutants and movement of pollutants into different water reservoirs of the water cycle. *Source: U.S. Geological Survey*

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. Point sources of pollution include animal "factory" farms that raise a large number and high density of livestock such as cows, pigs, and chickens (see Figure 7.14.8) and discharge 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 into surface waters (see Figure 7.14.9). Nonpoint sources of pollution include agricultural fields, cities, and abandoned mines. Rainfall runs over the land and through the ground, picking up pollutants such as herbicides, pesticides, and fertilizer from agricultural fields and lawns; oil, antifreeze, car detergent, animal waste, and road salt from urban areas; and acid and toxic elements from abandoned mines. Then, this pollution is carried into surface water bodies and groundwater. Nonpoint source pollution, which is the leading cause of water pollution in the U.S., is

usually much more difficult and expensive to control than point source pollution because of its low concentration, multiple sources, and much greater volume of water.



Figure 7.14.8 A Commercial Meat Chicken Production House. This chicken factory farm is a possible major point source of water pollution. *Source: Larry Rana at Wikimedia Commons*

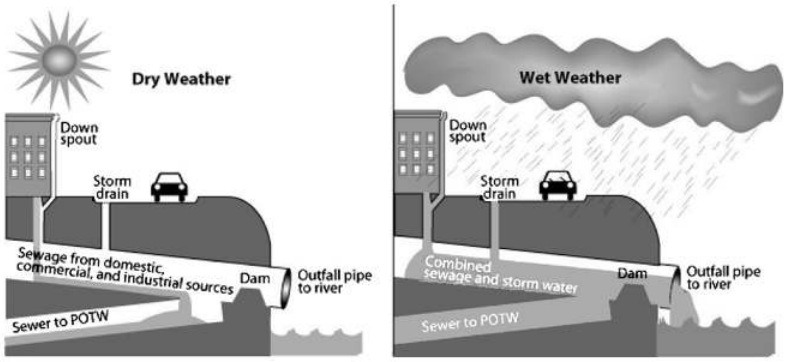


Figure 7.14.9 Combined Sewer System. A combined sewer system is a possible major point source of water pollution during heavy rain due to overflow of untreated sewage. During dry weather (and small storms), all flows are handled by the publicly owned treatment works (POTW). During large storms, the relief structure allows some of the combined stormwater and sewage to be discharged untreated to an adjacent water body. *Source: U.S. Environmental Protection Agency at Wikimedia Commons*

Types of Water Pollutants

Oxyden-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 (chemically represented in a simplified way as CH_2O) and remove dissolved oxygen (O_2) according to the following reaction:

$$CH2O + O2 \rightarrow CO2 + H2O$$

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 most natural waters is sewage; grass and leaves are smaller sources. An unpolluted water body with respect to oxygen 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, e.g., bass, trout, and perch dominate. 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, e.g., 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*, 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 P-based detergent) can cause cultural eutrophication, which involves the rapid growth of aquatic plants, particularly algae, called an *algal bloom*. Thick mats of floating and rooted green or sometimes red algae (see Figure **7.14.10**) create water pollution, damage the ecosystem by clogging fish gills and blocking sunlight, and damage lake aesthetics by making recreation difficult and creating an eyesore. A small percentage of algal species produce toxins that can kill fish, mammals, and birds, and may cause human illness; explosive growths of these algae are called *harmful algal blooms* (see Figure **7.14.11**). When the prolific algal layer dies, it becomes oxygen-demanding waste, which can create very low O_2 water (<~2 ppm O_2), called hypoxia or dead zone because it causes death 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 (see Figure **7.14.12**), 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.



Figure **7.14.10** Algal Bloom in River in Sichuan, China. Algal blooms can present problems for ecosystems and human society. *Source: Felix Andrews via Wikimedia Commons*



Figure **7.14.11** Harmful Algal Bloom. Harmful algal bloom with deep red color. *Source: Kai Schumann via National Oceanic and Atmospheric Administration*

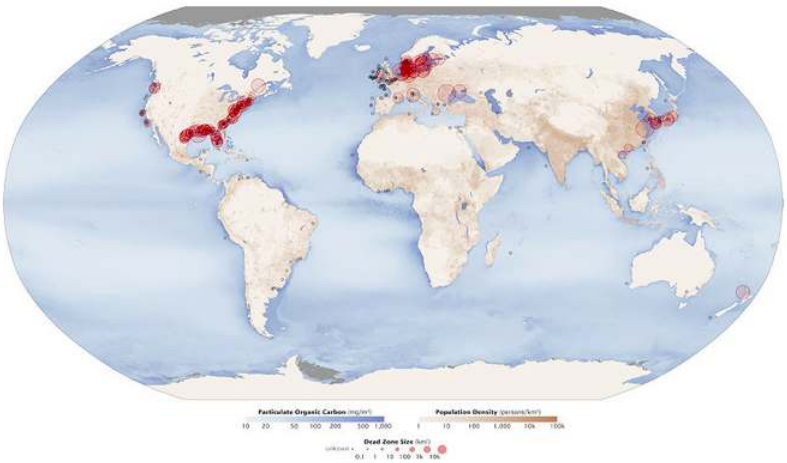


Figure **7.14.12** Aquatic Dead Zones. Zones of hypoxia shown as red circles. Black dots show hypoxia zones of unknown size, brown shading shows population density, and blue shading shows density of particulate organic carbon, an indicator of organic productivity. *Source: Robert Simmon & Jesse Allen at NASA Earth Observatory via Wikimedia Commons*

Pathogens are disease-causing microorganisms, e.g., viruses, bacteria, parasitic worms, and protozoa, which cause a variety of intestinal diseases such as dysentery, typhoid fever, hepatitis, and cholera. Pathogens are the major cause of the water pollution crisis discussed at the beginning of this section. Unfortunately nearly a billion people around the world are exposed to waterborne pathogen pollution daily and around 1.5 million children mainly in underdeveloped countries die every year of waterborne diseases from pathogens (see Figure **7.14.3**). Pathogens enter water primarily from human and animal fecal waste due to inadequate sewage treatment. In many underdeveloped countries, sewage is discharged into local waters either untreated or after only rudimentary treatment. In developed countries untreated sewage discharge can occur from overflows of combined sewer systems, poorly managed livestock factory farms, and leaky or broken sewage collection systems (see Figure **7.14.14**). Water with pathogens can be remediated by adding chlorine or ozone, by boiling, or by treating the sewage in the first place.



Figure **7.14.13** Overflowing Sanitary Sewer. A manhole cover blown off by a June 2006 sanitary sewer overflow in Rhode Island. *Source: U.S. Environmental Protection Agency via Wikimedia Commons*

Oil spills are another kind of organic pollution. Oil spills can result from supertanker accidents such as the Exxon Valdez in 1989, which spilled 10 million gallons of oil into the rich ecosystem of offshore south Alaska and killed massive numbers of animals. The largest marine oil spill was the Deepwater Horizon disaster, which began with a natural gas explosion (see Figure **7.14.14**) at an oil well 65 km offshore of Louisiana and flowed for 3 months in 2010, releasing an estimated 200 million gallons of oil. The worst oil spill ever occurred during the Persian Gulf war of 1991, when Iraq deliberately dumped approximately 200 million gallons of oil in offshore Kuwait and set more than 700 oil well fires that released enormous clouds of smoke and acid rain for over nine months. During an oil spill on water, oil floats to the surface because it is less dense than water, and the lightest hydrocarbons evaporate, decreasing the size of the spill but polluting the air. Then, bacteria begin to decompose the remaining oil, in a process that can take many years. After several months only about 15% of the original volume may remain, but it is in thick asphalt lumps, a form that is particularly harmful to birds, fish, and shellfish. Cleanup operations can include *skimmer ships* that vacuum oil from the water surface (effective only for small spills), controlled *burning* (works only in early stages before the light, ignitable part evaporates but also pollutes the air), *dispersants* (detergents that break up oil to accelerate its decomposition, but some dispersants may be toxic to the ecosystem), and *bioremediation* (adding microorganisms that specialize in quickly decomposing oil, but this can disrupt the natural ecosystem).



Figure **7.14.14** Deepwater Horizon Explosion. Boats fighting the fire from an explosion at the Deepwater Horizon drilling rig in Gulf of Mexico offshore Louisiana on April 20, 2010. *Source: United States Coast Guard via Wikimedia Commons*

Toxic chemicals involve many different kinds and sources, primarily from industry and mining. General kinds of toxic chemicals include *hazardous chemicals*, which are a wide variety of synthetic organic and inorganic chemicals such as acids, bases, cyanide,

and a class of compounds called persistent organic pollutants that includes *DDT* (pesticide), *dioxin* (herbicide by-product), and *PCBs* (polychlorinated biphenyls, which were used as a liquid insulator in electric transformers). Persistent organic pollutants are long-lived in the environment, accumulate through the food chain (bioaccumulation), and can be toxic. Another category of toxic chemicals includes *radioactive materials* such as cesium, iodine, uranium, and radon gas, which can result in long-term exposure to radioactivity if it gets into the body. A final group of toxic chemicals is heavy metals such as lead, mercury, arsenic, cadmium, and chromium, which can accumulate through the food chain. Heavy metals are commonly produced by industry and at metallic ore mines. Arsenic and mercury are discussed in more detail below. The US EPA regulates 83 contaminants in drinking water to ensure a safe public water supply. Similarly, at the international level the World Health Organization has drinking water standards for a variety of contaminants.

Arsenic (As) has been famous as an agent of death for many centuries. In large doses arsenic causes cancer and can be fatal. Only recently have scientists recognized that health problems can be caused by drinking small arsenic concentrations in water over a long time. It attacks the central nervous system and can damage the respiratory system, bladder, lungs, liver, and kidneys. It enters the water supply naturally from weathering of As-rich minerals and from human activities such as coal burning and smelting of metallic ores. The worst case of arsenic poisoning occurred in the densely populated impoverished country of Bangladesh, which had experienced 100,000s of deaths from diarrhea and cholera each year from drinking surface water contaminated with pathogens due to improper sewage treatment. In the 1970s the United Nations provided aid for millions of shallow water wells, which resulted in a dramatic drop in pathogenic diseases. Unfortunately, many of the wells produced water naturally rich in arsenic. Tragically, there are an estimated 77 million people (about half of the population) who inadvertently may have been exposed to toxic levels of arsenic in Bangladesh as a result. The World Health Organization has called it the largest mass poisoning of a population in history.

Mercury (Hg) is used in a variety of electrical products, such as dry cell batteries, fluorescent light bulbs, and switches, as well as in the manufacture of paint, paper, vinyl chloride, and fungicides. In the methylmercury form (CH_3Hg^+) it is highly toxic; ≥ 1 ppb of methylmercury represents water contaminated with mercury. Mercury concentrates in the food chain, especially in fish, in a process caused biomagnification (see Sidebar Biomagnification). It acts on the central nervous system and can cause loss of sight, feeling, and hearing as well as nervousness, shakiness, and death. Like arsenic, mercury enters the water supply naturally from weathering of Hg-rich minerals and from human activities such as coal burning and metal processing. A famous mercury poisoning case in Minamata, Japan involved methylmercury-rich industrial discharge that caused high Hg levels in fish. People in the local fishing villages ate fish up to three times per day for over 30 years, which resulted in over 2,000 deaths. During that time the responsible company and national government did little to mitigate, help alleviate, or even acknowledge the problem.

Biomagnification

Biomagnification represents the processes in an ecosystem that cause greater concentrations of a chemical, such as methylmercury, in organisms higher up the food chain. Mercury and methylmercury are present in only very small concentrations in seawater; however, at the base of the food chain algae absorb methylmercury. Then, small fish eat the algae, large fish and other organisms higher in the food chain eat the small fish, and so on. Fish and other aquatic organisms absorb methylmercury rapidly but eliminate it slowly from the body. Therefore, each step up the food chain increases the concentration from the step below (see Figure **7.14.15**). Largemouth bass can concentrate methylmercury up to 10 million times over the water concentration and fish-eating birds can concentrate it even higher. Other chemicals that exhibit biomagnification are DDT, PCBs, and arsenic.



Figure **7.14.15** An illustrative example of biomagnification of mercury from water through the food chain and into a bird's egg. *Source:* *U.S. Geological Survey*

Other water pollutants include sediment and heat. Muddy water is bad for drinking but even worse for underwater plants that need sunlight for photosynthesis. Much of the sediment in water bodies is derived from the erosion of soil, so it also represents a loss of agricultural productivity. Thermal pollution involves the release of heated waters from power plants and industry to surface water, causing a drop in the dissolved O_2 content, which can stress fish.

Hard water contains abundant calcium and magnesium, which reduces its ability to develop soapsuds and enhances *scale* (calcium and magnesium carbonate minerals) formation on hot water equipment. Water softeners remove calcium and magnesium, which allows the water to lather easily and resist scale formation. Hard water develops naturally from the dissolution of calcium and magnesium carbonate minerals in soil; it does not have negative health effects in people.

Groundwater pollution can occur from underground sources and all of the pollution sources that contaminate surface waters. Common sources of groundwater pollution are leaking underground storage tanks for fuel, septic tanks, agricultural activity, and landfills. Common groundwater pollutants include nitrate, pesticides, volatile organic compounds, and petroleum products. Polluted groundwater can be a more serious problem than polluted surface water because the pollution in groundwater may go undetected for a long time because usually it moves very slowly. As a result, the pollution in groundwater may create a contaminant plume, a large body of flowing polluted groundwater (see Figure **7.14.15**), making cleanup very costly. By the time groundwater contamination is detected, the entity responsible for the pollution may be bankrupt or nonexistent. Another troublesome feature of groundwater pollution is that small amounts of certain pollutants, e.g., petroleum products and organic solvents, can contaminate large areas. In Denver, Colorado 80 liters of several organic solvents contaminated 4.5 trillion liters of groundwater and produced a 5 km long contaminant plume. Most groundwater contamination occurs in shallow, unconfined aquifers located near the contamination source. Confined aquifers are less susceptible to pollution from the surface because of protection by the confining layer. A major threat to groundwater quality is from underground fuel storage tanks. Fuel tanks commonly are stored underground at gas stations to reduce explosion hazards. Before 1988 in the U.S. these storage tanks could be made of metal, which can corrode, leak, and quickly contaminate local groundwater. Now, leak detectors are required and the metal storage tanks are supposed to be protected from corrosion or replaced with fiberglass tanks. Currently there are around 600,000 underground fuel storage tanks in the U.S. and over 30% still do not comply with EPA regulations regarding either release prevention or leak detection.

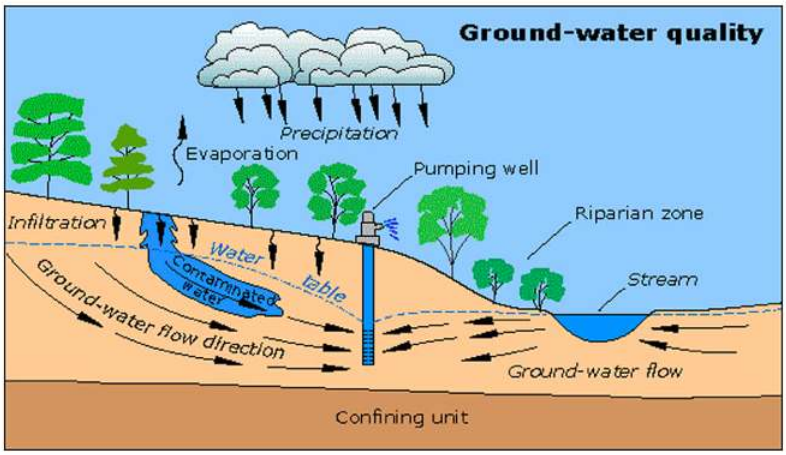


Figure **7.14.16** Contaminant Plume in Groundwater. Mapping how a contaminant plume will migrate once it reaches groundwater requires understanding of the pollutant's chemical properties, local soil characteristics, and how permeable the aquifer is. *Source:* United States Geological Survey

Sustainable Solutions to the Water Pollution Crisis?

Resolution of the global water pollution crisis described at the beginning of this section requires multiple approaches to improve the quality of our fresh water and move towards sustainability. The most deadly form of water pollution, pathogenic microorganisms that cause waterborne diseases, kills almost 2 million people in underdeveloped countries every year. The best strategy for addressing this problem is proper sewage (wastewater) treatment. Untreated sewage is not only a major cause of pathogenic diseases, but also a major source of other pollutants, including oxygen-demanding waste, plant nutrients (N and P), and toxic heavy metals. Wastewater treatment is done at a sewage treatment plant in urban areas and through a septic tank system in rural areas.

The main purpose of a sewage treatment plant is to remove organic matter (oxygen-demanding waste) and kill bacteria; special methods also can be used to remove plant nutrients and other pollutants. The numerous processing steps at a conventional sewage treatment plant (see Figure **7.14.17**) include *pretreatment* (screening and removal of sand and gravel), *primary treatment* (settling or floatation to remove organic solids, fat, and grease), *secondary treatment* (aerobic bacterial decomposition of organic solids), *tertiary treatment* (bacterial decomposition of nutrients and filtration), *disinfection* (treatment with chlorine, ozone, ultraviolet light, or bleach), and either *discharge* to surface waters (usually a local river) or *reuse* for some other purpose, such as irrigation, habitat preservation, and artificial groundwater recharge. The concentrated organic solid produced during primary and secondarytreatment is called sludge, which is treated in a variety of ways including landfill disposal, incineration, use as fertilizer, and anaerobic bacterial decomposition, which is done in the absence of oxygen. Anaerobic decomposition of sludge produces methane gas, which

can be used as an energy source. To reduce water pollution problems, separate sewer systems (where street runoff goes to rivers and only wastewater goes to a wastewater treatment plant) are much better than combined sewer systems, which can overflow and release untreated sewage into surface waters during heavy rain. Some cities such as Chicago, Illinois have constructed large underground caverns and also use abandoned rock quarries to hold storm sewer overflow. After the rain stops, the stored water goes to the sewage treatment plant for processing.

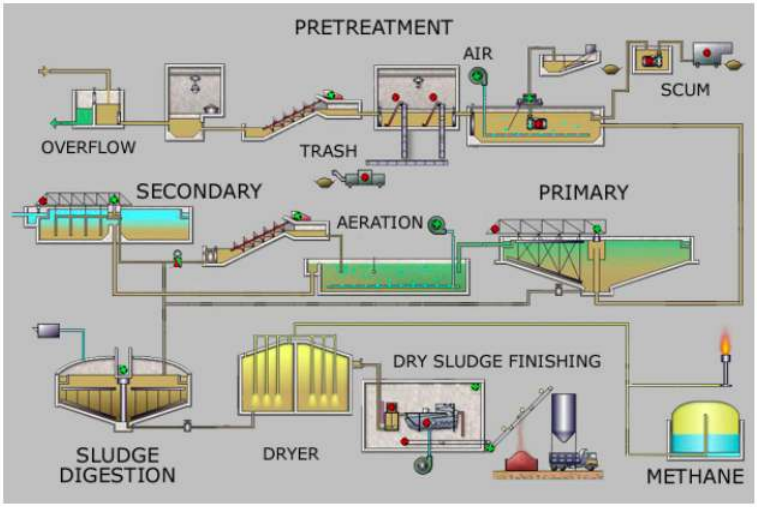


Figure **7.14.17** Steps at a Sewage Treatment Plant. The numerous processing steps at a conventional sewage treatment plant include pretreatment (screening and removal of sand and gravel), primary treatment (settling or floatation to remove organic solids, fat, and grease), secondary treatment (aerobic bacterial decomposition of organic solids), tertiary treatment (bacterial decomposition of nutrients and filtration), disinfection (treatment with chlorine, ozone, ultraviolet light, or bleach), and either discharge to surface waters (usually a local river) or reuse for some other purpose, such as irrigation, habitat preservation, and artificial groundwater recharge. *Source: Leonard G.via Wikipedia*

A septic tank system is an individual sewage treatment system for homes in rural and even some urban settings. The basic components of a septic tank system (see Figure **7.14.18**) include a sewer line from the house, a *septic tank* (a large container where sludge settles to the bottom and microorganisms decompose the organic solids anaerobically), and the drain field (network of perforated pipes where the clarified water seeps into the soil and is further purified by bacteria). Water pollution problems occur if the septic tank malfunctions, which usually occurs when a system is established in the wrong type of soil or maintained poorly.

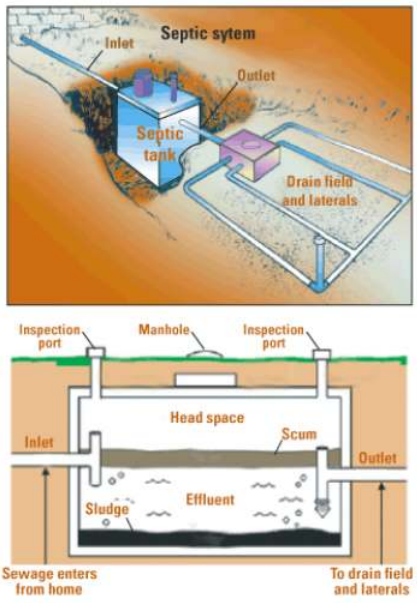


Figure **7.14.18** Septic System. Septic tank system for sewage treatment. *Source: United States Geological Survey*

For many developing countries, financial aid is necessary to build adequate sewage treatment facilities; however, the World Health Organization estimates an estimated cost savings of between $3 and $34 for every $1 invested in clean water delivery and sanitation (Water for Life, 2005). The cost savings are from health care savings, gains in work and school productivity, and deaths prevented. Simple and inexpensive techniques for treating water at home include chlorination, filters, and solar disinfection. Another alternative is to use constructed wetlands technology (marshes built to treat contaminated water), which is simpler and cheaper than a conventional sewage treatment plant.

Bottled water is *not* a sustainable solution to the water crisis, despite exponential growth in popularity in the U.S. and the world. Bottled water is not necessarily any safer than the U.S. public water supply, it costs on average about 700 times more than U.S. tap water, and every year it uses approximately 200 billion plastic and glass bottles that have a relatively low rate of recycling. Compared to tap water, it uses much more energy, mainly in bottle manufacturing and long-distance transportation. If you don't like the taste of your tap water, then please use a water filter instead of bottled water!



Figure **7.14.19** Storm Drain. Curbside storm drain receiving urban runoff. *Source: By Robert Lawton via Wikimedia Commons*

Additional sustainable solutions to the water pollution crisis include legislation to eliminate or greatly reduce point sources of water pollution. In the U.S., the Clean Water Act of 1972 and later amendments led to major improvements in water quality (see Sidebar Clean Water Act). Nonpoint sources of water pollution, e.g., agricultural runoff and urban runoff (see Figure **7.14.19**), are much harder to regulate because of their widespread, diffuse nature. There are many construction and agricultural practices that reduce polluted runoff including no-till farming and sediment traps. Artificial aeration or mechanical mixing can remediate lakes with oxygen depletion. Specific things that we can do to reduce urban runoff include the following: keep soil, leaves, and grass clippings off driveways, sidewalks, and streets; don't pour used motor oil, antifreeze, paints, pesticides, or any household hazardous chemical down the storm sewer or drain; recycle used motor oil; use hazardous waste disposal programs offered by the community; compost your organic waste; don't use fertilizers and herbicides on your lawn; and flush pet waste down the toilet.

Clean Water Act

During the early 1900s rapid industrialization in the U.S. resulted in widespread water pollution due to free discharge of waste into surface waters. The Cuyahoga River in northeast Ohio caught fire numerous times (see Figure **7.14.20**), including a famous fire in 1969 that caught the nation's attention. In 1972 Congress passed one of the most important environmental laws in U.S. history, the Federal Water Pollution Control Act, which is more commonly called the Clean Water Act. The purpose of the Clean Water Act and later amendments is to maintain and restore water quality, or in simpler terms to make our water swimmable and fishable. It became illegal to dump pollution into surface water unless there was formal permission and U.S. water quality improved significantly as a result. More progress is needed because currently the EPA considers over 40,000 U.S. water bodies as impaired, most commonly due to pathogens, metals, plant nutrients, and oxygen depletion. Another concern is protecting groundwater quality, which is not yet addressed sufficiently by federal law.


Figure 7.14.20 Cuyahoga River on Fire. *Source: National Oceanic and Atmospheric*

Sometimes slow flow through a soil can naturally purify groundwater because some pollutants, such as P, pesticides, and heavy metals, chemically bind with surfaces of soil clays and iron oxides. Other pollutants are not retained by soil particles: These include N, road salt, gasoline fuel, the herbicide atrazine, tetrachloroethylene (a carcinogenic cleaning solvent used in dry cleaning), and vinyl chloride. In other cases, slow groundwater flow can allow bacteria to decompose dead organic matter and certain pesticides. There are many other ways to remediate polluted groundwater. Sometimes the best solution is to stop the pollution source and allow natural cleanup. Specific treatment methods depend on the geology, hydrology, and pollutant because some light contaminants flow on top of groundwater, others dissolve and flow with groundwater, and dense contaminants can sink below groundwater. A common cleanup method called pump and treat involves pumping out the contaminated groundwater and treating it by oxidation, filtration, or biological methods. Sometimes soil must be excavated and sent to a landfill. In-situ treatment methods include adding chemicals to immobilize heavy metals, creating a permeable reaction zone with metallic iron that can destroy organic solvents, or using bioremediation by adding oxygen or nutrients to stimulate growth of microorganisms.

Review Questions

1. What are the major kinds of water pollutants and how do they degrade water quality?
2. How would you rank the water pollution problems described in this chapter? Why?
3. Why is untreated sewage such an important water pollutant to remediate?
4. What should society learn from the case history of Love Canal?
5. Why are people facing a crisis involving water pollution and how can we solve it?

References

Water for Life: Making it Happen (2005) World Health Organization and UNICEF. Retrieved from http://www.who.int/water_sanitation_health/waterforlife.pdf

World Health Statistics (2010) World Health Organization. Retrieved from http://www.who.int/whosis/whostat/EN_WHS10_Full.pdf

Glossary

arsenic

A type of water pollutant that can be fatal in large doses and can cause health problems in small doses over a long time.

biochemical oxygen demand

The amount of oxygen used by aerobic (in presence of oxygen) bacterial decomposition of organic matter.

bioremediation

Method of groundwater remediation involving the addition oxygen or nutrients. to stimulate growth of microorganisms, which decompose an organic pollutant.

bottled water

Drinking water packaged in plastic bottles or glass bottles, bottled water is not a sustainable solution to the water crisis because of the nonrenewable energy and material resources involved in manufacturing and transporting it.

combined sewer systems

A single set of underground pipes used to collect both sewage and. storm water runoff from streets for wastewater treatment.

constructed wetland

Marsh built to treat contaminated water.

contaminant plume

A large body of flowing polluted groundwater.

cultural eutrophication

Rapid aquatic plant growth, particularly algae, in a surface water body.

excessive plant nutrient

A type of water pollutant involving a limiting plant nutrient that usually is present in water at low concentrations and therefore restricts the total amount of plant growth, examples include nitrogen and phosphorous.

hard water

Water with abundant calcium and magnesium, which reduces its ability to develop soapsuds and enhances scale; hard water does not have negative health effects in people.

heat

A type of water pollutant that causes a drop in the dissolved oxygen content, which can stress fish.

heavy metal

A type of water pollutant involving elements such as lead, mercury, arsenic, cadmium, and chromium, which can accumulate through the food chain.

hypoxia

Very low oxygen water due to prolific growth of algae, algal death, and then decomposition, also called dead zone.

mercury

A type of water pollutant that acts on the central nervous system and can cause loss of sight, feeling, and hearing as well as nervousness, shakiness, and death.

nonpoint source (of water pollution)

Large and diffuse location where a pollution source occurs.

oil spill

A type of organic water pollutant involving the release of liquid petroleum into the environment due to human activity.

oxygen-demanding waste

A type of water pollutant involving abundant dead organic matter.

pathogens

Disease-causing microorganisms, e.g., viruses, bacteria, parasitic worms, and protozoa, which cause a variety of intestinal diseases such as dysentery, typhoid fever, hepatitis, and cholera.

persistent organic pollutant

A group of organic water pollutants that are long-lived in the environment, accumulate through the food chain, and can be toxic.

point source (of water pollution)

Readily identifiable and relatively small location where a pollution source occurs.

sediment

A type of water pollutant that degrades drinking water and can kill underwater plants that need sunlight for photosynthesis.

septic tank system

An individual sewage treatment system for homes in rural and even some urban settings.

sewage treatment plant

A facility that processes wastewater with the main goal of removing organic matter (oxygen-demanding waste) and killing bacteria.

sludge

Concentrated organic solid produced during primary and secondarytreatment of sewage treatment.

solvent

Capacity of a liquid such as water to dissolve soluble minerals.

total dissolved solids

Total amount of dissolved material in water, typically reported in parts per million (ppm) units.

toxic chemical

A type of organic water pollutant involving chemicals with a severe human health risk.

underground fuel storage tank

A type of water pollutant if it leaks.

water pollution

Contamination of water by an excess amount of a substance that can cause harm to human beings and the ecosystem.

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7.15: Water Treatment

Water treatment is` a process of making water suitable for its application or returning its natural state. Thus, water treatment required before and after its application. The required treatment depends on the application. For example, treatment of greywater (from bath, dish and wash water) differs from the black water (from flush toilets). Composting toilet is not allowed in urban dwelling. Yet, composting toilets are used in a 30,000-square-foot office complex at the Institute of Asian Research, University of British Columbia.

Water treatment involves science, engineering, business, and art. The treatment may include mechanical, physical, biological, and chemical methods. As with any technology, science is the foundation, and engineering makes sure that the technology works as designed. The appearance and application of water is an art.

In terms of business, RGF Environmental, Water Energy Technologies, Aquasana Store, Vitech, Recalyx Industrial SDN BHD and PACE Chemicals ltd are some of many companies that offer various processes for water treatment. Millipore, a Fisher Scientific partner, offers many lines of products to produce *ultrapure water*, using a combination of active charcoal membranes, and reverse osmosis filter. Internet sites of these companies offer useful information regarding water.

An environmental scientist or consultant matches the service provider, modify if necessary, with the requirement.

- Natural Water includes some discussion on hard and soft water. Softening hard water for boiler, cooler, and domestic application is discussed therein. These treatments prepare water so that it is suitable for the applications.
- Water Biology deals with water and biology. Drinking water is part of making water suitable for living. Thus, this link gives some considerations to drinking water problems.
- There are many different industry types, and waters from various sources are usually treated before and after their applications. Pre-application treatment and wastewater treatment offer a special opportunity or challenge. Only a general consideration will be given to some industrial processes.
- General municipal and domestic wastewater treatment converts used water (waste) into environmentally acceptable water or even drinking water. Every urban centre requires such a facility.

General Wastewater Treatment

Water is a renewable resource. All water treatments involve the removal of solids, bacteria, algae, plants, inorganic compounds, and organic compounds. Removal of solids is usually done by filtration and sediment. Bacteria digestion is an important process to remove harmful *pollutants*. Converting used water into environmentally acceptable water or even drinking water is **wastewater treatment**

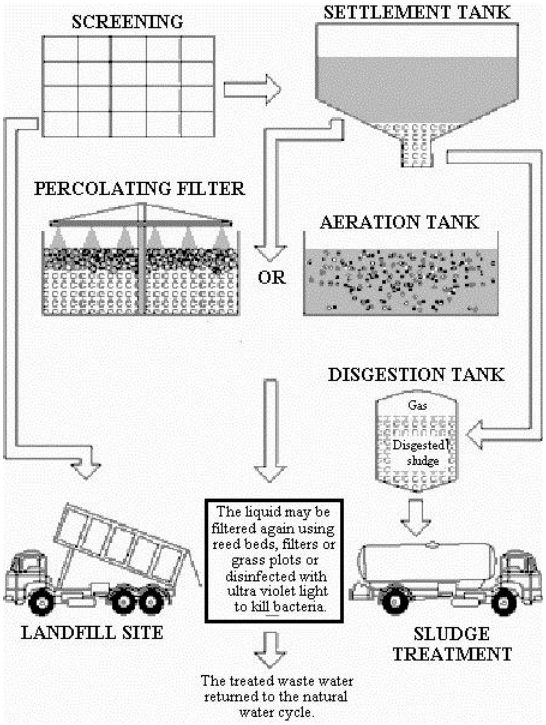
Water in the Great Lakes Region is an organization dealing with the water resources. Ontario Clean Water Agency (OCWA) is a provincial Crown corporation in business to provide environmentally responsible and cost-efficient water and wastewater services. It currently operates more than 400 facilities for 200 municipalities. This web site provides information on water and water treatment.

In April 1993, 403,000 people in Milwaukee were ill as a result of *cryptosporidium* contaimination of water due to spring run off. This outbreak caused the more stringent regulations to be implemented in the public dringking water system. The measures were aimed at removing cryptosporidium.

In May 2000, due to torrential downpour surface water got into shallow wells in a small town Walkerton, Ontario, Canada. On May 17, some residents complained of fever, bloody diarrhea and vomiting. This was know as the Walkerton *E. Coli* Outbreak. Nearly half of the population of the town fell ill, and several people died due to the E. Coli O157:H7 infection. A public inquiry recommended many measures to prevent similar outbreaks. These measures were aimed at eliminating E. Coli.

Sewage Treatment

As a general discussion, let us look at a typical process in sewage treatment. A flow diagram for a general sewage treatment plant from Water Education, Department of Computer Science, University of Exeter, U.K., is shown below:



Sewage is SCREENED to remove large solid chunks, which are disposed in LANDFILL SITE. It flows over to the SETTLEMENT TANK to let the fine particles to settle. The settlement is called the activated SLUDGE. The supernatant is then PERCOLATING FILTERED and/or AERATED. The water can be filtered again, and then disinfected (chlorinated in most cases). When there is no other complication, the water is returned to nature back to the ecological cycle.

The SLUDGE removed from the settlement is composed of living biological material. A portion of it may be returned to the AERATION TANK, but the raw SLUDGE is digested by both microorganism. Anaerobic (without oxygen) and aerobic (with air) bacteria digestions are used. At the digestion stage, carbon dioxide, ammonia, and methane gases are evolved. Volume of the digested sludge is reduced, and it is acceptable as a fertilizer supplement in farming.

Wastewater Treatment

Although the sewage water may be discharged back to the ecological system after AERATED DIGESTION and PERCOLATING FILTRATION, but in some cases, further treatment is required. Some general consideration of water treatment is given below.

A rather recent book, *Chemistry of Water Treatment* by S.D. Faust and O.M. Aly, 2nd Ed. (1998) [TD433 F38 1998], addresses the problem of quality natural and treated water.

The first three chapters discuss the criteria and standards for drinking water quality, organic compounds in waters, taste and order of water. Understandably, the standards change over the years. So are the standards of treated waters. Guidelines are available from government agencies such as Environment Canada which is equivalent to U.S. Public Health Service and the Environment Protection Agency (EPA). We have talked about drinking water in Water Biology.

Next seven chapters deal with the removal of the following:

- organics and inorganics by activated carbon
- particulate matter by coagulation
- particulate matter by filtration and sedimentation
- hardness and other scale-forming substances
- inorganic contaminants
- corrosive substances
- pathogenic (disease producing) bacteria, viruses, and protozoans (microorganisms).

There is a chapter dealing with **aeration**

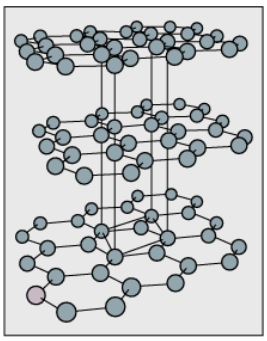
These items cover the chemistry, biology, and physics involved in the treatment of water. Some of these topics have been discussed in chemistry of water, physical properties of water, biology of water, and natural water. Introductions are going to be given to some selected topics below.

Treatment by Activated Carbon

Treatment by activated carbon is mostly due to adsorption or absorption. When a chemical species is adhered to the surface of a solid, it is an **adsorption**. When partial chemical bonds are formed between adsorbed species or when the absorbate got into the channels of the solids, we call it **absorption**. However, these two terms are often used to mean the same, because to distinguish one from type from the other is very difficult.

Application of activated charcoal for the removal of undesirable order and taste in drinking water has been recognized at the dawn of civilization. Using bone char and charred vegetation, gravel, and sand for the filtration of water for domestic application has been practised for thousands of years. Active research and production of activated charcoal was accelerated during the two world wars. The use of poison gas prompted the development of masks. They are still in use today.

Charcoal absorbs many substances, ranging from colored organic particulates to inorganic metal ions. Charcoal has been used to remove the colour of raw sugar from various sources.



Charcoal consists of microcrystallites of graphite. The particles are so small in charcoal that they were considered amorphous. The crystal structure of graphite consists of layers of hexagonal networks, stacked on top of each other. Today, making activated carbon is a new and widely varied industry. Other molecules attach themselves to the porous surface and dangling carbons in these microcrystallites.

Carbon containing substances are charred at less than 900 K to produce carbon in the manufacture of **activated carbon**. However, the carbon is activated at 1200 K using oxidizing agent to selectively oxidize portions of the char to produce pores in the material. Because of the special process to produce used, these materials with high surface to mass ratio, they are called **activated carbon** rather than activated charcoal. Factors affecting the absorption are particle size, surface area, pore structure, acidity (pH), temperature, and the nature of the material to be absorbed. Usually, adsorption (absorption) equilibria and rate of adsorption must be considered for effective applications.

Coagulation, Flucculation and Sedimentation

Natural and wastewater containing small particulates. They are suspended in water forming a **colloid**. These particles carry the same charges, and repulsion prevents them from combining into larger particulates to settle. Thus, some chemical and physical techniques are applied to help them settle. The phenomenon is known as **coagulation**. A well known method is the addition of electrolyte. Charged particulates combine with ions neutralizing the charges. The neutral particulates combine to form larger particles, and finally settle down.

Another method is to use high-molecular-weight material to attract or trap the particulates and settle down together. Such a process is called **flocculation**. Starch and multiply charged ions are often used.

Historically, dirty water is cleaned by treating with alum, $Al_2(SO_4)_3 \cdot 12\ H_2O$, and lime, $Ca(OH)_2$. These electrolytes cause the pH of the water to change due to the following reactions:

$$Al2(SO4)3.12\ H2O, \rightarrow Al(aq)3+ + 3\ SO4(aq)2- + 12\ H2O$$
$$SO4(aq)2- + H2O \rightarrow HSO4(aq)- + OH- \text{ (causing pH change)}$$
$$Ca(OH)2 \rightarrow Ca(aq)2+ + 2\ OH- \text{ (causing pH change)}$$

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 let the water sit around to let the floculated or coagulated particles to settle out. It works best with relatively dense particles (e.g. silt and minerals), while flotation works better for lighter particles (e.g. algae, color). A settling tank should be big enough so that it takes a long time (ideally 4 hours +) to get through. Inlets and outlets are designed so the water moves slowly in the tank. Long and narrow channels are installed to let the water to snake its way through the tank. The settled particles, sludge, must occasionally be removed from the tanks. The water is next ready to be filtered. Sedimentation is used in pre-treatment and wastewater treatment.

Filtration

Filtration is the process of removing solids from a fluid by passing it through a porous medium. Coarse, medium, and fine porous media have been used depending on the requirement. The filter media are artificial membranes, nets, sand filter, and high technological filter systems. The choice of filters depends on the required filtering speed and the *cleanness* requirement. The flow required for filtration can be achieved using gravity or pressure. In pressure filtration, one side of the filter medium is at higher pressure than that of the other so that the filter plane has a pressure drop. Some portion of this filter type must be enclosed in a container.

The process of removing the clogged portion of the filter bed by reversing the flow through the bed and washing out the solid is called **back washing**. During this process, the solid must be removed out of the system, but otherwise the filters must be either replaced or taken out of service to be cleaned.

Aqua-Rain manufacture water filters as shown here. This unit consists of four filters. Regarding the filtering system, its technical info gave the following statement.



> *At the heart of the AquaRain™ Water Filtration System are Marathonr State-of-the-art ceramic elements utilizing a long-proven filtration process that is over 100 years old which will safely remove dangerous waterborne pathogens such as cysts (Cryptospordium, Giardia lamblia) and bacteria (E. coli, Samonelli typhus, etc...). These innovative Marathonr ceramic elements are also filled with a high grade silvered granulated activated carbon (GAC). The GAC reduces pesticides, chemicals, chlorine, tastes & odors, while leaving the naturally occurring minerals found in the water unaffected.*

The units are designed for emergency and perhaps undeveloped countries.

AquaSelect of Mississauga has a pitcher water filter system, and *its cartridge contains hundreds of high efficiency activated carbon and ion exchange beads*, its web site claims. Brita filters is very popular.

Aeration

Bringing air into intimate contact with water for the purpose of exchanging certain components between the two phases is called **aeration**. Oxygenation is one of the purposes of aeration. Others are removal of volatile organic substances, hydrogen sulfide, ammonia, and volatile organic compounds.

A gas or substance dissolved in water may further react with water. Such a reaction is called **hydration**. Ionic substance dissolve due to hydration, for example:

$$\mathrm{HCl(g) + \mathit{x}\,H_2O \rightleftharpoons H(H_2O)_{\mathit{x}}{}^{+} + Cl(aq)^{-}} \tag{7.15.1}$$

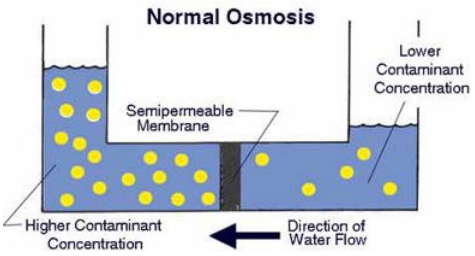
$$\mathrm{H_2S \rightleftharpoons H^{+}(aq) + HS^{-}(aq)} \tag{7.15.2}$$

These reactions are reversible, and aeration may also causes dehydration resulting in releasing the gas from water. Henry's law is applicable to this type of equilibrium for consideration. Methods of aeration are

- Diffused aeration - Air bubbles through water.

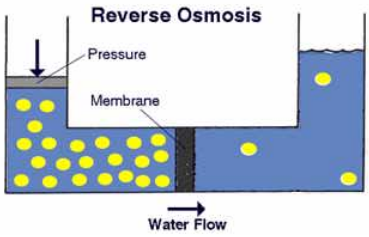
- Spray aeration - Water is sprayed through air.
- Multiple-tray aeration - Water flows through several trays to mix with air.
- Cascade aeration - Water flows downwards over many steps in the form of thin water falls.
- Air stripping - A combination of multiple tray and cascade technique plus random packed blocks causing water to mix thoroughly with air.

Reverse Osmosis Water Filter System



In the following discussion, a dilute solution and a concentrated solution are considered. The dilute solution can be a clean water whereas the concentrated solution contains undesirable solute (electrolyte or others).

When a compartment containing a dilute solution is connected to another compartment containing a concentrated solution by a semipermeable membrane, water molecules move from the dilute solution to concentrated solution. This phenomenon is called osmosis. Pig bladders are natural semipermeable membranes. As the water molecules migrate through the semipermeable membrane, water level in the solution will increase until the (osmotic) pressure prevents a net migration of water molecules in one direction. A pressure equivalent to the height difference is called the osmotic pressure. The illustration given on the right is from the PurePro, one of the many companies that manufacture reverse osmosis water filter devices. Millipore also use this technique.



By applying pressure in the higher concentration solution, water molecules migrate from a high concentration solution to a low concentration solution. This method is called **reverse osmosis water filter system**. The concept of reverse osmosis is illustrated in the diagram here from PurePro.

In this technique, the membrane must be able to tolerate the high pressure, and prevent solute molecules to pass through. Regarding membranes, PurePro made the following statement:

> *Semipermeable membranes have come a long way from the natural pig bladders used in the earlier osmosis experiments. Before the 1960's, these membranes were too inefficient, expensive, and unreliable for practical applications outside the laboratory. Modern advances in synthetic materials have generally solved these problems, allowing membranes to become highly efficient at rejecting contaminants, and making them tough enough to withstand the greater pressures necessary for efficient operation.*

This technology certainly works, and it has been used to convert salt (ocean or sea) water into fresh water. With this technique, the water with higher concentration is discharged. Thus, this technology is costly in regions where the water cost is high. Free Drinking Water also uses reverse osmosis filter system for domestic applications.

Industrial Wastewater Treatment

The Environment Canada's (Atlantic Region) Waste Management and Remediation Section contains links to detailed information on the programs and activities relating to Petroleum and Allied Petroleum Storage Tank Systems, Ocean Disposal, Contaminated Sites, and Hazardous Waste Disposal Advice for the provinces of Newfoundland and Labrador, Nova Scotia, New Brunswick and Prince Edward Island. It also contains links to sites within and outside of Environment Canada with related information.

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7.16: Case Study: The Love Canal Disaster

One of the most famous and important examples of groundwater pollution in the U.S. is the Love Canal tragedy in Niagara Falls, New York. It is important because the pollution disaster at Love Canal, along with similar pollution calamities at that time (Times Beach, Missouri and Valley of Drums, Kentucky), helped to create Superfund, a federal program instituted in 1980 and designed to identify and clean up the worst of the hazardous chemical waste sites in the U.S.

Love Canal is a neighborhood in Niagara Falls named after a large ditch (approximately 15 m wide, 3–12 m deep, and 1600 m long) that was dug in the 1890s for hydroelectric power. The ditch was abandoned before it actually generated any power and went mostly unused for decades, except for swimming by local residents. In the 1920s Niagara Falls began dumping urban waste into Love Canal, and in the 1940s the U.S. Army dumped waste from World War II there, including waste from the frantic effort to build a nuclear bomb. Hooker Chemical purchased the land in 1942 and lined it with clay. Then, the company put into Love Canal an estimated 21,000 tons of hazardous chemical waste, including the carcinogens benzene, dioxin, and PCBs in large metal barrels and covered them with more clay. In 1953, Hooker sold the land to the Niagara Falls school board for $1, and included a clause in the sales contract that both described the land use (filled with chemical waste) and absolved them from any future damage claims from the buried waste. The school board promptly built a public school on the site and sold the surrounding land for a housing project that built 200 or so homes along the canal banks and another 1,000 in the neighborhood (see Figure **7.16.1**). During construction, the canal's clay cap and walls were breached, damaging some of the metal barrels.



Figure **7.16.1** Love Canal. *Source:* *US Environmental Protection Agency*

Eventually, the chemical waste seeped into people's basements, and the metal barrels worked their way to the surface. Trees and gardens began to die; bicycle tires and the rubber soles of children's shoes disintegrated in noxious puddles. From the 1950s to the late 1970s, residents repeatedly complained of strange odors and substances that surfaced in their yards. City officials investigated the area, but did not act to solve the problem. Local residents allegedly experienced major health problems including high rates of miscarriages, birth defects, and chromosome damage, but studies by the New York State Health Department disputed that. Finally, in 1978 President Carter declared a state of emergency at Love Canal, making it the first human-caused environmental problem to be designated that way. The Love Canal incident became a symbol of improperly stored chemical waste. Clean up of Love Canal, which was funded by Superfund and completely finished in 2004, involved removing contaminated soil, installing drainage pipes to capture contaminated groundwater for treatment, and covering it with clay and plastic. In 1995, Occidental Chemical (the modern name for Hooker Chemical) paid $102 million to Superfund for cleanup and $27 million to Federal Emergency

Management Association for the relocation of more than 1,000 families. New York State paid $98 million to EPA and the US government paid $8 million for pollution by the Army. The total clean up cost was estimated to be $275 million. The only good thing about the Love Canal tragedy is that it helped to create Superfund, which has analyzed tens of thousands of hazardous waste sites in the U.S. and cleaned up hundreds of the worst ones. Nevertheless, over 1,000 major hazardous waste sites with a significant risk to human health or the environment are still in the process of being cleaned.

Glossary

superfund

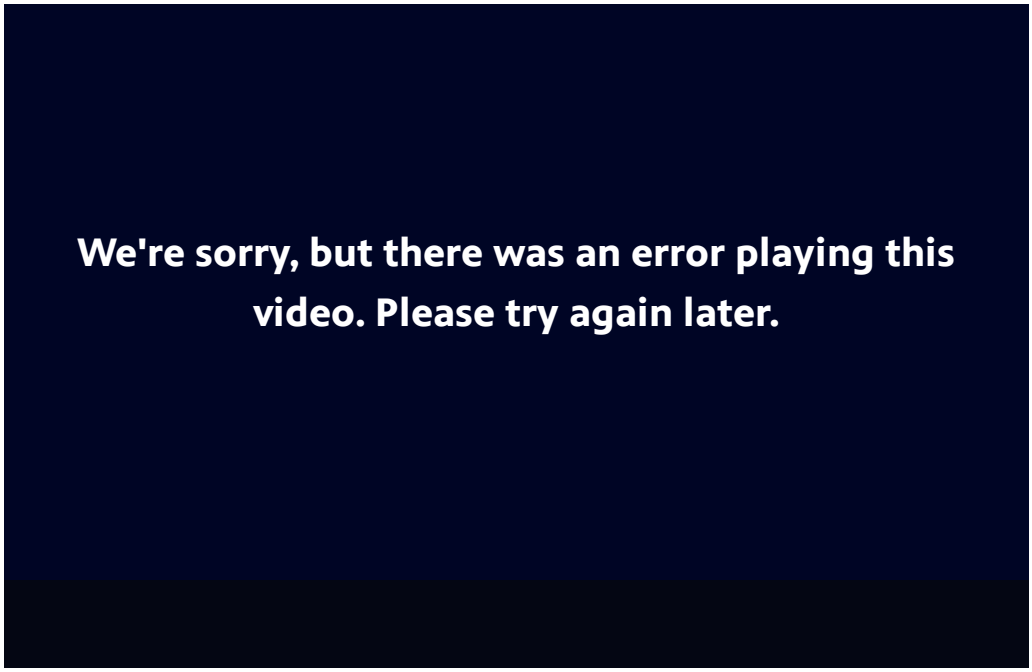
A federal program created in 1980 and designed to identify and clean up the worst of the hazardous chemical waste sites in the U.S.

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7.17 The Bull Run River

The Bull Run River supplies water to the city of Portland.



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8.1: Energy

Learning Objectives

- Define *energy*.
- Know the units of energy.
- Understand the law of conservation of energy.

Energy is the ability to do work. Think about it: when you have a lot of energy, you can do a lot of work; but if you're low on energy, you do not want to do much work. Work (w) itself is defined as a force (F) operating over a distance (Δx):

$$w = F \times \Delta x$$

In SI, force has units of newtons (N), while distance has units of meters. Therefore, work has units of N·m. This compound unit is redefined as a **Joule (J)**.

$$1 \text{ joule} = 1 \text{ newton·meter} 1 \text{ J} = 1 \text{ N·m}$$

Because energy is the ability to do work, energy is also measured in joules. This is the primary unit of energy we will use here.

How much is 1 J? It is enough to warm up about one-fourth of a gram of water by 1°C. It takes about 12,000 J to warm a cup of coffee from room temperature to 50°C. So a joule is not a lot of energy. It will not be uncommon to measure energies in thousands of joules, so the kilojoule (kJ) is a common unit of energy, with 1 kJ equal to 1,000 J.

An older—but still common—unit of energy is the *calorie*. The calorie (cal) was originally defined in terms of warming up a given quantity of water. The modern definition of calorie equates it to joules:

$$1 \text{ cal} = 4.184 \text{ J}$$

The calorie is used when considering nutrition. Energy contents of foods are often expressed in calories. However, the calorie unit used for foods is actually the kilocalorie (kcal). Most foods indicate this by spelling the word with a capital C—Calorie. Figure 8.1.1 - Calories on Food Labels, shows one example.



Nutrition Facts

Serving Size 3 oz. (85g)

Amount Per Serving	As Served
Calories 38	**Calories from Fat** 0
	% Daily Value
Total Fat 0g	0%
Saturated Fat 0g	0%
Cholesterol 0g	0%
Sodium 0g	2%
Total Carbohydrate 0g	3%
Dietary Fiber 0g	8%
Sugars 0g	
Protein 0g	

Vitamin A 270% • **Vitamin C** 10%

Calcium 2% • **Iron** 0%

Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs:

	Calories	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	80g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

Figure 8.1.1 Calories on Food Labels © Thinkstock. *This label expresses the energy content of the food, but in Calories (which are actually kilocalories).*

Example 8.1.1

The label in Figure **8.1.1** states that the serving has 38 Cal. How many joules is this?

Solution

We recognize that with a capital C, the Calories unit is actually kilocalories. To determine the number of joules, we convert first from kilocalories to calories (using the definition of the *kilo-* prefix) and then from calories to joules (using the relationship between calories and joules). So

$$38 \cancel{kcal} \times \frac{1000 \cancel{cal}}{1 \cancel{kcal}} \times \frac{4.184\ J}{1 \cancel{cal}} = 160,000\ J$$

Exercise 8.1.1

A serving of breakfast cereal usually has 110 Cal. How many joules of energy is this?

Answer

460,000 J

In the study of energy, we use the term **system** to describe the part of the universe under study: a beaker, a flask, or a container whose contents are being observed and measured. An **isolated system** is a system that does not allow a transfer of energy or matter into or out of the system. A good approximation of an isolated system is a closed, insulated thermos-type bottle. The fact that the thermos-type bottle is closed keeps matter from moving in or out, and the fact that it is insulated keeps energy from moving in or out.

One of the fundamental ideas about the total energy of an isolated system is that it does not increase or decrease. When this happens to a quantity, we say that the quantity is *conserved*. The **law of conservation of energy** states that the total energy of an isolated system does not change. As a scientific law, this concept occupies the highest level of understanding we have about the natural universe.

Key Takeaways

- Energy is the ability to do work and uses the unit joule.
- The law of conservation of energy states that the total energy of an isolated system does not increase or decrease.



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8.2: Electricity Generation

Electricity is produced at a an electric power plant. Some fuel source, such as coal, oil, natural gas, or nuclear energy produces heat, which is used to boil water to create steam. The steam under high pressure is then used to spin a turbine that interacts with a system of magnets to produce electricity. The electricity is transmitted as moving electrons through a series of wires to homes and business.

Introduction

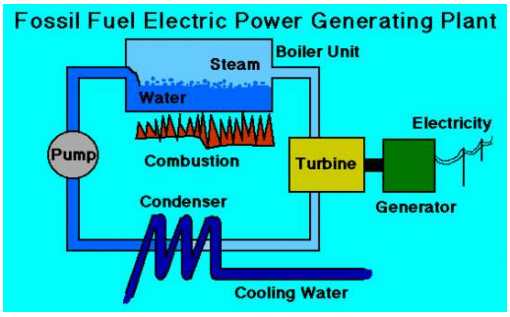
This is a typical electric power plant located in Shawville, Pennsylvania.



Notice the large pile of coal on the left side of the plant and the three smokestacks, each one taller than the previous. The tallest stack was built to cut down on the local air pollution, where the sulfur oxides are emitted higher into the atmosphere. This has not proven to be a solution to the probelm. As a result the sulfur oxides now travel great distances before coming down in the form of acid rain.

Electric Power Plants

Electric Power Plants have a number of components in common and are an interesting study in the various forms and changes of energy necessary to produce electricity.



- **Boiler Unit:** Almost all of power plants operate by heating water in a boiler unit into super heated steam at very high pressures. The source of heat from combustion reactions may vary in fossil fuel plants from the source of fuels such as coal, oil, or natural gas. Biomass or waste plant parts may also be used as a source of fuel. In some areas solid waste incinerators are also used as a source of heat. All of these sources of fuels result in varying amounts of air pollution, as well as, the carbon dioxide (a gas implicated in global warming problems).
- **Turbine-Generator**: The super heated steam is used to spin the blades of a turbine, which in turn is used in the generator to turn a coil of wires within a circular arrangements of magnets. The rotating coil of wire in the magnets results in the generation of electricity.
- **Cooling Water**: After the steam travels through the turbine, it must be cooled and condensed back into liquid water to start the cycle over again. Cooling water can be obtained from a nearby river or lake. The water is returned to the body of water 10-20 °C higher in temperature than the intake water. Alternate method is to use a very tall cooling tower, where the evaporation of water falling through the tower provides the cooling effect.
- In a **nuclear power plant**, the fission chain reaction of splitting nuclei provides the source of heat.

Creating Electricity using a Generator

If a magnetic field can create a current then we have a means of generating electricity. Experiments showed that a magnetic just sitting next to a wire produced no current flow through that wire. However, if the magnet is moving, a current is induced in the wire. The faster the magnet moves, the greater the induced current. This is the principal behind simple electric generators in which a wire loop is rotated between to stationary magnetics. This produces a continuously varying voltage which in turn produces an alternating current .

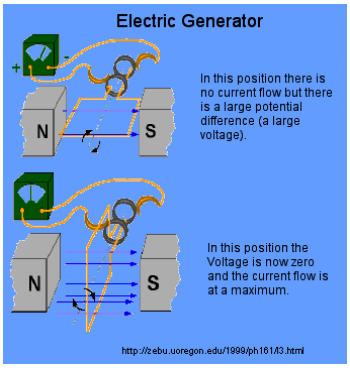


Diagram of a simple electric generator is shown above. To generate electricty then, some (mechanical) mechanism is used to turn a crank that rotates a loop of wire between stationary magnets. The faster the crank turns, the more current that is generated. In hydroelectric, the falling water turns the turbine. The wind can also turn the turbine. In fossil fuel plants and nuclear plants, water is heated to steam which turns the turbine.

Outside Links

- http://zebu.uoregon.edu/1999/ph161/l3.html

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8.3: Fossil Fuels

Fossil fuels is the term given to energy sources with a high **hydrocarbon** content (see Chapter 1 for a review of hydrocarbon molecules) found in the Earth's crust that formed in the geologic past and can be burned to release their energy. They were formed from prehistoric plants and animals that lived hundreds of millions of years ago (100 – 500 million years ago). When these ancient living organisms died they were quickly buried and subjected to immense pressure from overlying earth materials including layers of mud, rock, sand, and sometimes surface water bodies such as oceans and lakes.

During the millions of years that passed, the dead plants and animals slowly decomposed in **anaerobic** (very low to no oxygen) conditions and their chemical energy became concentrated. The organic compounds that once made up tissues of these organisms were chemically changed under high pressures and temperatures. While some fossil fuels may be in the process of formation today, the amount of time required for usable quantities to form is measured in millions of years, so these fuels will never be available for us. Thus for all practical purposes we consider fossil fuels to be finite and **non-renewable**.

4.2.1: Fossil Fuel Types and Formation

There are three main types of fossil fuels – natural gas, oil, and coal – and the specific type formed depends on the combination of organic matter that was present, how long it was buried and what temperature and pressure conditions existed when they were decomposing. **Oil** and **natural gas** were created from organisms that lived in water and were buried under ocean or river sediments. Long after the great prehistoric seas and rivers vanished, heat, pressure, and bacteria combined to compress and transform the organic material under layers of silt or shale rock (Figure **8.3.1**). In most areas, a thick liquid called oil formed first, but in deeper, hot regions underground, the transformation process continued until natural gas was formed. Over time, some of this oil and natural gas began working its way upward through the earth's crust until they ran into rock formations called "**caprocks**" that are dense enough to prevent them from seeping to the surface. It is from under these caprocks that most oil and natural gas is retrieved today.

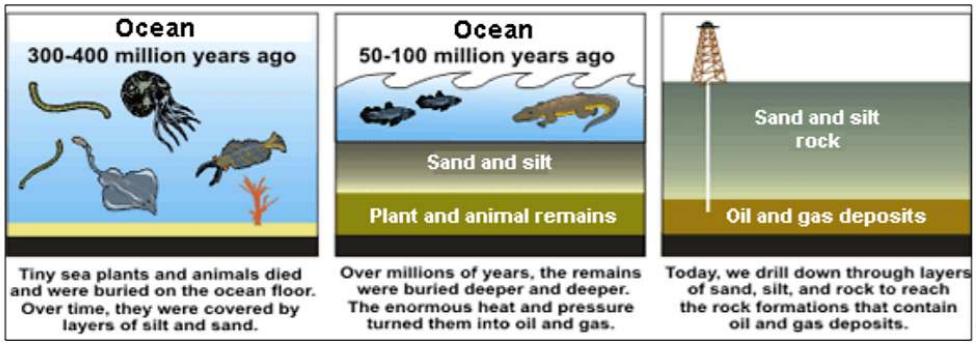


Figure **8.3.1**: Oil and natural gas (petroleum) formation. Source: U.S. Energy Information Administration. http://www.eia.gov/energyexplained/i...tural_gas_home

Coal is a fossil fuel that formed from the remains of trees, ferns, and other plants that lived 300 to 400 million years ago (Figure **8.3.2**). In some areas, such as portions of what is now the eastern United States, coal was formed from swamps covered by sea water. The sea water contained a large amount of sulfur, and as the seas dried up, the sulfur was left behind in the coal. Scientists are working on ways to take the sulfur out of coal because when coal burns, the sulfur is released in to the atmosphere as an air pollutant (see Chapter 6). Some coal deposits, however, were formed from freshwater swamps which had very little sulfur in them. These coal deposits, located largely in the western part of the United States, have much less sulfur in them.

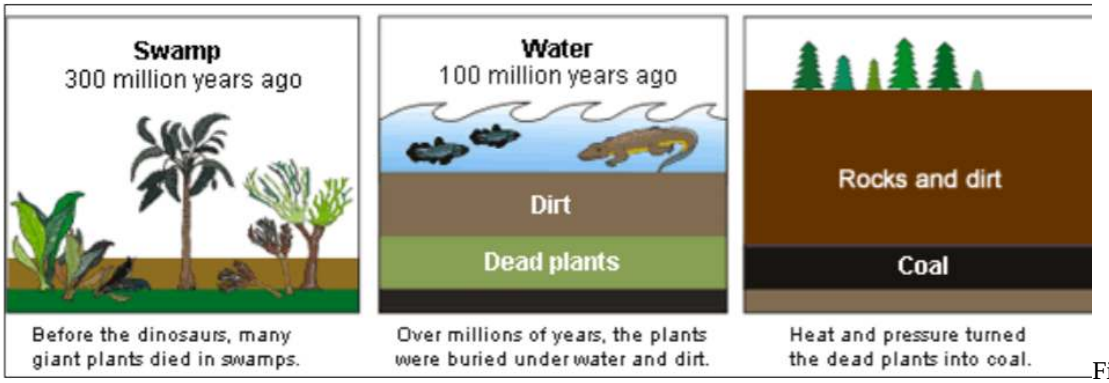


Figure **8.3.2**: The process of coal formation. Source: U.S. Energy Information Administration.

4.2.2: Consumption

Historically, human prosperity has been directly correlated with energy use. The health and vitality of world societies critically depends on energy, most of which comes from fossil fuels (Figure **8.3.3**). Energy resources, however, are unevenly distributed throughout the world, and so are the consumption rates. Developed regions generally consume far more energy than the developing regions. For example, the United States has only about 5% of the world's population but constitutes over 20% of the world's energy consumption. Additionally, developing countries devote a larger proportion of energy consumption to subsistence activities such as growing and preparing food, and heating homes. Industrialized nations rely more on mechanized equipment and technology and, therefore, a greater proportion of their energy consumption goes to transportation and industry.

Fossil fuels can be utilized without being converted or transformed to another form of energy, this is referred to as **primary energy** consumption. In their primary form, fossil fuels can be used for transportation, heating and cooking, or used to generate electricity. The use of **electricity** is a form of **secondary energy** consumption. Transforming fossil fuel energy into electricity allows for easier transportation over long distances and application to a variety of uses. Additionally, there are four major sectors that consume energy: 1) The industrial sector which includes facilities and equipment used for manufacturing, agriculture, mining, and construction; 2) The transportation sector includes vehicles that transport people or goods including cars, trucks, buses, motorcycles, trains, aircraft, boats, barges, and ships; 3) The residential sector consists of homes and apartments; 4) The commercial sector includes offices, malls, stores, schools, hospitals, hotels, warehouses, restaurants, places of worship, and more. Each of these sectors also consumes electricity produced by the electric power sector.

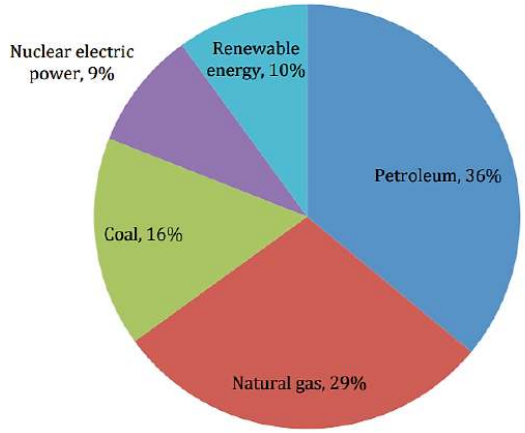


Figure **8.3.3**: U.S. primary energy consumption by source (all sectors), 2015, showing that about 80 % of our energy consumption comes from fossil fuels. Data from U.S. Energy Information Administration, April 2016.

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8.4: Hydrocarbons

Learning Objectives

- Identify alkanes, alkenes, alkynes, and aromatic compounds.
- List some properties of hydrocarbons.

The simplest organic compounds are those composed of only two elements: carbon and hydrogen. These compounds are called **hydrocarbons**. Hydrocarbons themselves are separated into two types: aliphatic hydrocarbons and aromatic hydrocarbons.

Aliphatic hydrocarbons are hydrocarbons based on chains of C atoms. There are three types of aliphatic hydrocarbons. **Alkanes** are aliphatic hydrocarbons with only single covalent bonds. **Alkenes** are hydrocarbons that contain at least one C–C double bond, and **Alkynes** are hydrocarbons that contain a C–C triple bond. Occasionally, we find an aliphatic hydrocarbon with a ring of C atoms; these hydrocarbons are called *cycloalkanes* (or *cycloalkenes* or *cycloalkynes*).

Aromatic hydrocarbons have a special six-carbon ring called a *benzene* ring. Electrons in the benzene ring have special energetic properties that give benzene physical and chemical properties that are markedly different from alkanes. Originally, the term *aromatic* was used to describe this class of compounds because they were particularly fragrant. However, in modern chemistry the term *aromatic* denotes the presence of a six-membered ring that imparts different and unique properties to a molecule.

The simplest alkanes have their C atoms bonded in a straight chain; these are called *normal* alkanes. They are named according to the number of C atoms in the chain. The smallest alkane is methane:

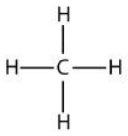


Figure **8.4.1** - Three-Dimensional Representation of Methane.

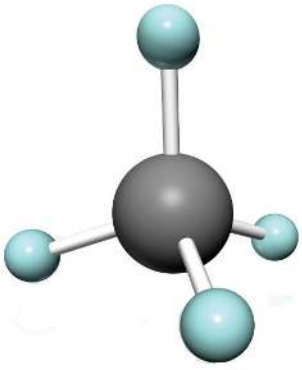
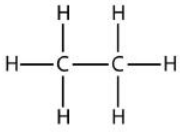
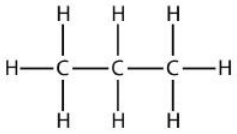


Figure **8.4.1** Three-Dimensional Representation of Methane © Thinkstock. The methane molecule is three dimensional, with the H atoms in the positions of the four corners of a tetrahedron.

The next-largest alkane has two C atoms that are covalently bonded to each other. For each C atom to make four covalent bonds, each C atom must be bonded to three H atoms. The resulting molecule, whose formula is C_2H_6, is ethane:



Propane has a backbone of three C atoms surrounded by H atoms. You should be able to verify that the molecular formula for propane is C_3H_8:



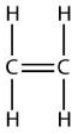
The diagrams representing alkanes are called **structural formulas** because they show the structure of the molecule. As molecules get larger, structural formulas become more and more complex. One way around this is to use a **condensed structural formula**, which lists the formula of each C atom in the backbone of the molecule. For example, the condensed structural formula for ethane is CH_3CH_3, while for propane it is $CH_3CH_2CH_3$. Table **8.4.1** - The First 10 Alkanes, gives the molecular formulas, the condensed structural formulas, and the names of the first 10 alkanes.

Table **8.4.1** The First 10 Alkanes

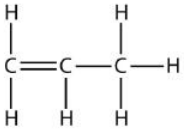
Molecular Formula	Condensed Structural Formula	Name
CH_4	CH_4	methane
C_2H_6	CH_3CH_3	ethane
C_3H_8	$CH_3CH_2CH_3$	propane
C_4H_{10}	$CH_3CH_2CH_2CH_3$	butane
C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$	pentane
C_6H_{14}	$CH_3(CH_2)_4CH_3$	hexane
C_7H_{16}	$CH_3(CH_2)_5CH_3$	heptane
C_8H_{18}	$CH_3(CH_2)_6CH_3$	octane
C_9H_{20}	$CH_3(CH_2)_7CH_3$	nonane
$C_{10}H_{22}$	$CH_3(CH_2)_8CH_3$	decane

Because alkanes have the maximum number of H atoms possible according to the rules of covalent bonds, alkanes are also referred to as **saturated hydrocarbons**.

Alkenes have a C–C double bond. Because they have less than the maximum number of H atoms possible, they are **unsaturated hydrocarbons**. The smallest alkene—ethene—has two C atoms and is also known by its common name ethylene:

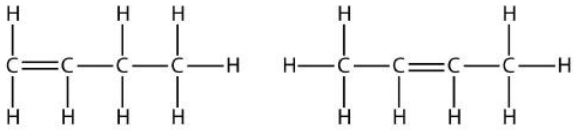


The next largest alkene—propene—has three C atoms with a C–C double bond between two of the C atoms. It is also known as propylene:



What do you notice about the names of alkanes and alkenes? The names of alkenes are the same as their corresponding alkanes except that the ending is *-ene*, rather than *-ane*. Using a stem to indicate the number of C atoms in a molecule and an ending to represent the type of organic compound is common in organic chemistry, as we shall see.

With the introduction of the next alkene, butene, we begin to see a major issue with organic molecules: choices. With four C atoms, the C–C double bond can go between the first and second C atoms or between the second and third C atoms:



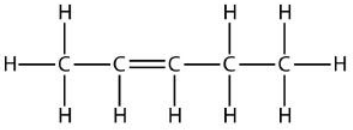
2 structural formulas for butene, with the first butene having the double bond on the first and second carbon from the left and the latter having its double bond on the second and third carbon from the left.

(A double bond between the third and fourth C atoms is the same as having it between the first and second C atoms, only flipped over.) The rules of naming in organic chemistry require that these two substances have different names. The first molecule is named *1-butene*, while the second molecule is named *2-butene*. The number at the beginning of the name indicates where the double bond originates. The lowest possible number is used to number a feature in a molecule; hence, calling the second molecule 3-butene would be incorrect. Numbers are common parts of organic chemical names because they indicate which C atom in a chain contains a distinguishing feature.

The compounds 1-butene and 2-butene have different physical and chemical properties, even though they have the same molecular formula—C_4H_8. Different molecules with the same molecular formula are called **isomers**. Isomers are common in organic chemistry and contribute to its complexity.

Example 8.4.1

Based on the names for the butene molecules, propose a name for this molecule.



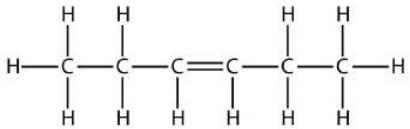
A structural formula of a five carbon molecule with a double bond on the third and fourth carbon from the left. There are ten hydrogen atoms in total.

Solution

With five C atoms, we will use the *pent-* stem, and with a C–C double bond, this is an alkene, so this molecule is a pentene. In numbering the C atoms, we use the number *2* because it is the lower possible label. So this molecule is named 2-pentene.

Exercise 8.4.1

Based on the names for the butene molecules, propose a name for this molecule.



A structural formula of a six carbon molecule with a double bond on the third and fourth carbon from the left. There are twelve hydrogen atoms in total.

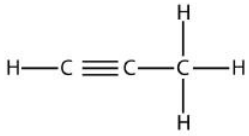
Answer

3-hexene

Alkynes, with a C–C triple bond, are named similarly to alkenes except their names end in *-yne*. The smallest alkyne is ethyne, which is also known as acetylene:

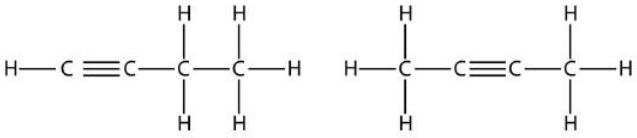


Propyne has the structure



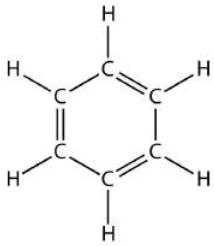
Structural formula showing three carbon molecules with a triple bond present between the first and second carbon atom. The appropriate number of hydrogen atoms is attached to each carbon atom.

With butyne, we need to start numbering the position of the triple bond, just as we did with alkenes:



Two structural formula of butyne. One butyne has a triple bond between the first and second carbon atom, while two butyne has the triple bond between the second and third carbon atom.

Aromatic compounds contain the benzene unit. Benzene itself is composed of six C atoms in a ring, with alternating single and double C–C bonds:



The six carbons are arranged in a hexagon pattern with one hydrogen atom emerging outwards from each carbon atom. The presence of a double bond is alternated between every other carbon atom.

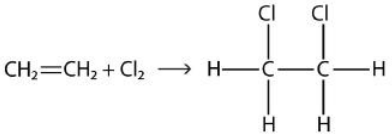
The alternating single and double C–C bonds give the benzene ring a special stability, and it does not react like an alkene as might be suspected. Benzene has the molecular formula C_6H_6; in larger aromatic compounds, a different atom replaces one or more of the H atoms.

As fundamental as hydrocarbons are to organic chemistry, their properties and chemical reactions are rather mundane. Most hydrocarbons are nonpolar because of the close electronegativities of the C and H atoms. As such, they dissolve only sparingly in H_2O and other polar solvents. Small hydrocarbons, such as methane and ethane, are gases at room temperature, while larger hydrocarbons, such as hexane and octane, are liquids. Even larger hydrocarbons are solids at room temperature and have a soft, waxy consistency.

Hydrocarbons are rather unreactive, but they do participate in some classic chemical reactions. One common reaction is substitution with a halogen atom by combining a hydrocarbon with an elemental halogen. Light is sometimes used to promote the reaction, such as this one between methane and chlorine:

$$CH_4 + Cl_2 \overset{light}{\rightarrow} CH_3Cl + HCl$$

Halogens can also react with alkenes and alkynes, but the reaction is different. In these cases, the halogen reacts with the C–C double or triple bond and inserts itself onto each C atom involved in the multiple bonds. This reaction is called an **addition reaction**. One example is



Structural formula showing the reaction of ethylene with a chlorine molecule to form ethylene dichloride.

The reaction conditions are usually mild; in many cases, the halogen reacts spontaneously with an alkene or an alkyne.

Hydrogen can also be added across a multiple bond; this reaction is called a **hydrogenation reaction**. In this case, however, the reaction conditions may not be mild; high pressures of H_2 gas may be necessary. A platinum or palladium catalyst is usually employed to get the reaction to proceed at a reasonable pace:

$$CH_2 = CH_2 + H_2 \overset{metal\ catalyst}{\rightarrow} CH_3CH_3$$

By far the most common reaction of hydrocarbons is **combustion**, which is the combination of a hydrocarbon with O_2 to make CO_2 and H_2O. The combustion of hydrocarbons is accompanied by a release of energy and is a primary source of energy production in our society (Figure 8.4.2 - Combustion). The combustion reaction for gasoline, for example, which can be represented by C_8H_{18}, is as follows:

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O + \sim 5060kJ$$



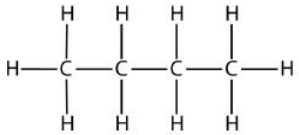
Figure 8.4.2 Combustion © Thinkstock. The combustion of hydrocarbons is a primary source of energy in our society.

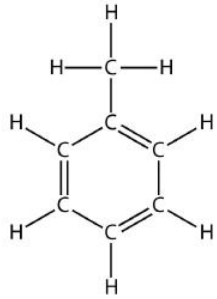
Key Takeaways

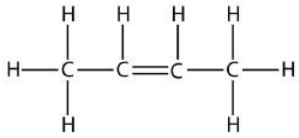
- The simplest organic compounds are hydrocarbons and are composed of carbon and hydrogen.
- Hydrocarbons can be aliphatic or aromatic; aliphatic hydrocarbons are divided into alkanes, alkenes, and alkynes.
- The combustion of hydrocarbons is a primary source of energy for our society.

Exercise 8.4.2

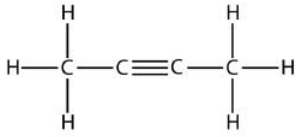
1. Define *hydrocarbon*. What are the two general types of hydrocarbons?
2. What are the three different types of aliphatic hydrocarbons? How are they defined?
3. Indicate whether each molecule is an aliphatic or an aromatic hydrocarbon; if aliphatic, identify the molecule as an alkane, an alkene, or an alkyne.

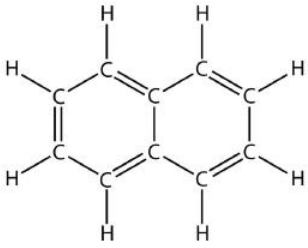
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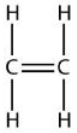
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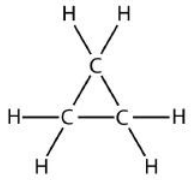
4. Indicate whether each molecule is an aliphatic or an aromatic hydrocarbon; if aliphatic, identify the molecule as an alkane, an alkene, or an alkyne.

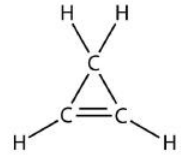
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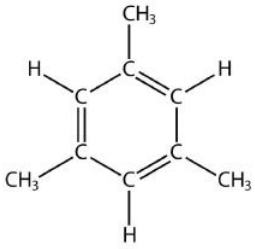
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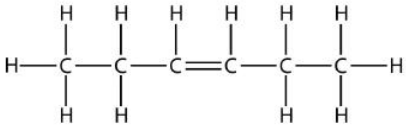
5. Indicate whether each molecule is an aliphatic or an aromatic hydrocarbon; if aliphatic, identify the molecule as an alkane, an alkene, or an alkyne.

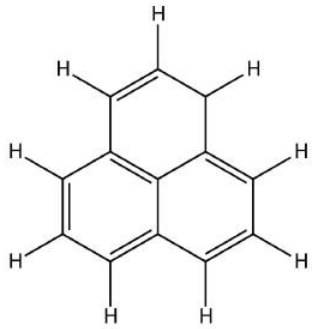
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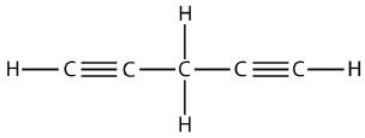
b. 

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6. Indicate whether each molecule is an aliphatic or an aromatic hydrocarbon; if aliphatic, identify the molecule as an alkane, an alkene, or an alkyne.

a. 

b. 

c. 

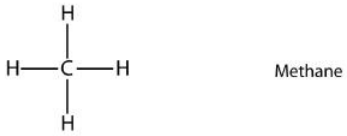
7. Name and draw the structural formulas for the four smallest alkanes.
8. Name and draw the structural formulas for the four smallest alkenes.
9. What does the term *aromatic* imply about an organic molecule?
10. What does the term *normal* imply when used for alkanes?
11. Explain why the name *1-propene* is incorrect. What is the proper name for this molecule?
12. Explain why the name *3-butene* is incorrect. What is the proper name for this molecule?
13. Name and draw the structural formula of each isomer of pentene.
14. Name and draw the structural formula of each isomer of hexyne.
15. Write a chemical equation for the reaction between methane and bromine.
16. Write a chemical equation for the reaction between ethane and chlorine.

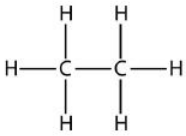
17. Draw the structure of the product of the reaction of bromine with propene.
18. Draw the structure of the product of the reaction of chlorine with 2-butene.
19. Draw the structure of the product of the reaction of hydrogen with 1-butene.
20. Draw the structure of the product of the reaction of hydrogen with 1-butene.
21. Write the balanced chemical equation for the combustion of heptane.
22. Write the balanced chemical equation for the combustion of nonane.

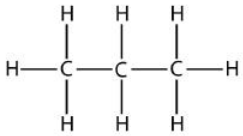
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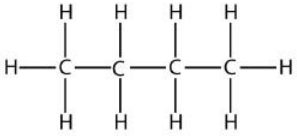
Answers

1. an organic compound composed of only carbon and hydrogen; aliphatic hydrocarbons and aromatic hydrocarbons
2.
3. a. aliphatic; alkane
   b. aromatic
   c. aliphatic; alkene
4.
5. a. aliphatic; alkane
   b. aliphatic; alkene
   c. aromatic
6.

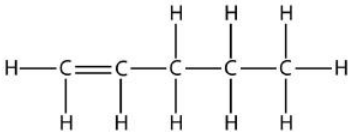

Methane


Ethane

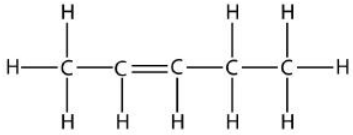

Propane


Butane

7.
8.
9. Aromatic means that the molecule has a benzene ring.
10.
11. The 1 is not necessary. The name of the compound is simply *propene*.
12.



1-Pentene



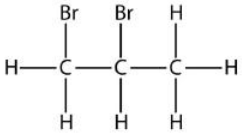
2-Pentene

13.

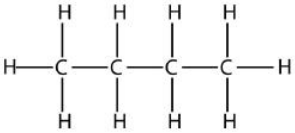
14.

15. $CH_4 + Br_2 \rightarrow CH_3Br + HBr$

16.

17. 

18.

19. 

20.

21. $C_7H_{16} + 11O_2 \rightarrow 7CO_2 + 8H_2O$

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8.5: Branched Hydrocarbons

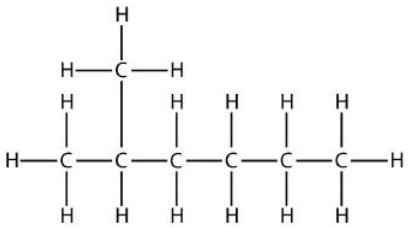
Learning Objectives

- Name a branched hydrocarbon from its structure.
- Draw the structural formula of a branched hydrocarbon from its name.

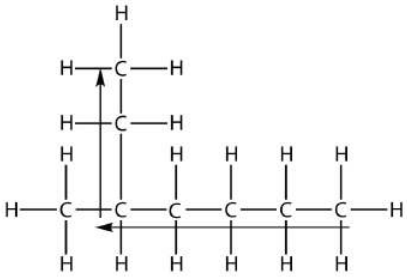
Not all hydrocarbons are straight chains. Many hydrocarbons have branches of C atoms attached to a chain. These branched alkanes are isomers of straight-chain alkanes having the same number of C atoms. However, they are different compounds with different physical and chemical properties. As such, they need different names. How do we name **branched hydrocarbons**?

There are a series of rules for naming branched alkanes (and, ultimately, for all organic compounds). These rules make up the system of **nomenclature** for naming organic molecules. Worldwide, the International Union of Pure and Applied Chemistry (IUPAC) has developed the system of nomenclature for organic compounds. So these rules are sometimes called the *IUPAC rules of nomenclature*. By learning and applying these rules, you can name any organic compound when given its structure or determine the unique structure of a molecule from its name. You have already learned the basics of nomenclature—the names of the first 10 normal hydrocarbons. Here, we will add some steps to the procedure so you can name branched hydrocarbons.

First, given the structure of an alkane, identify the longest continuous chain of C atoms. Note that the longest chain may not be drawn in a straight line. The longest chain determines the parent name of the hydrocarbon. For example, in the molecule



the longest chain of carbons has six C atoms. Therefore, it will be named as a hexane. However, in the molecule



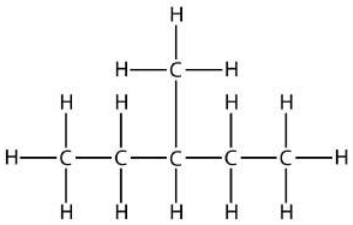
the longest chain of C atoms is not six, but seven, as shown. So this molecule will be named as a heptane.

The next step is to identify the branches, or **substituents,** on the main chain. The names of the substituents, or *alkyl groups*, are derived from the names of the parent hydrocarbons; however, rather than having the ending *-ane*, the substituent name has the ending *-yl*. Table 8.5.1 - Substituent Names, lists the substituent names for the five smallest substituents.

Table **8.5.1** Substituent Names

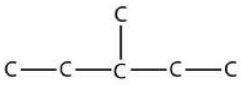
Substituent Formula	Number of C Atoms	Name of Substituent
CH_3	1	*methyl-*
CH_3CH_2	2	*ethyl-*
$CH_3CH_2CH_2$	3	*propyl-*
$CH_3CH_2CH_2CH_2$	4	*butyl-*
$CH_3CH_2CH_2CH_2CH_2$	5	*pentyl-*
and so forth	and so forth	and so forth

In naming the branched hydrocarbon, the name of the substituent is combined with the parent name of the hydrocarbon without spaces. However, there is likely one more step. The longest chain of the hydrocarbon must be numbered, and the numerical position of the substituent must be included to account for possible isomers. As with double and triple bonds, the main chain is numbered to give the substituent the lowest possible number. For example, in this alkane



the longest chain is five C atoms long, so it is a pentane. There is a one-carbon substituent on the third C atom, so there is a methyl group at position 3. We indicate the position using the number, which is followed by a hyphen, the substituent name, and the parent hydrocarbon name—in this case, 3-methylpentane. That name is specific to that particular hydrocarbon and no other molecule. Organic chemistry nomenclature is very specific!

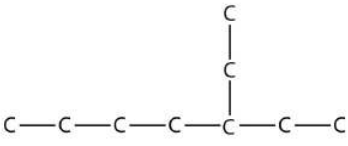
It is common to write the structural formula of a hydrocarbon without the H atoms, for clarity. So we can also represent 3-methylpentane as



where it is understood that any unwritten covalent bonds are bonds with H atoms. With this understanding, we recognize that the structural formula for 3-methylpentane refers to a molecule with the formula of C_6H_{14}.

Example 8.5.1

Name this molecule.

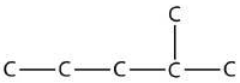


Solution

The longest continuous carbon chain has seven C atoms, so this molecule will be named as a heptane. There is a two-carbon substituent on the main chain, which is an ethyl group. To give the substituent the lowest numbering, we number the chain from the *right* side and see that the substituent is on the third C atom. So this hydrocarbon is 3-ethylheptane.

Exercise 8.5.1

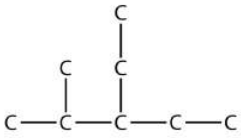
Name this molecule.



Answer

2-methylpentane

Branched hydrocarbons may have more than one substituent. If the substituents are different, then give each substituent a number (using the smallest possible numbers) and list the substituents in alphabetical order, with the numbers separated by hyphens and with no spaces in the name. So the molecule



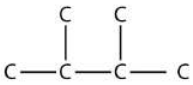
is 3-ethyl-2-methylpentane.

If the substituents are the same, then use the name of the substituent only once, but use more than one number, separated by a comma. Also, put a numerical prefix before the substituent name that indicates the number of substituents of that type. The numerical prefixes are listed in Table **8.5.2** - Numerical Prefixes to Use for Multiple Substituents. The number of the position values must agree with the numerical prefix before the substituent.

Table **8.5.2**: Numerical Prefixes to Use for Multiple Substituents

Number of Same Substituent	Numerical Prefix
2	*di-*
3	*tri-*
4	*tetra-*
5	*penta-*
and so forth	and so forth

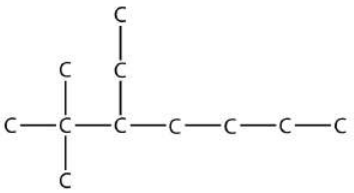
Consider this molecule:



The longest chain has four C atoms, so it is a butane. There are two substituents, each of which consists of a single C atom; they are methyl groups. The methyl groups are on the second and third C atoms in the chain (no matter which end the numbering starts from), so we would name this molecule 2,3-dimethylbutane. Note the comma between the numbers, the hyphen between the numbers and the substituent name, and the presence of the prefix *di-* before the *methyl*. Other molecules—even with larger numbers of substituents—can be named similarly.

Example 8.5.2

Name this molecule.

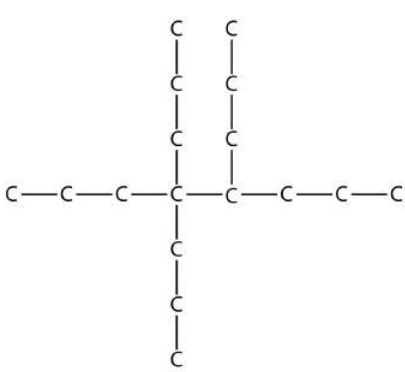


Solution

The longest chain has seven C atoms, so we name this molecule as a heptane. We find two one-carbon substituents on the second C atom and a two-carbon substituent on the third C atom. So this molecule is named 3-ethyl-2,2-dimethylheptane.

? Exercise 8.5.2

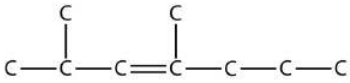
Name this molecule.



Answer

4,4,5-tripropyloctane

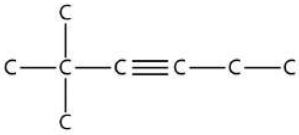
Alkenes and alkynes are named in a similar fashion. The biggest difference is that when identifying the longest carbon chain, it *must* contain the C–C double or triple bond. Furthermore, when numbering the main chain, the double or triple bond gets the lowest possible number. This means that there may be longer or higher-numbered substituents than may be allowed if the molecule were an alkane. For example, this molecule



is 2,4-dimethyl-3-heptene (note the number and the hyphens that indicate the position of the double bond).

✓ Example 8.5.3

Name this molecule.



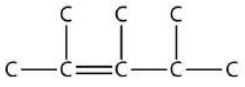
Solution

The longest chain that contains the C–C triple bond has six C atoms, so this is a hexyne molecule. The triple bond starts at the third C atom, so this is a 3-hexyne. Finally, there are two methyl groups on the chain; to give them the lowest possible number,

we number the chain from the left side, giving the methyl groups the second position. So the name of this molecule is 2,2-dimethyl-3-hexyne.

? Exercise 8.5.3

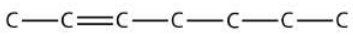
Name this molecule.



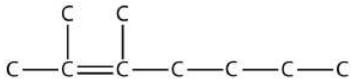
Answer

2,3,4-trimethyl-2-pentene

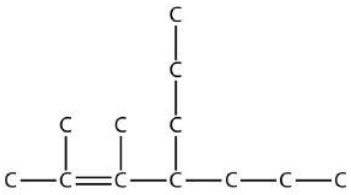
Once you master naming hydrocarbons from their given structures, it is rather easy to draw a structure from a given name. Just draw the parent chain with the correct number of C atoms (putting the double or triple bond in the right position, as necessary) and add the substituents in the proper positions. If you start by drawing the C atom backbone, you can go back and complete the structure by adding H atoms to give each C atom four covalent bonds. From the name 2,3-dimethyl-4-propyl-2-heptene, we start by drawing the seven-carbon parent chain with a double bond starting at the third carbon:



We add two one-carbon substituents to this structure on the second and third C atoms:



We finish the carbon backbone by adding a three-carbon propyl group to the fourth C atom in the parent chain:



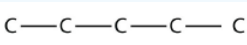
If we so choose, we can add H atoms to each C atom to give each carbon four covalent bonds, being careful to note that the C atoms in the double bond already have an additional covalent bond. (How many H atoms do you think are required? There will need to be 24 H atoms to complete the molecule.)

✓ Example 8.5.4

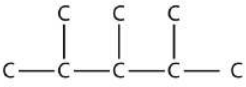
Draw the carbon backbone for 2,3,4-trimethylpentane.

Solution

First, we draw the five-carbon backbone that represents the pentane chain:



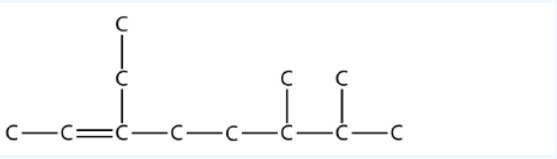
According to the name, there are three one-carbon methyl groups attached to the second, the third, and the fourth C atoms in the chain. We finish the carbon backbone by putting the three methyl groups on the pentane main chain:



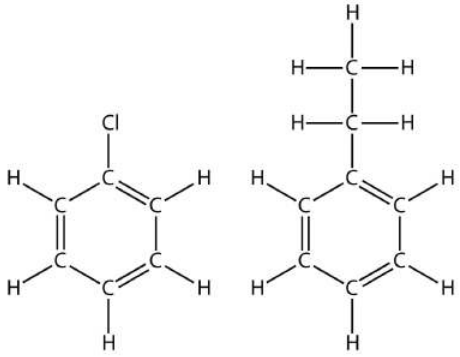
? Exercise 8.5.4

Draw the carbon backbone for 3-ethyl-6,7-dimethyl-2-octene.

Answer

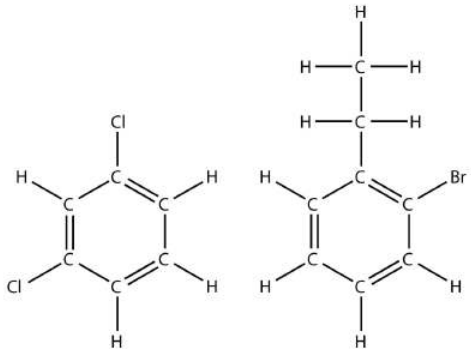


Naming substituted benzene molecules is straightforward. If there is only one substituent, the substituent is named as a side chain on a benzene molecule, like this:



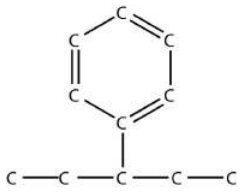
Chlorobenzene Ethylbenzene

If there are two or more substituents on a benzene molecule, the relative positions must be numbered, just as an aliphatic chain of C atoms is numbered. The substituent that is first alphabetically is assigned position 1, and the ring is numbered in a circle to give the other substituents the lowest possible number(s).



1,3-Dichlorobenzene 1-Bromo-2-ethylbenzene

If a benzene ring is treated as a substituent, it is given the name *phenyl-*. The following molecule is 3-phenylpentane:



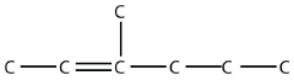
where the H atoms have been omitted for clarity.

Summary

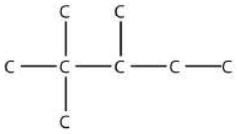
A unique name can be given to branched hydrocarbons. A unique structure can be drawn for the name of a hydrocarbon.

Exercise 8.5.5

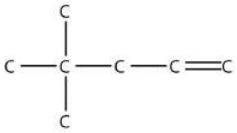
1. How does a branched hydrocarbon differ from a normal hydrocarbon?
2. How does a substituent get its unique name?
3. Name this molecule.



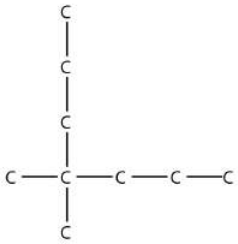
4. Name this molecule.



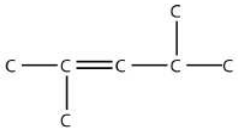
5. Name this molecule.



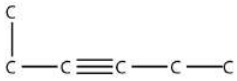
6. Name this molecule.



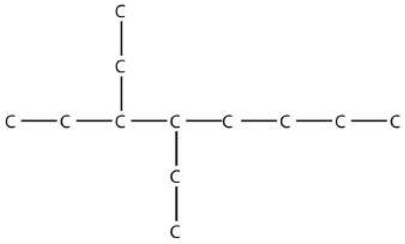
7. Name this molecule.



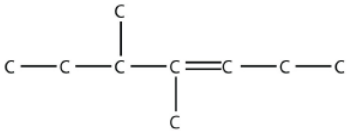
8. Name this molecule.



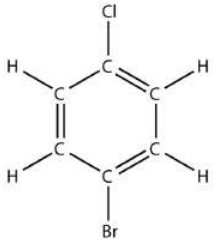
9. Name this molecule.



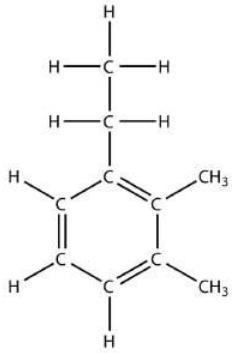
10. Name this molecule.



11. Name this molecule.



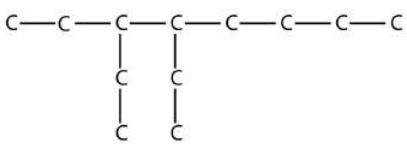
12. Name this molecule.

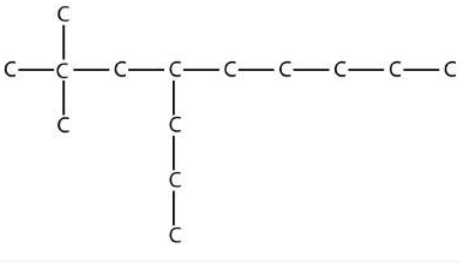


13. Draw the carbon backbone for each molecule.
    a. 3,4-diethyloctane
    b. 2,2-dimethyl-4-propylnonane
14. Draw the carbon backbone for each molecule.
    a. 3-ethyl-4-methyl-3-heptene
    b. 3,3-diethyl-1-pentyne
15. Draw the carbon backbone for each molecule.
    a. 4-ethyl-4-propyl-2-octyne
    b. 5-butyl-2,2-dimethyldecane
16. Draw the carbon backbone for each molecule.
    a. 3,4-diethyl-1-hexyne
    b. 4-propyl-3-ethyl-2-methyloctane
17. The name 2-ethylhexane is incorrect. Draw the carbon backbone and write the correct name for this molecule.
18. The name 3-butyl-7-methyloctane is incorrect. Draw the carbon backbone and write the correct name for this molecule.

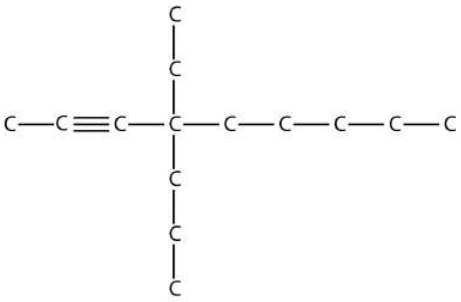
Answers

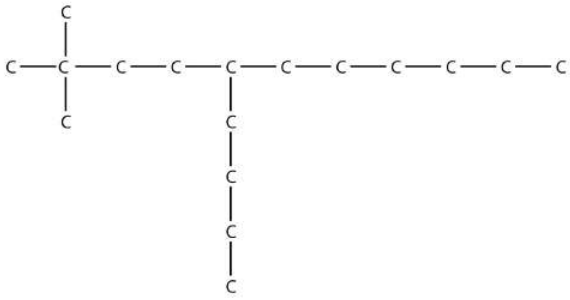
1. A branched hydrocarbon does not have all of its C atoms in a single row.
2.
3. 3-methyl-2-hexene
4.
5. 4,4-dimethyl-1-pentene
6.
7. 2,4-dimethyl-2-pentene
8.
9. 3,4-diethyloctane
10.
11. 1-bromo-4-chlorobenzene
12.

13. a. 

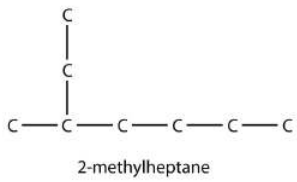
b. 

14.

15. a. 

b. 

16.

17. 

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8.6: Coal

Coal is a combustible black or brownish-black sedimentary rock with a high amount of carbon and hydrocarbons. Coal is classified into four main types, or ranks depending on the types and amounts of carbon present and on the amount of heat energy the coal can produce, including anthracite, bituminous, subbituminous, and lignite (highest to lowest ranked, pictured in Figure 8.6.1). For us to use the potential energy stored in coal, it first must be mined from the ground. This process in itself uses a great deal of resources and has its own environmental impacts. Coal then typically undergoes processing to make it suitable for use in **coal-fire power plants**. Finally, the processed coal is burned in these power plants, and the kinetic energy released from its combustion is harnessed for electricity generation or other purposes. We will investigate each of these steps individually below.

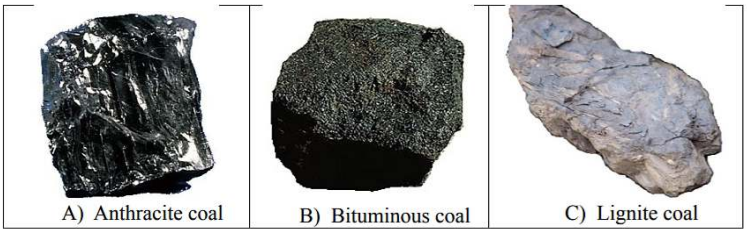


Figure 8.6.1: Different types of coal. Images obtained from Wikimedia Commons.

4.3.1: Coal Mining and Processing, and Electricity Generation

There are two primary methods of coal mining: **strip mining** and **underground mining**. Strip-, or surface-, mining uses large machines to remove the soil and layers of rock known as **overburden** to expose coal seams. It is typically used when the coal is less than 200 feet underground. **Mountaintop removal** is a form of surface mining where the tops of mountains are blasted with dynamite and removed to access coal seams. After the mining is finished, the disturbed area can be re-covered with topsoil, and the area is replanted. However, the topography of the mountain is permanently altered.

Underground mining, sometimes called deep mining, is used when the coal is several hundred feet below the surface. Some underground mines are thousands of feet deep, and extend for miles. Miners ride elevators down deep mine shafts and travel on small trains in long tunnels to get to the coal. The miners use large machines that dig out the coal.

Once mined, coal may go to a preparation plant located near the mining site where it is cleaned and processed to remove impurities such as rocks and dirt, ash, sulfur, and other unwanted materials. This process increases the amount of energy that can be obtained from a unit of coal, known as its **heating value**.

Finally, the mined and processed coal must be transported. Transportation can be more expensive than mining the coal. Nearly 70% of coal delivered in the United States is transported, for at least part of its trip, by train. Coal can also be transported by barge, ship, or truck. Coal can also be crushed, mixed with water, and sent through a slurry **pipeline**. Sometimes, coal-fired electric power plants are built near coal mines to lower transportation costs.

Once at the power plant, coal is first pulverized into a fine powder then mixed with hot air and blown into a **furnace**, allowing for the most complete combustion and maximum heat possible. Purified water, pumped through pipes inside a **boiler**, is turned into steam by the heat from the combustion of coal. The high pressure of the steam pushing against a series of giant **turbine** blades turns the turbine shaft. The turbine shaft is connected to the shaft of the **generator**, where magnets spin within wire coils to produce electricity. After doing its work in the turbine, the steam is drawn into a **condenser**, a large chamber in the basement of the power plant. In this important step, millions of gallons of cool water from a nearby source (such as a river or lake) are pumped through a network of tubes running through the condenser. The cool water in the tubes converts the steam back into water that can be used over and over again in the plant. The cooling water is returned to its source without any contamination except at a higher temperature than when first extracted from the river or lake. Figure 8.6.2 below is a schematic diagram showing a typical layout of a coal-fire power plant. You can also watch a short video of a virtual tour of a coal power plant at the URL provided below.

Coal Power Plant Virtual Tour Video https://www.youtube.com/watch?v=2IKECt4Y3RI

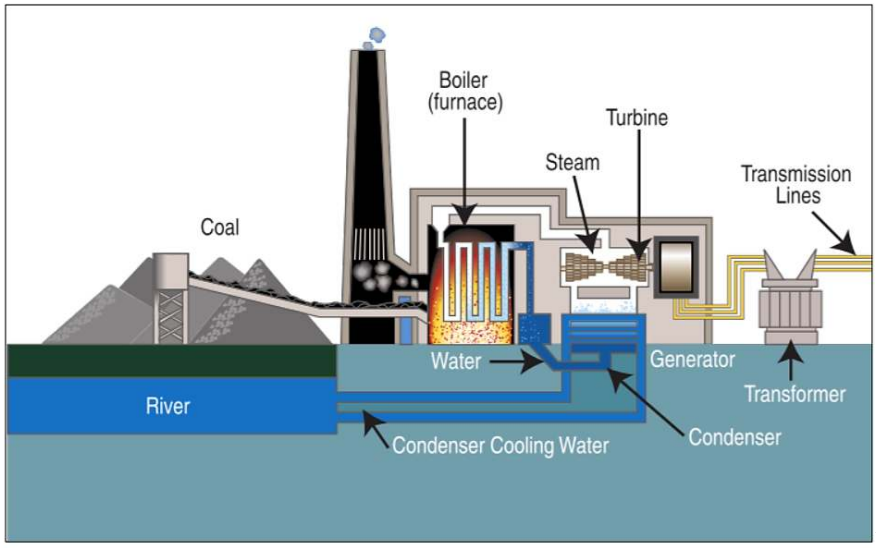


Figure **8.6.2**: Diagram of a typical steam-cycle coal power plant (proceeding from left to right). Image by US Tennessee Valley Authority – Public domain. www.tva.com

4.3.2: Impacts of coal mining and burning

Impacts of coal mining on the environment

A majority of the coal mined in the United States (about 66%) is from surface, or strip mines which leave highly visible impacts at the surface. Strip mining operations generally involve removing soils, rock, and other material to access shallow deposits of coal and therefore leave permanent scars on the landscape. It also involves the destruction of substantial amounts of forests and other ecosystems, destroying natural habitats and threatening biodiversity.

Mountaintop removal, the extreme form of strip mining, has affected large areas of the Appalachian Mountains in West Virginia and Kentucky. The tops of mountains are removed using a combination of explosives and mining equipment and the material is deposited into nearby valleys. This technique not only alters the landscape (Figure **8.6.3**) but affects the health and quality of nearby streams by depositing rocks, dirt, and pollutants that can harm aquatic wildlife. While mountaintop removal mining has existed since the 1970s, its use became more widespread and controversial beginning in the 1990s. U.S. laws require that dust and water runoff from areas affected by coal mining operations be controlled, and that the area be **reclaimed**, and returned to close to its original condition.



Figure **8.6.3**: A) Strip mine for lignite coal at Garzweiler near Koln, Germany. Image from Wikimedia commons https://commons.wikimedia.org/wiki/F...strip.mine.jpg B) Mountaintop removal in Eunice WV, photo by Roston, obtained from https://commons.wikimedia.org/wiki/F...niceblast3.JPG

One of the largest environmental impacts of underground mining may be the **methane** gas that must be vented out of mines to make the mines a safe place to work. Methane is a **greenhouse gas**, meaning that it enhances the greenhouse effect naturally occurring in our atmosphere, and contributes to global warming and global climate change. Its global warming potential, or relative capacity to produce the greenhouse effect, is higher than that of carbon dioxide (see chapter 7). Other impacts of underground mining include ground collapse above mine tunnels and draining of acidic water from abandoned mines into nearby streams. Acidic

water lowers the pH (resulting in increased **acidity**), which is detrimental to aquatic organisms. This **acid mine drainage** is an environmental impact associated with both underground mining and strip mining.

Impacts of coal burning on the environment and human health

In the United States and most of the world, most of the coal consumed is used as a fuel to generate electricity. Burning coal produces emissions such as **sulfur dioxide** (SO_2) and **nitrogen oxides** (NO_x) that are associated with acid rain (more on this in chapter 6). **Carbon dioxide** (CO_2), another emission resulting from burning coal, is a major greenhouse gas that is associated with global warming (see chapter 7).

Ash (including fly ash and bottom ash) is a residue created when coal is burned at power plants. In the past, fly ash was released into the air through the smokestack, where it would contribute to **particulate matter** air pollution (see chapter 6). Laws now require that much of the fly ash now must be be captured by pollution control devices, like scrubbers. In the United States, fly ash is generally stored at coal power plants or placed in landfills. Pollution leaching from ash storage and landfills into groundwater and the rupture of several large impoundments of ash are environmental concerns.

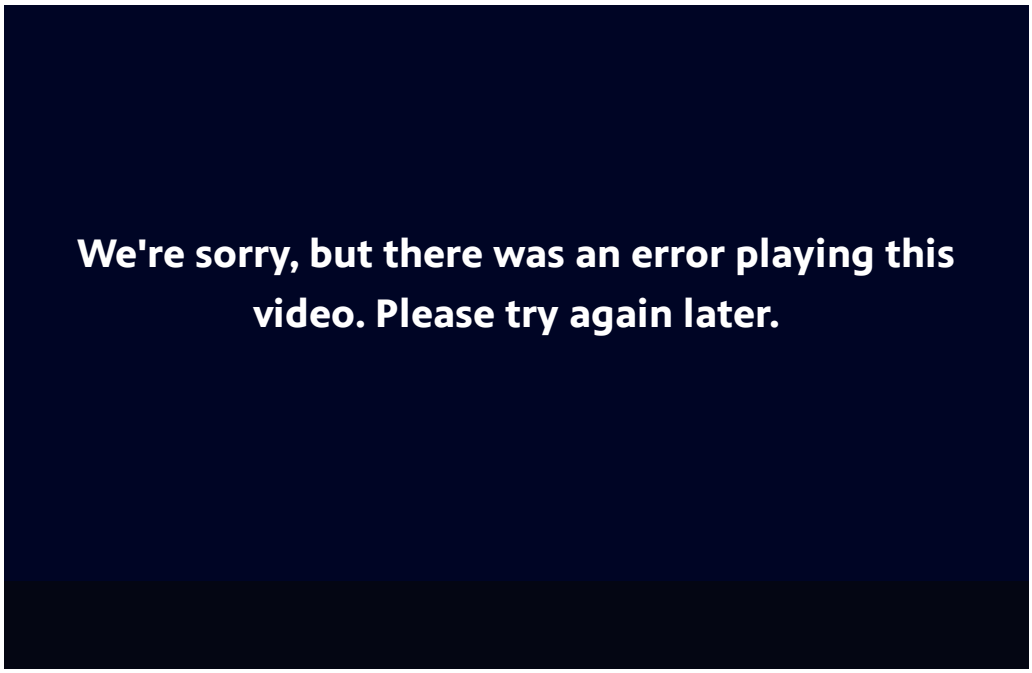
Burning coal produces emissions that also impact human health. Emissions such as sulfur dioxide (SO_2), nitrogen oxides (NO_x) and particulates contribute to respiratory illnesses. Particulates also contribute to a condition among coal miners and other coal workers known as coal workers' pneumoconiosis (CWP) or black lung disease, which results from long exposure to coal dust. Inhaled coal dust progressively builds up in the lungs and is unable to be removed by the body; this leads to inflammation, fibrosis, and in worse cases, tissue death (necrosis).

Coal is the largest source of **mercury** and also a source of other **heavy metals**, many of which have been linked to both neurological and developmental problems in humans and other animals. Mercury concentrations in the air usually are low and of little direct concern. However, when mercury enters water, either directly or through deposition from the air, biological processes transform it into **methylmercury**, a highly toxic chemical that accumulates in fish and the animals (including humans) that eat fish.

4.3.3: Reducing the environmental impacts of coal use

Regulations such as the **Clean Air Act** and the **Clean Water Act** require industries to reduce pollutants released into the air and water. Below are some actions that have been taken to reduce the negative impacts of coal on human and environmental health:

- **Clean coal technology**: Industry has found several ways to reduce sulfur, NO_x, and other impurities from coal before burning.
- Coal consumers have shifted toward greater use of low sulfur coal.
- Power plants use **scrubbers**, to clean SO_2, NO_x, particulate matter, and mercury from the smoke before it leaves their smokestacks. In addition, industry and the U.S. government have cooperated to develop technologies that make coal more energy-efficient so less needs to be burned.
- Research is underway to address emissions of carbon dioxide from coal combustion. **Carbon capture & sequestration** separates CO_2 from emissions sources and recovers it in a concentrated stream. The CO_2 can then be sequestered, which puts CO_2 into storage, possibly underground, where it will remain permanently (see chapter 7)
- Reuse and recycling can also reduce coal's environmental impact. Land that was previously used for coal mining can be reclaimed and used for airports, landfills, and golf courses. Waste products captured by scrubbers can be used to produce products like cement and synthetic gypsum for wallboard.



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8.7: Acid Rain

"Normal" rainfall is slightly acidic because of the presence of dissolved carbonic acid. Carbonic acid is the same as that found in soda pop.

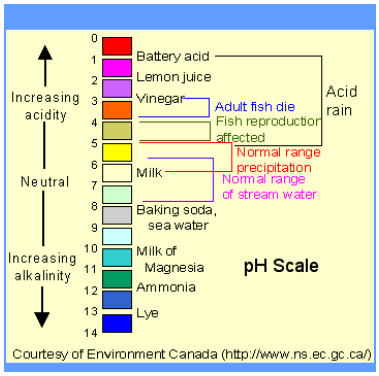
Introduction

The pH of "normal" rain has traditionally been given a value of 5.6. However scientists now believe that the pH of rain may vary from 5.6 to a low of 4.5 with the average value of 5.0. Acid rain or acid snow is a direct result of the method that the atmosphere cleans itself. The tiny droplets of water that make up clouds, continuously capture suspended solid particles and gases in the atmosphere. The gases of sulfur oxides and nitrogen oxides are chemically converted into sulfuric and nitric acids. The non-metal oxide gases react with water to produce acids (ammonia produces a base).

$$SO_2 + HOH \rightarrow H_2SO_3 \tag{8.7.1}$$

$$2NO_2 + HOH \rightarrow HNO_2 + HNO_3 \tag{8.7.2}$$

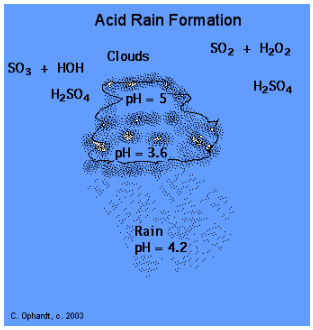
When enough of the tiny cloud droplets clump together to form a larger water drop it may fall to the earth as "wet" acid precipitation including rain, snow, ice, sleet, or fog. See Acids and Bases for more details.



Formation

The sulfuric and nitric acids formed from gaseous pollutants can easily make their way into the tiny cloud water droplets. These sulfuric acid droplets are one component of the summertime haze in the eastern United States. Some sulfuric acid is formed directly in the water droplets from the reaction of sulfur dioxide and hydrogen peroxide. Some of these sulfuric acid particles drop to the earth as "dry" acid deposition. The tiny water droplets containing sulfuric acid provide a ready surface to attract more molecules of water to form a larger droplet of dilute sulfuric acid.

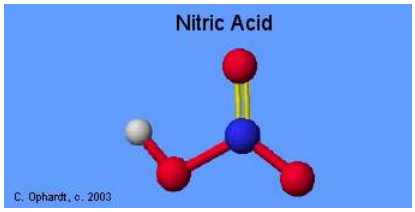
Water droplets collected from the base of clouds in the Eastern U.S.during the summer have an average pH of 3.6, with some values as low as pH 2.6. The pH in the upper portion of a cloud is much higher. The final rain droplet has an average pH of 4.2 in the Northeast U.S. In Los Angeles, the pH of fog has been measured at 2.0 - about the acidity of lemon juice.



Contrary View

The benchmark of "natural rain" is 5.6. Acid precipitation in the range of 4.2-5.0 has been recorded in most of the Eastern United States and Canada. EPRI (Electric Power Research Institute) likes to compare these values to familiar objects to give the impression that these pH values are not harmful. Examples: Carrots = 5.0, Bananas = 4.6, Tomatoes = 4.2, Apples and soft drinks = 3.0, Lemon juice = 2.0. EPRI also contends that pH 5.6 may or may not be a valid reference point. It should not be considered the "background" or "natural" acidity of precipitation.

Even without man-made influences, there are natural sources of sulfur oxides, nitrogen oxides, and other species important to determining the precipitation acidity at any given time. Hence, trying to quantify man's contribution to the natural condition will never be possible, since the "natural background" condition cannot be known.



In the forest areas of Brazil at the headlands of the Amazon River, an area remote from civilization, the monthly average of 100 rain events in the 1960s ranged from pH 4.3 to pH 5.2, with the median value of pH 4.6 and one reading as low as pH 3.6. On the island of Hawaii, remote from all industrial activity, the weighted average of precipitation over a 4 year period was pH 5.3, with a minimum value of pH 3.8.

Outside Links

- www.epa.gov/acidrain/what/index.html
- Bretherick, Leslie. "Nitric acid, nitrates, and nitro compounds: Energetic servants, but overwhelming masters (SAFETY)." *J. Chem. Educ.* **1989**, *66*, A220.
- Driscoll, Jerry A. "Acid Rain Demonstration: The Formation of Nitrogen Oxides as a By-Product of High-Temperature Flames in Connection with Internal Combustion Engines." *J. Chem. Educ.* **1997** *74* 1424.
- Johns, Nicholas; Longstaff, Stephen J. "A convenient, low-cost method for determining sulfate in acid rain (F&R)." *J. Chem. Educ.* **1987**, *64*, 449.

Contributors

- Charles Ophardt, Professor Emeritus, Elmhurst College
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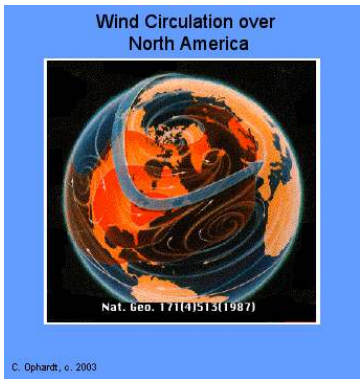
8.8: Acid Rain Transport

The reactions of sulfur oxides to form sulfuric acid are quite slow. Sulfur dioxide may remain airborne for 3-4 days.As a consequence acid rain derived from sulfur oxides may travel for hundreds of miles or even a thousand miles. Nitrogen oxides may persist for only about one half day and therefore may travel only tens or hundreds of miles.

Introduction

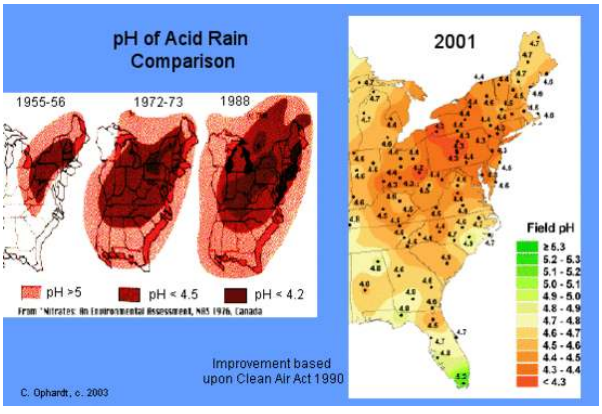
Once airborne, the sulfur and nitrogen oxides eventually come down in one form or another. Where they come down depends on the height of the smokestack and the prevailing weather conditions. In general, prevailing winds in North America transport pollutants from west to east or northeast. The nine largest coal burning states are in the Midwest and the Ohio River valley. It is estimated that two thirds of the acid rain in the Northeast and Eastern Canada comes from these sources.

Blue arrow shows the upper winds that travel from the west to the east or northeast. Winds travel from the mid-west to the northeast. In addition, a copper-nickel smelter in Sudbury, Ontario, just north of Lake Huron is the most significant sulfur oxide source in Canada. The winds may also carry the sulfur oxide clouds to the Northeast in the U.S. where it may be converted to acid rain.



pH of Acid Rain over a Period of 50 years

These maps show how the areas of lower pH have spread in a 30 years 1955-1988. Darkest area is lowest pH. Since the Clean Air Act Amendments of 1990, there have been significant decreases in the amount of sulfur oxides escaping from the electric power plants. As a result there has been a measurable reduction in the amount of acid rain, which is actually translated as an increase in the pH levels (higher pH means less acid).



Contributors

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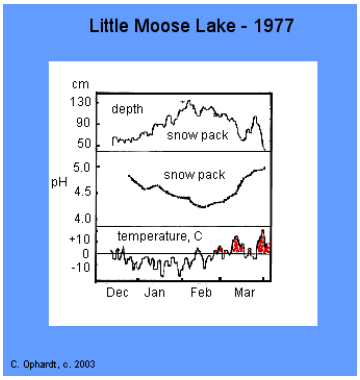
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8.9: Acid Snow

The impact of acid precipitation on aquatic ecosystems may be intensified by melting snow. When snow melts rapidly in the spring, the stream or lake may be "shocked" with an excessive amount of acid. In the spring, at the time of acid snow melting, the various aquatic organisms are reproducing and are the most sensitive increases in acid.

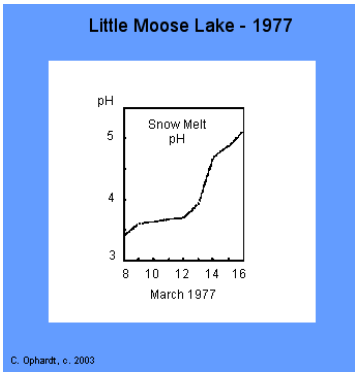
Little Moose Lake

During the winter of 1976-77, there was no significant snow melt from mid-December through February at Little Moose Lake, New York. The highest depth was 130 cm. (top graph). The snow had an average pH of 4.4 in Feb. (middle graph). The snow started to melt in early March.(bottom graph).



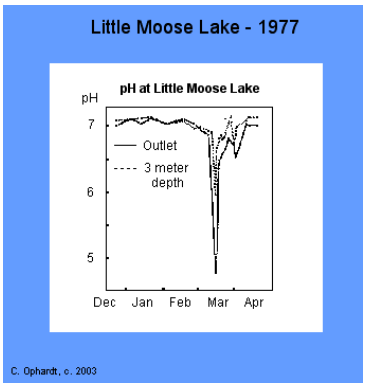
Snow Thaw at Little Moose Lake

The first major thaw in early March resulted in the release of 80% of the stored acid in a one week period. During the first couple of days, the snow melt water pH was 3.4-3.6. The pH gradually rose over the eight day period to pH 5.0.



pH at Little Moose Lake

The pH levels in Little Moose lake are normally about 7.0. During the snow melting, in early March, the lake pH dropped to 6.0. An outlet stream from the lake reached a low pH of 4.8. A small brook nearby hit a low pH of 4.6 during the snow melt period. The average pH in this brook during the rest of the year is about 5.4. There was a definite relationship of the amount of aluminum in the brook vs. pH. As the pH decreased, the aluminum concentration increased.



Captive populations of adult, year old, and larval brook trout had been maintained over the winter in Little Moose Lake water without any problems. During the snow melt in March, 3 adult brook trout died, 25 one year old trout died, and about 50,000 recently hatched trout died. Various other abnormal behaviors were also observed.

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8.10: Oil

Petroleum Oil is currently the most widely used fossil fuel and accounts for about one third of global energy consumption. Unlike coal, which is primarily used as a fuel for electricity generation, oil is primarily used as a fuel for **transportation**. Oil is also used to manufacture plastics and other synthetic compounds ubiquitous to our everyday life. **Crude** (unprocessed) oil varies greatly in appearance depending on its composition. It is usually black or dark brown (although it may be yellowish, reddish, or even greenish). In the reservoir it is usually found in association with natural gas, which being lighter forms a gas cap over the oil.

Oil is made up of hydrocarbons which are molecules that contain hydrogen and carbon in various lengths and structures, from straight chains to branching chains to rings. Hydrocarbons contain a lot of energy and many of the things derived from crude oil like gasoline, diesel fuel, paraffin wax and so on take advantage of this energy.

4.4.1: Extraction

Oil is mainly obtained by drilling either on land (**onshore**) or in the ocean (**offshore**). Early offshore drilling was generally limited to areas where the water was less than 300 feet deep. Oil and natural gas drilling rigs now operate in water as deep as two miles. Floating platforms are used for drilling in deeper waters. These self-propelled vessels are attached to the ocean floor using large cables and anchors. Wells are drilled from these platforms which are also used to lower production equipment to the ocean floor. Some drilling platforms stand on stilt-like legs that are embedded in the ocean floor. These platforms hold all required drilling equipment as well as housing and storage areas for the work crews.

Offshore oil producers are required to take precautions to prevent pollution, spills, and significant changes to the ocean environment. Offshore rigs are designed to withstand hurricanes. Offshore production is much more expensive than land-based production. When offshore oil wells are no longer productive enough to be economical, they are sealed and abandoned according to applicable regulations.

4.4.2: Processing and Refining

When extracted, crude oil consists of many types of hydrocarbons as well as some unwanted substances such as sulfur, nitrogen, oxygen, dissolved metals, and water all mixed together. Unprocessed crude oil is therefore, not generally useful in industrial applications and must first be separated into different useable products at a **refinery** (Figure 8.10.1). All refineries perform three basic steps: separation, conversion, and treatment in the processing and refining of crude oil.



Figure 8.10.1: Tesoro Corporation Oil Refinery in Anacortes, Washington. Photo by Walter Siegmund. https://commons.wikimedia.org/wiki/O...nery_31904.JPG

During **separation**, the various products (hydrocarbons) are separated in to different components (called *fractions*), by taking advantage of the differences in boiling temperature of the components. This process is called *fractional distillation* and involves heating up the crude, letting it vaporize and then condensing the vapor. The lightest components have the lowest boiling temperature and rise to the top while the heaviest which also have the highest boiling temperature remain at the bottom.

Conversion is the chemical processing in which some of the fractions are transformed in to other products, for example, a refinery can turn diesel fuel into gasoline depending on the demand for gasoline. Conversion can involve breaking larger hydrocarbon chains into smaller ones (cracking), combining smaller chains into larger ones (unification) or rearranging the molecules to created desired products (alteration).

Treatment is done to the fractions to remove impurities such as sulfur, nitrogen and water among others. Refineries also combine the various fractions (processed and unprocessed) into mixtures to make desired products. For example, different mixtures of hydrocarbon chains can create gasolines with different octane ratings, with and without additives, lubricating oils of various weights and grades (e.g., WD-40, 10W-40, 5W-30, etc.), heating oil and many others. The products are stored on-site until they can be delivered to various markets such as gas stations, airports and chemical plants.

A 42 U.S. gallon barrel of crude oil yields about 45 gallons of petroleum products because of refinery processing gain. This increase in volume is similar to what happens to popcorn when it is popped. Gasoline makes up the largest fraction of all petroleum products obtained (Figure 8.10.2). Other products include diesel fuel and heating oil, jet fuel, petrochemical feedstocks, waxes, lubricating oils, and asphalt.

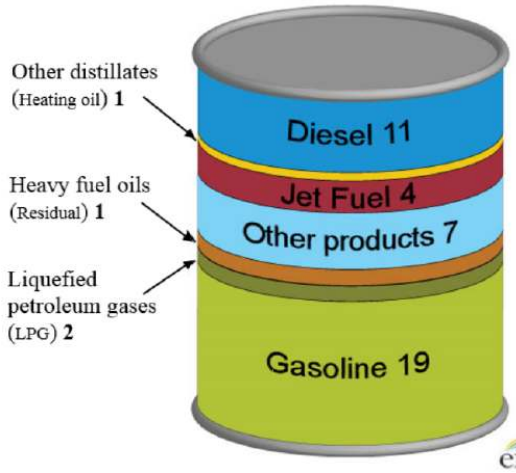


Figure 8.10.2: Main products (measured in gallons) made from a barrel of crude oil, 2013. Source: U.S. Energy information administration http://www.eia.gov/energyexplained/i...?page=oil_home

4.4.3: Fracking for oil

Hydraulic fracturing, informally referred to as "fracking," is an oil well development process that typically involves injecting water, sand, and chemicals under high pressure into a bedrock formation via the well. This process is intended to create new fractures in the rock as well as increase the size, extent, and connectivity of existing fractures. Hydraulic fracturing is a well-stimulation technique used commonly in low-permeability rocks like tight sandstone, shale, and some coal beds to increase oil flow to a well from petroleum-bearing rock formations (Figure 8.10.3).

Energy development often requires substantial amounts of water, and hydraulic fracturing is no exception. Water is needed not only for the traditional drilling process, but also for the actual fracturing as well. Water is first mixed with chemicals and fine sands, then pumped at extremely high pressure into the shale rock to fracture it, forming pathways for the oil and gas to reach the well. The water is then recovered, along with the oil and gas.

There are concerns regarding the potential contamination of fresh groundwater resources from oil and gas extraction wells that use hydraulic fracturing; either from the petroleum resource being produced or from the chemicals introduced in the fracturing process. Fracking fluid flowback – the fluid pumped out of the well and separated from oil and gas – not only contains the chemical additives used in the drilling process but also contains heavy metals, radioactive materials, **volatile organic compounds** (VOCs) and hazardous air pollutants such as benzene, toluene, ethylbenzene and xylene. In some cases, this contaminated water is sent to water treatment plants that are not equipped to deal with some of these classes of contamination.

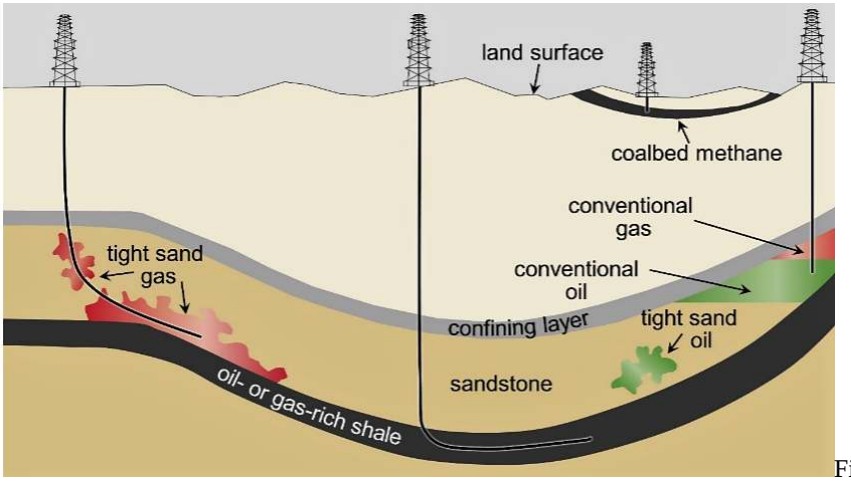


Figure 8.10.3: Schematic cross-section of general types of oil and gas resources and the orientations of production wells used in hydraulic fracturing. Source: US EPA (Public Domain)

4.4.4: Environmental Impacts of Oil

Burning petroleum oil products releases emission such as carbon monoxide (CO), sulfur dioxide (SO_2), nitrogen oxides (NO_x), and particulate material all of which are air pollutants that impact the environment as well as human health (see more on air pollution in Chapter 6). Petroleum also emits carbon dioxide which is a greenhouse gas Exploring and drilling for oil may disturb land and ocean habitats. On land, extensive infrastructure such as road networks, transport pipelines and housing for workers are needed to support a full-scale drilling operation. These can pollute soil and water, fragment habitats, and disturb wildlife.

Human-caused **oil spills** in rivers and oceans harm ecosystems. Natural oil seepages do occur and may be a significant source of oil that enters the environment globally, but they are slow, small, and spread out over large areas, and the ecosystem has adapted to them. Spills from tankers or well spills have more catastrophic impacts. The quantity of oil spilled during accidents has ranged from a few hundred tons to several hundred thousand tons but even small spills have been shown to have a great impact on ecosystems.

Oil spills at sea are generally much more damaging than those on land, since they can spread for hundreds of nautical miles in a thin oil slick which can cover beaches with a thin coating of oil. This can kill sea birds, mammals, shellfish and other organisms it coats. Oil spills on land are more readily containable if a makeshift earth dam can be rapidly bulldozed around the spill site before most of the oil escapes, and land animals can avoid the oil more easily. The amount of oil spilled from ships dropped significantly during the 1990s partly because new ships were required to have a double-hull lining to protect against spills.

Leaks also happen when we use petroleum products on land. For example, gasoline sometimes drips onto the ground when people are filling their gas tanks, when motor oil gets thrown away after an oil change, or when fuel escapes from a leaky storage tank. When it rains, the spilled products get washed into the gutter and eventually flow to rivers and into the ocean. Another way that oil sometimes gets into water is when fuel is leaked from motorboats and jet skis.

When a leak in a storage tank or pipeline occurs, petroleum products can also get into the ground, and the ground must be cleaned up. To prevent leaks from underground storage tanks, all buried tanks are supposed to be replaced by tanks with a double lining.



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8.11: Natural Gas

4.5.1: Introduction

Crude oil is frequently found in reservoirs along with natural gas. In the past, natural gas was either burned or allowed to escape into the atmosphere. Now, technology has been developed to capture the natural gas and either reinject it into the well or compress it into **liquid natural gas** (LNG).

Natural gas is predominately composed of methane (CH_4). Some of the gases that are produced along with methane, such as butane and propane (by-products), are separated and cleaned at a gas processing plant. The by-products, once removed, are used in a number of ways. For example, propane can be used for cooking on gas grills. Natural gas withdrawn from a well may contain liquid hydrocarbons and nonhydrocarbon gases. This is called "wet" natural gas. The natural gas is separated from these components near the site of the well or at a processing plant. The gas is then considered "dry" and is sent through pipelines to a local distribution company, and, ultimately, to the consumer.

Most of the natural gas consumed in the United States is produced in the United States. Some is imported from Canada and shipped to the United States in pipelines. A small amount of natural gas is shipped to the United States as LNG. We can also use machines called **digesters** that turn today's organic material (plants, animal wastes, etc.) into natural gas through the process of anaerobic decomposition. This process replaces waiting for millions of years for the gas to form naturally. The natural gas produced by these digesters is not a fossil fuel, but is rather a renewable source of bioenergy (see chapter 5).

4.5.2: Fracking for Gas

Conventional natural gas is found in permeable reservoirs, typically composed of sandstone or limestone, where extraction is relatively straightforward because the gas generally flows freely. Unconventional gas is found in rocks with extremely low permeability, which makes extracting it much more difficult. Such gas is extracted by employing so-called “unconventional” techniques such as hydraulic fracturing (fracking), which has been in use since the late 1940s. In recent decades, fracking technology has greatly improved, and its use has been expanded. The process of fracking for gas is very similar to that of fracking for oil, and the environmental impacts are similar also.

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8.12: Outdoor Air Pollution

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 8.12.1). 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.

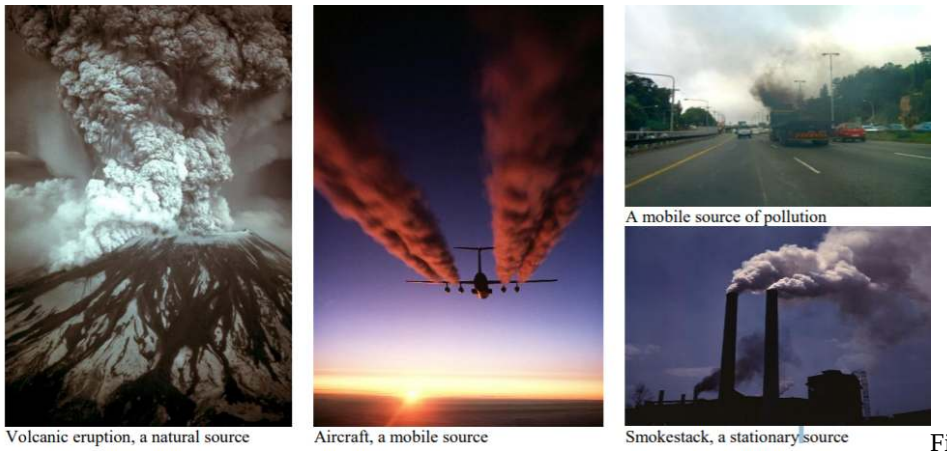


Figure 8.12.1: Images showing various sources of air pollution, including natural and anthropogenic, stationary and mobile. Source: All images obtained from Wiki Commons commons.wikimedia.org/wiki/Air_pollution

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.

6.3.1: Criteria pollutants

Under the Clean Air Act (section 6.7.1), 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_2emissions 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 **8.12.2**). 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.

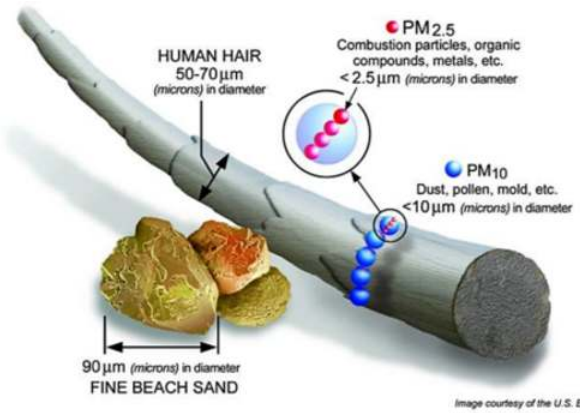


Figure **8.12.2**: Graphic showing size comparison of particulate material (PM_{10} and $PM_{2.5}$) compared to fine beach sand and human hair. Source: US EPA www.epa.gov/air/particlepollution/basic.html

6.3.2: 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.

6.3.3: 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 **8.12.3** and **8.12.4**). 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.



Figure **8.12.3**: Smog over Almaty city, Kazakhstan. Photo by Igors Jefimovs. Source: Wikicommons commons.wikimedia.org/wiki/C...dia/File:Smog_ over_Almaty.jpg



Figure **8.12.4**: Smog over Santiago in Chile. Source: German Wikipedia commons.wikimedia.org/wiki/C...#/media/File:S antiago30std.jpg

6.3.4: 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.

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8.13: 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 **8.13.1**) 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 **8.13.1**: 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.

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CHAPTER OVERVIEW

9: Renewable Energy

9.1: Introduction
9.2: Solar Energy
9.3: Biomass Energy
9.4: Wind Power
9.5: Geothermal Power
9.6: Hydroelectric Power (Hydropower)
9.7: Other Alternative Renewable Energy Sources

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9.1: Introduction

Learning Objectives

By the end of this chapter, students should be able to.

- Describe arguments for alternative energy
- Explain the following aspects of solar energy:
  - How passive solar energy works and provide examples of its use.
  - How solar panels (photovoltaic cells) work.
  - The limitations and environmental costs associated with solar energy.
- Explain the following aspects of biofuels / biomass energy:
  - Describe what is meant by the term “carbon neutral” and explain how biomass energy can and cannot be carbon neutral.
  - Describe current achievements in biofuels and potential of this area for growth.
- Describe wind energy and explain the advantages and disadvantages.
- Describe geothermal energy and explain the advantages and disadvantages
- Describe hydroelectric energy, advantages and disadvantages

Energy sources that are more or less continuously made available in a time frame useful to people are called **renewable energy.** Renewable energy sources are often considered alternative sources because, in general, most industrialized countries do not rely on them as their main energy source. Instead, they tend to rely on the conventional energy sources such as fossil fuels or nuclear power that are non-renewable. Because of the energy crisis in the United States during the 1970s, dwindling supplies of fossil fuels and hazards associated with nuclear power, use of renewable energy sources such as solar energy, hydroelectric, wind, biomass, and geothermal has grown. Renewable energy comes from the sun (considered an "unlimited" supply) or other sources that can theoretically be renewed at least as quickly as they are consumed. If used at a sustainable rate, these sources will be available for consumption for thousands of years or longer. Renewable alternatives derive from wind, water, solar or biomass (Figure **9.1.1**). Note that wind, water and biomass energy sources are indirect sources of solar energy. One limitation currently associated with most forms of renewable energy is that the energy is not concentrated and not easily portable.



Figure **9.1.1**: A variety of voltage sources (clockwise from top left): the Brazos Wind Farm in Fluvanna, Texas (credit: Leaflet, Wikimedia Commons); the Krasnoyarsk Dam in Russia (credit: Alex Polezhaev); a solar farm (credit: U.S. Department of Energy); and a group of nickel metal hydride batteries (credit: Tiaa Monto). The voltage output of each depends on its construction and load, and equals emf only if there is no load.

Energy is an important ingredient in all phases of society. We live in a very interdependent world, and access to adequate and reliable energy resources is crucial for economic growth and for maintaining the quality of our lives. However, current levels of energy consumption and production are not sustainable because of the heavy reliance on non-renewable energy sources. The principal energy resources used in the world are shown in Figure **9.1.2**. The fuel mix has changed over the years but now is dominated by oil, although natural gas and solar contributions are increasing. About 80 % of our energy comes from nonrenewable fossil fuels and nuclear. The link between global warming and fossil fuel use, with its production of carbon dioxide through

combustion, has made, in the eyes of many scientists, a shift to non-fossil fuels of utmost importance – but it will not be easy. About 40 % of the world's energy comes from oil, and much of that goes to transportation uses.

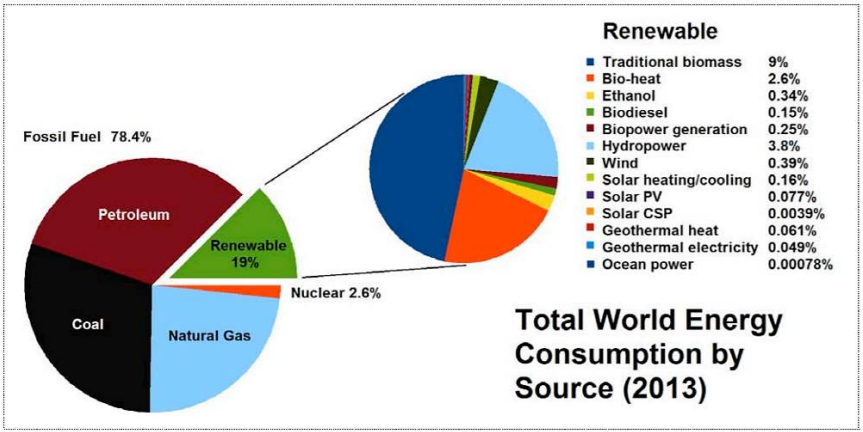


Figure 9.1.2: World energy consumption by source, 2013, in percent. Notice that when all forms of biomass energy are combined, biomass makes up the largest proportion of renewable energy used worldwide. Source: Wikimedia commons - By Delphi234 - Own work, CC0, https://commons.wikimedia.org/w/inde...curid=33567170

5.1.1: The World's Growing Energy Needs

World energy consumption continues to rise especially in countries like China where the economy is improving. Global demand for energy has tripled in the past 50 years and might triple again in the next 30 years (Figure 5.1.3). While much of this growth will come from the rapidly booming economies of China and India, many of the industrialized countries, especially those in Europe, are hoping to meet their energy needs by expanding the use of renewable sources. Although presently only a small percentage, renewable energy is growing very fast, especially wind energy. For example, Germany plans to meet 20% of its electricity and 10% of its overall energy needs with renewable resources by the year 2020. Energy is a key constraint in the rapid economic growth of China and India. In 2003, China surpassed Japan as the world's second largest consumer of oil. However, over 1/3 of this oil is imported. Unlike most Western countries, coal dominates the commercial energy resources of China, accounting for 2/3 of its energy consumption. In 2009 China surpassed the United States as the largest emitter of CO_2. In India, the main energy resources are biomass (wood and dung) and coal. Half of India's oil is imported. About 70% of India's electricity is generated by highly polluting coal. Yet there are sizeable strides being made in renewable energy. India has a rapidly growing wind energy base, and it has the largest solar cooking program in the world. While non-renewable sources dominate, some countries get a sizeable percentage of their electricity from renewable resources. For example, about 67% of New Zealand's electricity demand is met by hydroelectric. Renewable resources, primarily hydroelectric, generate only 10% of the U.S. electricity.

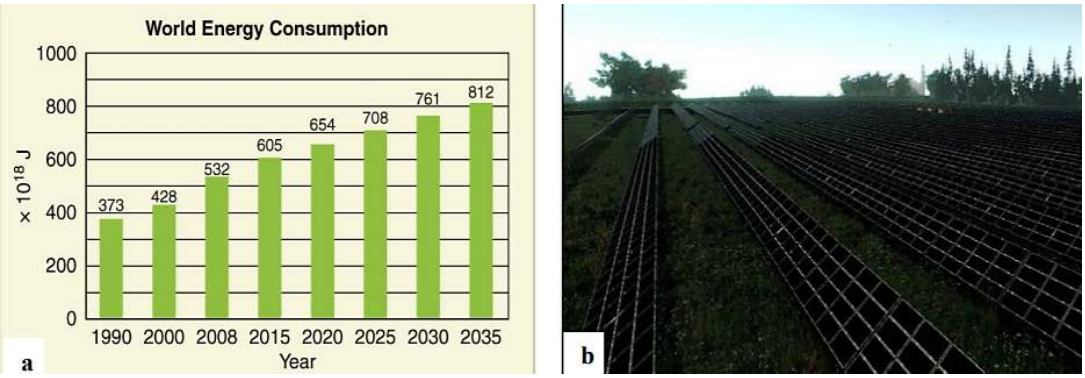


Figure 9.1.3: (*a) Past and projected world energy use (source: Based on data from U.S. Energy Information Administration, 2011). (b) Solar cell arrays at a power plant in Steindorf, Germany (credit: Michael Betke, Flickr).*

5.1.2: Why Use Renewable Energy Sources?

Majority of renewable energy sources including solar, wind, water, and biomass can be directly or indirectly attributed to the sun. The fact that the sun will continue burning for another 4-5 billion years makes it inexhaustible as an energy source for human

civilization. With appropriate technology, renewable energy sources allow for local, decentralized control over power. Homes, businesses, and isolated communities can use sources such as solar to produce electricity without being near a power plant or being connected to a grid. This eliminates problems such as spills associated with extraction and transportation of fossil fuels that is needed in order to supply these fossil fuels to those areas that are lacking. Most renewable energy sources do not pollute the air with greenhouse gas emissions and other air pollutants associated with fossil fuels. This is especially important in combating climate change.

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9.2: Solar Energy

Solar energy is the ultimate energy source driving life on earth and many human activities. Though only one billionth of the energy that leaves the sun (Figure **9.2.1**) actually reaches the earth's surface, this is more than enough to meet the world's energy requirement. In fact, all other sources of energy, renewable and non-renewable, are actually stored forms of solar energy. The process of directly converting solar energy to heat or electricity is considered a renewable energy source. Solar energy represents an essentially unlimited supply of energy as the sun will long outlast human civilization on earth. The difficulties lie in harnessing the energy. Solar energy has been used for centuries to heat homes and water, and modern technology (photovoltaic cells) has provided a way to produce electricity from sunlight. There are two basic forms of solar energy collectors are passive and active.

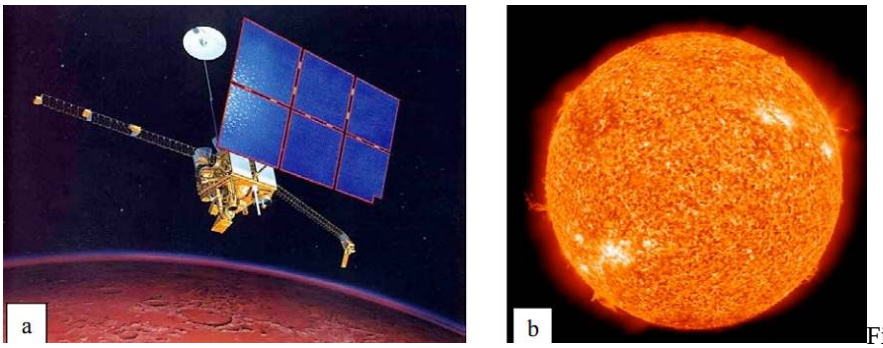


Figure **9.2.1**: a) Mars Observer in Mars Orbit showing the solar panel. (Picture: Wikimedia Commons www.jpl.nasa.gov/images/space...t1-browse.jpg,). b) The Sun photographed at 304 angstroms by the Atmospheric Imaging Assembly (AIA 304) of NASA's Solar Dynamics Observatory (SDO). This is a false-color image of the Sun observed in the extreme ultraviolet region of the spectrum. sdo.gsfc.nasa.gov/assets/img/..._4096_0304.jpg, NASA/SDO (AIA)

5.2.1: Passive and Active Solar Energy

Passive solar energy uses heating and cooling strategies that have been used historically such as natural ventilation, solar heat gain, solar shading and efficient insulation. Passive solar space heating happens when the sun shines through the windows of a building and warms the interior (Figure **9.2.2**). Building designs that optimize passive solar heating usually have south facing windows that allow the sun to shine on solar heat-absorbing walls or floors during the winter. The solar energy heats the building by natural radiation and convection and the heat is trapped, absorbed and stored by materials with high thermal mass (usually bricks or concrete) inside the house. The heat is released at night when needed to warm up the building as it loses heat to the cooler outdoors. Window overhangs or shades block the sun from entering the windows during the summer to keep the building cool.

Active solar energy systems require the input of some energy to pump a heat-absorbing fluid medium through a collector to store and distribute the energy. Fans or pumps circulate air or heat-absorbing liquids through collectors and then transfer the heated fluid directly to a room or to a heat storage system. The solar collectors are either concentrating or non-concentrating. In the **non-concentrating collectors**, the surface area that intercepts the solar radiation is the same as the area absorbing the radiation. Flat-plate collectors are the most common type of nonconcentrating collectors and are used when temperatures lower than 200°F are sufficient. The collectors absorb and transfer heat to a fluid (water or air) which is then circulated to provide heating to a building. In **concentrating collectors** the surface area intercepting the solar radiation is greater, sometimes hundreds of times greater, than the absorber area. The collector focuses or concentrates solar energy onto an absorber. The collector usually moves so that it maintains a high degree of concentration on the absorber.


Figure **9.2.2**: Passive solar building near Crestone, Colorado (image from Wiki Commons)

Photovoltaic (PV) Cells

Solar photovoltaic (PV) devices, or solar cells, change sunlight directly into electricity. PV uses semiconducting materials such as silicon to produce electricity from sunlight: when light hits the cells, the material produces free electrons that migrate across the cell, creating an electric current. Small PV cells can power calculators, watches, and other small electronic devices. Arrangements of many solar cells in PV panels and arrangements of multiple PV panels in PV arrays can produce electricity for an entire house (Figure **9.2.3**a). Some PV power plants have large arrays that cover many acres to produce electricity for thousands of homes.

Hundreds of thousands of houses and buildings around the world have PV systems on their roofs. Many multi-megawatt PV power plants have also been built. Covering 4% of the world's desert areas with photovoltaics could supply the equivalent of all of the world's electricity. The Gobi Desert alone could supply almost all of the world's total electricity demand.

Solar Thermal Power Plants

Solar thermal power plants, like the one in Figure **9.2.3**b use concentrating solar collector systems to collect and concentrate sunlight to produce the high temperature heat needed to generate electricity. All solar thermal power systems have solar energy collectors with two main components: reflectors (mirrors) that capture and focus sunlight onto a receiver. In most types of systems, a heat-transfer fluid is heated and circulated in the receiver and used to produce steam. The steam is converted into mechanical energy in a turbine, which powers a generator to produce electricity. Solar thermal power systems have tracking systems that keep sunlight focused onto the receiver throughout the day as the sun changes position in the sky.

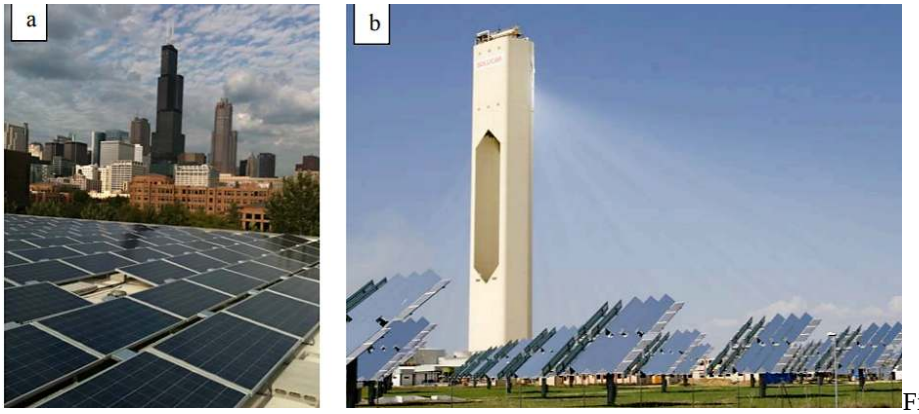
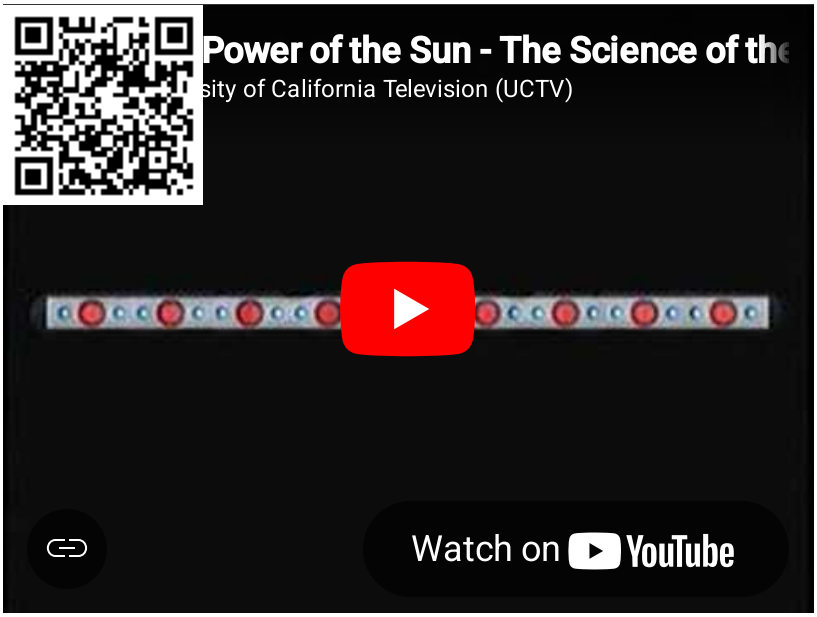


Figure **9.2.3**: a) Rooftop Solar Installations on Douglas Hall at the University of Illinois at Chicago has no effect on land resources, while producing electricity with zero emissions. Source: Office of Sustainability. b) Solucar PS10 solar power tower in Andalusia, Spain, is a solar thermal power plant that generates electricity commercially. (Photo by Afloresm Solucar PS10 CC BY 2.0)

5.2.3: Environmental Impacts of solar energy

Solar power has minimal impact on the environment, depending on where it is placed. In 2009, one percent of the renewable energy generated in the United States was from solar power (1646 MW) out of the eight percent of the total electricity generation that was from renewable sources. The manufacturing of photovoltaic (PV) cells generates some hazardous waste from the

chemicals and solvents used in processing. Often solar arrays are placed on roofs of buildings or over parking lots or integrated into construction in other ways. However, large systems may be placed on land and particularly in deserts where those fragile ecosystems could be damaged if care is not taken. Some solar thermal systems use potentially hazardous fluids (to transfer heat) that require proper handling and disposal. Concentrated solar systems may need to be cleaned regularly with water, which is also needed for cooling the turbine-generator. Using water from underground wells may affect the ecosystem in some arid locations.



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9.3: Biomass Energy

Biomass energy is from the energy stored in materials of biological origin such as plants and animals. Biomass energy is the oldest energy source used by humans. Until the Industrial Revolution prompted a shift to fossil fuels in the mid-18th century, biomass energy was the world's dominant fuel source. It includes *direct combustion* of solid biomass to provide energy for heating, cooking, and even generating electricity. The most common source for direct combustion is wood, but energy can also be generated by burning animal manure (dung), herbaceous plant material (non-wood), peat (partially decomposed plant and animal tissues), or converted biomass such as charcoal (wood that has been partially burned to produce a coal-like substance). Biomass can also be converted into a liquid biofuels used to power vehicles such as ethanol from corn, sugarcane residue and soybeans or even used cooking oil for biodiesel. Biomass energy can also be in gaseous form such as methane. Currently, about 12 percent of the world's energy comes from biomass (Figure **9.3.2**). Biomass is most frequently used as a fuel source in many developing nations, but with the decline of fossil fuel availability and the increase in fossil fuel prices, biomass is increasingly being used as a fuel source even in developed nations.

5.3.1: Direct combustion of solid biomass

Using wood and charcoal made from wood, for heating and cooking can replace fossil fuels and may result in lower CO_2 emissions. If wood is harvested from forests or woodlots that have to be thinned or from urban trees that fall down or needed be cut down anyway, then using it for biomass does not impact those ecosystems. However, wood smoke contains harmful pollutants like carbon monoxide and particulate matter. For home heating, it is most efficient and least polluting when using a modern wood stove or replace insert that are designed to release small amounts of particulates. However, in places where wood and charcoal are major cooking and heating fuels such as in industrializing countries, the wood may be harvested faster than trees can grow resulting in deforestation.

9Biomass is also being used on a larger scale, where small power plants are powered by biomass such as woodchips (Figure **9.3.1**). For instance, Central State Hospital, Milledgeville, GA has a woodchip burning plant that was the most advanced system available for its time and operating today. Colgate University in Hamilton, New York, has had a wood-burning boiler since the mid-1980's that processes about 20,000 tons of locally and sustainably harvested wood chips, the equivalent of 1.17 million gallons (4.43 million liters) of fuel oil, avoiding 13,757 tons of emissions, and saving the university over $1.8 million in heating costs. The University's steam-generating wood-burning facility now satisfies more than 75 percent of the campus's heat and domestic hot water needs. For more information about this, click here Colgate University.



Figure **9.3.1**: A biomass (woodchip) power plant in Germany. Image from Wikimedia Public Domain.

Waste products of various industries and processes such as lumber mill sawdust, paper mill sludge, yard waste, oat hulls from an oatmeal processing plant, woody debris from logging, organic waste from feedlots, and residue from crops can also be used for energy. Waste to energy processes are gaining renewed interest as they can solve two problems at once: 1) disposal of waste as landfill capacity decreases, and 2) production of energy from a renewable resource. In the United States, several plants have been constructed to burn urban biomass waste and use the energy to generate electricity. Many of the environmental impacts are similar to those of a coal plant like air pollution, ash generation, etc. Since the fuel source is less standardized than coal and hazardous materials may be present in municipal solid waste (MSW), or garbage, incinerators and waste-to-energy power plants need to clean

the stack gases of harmful materials. The U.S. EPA regulates these plants very strictly and requires anti-pollution devices to be installed. If not contained, waste is distributed in many ecosystems. The ash from these plants may contain high concentrations of various metals that were present in the original waste. If ash is clean enough it can be recycled as a MSW landfill cover or to build roads, cement block and artificial reefs. Also, while incinerating at high temperature many of the toxic chemicals may break down into less harmful compounds.

5.3.2: Gaseous Biomass

Organic material can be converted to methane, the main component of natural gas, by anaerobic decomposition or fermentation, a process that utilizes anaerobic bacteria. Methane is a relatively clean fuel that burns efficiently. It can be generated from any kind of organic waste such as municipal sewage and garbage, livestock manure, kitchen, and garden scraps. In fact, municipal landfills are active sites of methane production contributing annually to methane in the atmosphere and to global warming. This gas can and is currently being captured by numerous landfills around the United States that burn it to generate electricity at power plants or supply it to homes for heating. The electricity may replace electricity produced by burning fossil fuels and result in a net reduction in CO_2 emissions. Burning methane produced from manure provides more heat than burning the dung itself, and the sludge left over from bacterial digestion is a rich fertilizer, containing healthy bacteria as well as most of the nutrients originally in the dung. The main environmental impacts are from the construction of the plant itself. Burning methane releases CO_2 and although CO_2 is a greenhouse gas, its global warming potential is much lower than that of methane (see chapter 7). Also, since this methane if from organic waste resulting from ongoing photosynthetic processes, it is considered carbon-neutral, unlike CO_2 from fossil fuels.

5.3.3: Liquid Biofuels

Biofuels are transportation fuels produced from plant sources and used to power vehicles. The most common ones are ethanol and biodiesel. **Ethanol**, also known as ethyl alcohol or grain alcohol, is produced by fermenting crops such as corn, sugarcane, and other crops and then mixed with conventional gasoline. As an additive, ethanol lowers reliance on conventional oil and increases the combustion efficiency of gasoline, reducing pollutant emissions. In Brazil, which has a sizeable ethanol industry based on sugarcane, all gasoline sold contains 25 percent alcohol, and over 70 percent of the cars sold each year can run on either ethanol or gasoline. Ethanol-gasoline mixtures burn cleaner than pure gasoline but are more volatile and thus have higher "evaporative emissions" from fuel tanks and dispensing equipment. These emissions contribute to the formation of harmful, ground level ozone and smog. Gasoline requires extra processing to reduce evaporative emissions before it is blended with ethanol.

Biodiesel which is essentially vegetable oil, can also be derived from a wide range of plant sources, including rapeseed, sunflowers, and soybeans, and can be used in most conventional diesel engines. Biodiesel can also be made from used vegetable oil and has been produced on a very local basis. Because it burns more cleanly than its petroleum-based counterpart, biodiesel can reduce pollution from heavy-duty vehicles such as trucks and buses. Compared to petroleum diesel, biodiesel combustion produces less sulfur oxides, particulate matter, carbon monoxide, and unburned and other hydrocarbons, but more nitrogen oxide.

Calculating the net energy or CO_2 generated or reduced in the process of producing the biofuel is crucial to determining its environmental impact. Biofuels may be derived from parts of plants not used for food (cellulosic biomass) thus reducing that impact. Cellulosic ethanol feedstock includes native prairie grasses, fast growing trees, sawdust, and even waste paper. Also, in some parts of the world, large areas of natural vegetation and forests have been cut down to grow sugarcane for ethanol and soybeans and palm-oil trees to make biodiesel. This is not sustainable land use.

Biomass energy may be considered to be **carbon-neutral** because the plants that are used to make them (such as corn and sugarcane for ethanol, or soy beans and palm oil trees for biodiesel) take up CO_2 from the atmosphere through photosynthesis as they grow and may offset the CO_2 produced when burned. If the biomass is not burned for energy generation, the carbon contained in the biomass would still be returned to the atmosphere as CO_2 when the organisms die and decompose to complete the cycle. Burning biomass, therefore, does not result in new additional CO_2 in the atmosphere but rather returns carbon that came from the atmosphere and used to produce the biomass in the first place. On the other hand, burning fossil fuels introduces new carbon into the atmosphere that was previously stored deep in the Earth's crust and would have stayed there if we (humans) did not extract it.

5.3.4: Environmental Impacts of Biomass Energy

A major challenge of biomass is determining if it is really a more sustainable option. The energy content of some biomass energy sources may not be as high as fossil fuels so more must be burned to generate the same energy. It often takes energy to make energy and biomass is one example where the processing to make it may not be offset by the energy it produces. If traditional monoculture crops like corn or soybeans are used, they require major quantities of fossil fuel to manufacture fertilizer, run farm machines, and ship the fuel to markets, so these biofuels do not always offer significant net energy savings over gasoline and diesel fuel. In such instances, biofuels may not be carbon-neutral because the process of producing the biofuels results in more CO_2 added to the atmosphere than that removed by the growing crops. Even if the environmental impact is net positive, for example if renewable energy sources are used to make the biofuels, the economic and social effects of growing plants for fuels need to be considered. The land, fertilizers, water, and energy used to grow biofuel crops could be used to grow food crops instead. The competition of land for fuel vs. food can increase the price of food, which has a negative effect on society. It could also decrease the food supply and increase malnutrition and starvation globally.

Trees that are cut for firewood are frequently not replanted. In order to be used sustainably, one tree must be planted for every one cut down. If too much biomass is taken it can reduce forest and grassland contributions to ecosystem services. Forests and grasslands help take CO_2 out of the atmosphere through photosynthesis and the loss of photosynthetic activity results in increased amounts of CO_2 in the atmosphere and contribute to global warming since CO_2 is a greenhouse gas. Burning biomass directly (wood, manure, etc.) produces high particulate material pollution (see chapter 6 on Air Pollution), produces CO_2 and deprives the soil of nutrients it normally would have received from the decomposition of the organic matter. Each type of biomass energy source, therefore, must be evaluated for its full life-cycle impact in order to determine if it is really advancing sustainability and reducing environmental impacts.

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9.4: Wind Power

Wind is a renewable energy source that uses the power of moving air to generate electricity. Wind turbines use blades to collect the wind's kinetic energy. Wind flows over the blades creating lift (similar to the effect on airplane wings), which causes the blades to turn. The blades are connected to a drive shaft that turns an electric generator, which produces electricity (Figure **9.4.1**). Wind turbines are becoming a more prominent sight across the United States, even in regions that are considered to have less wind potential. Wind turbines (often called windmills) do not release emissions that pollute the air or water (with rare exceptions), and they do not require water for cooling. The U.S. wind industry had 40,181 MW of wind power capacity installed at the end of 2010, with 5,116 MW installed in 2010 alone, providing more than 20 % of installed wind power around the globe. According to the American Wind Energy Association, over 35 % of all new electrical generating capacity in the United States since 2006 was due to wind, surpassed only by natural gas.

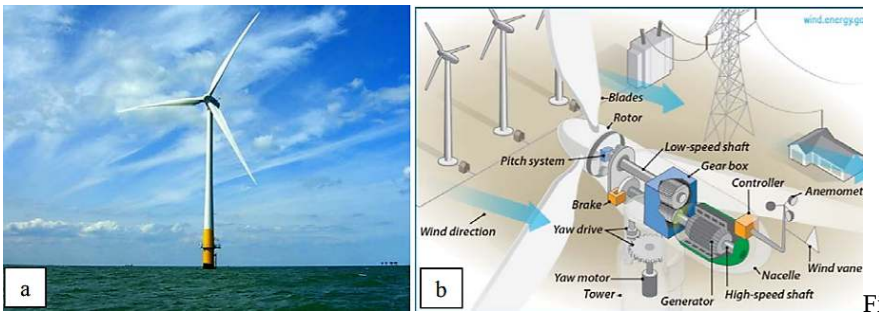


Figure **9.4.1**: a) This wind turbine in the Thames Estuary in the UK is an example of induction at work, (credit: Phault, Flickr). Wind pushes the blades of the turbine, spinning a shaft attached to magnets, see image b for parts. The magnets spin around a conductive coil, inducing an electric current in the coil, and eventually feeding the electrical grid. Follow this link for an Animation.

Most windmills generate about 1kW of electricity, which is only practical for decentralized power generation. California has about 17,000 windmills with a capacity of about 1,400 MW. This is about 80% of all windmills in the U.S. In West Europe windmill generators are quite common. Since a wind turbine (Figure **9.4.1**) has a small physical footprint relative to the amount of electricity it produces, many wind farms are located on crop, pasture, forest land, or coastal areas. They contribute to economic sustainability by providing extra income to farmers and ranchers, allowing them to stay in business and keep their property from being developed for other uses. For example, energy can be produced by installing wind turbines in the Appalachian Mountains of the United States instead of engaging in mountain top removal for coal mining.

5.4.1: Environmental Impacts of Wind Energy

Offshore wind turbines on lakes or the ocean may have smaller environmental impacts than turbines on land. Wind turbines do have a few environmental challenges. There are aesthetic concerns to some people when they see them on the landscape. A few wind turbines have caught on fire, and some have leaked lubricating fluids, though this is relatively rare. Some people do not like the sound that wind turbine blades make. Turbines have been found to cause bird and bat deaths particularly if they are located along their migratory path. This is of particular concern if these are threatened or endangered species. There are ways to mitigate that impact and it is currently being researched. There are some small impacts from the construction of wind projects or farms, such as the construction of service roads, the production of the turbines themselves, and the concrete for the foundations. However, overall life cycle analysis has found that turbines make much more energy than the amount used to make and install them.

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9.5: Geothermal Power

Geothermal energy uses heat from the Earth's internal geologic processes in order to produce electricity or provide heating. The subsurface temperature of the Earth provides an endless energy resource. One source of geothermal energy is steam. Groundwater percolates down through cracks in the subsurface rocks until it reaches rocks heated by underlying magma, and the heat converts the water to steam. Sometimes this steam makes its way back to the surface in the form of a geyser or hot spring. Wells can be dug to tap the steam reservoir and bring it to the surface, to drive generating turbines and produce electricity (Figure 9.5.1a). Hot water can be circulated to heat buildings. Regions near tectonic plate boundaries have the best potential for geothermal activity.

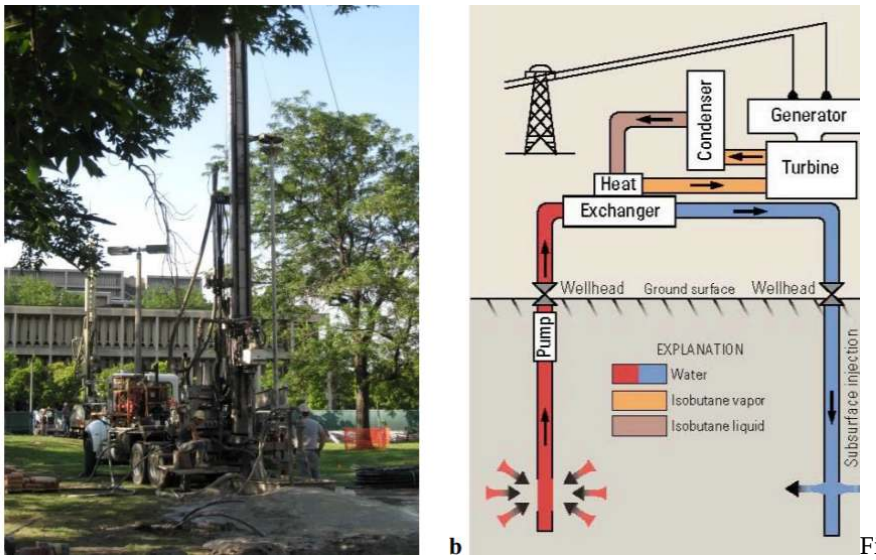


Figure 9.5.1: a) Installing a Geothermal Pipe System Drilling to install geothermal ground source pipe system. Source: Office of Sustainability. b) Electricity generation at a moderate-temperature hydrothermal system. The geothermal water is used to boil a second fluid (isobutane in this example) whose vapor then drives a turbine generator. The wastewater is injected back into the subsurface. Source: USGS.

5.5.1: Environmental Impacts of Geothermal Energy

The environmental impact of geothermal energy depends on how it is being used. Direct use and heating applications have almost no negative impact on the environment. Geothermally heated water can release dissolved gases, including carbon dioxide, methane, ammonia and hydrogen sulfide, although these are usually in very small quantities when compared to those released from fossil fuel plants. In addition, geothermal plants use scrubber systems to clean the air of hydrogen sulfide that is naturally found in the steam and hot water. They emit 97 % less acid rain-causing sulfur compounds than are emitted by fossil fuel plants.

Even though geothermal energy is renewable, not every plant built to capture this energy will be able to operate indefinitely because the energy relies on groundwater recharge. If the heated water is used faster than the recharge rate of groundwater, the plant will eventually run out of water. The Geysers in California started experiencing this and operators responded by injecting municipal wastewater into the ground to replenish the supply. Also, patterns of geothermal activity in the Earth’s crust naturally shift over time and an area that produces hot groundwater now may not always so do. The water of many hot springs is laced with salts and minerals that corrode equipment, shorten the lifetime of plants and increase maintenance costs.

Electrical power is restricted to regions where energy can be tapped from naturally heated groundwater but most areas of the world are not rich in naturally heated groundwater. Engineers are trying to overcome this by drilling deeply into dry rock, fracture the rock and pump in cold water which becomes heated and drawn up through an outlet well and used to generate power. However, this approach is said to trigger minor earthquakes.

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9.6: Hydroelectric Power (Hydropower)

This is the second largest source of renewable energy used, next to biomass energy. The majority of hydropower currently comes from dams built across a river to block the flow of river water. The water stored behind the dam contains potential energy (see chapter 4) and when released, the potential energy is converted to kinetic energy as the water rushes down. This energy is used to turn blades of turbines and causing a generator to generate electricity. Electricity generated in the powerhouse of a dam is transmitted to the electric grid by transmission lines while the water flows into the riverbed below the dam and continues down river. An alternative approach considered less disruptive involves diverting a portion of the river's water through a pipe or channel and passed through a powerhouse to generate electricity and returned to the river. Another approach involves pumping water from a lower reservoir to a higher reservoir and then allowed to flow downhill through a turbine, generating electricity. This approach, however, requires energy input to pump the water.

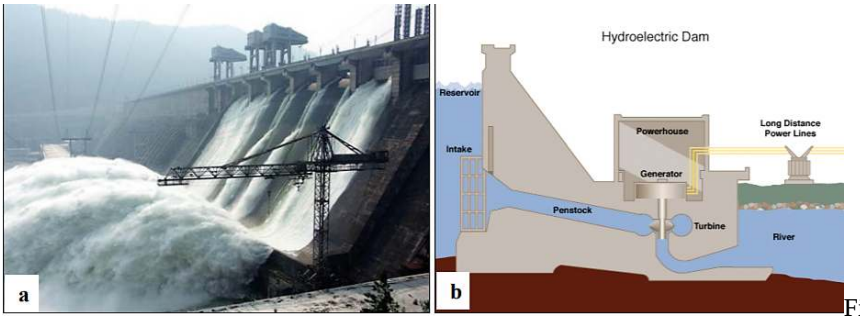


Figure **9.6.1**: a) Hydroelectric facility (credit: Denis Belevich, Wikimedia Commons); b) An illustration showing how water is funneled through a portion of the dam to rotate turbines and generate electricity. (Image from Wikimedia, public domain).

Figure **9.6.2** shows the Hoover Power Plant located on the Colorado River. In the U.S., hydroelectric plants account for about 10% of total production and account for about 35 % of the United States' renewable energy consumption. In 2003 capacity was at 96,000 MW and it was estimated that 30,000 MW capacity is undeveloped. Energy produced can be calculated and modeled as shown in Figure **9.6.3**.



Figure **9.6.2**: a) A view of the Hoover dam from above (credit Adam Kliczek). b) Hydroelectric generators at the Hoover dam (credit: Jon Sullivan).

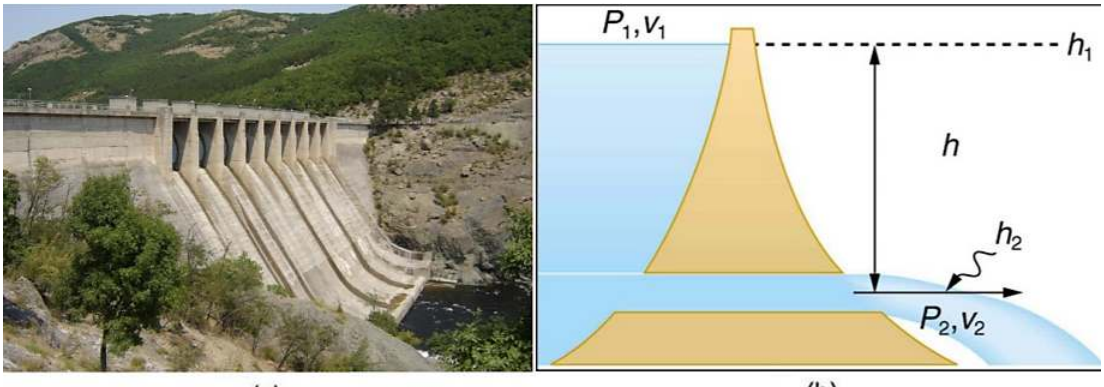


Figure **9.6.3**: (a) Water gushes from the base of the Studen Kladenetz dam in Bulgaria. (credit: Kiril Kapustin; http://www.ImagesFromBulgaria.com) (b)

In the absence of significant resistance, water flows from the reservoir with the same speed it would have if it fell the distance h without friction.

5.6.1: Environmental Impacts of Hydroelectric Power

Hydropower (hydro-electric) is considered a clean and renewable source of energy since it does not directly produce emissions of air pollutants and the source of power is regenerated. However, hydropower dams, reservoirs, and the operation of generators can have serious environmental impacts. A dam that is used to create a reservoir or to divert water to a run-ofriver hydropower plant can obstruct migration of fish to their upstream spawning areas in areas where salmon must travel upstream to spawn, such as along the Columbia River in Washington and Oregon. Hydro turbines kill and injure some of the fish that pass through the turbine although there are ways to reduce that effect. This problem can be partially alleviated by using 'fish ladders' that help the salmon get up the dams.

A reservoir and operation of the dam can affect the natural water habitat due to changes in water temperatures, chemistry, flow characteristics, and silt loads, all of which can lead to significant changes in the ecology and physical characteristics of the river upstream and downstream. Construction of reservoirs may cause natural areas, farms, and archeological sites to be covered and force populations to relocate and result in the loss of scenic rivers.

Carbon dioxide and methane may also form in reservoirs where water is more stagnant and be emitted to the atmosphere. The exact amount of greenhouse gases produced from hydropower plant reservoirs is uncertain. If the reservoirs are located in tropical and temperate regions, including the United States, those emissions may be equal to or greater than the greenhouse effect of the carbon dioxide emissions from an equivalent amount of electricity generated with fossil fuels (EIA, 2011 p. 333).

5.6.2: Potential of Tidal Power

Tidal power involves placing turbines in zones of the ocean with significant tides and currents, and using the power of flowing water to turn the blades of a turbine to generate electricity. Ocean power systems are still being researched and currently still experimental. For example, the Bay of Fundy, which has a 15 m tide, a dam constructed across the estuary would let water enter on the incoming tide, then release the water through turbines at low tide. The energy potential is great http://www.ialtenergy.com/tidal-power-news.html , and so is the environmental cost. Tapping tidal energy resources involves building major dams on inlets and estuaries that are prized for other purposes, so few tidal energy facilities have been developed. Harnessing waves and currents on a significant scale will involve designing turbine structures that are large, inexpensive, and can operate for long periods under the physical stresses and corrosive forces of ocean environments. Though proposed, a tidal power plant has not been constructed at Fundy. There is a 240,000 kW tidal plant at La Rance, France.

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9.7: Other Alternative Renewable Energy Sources

5.7.1: Hydrogen

Hydrogen gas may be an important clean fuel of the future. Hydrogen is considered an energy carrier, like electricity and batteries, it carries energy that can be converted for use later. Hydrogen gas does not tend to exist freely but rather hydrogen atoms bind to other atoms and molecules becoming incorporated in everything from water to organic compounds. Therefore, to obtain hydrogen gas for fuel, energy is needed to force these substances to release their hydrogen atoms. One such procedure is known as **electrolysis** in which an electric current is passed through water to decompose the water molecule into oxygen and hydrogen (Figure **9.7.1**). Hydrogen can also be produced from hydrocarbons such as natural gas and coal, fermentation of plant waste material, and using algae.

The nation of Iceland is attempting to become the first hydrogen-based energy economy using its abundant geothermal energy resources. Some energy experts believe that combining hydrogen fuel and electricity could serve as a basis for a clean, safe, and energy efficient energy system. Electricity generated from intermittent renewable sources such as wind and solar can be used to produce hydrogen fuel for fuel cells that would then generate electricity to power vehicles, computers, heat homes and many other uses. An energy system based on hydrogen would alleviate dependence on foreign fuels and help fight climate change. Hydrogen is the most abundant element in the universe and we will never run out of it.

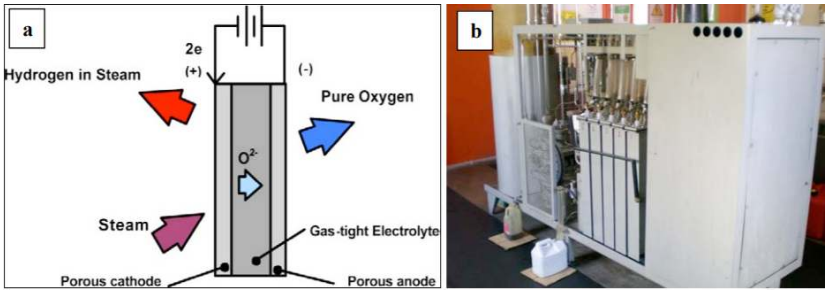


Figure **9.7.1**: a) An electric current passed through water resulting in the separation of hydrogen atoms from oxygen to produce the hydrogen fuel. b) An electrolyser used in the production of hydrogen by CambridgeBayWeather. After production the hydrogen is used to inflate weather balloons to which a radiosonde is attached. The oxygen produced is vented to the outside.

Fuel Cells

Fuel cells are highly efficient power plants that produce electricity using hydrogen fuel in a chemical reaction that is a reverse of the electrolysis process that produced the hydrogen fuel (Figure **9.7.2**). Energy is released by an exothermic electrochemical reaction that combines hydrogen and oxygen ions through an electrolyte material to generate electricity and heat. Experimental fuel cells that can power automobiles have been developed (Figure **9.7.2**).



Figure **9.7.2**: The electrochemical processes showing how hydrogen fuel is combined with oxygen generating heat energy along with water as a waste product. Some vehicles that use the fuel cell include buses (photo is a hydrogen fuel bus on Tower Bridge London, photo by Sludge G.), cars (The Toyota Fine N car based on fuel cell technology, photo by Chris 73, CC BY-SA 3.0) and motorcycles (Suzuki Burgman Fuel Cell cutaway model shown here, Photo by Mario CC BY-SA 3.0).

Challenges of Hydrogen

Currently, the infrastructure for using hydrogen fuel is lacking and converting a nation such as the United States to hydrogen would require massive and costly development of facilities to produce, store, transport, and provide the fuel. The environmental impact of hydrogen production itself depends on the source of material used to supply the hydrogen. For example, biomass and fossil fuel sources result in carbon-based emissions. Some research suggests that leakage of hydrogen from its production, transport, and use could potentially deplete stratospheric ozone. Research into this is still ongoing.

5.7.2: Electric cars

Electric cars are vehicles that solely depend on electricity to work. They are charged using fossil fuel energy sources (Figure **9.7.3** a), but they are very efficient. Although it might seem like science fiction the electric car is already here, and has been here for the last hundred years. The main problem with the electric car is that it cannot go very far before it needs to be recharged, something that takes between three to six hours. Different models have been explored (Figure **9.7.3**b), but for now most cars are equipped as a hybrid with a back up reservoir using fossil fuel.

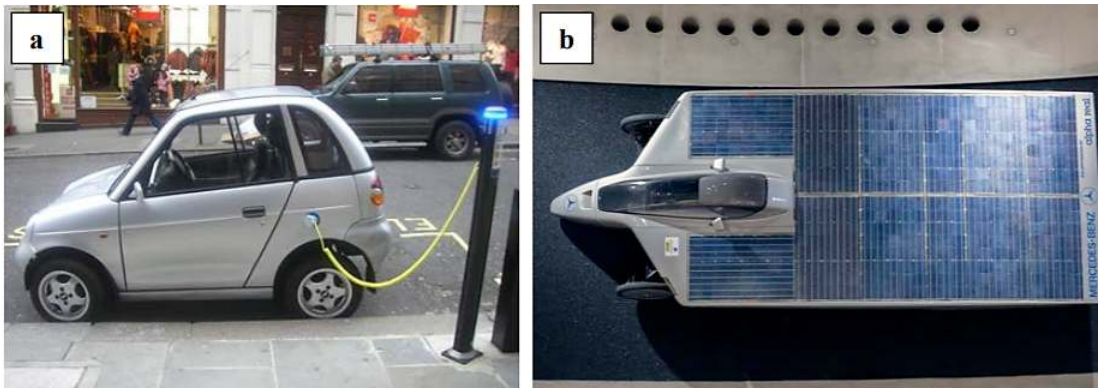


Figure 5.7.3: a) This REVAi, an electric car, gets recharged on a street in London. (credit: Frank Hebbert). b) 1985 Mercedes-Benz/Alpha Real "Tour de Sol" Solarmobile; built for the 1985 Tour de Sol Centennial Rally, solar car race from Lake Constance to Lake Geneva.

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CHAPTER OVERVIEW

10: Nuclear Energy

10.1: Prelude to Nuclear Chemistry
10.2: Radioactivity
10.3: Uses of Radioactive Isotopes
10.4: Nuclear Energy

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10.1: Prelude to Nuclear Chemistry

Most of us have at least one device in our homes that guards our safety and, at the same time, depends on radioactivity to operate properly. This device is a smoke detector.

A typical smoke detector contains an electric circuit that includes two metal plates about 1 cm apart. A battery in the circuit creates a voltage between the plates. Next to the plates is a small disk containing a tiny amount (∼0.0002 g) of the radioactive element americium. The radioactivity of americium ionizes the air between the plates, causing a tiny current to constantly flow between them. (This constant drain on the battery explains why the batteries in smoke detectors should be replaced regularly, whether the alarm has been triggered or not.) When particles of smoke from a fire enter the smoke detector, they interfere with the ions between the metal plates, interrupting the flow of current. When the current drops beneath a set value, another circuit triggers a loud alarm, warning of the possible presence of fire.



Figure **10.1.1**: Smoke Detector. (CC SA -BY 3.0; Tumi-1983).

Although radioactive, the americium in a smoke detector is embedded in plastic and is not harmful unless the plastic package is taken apart, which is unlikely. Nuclear chemistry is often thought of only in connection with the nuclear power industry and atomic bombs. Although many people have an unfounded fear of radioactivity, smoke detectors save thousands of lives every year. The applications of nuclear chemistry may be more widespread than you think.

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10.2: Radioactivity

Learning Objective

- Define and give examples of the major types of radioactivity.

We saw in Chapter 3 that atoms are composed of subatomic particles—protons, neutrons, and electrons. Protons and neutrons are located in the nucleus and provide most of the mass of an atom, while electrons circle the nucleus in shells and subshells and account for an atom's size.

In Chapter 2, we also introduced the notation for succinctly representing an isotope of a particular atom:

$$^{12}_{6}\mathrm{C}$$

The element in this example, represented by the symbol C, is carbon. Its atomic number, 6, is the subscript next to the symbol and is the number of protons in the atom. The mass number, the superscript next to the symbol, is the sum of the number of protons and neutrons in the nucleus of this particular isotope. In this case, the mass number is 12, which means that the number of neutrons in the atom is 12 − 6 = 6 (that is, the mass number of the atom minus the number of protons in the nucleus equals the number of neurons). Occasionally, the atomic number is omitted in this notation because the symbol of the element itself conveys its characteristic atomic number. The two isotopes of hydrogen—$^{2}\mathrm{H}$ and $^{3}\mathrm{H}$—are given their own names and symbols: deuterium (D) and tritium (T), respectively.

Atomic theory in the nineteenth century presumed that nuclei had fixed compositions. But in 1896, the French scientist Henri Becquerel found that a uranium compound placed near a photographic plate made an image on the plate, even if the compound was wrapped in black cloth. He reasoned that the uranium compound was emitting some kind of radiation that passed through the cloth to expose the photographic plate. Further investigations showed that the radiation was a combination of particles and electromagnetic rays, with its ultimate source being the atomic nucleus. These emanations were ultimately called, collectively, **radioactivity**.

There are three main forms of radioactive emissions. The first is called an **alpha particle**, which is symbolized by the Greek letter α. An alpha particle is composed of 2 protons and 2 neutrons and is the same as a helium nucleus. We often use:

$$^{4}_{2}\mathrm{He}$$

to represent an alpha particle. It has a 2+ charge. When a radioactive atom emits an alpha particle, the original atom's atomic number decreases by 2 (because of the loss of 2 protons), and its mass number decreases by 4 (because of the loss of 4 nuclear particles). We can represent the emission of an alpha particle with a chemical equation. For example, the alpha-particle emission of uranium-235 is as follows:

$$^{235}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{231}_{90}\mathrm{Th}$$

Rather than calling this equation a chemical equation, we call it a **nuclear equation** to emphasize that the change occurs in an atomic nucleus. How do we know that a product of this reaction is \(\ce{^{90}_{231}T}\)? We use the law of conservation of matter, which says that matter cannot be created or destroyed. This means that we must have the same number of protons and neutrons on both sides of the nuclear equation. If our uranium nucleus loses 2 protons, there are 90 protons remaining, identifying the element as thorium. Moreover, if we lose four nuclear particles of the original 235, there are 231 remaining. So we use subtraction to identify the isotope of the Th atom—in this case,

$$^{231}_{90}\mathrm{Th}$$

Chemists often use the names **parent isotope** and **daughter isotope** to represent the original atom and the product other than the alpha particle. In the previous example,

$$^{235}_{92}\mathrm{U}$$

is the parent isotope, and

$$^{231}_{90}\text{Th}$$

is the daughter isotope. When one element changes into another in this manner, it undergoes **radioactive decay**.

Example 10.2.1

Write the nuclear equation that represents the radioactive decay of radon-222 by alpha particle emission and identify the daughter isotope.

Solution

Radon has an atomic number of 86, so the parent isotope is represented as

$$^{222}_{86}\text{Rn}$$

.

We represent the alpha particle as

$$^{4}_{2}\text{He}$$

Use subtraction (222 − 4 = 218 and 86 − 2 = 84) to identify the daughter isotope as polonium:

$$^{222}_{86}\text{Rn} \rightarrow {}^{4}_{2}\text{He} + {}^{218}_{84}\text{Th}$$

Exercise 10.2.2

Write the nuclear equation that represents radioactive decay of polonium-208 by alpha particle emission and identify the daughter isotope.

Answer

$$^{208}_{80}\text{Po} \rightarrow {}^{4}_{2}\text{He} + {}^{204}_{82}\text{Pb}$$

$$^{204}_{82}\text{Pb}$$

The second major type of radioactive emission is called a **beta particle**, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus (not from the shells of electrons about the nucleus) and has a 1− charge. We can also represent a beta particle as e−10. The net effect of beta particle emission on a nucleus is that a neutron is converted to a proton. The overall mass number stays the same, but because the number of protons increases by one, the atomic number goes up by one. Carbon-14 decays by emitting a beta particle:

$$^{14}_{6}\text{C} \rightarrow {}^{14}_{7}\text{N} + {}^{0}_{-1}\text{e}$$

Again, the sum of the atomic numbers is the same on both sides of the equation, as is the sum of the mass numbers. (Note that the electron is assigned an "atomic number" of –1, equal to its charge.)

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called **gamma rays**, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed *ionizing radiation*). For example, in the radioactive decay of radon-222, both alpha and gamma radiation are emitted, with the latter having an energy of 8.2×10^{-14} J per nucleus decayed:

$$^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He} + \gamma$$

This may not seem like much energy, but if 1 mol of Rn atoms were to decay, the gamma ray energy would be 4.9×10^{7} kJ!

Example 10.2.3

Write the nuclear equation that represents the radioactive decay of boron-12 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is B-12,

$$^{12}_{5}\mathrm{B}$$

while one of the products is B-12,

$$^{0}_{-1}\mathrm{e}$$

So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 12, and its atomic number must be 6. The element having an atomic number of 6 is carbon. Thus the complete nuclear equation is as follows:

$$^{12}_{5}\mathrm{B} \rightarrow {}^{12}_{6}\mathrm{C} + {}^{0}_{-1}\mathrm{e} + \gamma$$

The daughter isotope is carbon-12.

Exercise 10.2.4

Answer

$$^{133}_{43}\mathrm{Tc} \rightarrow {}^{133}_{44}\mathrm{Ru} + {}^{0}_{-1}\mathrm{e} + \gamma$$

Alpha, beta, and gamma emissions have different abilities to penetrate matter. The relatively large alpha particle is easily stopped by matter (although it may impart a significant amount of energy to the matter it contacts). Beta particles penetrate slightly into matter, perhaps a few centimeters at most. Gamma rays can penetrate deeply into matter and can impart a large amount of energy into the surrounding matter. Table 10.2.1 summarizes the properties of the three main types of radioactive emissions.

Table 10.2.1 The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, H24	β $^{0}_{-1}$e	γ
identity	helium nucleus	electron	electromagnetic radiation
charge	2+	1−	none
mass number	4	0	0
penetrating power	minimal (will not penetrate skin)	short (will penetrate skin and some tissues slightly)	deep (will penetrate tissues deeply)

Occasionally, an atomic nucleus breaks apart into smaller pieces in a radioactive process called **spontaneous fission (or fission)**. Typically, the daughter isotopes produced by fission are a varied mix of products, rather than a specific isotope as with alpha and beta particle emission. Often, fission produces excess neutrons that will sometimes be captured by other nuclei, possibly inducing additional radioactive events. Uranium-235 undergoes spontaneous fission to a small extent. One typical reaction is

$$^{235}_{92}\mathrm{Tc} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 2\,{}^{1}_{0}\mathrm{n}$$

where $^{1}_{0}\mathrm{n}$ is a neutron. As with any nuclear process, the sums of the atomic numbers and mass numbers must be the same on both sides of the equation. Spontaneous fission is found only in large nuclei. The smallest nucleus that exhibits spontaneous fission is lead-208. (Fission is the radioactive process used in nuclear power plants and one type of nuclear bomb.)

Key Takeaways

- The major types of radioactivity include alpha particles, beta particles, and gamma rays.
- Fission is a type of radioactivity in which large nuclei spontaneously break apart into smaller nuclei.

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10.3: Uses of Radioactive Isotopes

Learning Objective

- Learn some applications of radioactivity.

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

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

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

Radioactive Dating

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

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

One isotope, carbon-14, is particularly useful in determining the age of once-living artifacts. A tiny amount of carbon-14 is produced naturally in the upper reaches of the atmosphere, and living things incorporate some of it into their tissues, building up to a constant, albeit very low, level. Once a living thing dies, it no longer acquires carbon-14; as time passes the carbon-14 that was in the tissues decays. (The half-life of carbon-14 is 5,370 y.) If a once-living artifact is discovered and analyzed many years after its death and the remaining carbon-14 is compared to the known constant level, an approximate age of the artifact can be determined. Using such methods, scientists determined that the age of the Shroud of Turin (Figure **10.3.1** purported by some to be the burial cloth of Jesus Christ and composed of flax fibers, a type of plant) is about 600–700 y, not 2,000 y as claimed by some. Scientists were also able to use radiocarbon dating to show that the age of a mummified body found in the ice of the Alps was 5,300 y.

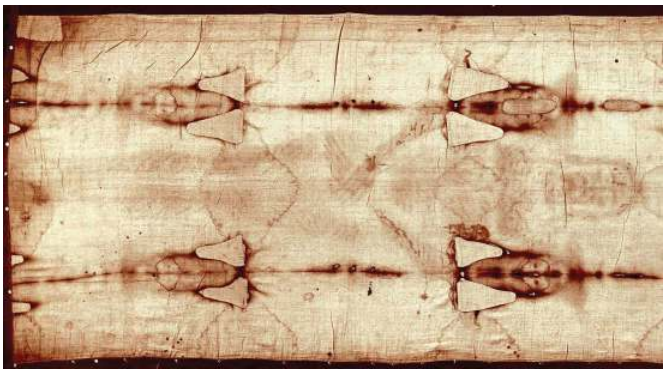


Figure **10.3.1**: Shroud of Turin. In 1989, several groups of scientists used carbon-14 dating to demonstrate that the Shroud of Turin was only 600–700 y. Many people still cling to a different notion, despite the scientific evidence.

Irradiation of Food

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

Medical Applications

Radioactive isotopes have numerous medical applications in diagnosing and treating illness and diseases. One example of a diagnostic application is using radioactive iodine-131 to test for thyroid activity (Figure **10.3.1**). The thyroid gland in the neck is one of the few places in the body with a significant concentration of iodine. To evaluate thyroid activity, a measured dose of ^{131}I is administered to a patient, and the next day a scanner is used to measure the amount of radioactivity in the thyroid gland. The amount of radioactive iodine that collects there is directly related to the activity of the thyroid, allowing trained physicians to diagnose both hyperthyroidism and hypothyroidism. Iodine-131 has a half-life of only 8 d, so the potential for damage due to exposure is minimal. Technetium-99 can also be used to test thyroid function. Bones, the heart, the brain, the liver, the lungs, and many other organs can be imaged in similar ways by using the appropriate radioactive isotope.

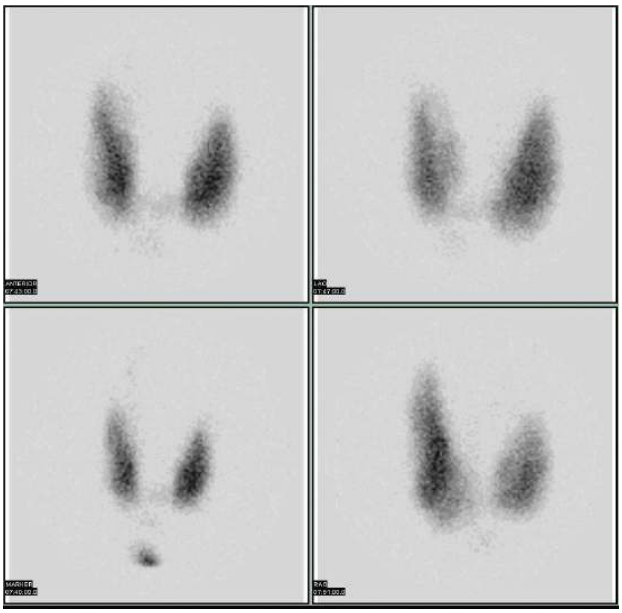


Figure **10.3.2**: Medical Diagnostics. Radioactive iodine can be used to image the thyroid gland for diagnostic purposes. Source: Scan courtesy of Myo Han, en.Wikipedia.org/wiki/File:Thyroid_scan.jpg.

Very little radioactive material is needed in these diagnostic techniques because the radiation emitted is so easy to detect. However, therapeutic applications usually require much larger doses because their purpose is to preferentially kill diseased tissues. For example, if a thyroid tumor were detected, a much larger infusion (thousands of rem, as opposed to a diagnostic dose of less than 40 rem) of iodine-131 could help destroy the tumor cells. Similarly, radioactive strontium is used to not only detect, but also to ease the pain of bone cancers. Table **10.3.1** lists several radioactive isotopes and their medical uses.

Table **10.3.1**: Some Radioactive Isotopes with Medical Applications

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

*The "m" means that it is a metastable form of this isotope of technetium.

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

Food and Drink Application: Radioactivity in Wines

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

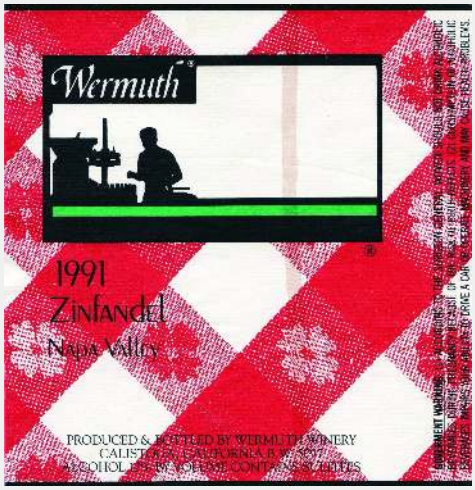


Figure **10.3.3**: A wine label. This wine label from a bottle of wine claims a vintage of 1991. Is the wine really from this vintage, or is it a fake? Radioactivity can help determine the answer. Source: Used by permission of Ralph E. Wermuth.

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

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

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

Summary

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

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10.4: Nuclear Energy

Learning Objectives

- Explain where nuclear energy comes from.
- Describe the difference between fission and fusion.

Nuclear changes occur with a simultaneous release of energy. Where does this energy come from? If we could precisely measure the masses of the reactants and products of a nuclear reaction, we would notice that the amount of mass drops slightly in the conversion from reactants to products. Consider the following nuclear equation, in which the molar mass of each species is indicated to four decimal places:

$$^{235}_{235.0439}\text{U} \rightarrow ^{139}_{138.9088}\text{Ba} + ^{94}_{93.9343}\text{Kr} + ^{2^1}_{2\times1.0087}\text{n}$$

If we compare the mass of the reactant (235.0439) to the masses of the products (sum=234.8605) we notice a mass difference of -0.1834 g or -0.0001834 kg. Where did this mass go?

According to Albert Einstein's theory of relativity, energy (E) and mass (m) are related by the following equation:

$$E = mc^2$$

where c is the speed of light, or

$$c = 3.00 \times 10^8\, m/s$$

In the course of the chemical reaction for uranium, the mass difference is converted to energy, which is given off by the reaction:

$$E = (-0.0001834\, kg)(3.00 \times 10^8\, m/s)^2 = -1.65 \times 10^{13} J = -1.65 \times 10^{10} kJ$$

(For the units to work out, mass must be expressed in units of kilograms.) That is, 16.5 billion kJ of energy is given off every time 1 mol of uranium-235 undergoes this nuclear reaction. This is an extraordinary amount of energy. Compare it to combustion reactions of hydrocarbons, which give off about 650kJ/mol of energy for every CH_2 unit in the hydrocarbon, on the order of *hundreds* of kilojoules per mole. Nuclear reactions give off *billions* of kilojoules per mole.

If this energy could be properly harvested, it would be a significant source of energy for our society. **Nuclear energy** involves the controlled harvesting of energy from fission reactions. The reaction can be controlled because the fission of uranium-235 (and a few other isotopes, such as plutonium-239) can be artificially initiated by injecting a neutron into a uranium nucleus. The overall nuclear equation, with energy included as a product, is then as follows:

$$^{235}\text{U} + {}^{1}\text{n} \rightarrow {}^{139}\text{Ba} + {}^{94}\text{Kr} + 3{}^{1}\text{n}$$

Thus by the careful addition of extra neutrons into a sample of uranium, we can control the fission process and obtain energy that can be used for other purposes. (Artificial or induced radioactivity, in which neutrons are injected into a sample of matter that subsequently cause fission, was first demonstrated in 1934 by Irène Joliot-Curie and Frédéric Joliot, the daughter and son-in-law of Marie Curie.)

Plutonium-239 can absorb a neutron and undergo a fission reaction to make an atom of gold-204 and an atom of phosphorus-31. Write the balanced nuclear equation for the process and determine the number of neutrons given off as part of the reaction.

Solution

Using the data given, we can write the following initial equation:

$$^{1}_{0}\text{n} + {}^{239}_{94}\text{Pu} \rightarrow {}^{204}_{79}\text{Au} + {}^{31}_{15}\text{P} + ?{}^{1}_{0}\text{n}$$

In balanced nuclear equations, the sums of the subscripts on each side of the equation are the same, as are the sums of the superscripts. The subscripts are already balanced:

0 + 94 = 94 and 79 + 15 = 94.

The superscripts on the left equal 240 (1 + 239) but equal 235 (204 + 31) on the right. We need five more mass number units on the right. Five neutrons should be products of the process for the mass numbers to balance. (Because the atomic number of a neutron is zero, including five neutrons on the right does not change the overall sum of the subscripts.) Thus the balanced nuclear equation is as follows:

$$^{1}_{0}\text{n} + {}^{239}_{94}\text{Pu} \rightarrow {}^{204}_{79}\text{Au} + {}^{31}_{15}\text{P} + 5{}^{1}_{0}\text{n}$$

We predict that the overall process will give off five neutrons.

? Exercise 10.4.1

Uranium-238 can absorb a neutron and undergo a fission reaction to produce an atom of cesium-135 and an atom of rubidium-96. Write the balanced nuclear equation for the process and determine the number of neutrons given off as part of the reaction.

Answer

$$^{1}_{0}\text{n} + {}^{238}_{92}\text{U} \rightarrow {}^{96}_{37}\text{Rb} + {}^{135}_{55}\text{Cs} + 8{}^{1}_{0}\text{n}$$

eight neutrons

One balanced nuclear reaction for the fission of plutonium-239 is as follows:

$$\underset{1.0087}{^{1}_{0}\text{n}} + \underset{239.0522}{^{239}_{94}\text{Pu}} \rightarrow \underset{203.9777}{^{204}_{79}\text{Au}} + \underset{30.9738}{^{31}_{15}\text{P}} + \underset{5\times1.0087}{5{}^{1}_{0}\text{n}}$$

The molar mass in grams of each species is given for each particle. What is the energy change of this fission reaction?

Solution

We start by adding the masses of all species on each side of the nuclear equation. Then we determine the difference in mass as the reaction proceeds and convert this to an equivalent amount of energy. The total mass of the reactants is as follows:

1.0087 + 239.0522 = 240.0609 g

The total mass of the products is as follows:

203.9777 + 30.9738 + (5 × 1.0087) = 239.9950 g

The change is mass is determined by subtracting the mass of the reactants from the mass of the products:

change in mass = 239.9950 − 240.0609 = −0.0659 g

This mass change must be converted into kilogram units:

$$-0.0659g \times \frac{1kg}{1,000g} = -0.0000659kg$$

Now we can use Einstein's equation to determine the energy change of the nuclear reaction:

E = (−0.0000659 kg)(3.00 × 10^{8} m/s)^{2} = −5.93 × 10^{12} J

This is almost 6 *trillion* joules given off.

? Exercise 10.4.2

The nuclear equation for the fission of uranium-238 is as follows:

$$\underset{1.0087}{^{1}_{0}\text{n}} + \underset{238.0508}{^{238}_{92}\text{U}} \rightarrow \underset{95.9342}{^{96}_{37}\text{Rb}} + \underset{134.9060}{^{135}_{55}\text{Cs}} + \underset{8\times1.0087}{8{}^{1}_{0}\text{n}}$$

The molar mass in grams of each species is given for each particle. What is the energy change of this fission reaction?

Answer

-1.35×10^{13} J

A **nuclear reactor** is an apparatus designed to carefully control the progress of a nuclear reaction and extract the resulting energy for useful purposes. Figure **10.4.1** shows a simplified diagram of a nuclear reactor. The energy from the controlled nuclear reaction converts water into high-pressure steam, which is used to run turbines that generate electricity.

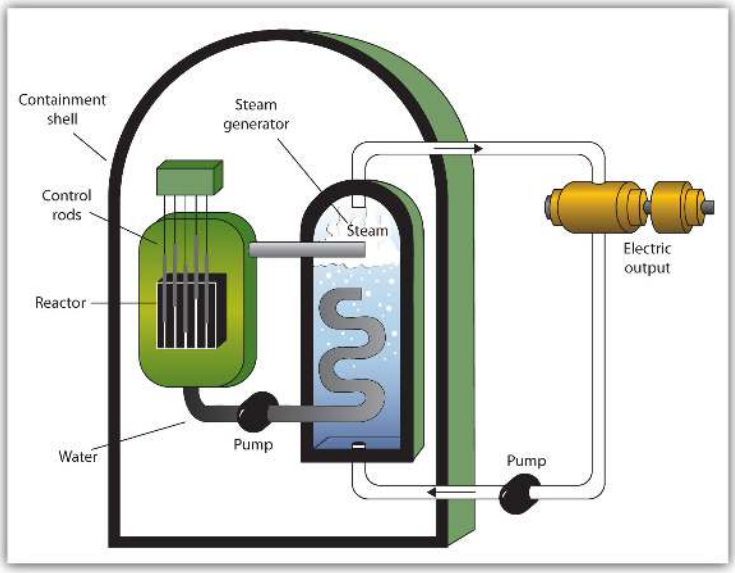


Figure **10.4.1**: A Diagram of a Nuclear Power Plant for Generating Electricity. The two main components of the power plant are the nuclear reactor itself and the steam-driven turbine and electricity generator.

Diagram of the nuclear power plant includes the containment shell, control rods, reactor, water, pump, steam generator, and electric output.

Although the fission of large nuclei can produce different products, on average the fission of uranium produces two more free neutrons than were present to begin with. These neutrons can themselves stimulate other uranium nuclei to undergo fission, releasing yet more energy and even more neutrons, which can in turn induce even more uranium fission. A single neutron can thus begin a process that grows exponentially in a phenomenon called a **chain reaction**.

$$1 \rightarrow 2 \rightarrow 4 \rightarrow 8 \rightarrow 16 \rightarrow 32 \rightarrow 64 \rightarrow 128 \rightarrow 256 \rightarrow 512 \rightarrow 1{,}024 \rightarrow 2{,}048 \rightarrow 4{,}096 \rightarrow 8{,}192 \rightarrow 16{,}384 \rightarrow \ldots$$

Because energy is produced with each fission event, energy is also produced exponentially and in an uncontrolled fashion. The quick production of energy creates an explosion. This is the mechanism behind the **atomic bomb**. (The first controlled chain reaction was achieved on December 2, 1942, in an experiment supervised by Enrico Fermi in a laboratory underneath the football stadium at the University of Chicago.)

Although fairly simple in theory, an atomic bomb is difficult to produce, in part because uranium-235, the isotope that undergoes fission, makes up only 0.7% of natural uranium; the rest is mostly uranium-238, which does not undergo fission. (Remember that the radioactive process that a nucleus undergoes is characteristic of the isotope.) To make uranium useful for nuclear reactors, the uranium in uranium-235 must be *enriched* to about 3%. The enrichment of uranium is a laborious and costly series of physical and chemical separations. To be useful in an atomic bomb, uranium must be enriched to 70% or more. At lesser concentrations, the chain reaction cannot sustain itself, so no explosion is produced.

Fusion is another nuclear process that can be used to produce energy. In this process, smaller nuclei are combined to make larger nuclei, with an accompanying release of energy. One example is hydrogen fusion, which makes helium:

$$4\,^1H \rightarrow\,^4 \text{He} + 2.58 \times 10^{12}\ J$$

Notice that the amount of energy given off per mole of reactant is only one-tenth of the amount given off by the fission of 1 mol of uranium-235. On a mass (per gram) basis, however, hydrogen fusion gives off 10 times more energy than fission does. In addition, the product of fusion is helium gas, not a wide range of isotopes (some of which are also radioactive) produced by fission.

Fusion occurs in nature: the sun and other stars use fusion as their ultimate energy source. Fusion is also the basis of very destructive weapons that have been developed by several countries around the world. However, one current goal is to develop a source of *controlled* fusion for use as an energy source. The practical problem is that to perform fusion, extremely high pressures and temperatures are necessary. Currently, the only known stable systems undergoing fusion are the interiors of stars. The conditions necessary for fusion can be created using an atomic bomb, but the resulting fusion is uncontrollable (and the basis for another type of bomb, a hydrogen bomb). Currently, researchers are looking for safe, controlled ways for producing useful energy using fusion.

Summary

- Nuclear energy comes from tiny mass changes in nuclei as radioactive processes occur.
- In fission, large nuclei break apart and release energy; in fusion, small nuclei merge together and release energy.



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