# 113 FUNDAMENTALS OF CHEMISTRY



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# 113 Fundamentals of Chemistry

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Detailed Licensing

# Licensing

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

# CHAPTER OVERVIEW

## 1: First Chapter


1.1: Prelude to Chemistry

1.2: Chemistry as a Science



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# 1.1: Prelude to Chemistry

If you are reading these words, you are likely starting a chemistry course. Get ready for a fantastic journey through a world of wonder, delight, and knowledge. One of the themes of this book is "chemistry is everywhere," and indeed it is; you would not be alive if it were not for chemistry, because your body is a big chemical machine.



Figure **1.1.1** © Thinkstock

6 images showing a volcanic eruption, large dinner spread, explosion, red wine, a pool of large fish, and molecules are laid out together.

If you do not believe it, do not worry. Every chapter in this book contains examples that will show you how chemistry is, in fact, everywhere. So enjoy the ride, and enjoy chemistry.

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# 1.2: 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 **1.2.1**)

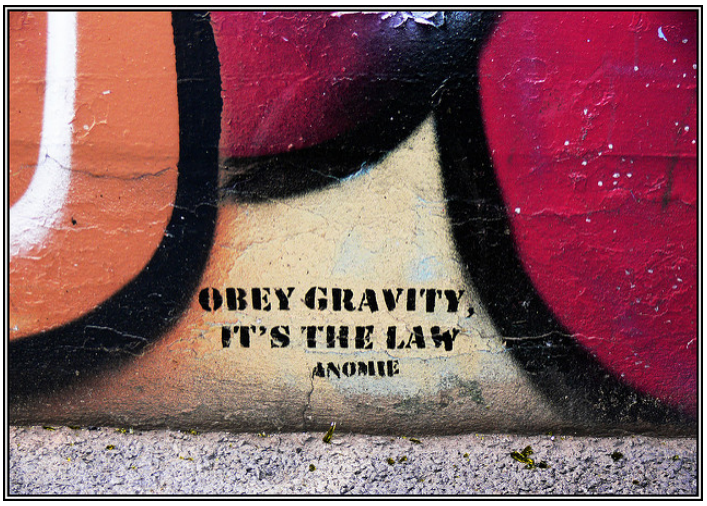



Figure **1.2.1**: Defining a law. Does this graffiti mean "law" the way science defines "law"? (CC BY-SA-NC-ND; Koppenbadger).

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 **1.2.2**) 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.

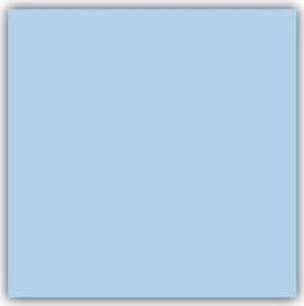


Figure 1.2.2: English Is Not Science. How would you describe this color? Blue or hard? Either way, you're not doing science.

### Example 1.2.1: 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 1.2.1

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 1.2.2: 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 1.2.2

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 **1.2.3**: 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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# CHAPTER OVERVIEW

## 2: Measurements and Calculations

In 1983, an Air Canada airplane had to make an emergency landing because it unexpectedly ran out of fuel; ground personnel had filled the fuel tanks with a certain number of pounds of fuel, not kilograms of fuel. In 1999, the Mars Climate Orbiter spacecraft was lost whilst attempting to orbit Mars because the thrusters were programmed in terms of English units, even though the engineers built the spacecraft using metric units. In 1993, a nurse mistakenly administered 23 units of morphine to a patient rather than the "2–3" units prescribed (the patient ultimately survived). These incidents occurred because people were not paying attention to quantities.

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


2.1: Prelude to Measurements

2.2: Expressing Numbers

2.3: Expressing Units

2.4: Significant Figures

2.5: Converting Units

2.6: Other Units - Temperature and Density

2.6.1: Temperature - Random Motion of Molecules and Atoms

2.E: Measurements (Exercises)


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# 2.1: Prelude to Measurements

Data suggest that a male child will weigh 50% of his adult weight at about 11 years of age. However, he will reach 50% of his adult height at only 2 years of age. It is obvious, then, that people eventually stop growing up but continue to grow out. Data also suggest that the average human height has been increasing over time. In industrialized countries, the average height of people increased 5.5 inches from 1810 to 1984. Most scientists attribute this simple, basic measurement of the human body to better health and nutrition.

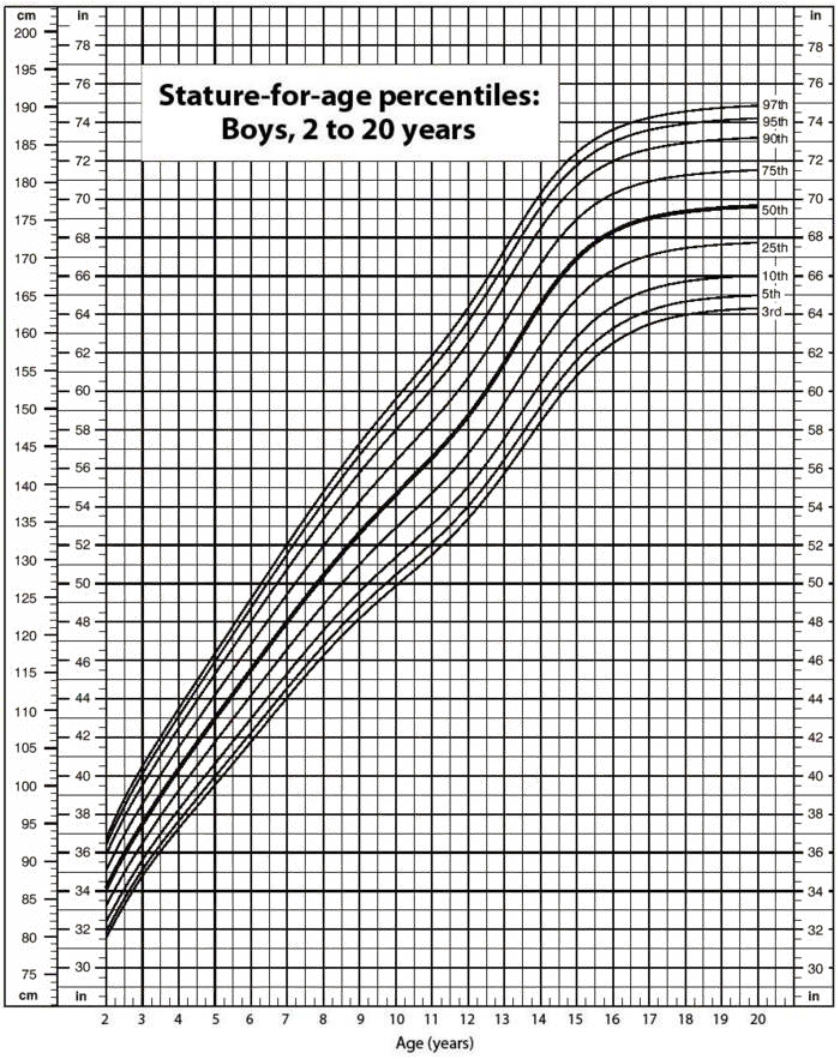



Figure **2.1.1**: Human body measurement. Source: Chart courtesy of Centers for Disease Control and Prevention(opens in new window).

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# 2.2: 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 $7.9345 \times 10^4$. 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 2.2.1).



Figure 2.2.1: "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 $2.9 \times 10^4$ *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 $9.3 \times 10^7$ *miles*. (Public Domain; NASA/Apollo 17 crew; taken by either Harrison Schmitt or Ron Evans.)

### Example 2.2.1: 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 $3.06 \times 10^5$.
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 $8.84 \times 10^{-3}$.
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 $2.76 \times 10^6$. 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 $5.59 \times 10^{-7}$.

### Exercise 2.2.1

Express these numbers in scientific notation.

a. 23,070
b. 0.0009706

**Answer a**

$2.307 \times 10^4$

**Answer b**

$9.706 \times 10^{-4}$

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 under 56,900; 1 2 3 4 5 under 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 **2.2.2**).

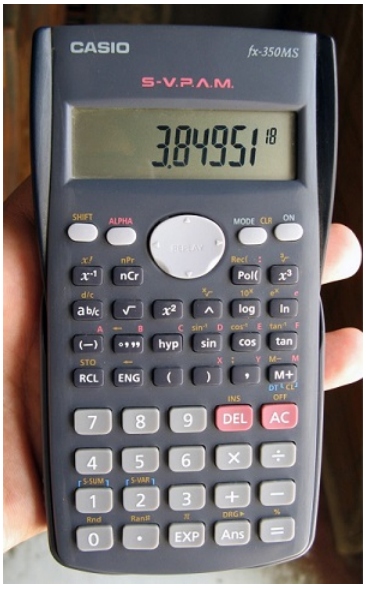



Figure **2.2.2**: This calculator shows only the coefficient and the power of 10 to represent the number in scientific notation. Thus, the number being displayed is $\mathbf{3.84951 \times 10^{18}}$, or $\mathbf{3,849,510,000,000,000,000}$. Source: "Casio" Asim Bijarani is licensed under Creative Commons 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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## 2.3: Expressing Units

**Learning Objective**

- Learn the units that go with various quantities
- Express units using their abbreviations
- Make new units by combining numerical prefixes with units

A number indicates "how much," but the unit indicates "of what." The "of what" is important when communicating a quantity. For example, if you were to ask a friend how close you are to Lake Erie and your friend says "six," then your friend isn't giving you complete information. Six *what*? Six miles? Six inches? Six city blocks? The actual distance to the lake depends on what units you use.

Chemistry, like most sciences, uses the International System of Units, or SI for short. (The letters *SI* stand for the French "le Système International d'unités.") SI specifies certain units for various types of quantities, based on seven fundamental units. We will use most of the fundamental units in chemistry. Initially, we will deal with three fundamental units. The **meter** (m) is the SI unit of length. It is a little longer than a yard (Figure **2.3.1**). The SI unit of mass is the **kilogram** (kg), which is about 2.2 pounds (lb). The SI unit of time is the **second** (s).

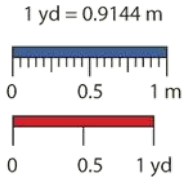



Figure **2.3.1**: The Meter. The SI standard unit of length, the meter, is a little longer than a yard.

To express a quantity, you need to combine a number with a unit. If you have a length that is 2.4 m, then you express that length as simply 2.4 m. A time of 15,000 s can be expressed as $1.5 \times 10^4$ s in scientific notation.

Sometimes, a given unit is not an appropriate size to easily express a quantity. For example, the width of a human hair is very small, and it doesn't make much sense to express it in meters. SI also defines a series of *numerical prefixes*, referring to multiples or fractions of a fundamental unit, to make a unit more conveniently sized for a specific quantity. Table **2.3.1** lists the prefixes, their abbreviations, and their multiplicative factors. Some of the prefixes, such as kilo-, mega-, and giga-, represent more than one of the fundamental unit, while other prefixes, such as centi-, milli-, and micro-, represent fractions of the original unit. Note, too, that once again we are using powers of 10. Each prefix is a multiple of or fraction of a power of 10.

Table **2.3.1**: Multiplicative Prefixes for SI Units

| Prefix | Abbreviation | Multiplicative Amount | Power of Ten |
|---|---|---|---|
| giga- | G | 1,000,000,000 × | $10^9$ × |
| mega- | M | 1,000,000 × | $10^6$ × |
| kilo- | k | 1,000 × | $10^3$ × |
| deci- | d | 1/10 × | × |
| centi- | c | 1/100 × | × |
| milli- | m | 1/1,000 × | × |
| micro- | μ* | 1/1,000,000 × | × |
| nano- | n | 1/1,000,000,000 × | × |
| pico- | p | 1/1,000,000,000,000 × | x |

* The letter $\mu$ is the Greek letter lowercase equivalent to an m and is called "mu" (pronounced "myoo").

To use the fractions to generate new units, simply combine the prefix with the unit itself; the abbreviation for the new unit is the combination of the abbreviation for the prefix and the abbreviation of the unit. For example, the kilometer (km) is 1,000 × meter, or 1,000 m. Thus, 5 kilometers (5 km) is equal to 5,000 m. Similarly, a millisecond (ms) is 1/1,000 × second, or one-thousandth of a second. Thus, 25 ms is 25 thousandths of a second. You will need to become proficient in combining prefixes and units. (You may recognize that one of our fundamental units, the kilogram, automatically has a prefix-unit combination. The word *kilogram* means 1,000 g.)

In addition to the fundamental units, SI also allows for derived units based on a fundamental unit or units. There are many derived units used in science. For example, the derived unit for area comes from the idea that area is defined as width times height. Because both width and height are lengths, they both have the fundamental unit of meter, so the unit of area is meter × meter, or $\text{meter}^2$ ($\text{m}^2$). This is sometimes spoken as "square meters." A unit with a prefix can also be used to derive a unit for area, so we can also have $\text{cm}^2$, $\text{mm}^2$, or $\text{km}^2$ as acceptable units for area.

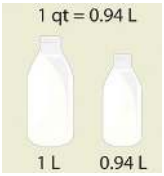



Figure **2.3.2**: The Liter. The SI unit of volume, the liter, is slightly larger than 1 quart.

Volume is defined as length times width times height, so it has units of meter × meter × meter, or $\text{meter}^3$ ($\text{m}^3$)—sometimes spoken as "cubic meters." The cubic meter is a rather large unit, however, so another unit is defined that is somewhat more manageable: the liter (L). A liter is 1/1,000th of a cubic meter and is a little more than 1 quart in volume (Figure **2.3.2**). Prefixes can also be used with the liter unit, so we can speak of milliliters (1/1,000th of a liter; mL) and kiloliters (1,000 L; kL).

Another definition of a liter is one-tenth of a meter cubed. Because one-tenth of a meter is 10 cm, then a liter is equal to 1,000 $\text{cm}^3$ (Figure **2.3.3**). Because 1 L equals 1,000 mL, we conclude that 1 mL equals 1 $\text{cm}^3$; thus, these units are interchangeable.

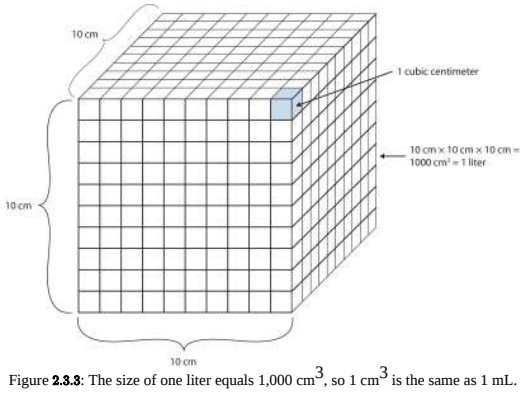



Figure **2.3.3**: The size of one liter equals 1,000 $cm^3$, so 1 $cm^3$ is the same as 1 mL.

Units are not only multiplied together—they can also be divided. For example, if you are traveling at one meter for every second of time elapsed, your velocity is 1 meter per second, or 1 m/s. The word *per* implies division, so velocity is determined by dividing a distance quantity by a time quantity. Other units for velocity include kilometers per hour (km/h) or even micrometers per nanosecond (μm/ns). Later, we will see other derived units that can be expressed as fractions.

**Example 2.3.1**

a. A human hair has a diameter of about $6.0 \times 10^{-5}$ m. Suggest an appropriate unit for this measurement and write the diameter of a human hair in terms of that unit.
b. What is the velocity of a car if it goes 25 m in 5.0 s?

**Solution**

a. The scientific notation $10^{-5}$ is close to $10^{-6}$, which defines the micro- prefix. Let us use micrometers as the unit for hair diameter. The number $6.0 \times 10^{-5}$ can be written as $60 \times 10^{-6}$, and a micrometer is $10^{-6}$ m, so the diameter of a human hair is about 60 μm.
b. If velocity is defined as a distance quantity divided by a time quantity, then velocity is 25 meters/5.0 seconds. Dividing the numbers gives us 25/5.0 = 5.0, and dividing the units gives us meters/second, or m/s. The velocity is 5.0 m/s.

**Exercise 2.3.1**

a. Express the volume of an Olympic-sized swimming pool, 2,500,000 L, in more appropriate units.
b. A common garden snail moves about 6.1 m in 30 min. What is its velocity in meters per minute (m/min)?

**Answer a**

- 2.5 ML

**Answer b**

- 0.203 m/min

## Key Takeaways

- Numbers tell "how much," and units tell "of what."
- Chemistry uses a set of fundamental units and derived units from SI units.
- Chemistry uses a set of prefixes that represent multiples or fractions of units.
- Units can be multiplied and divided to generate new units for quantities.

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# 2.4: Significant Figures

**Learning Objective**

- Apply the concept of significant figures to limit a measurement to the proper number of digits.
- Recognize the number of significant figures in a given quantity.
- Limit mathematical results to the proper number of significant figures.

If you use a calculator to evaluate the expression 337/217, you will get the following:

$$337 \div 217 = 1.5529953917$$

and so on for many more digits. Although this answer is correct, it is somewhat presumptuous. You start with two values that each have three digits, and the answer has *twelve* digits? That does not make much sense from a strict numerical point of view.

Consider using a ruler to measure the width of an object, as shown in Figure **2.4.1**. The object is definitely more than 1 cm long, so we know that the first digit in our measurement is 1. We see by counting the tick marks on the ruler that the object is at least three ticks after the 1. If each tick represents 0.1 cm, then we know the object is at least 1.3 cm wide. But our ruler does not have any more ticks between the 0.3 and the 0.4 marks, so we can't know exactly how much the next decimal place is. But with a practiced eye we can estimate it. Let us estimate it as about six-tenths of the way between the third and fourth tick marks, which estimates our hundredths place as 6, so we identify a measurement of 1.36 cm for the width of the object.

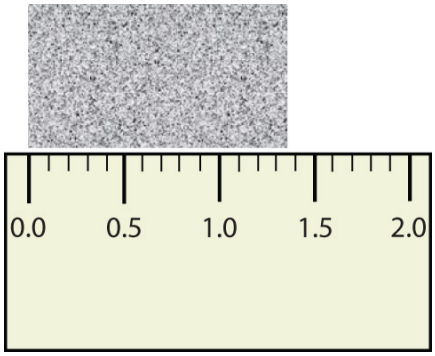


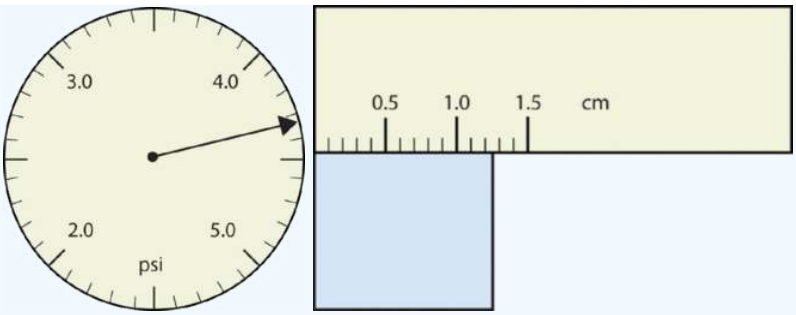

Figure **2.4.1**: Expressing Width. What is the proper way to express the width of this object?

Does it make any sense to try to report a thousandths place for the measurement? No, it doesn't; we are not exactly sure of the hundredths place (after all, it was an estimate only), so it would be fruitless to estimate a thousandths place. Our best measurement, then, stops at the hundredths place, and we report 1.36 cm as proper measurement.

This concept of reporting the proper number of digits in a measurement or a calculation is called significant figures. Significant figures (sometimes called significant digits) represent the limits of what values of a measurement or a calculation we are sure of. The convention for a measurement is that the quantity reported should be all known values and the first estimated value. The conventions for calculations are discussed as follows.

**Example 2.4.1**

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

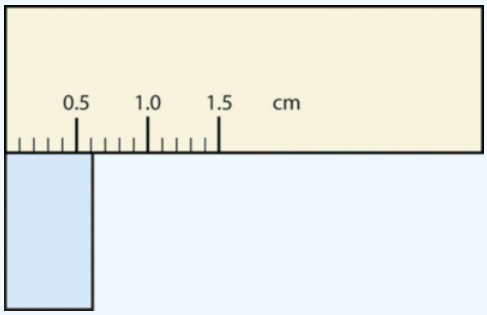



**Solution**

1. 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. Combining the digits, we have a measurement of 4.33 psi (psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire). We say that the measurement is reported to three significant figures.
2. 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. Again, the measurement is reported to three significant figures.

**? Exercise 2.4.1**

What would be the reported width of this rectangle?




**Answer**

0.63 cm

In many cases, you will be given a measurement. How can you tell by looking what digits are significant? For example, the reported population of the United States is 306,000,000. Does that mean that it is *exactly* three hundred six million, or is some estimation occurring?

The following conventions dictate which numbers in a reported measurement are significant and which are not significant:

1. Any nonzero digit is significant.
2. Any zeros between nonzero digits (i.e., embedded zeros) are significant.
3. Zeros at the end of a number without a decimal point (i.e., trailing zeros) are not significant; they serve only to put the significant digits in the correct positions. However, zeros at the end of any number with a decimal point are significant.
4. Zeros at the beginning of a decimal number (i.e., leading zeros) are not significant; again, they serve only to put the significant digits in the correct positions.

So, by these rules, the population figure of the United States has only three significant figures: the 3, the 6, and the zero between them. The remaining six zeros simply put the 306 in the millions position. (See Figure **2.4.2** for another example.)




Figure **2.4.2**: Significant Figures. How many of the numbers in this display are actually significant? (Public Domain.)

### Example 2.4.2

Give the number of significant figures in each measurement.

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

**Solution**

a. By rule 1, all nonzero digits are significant, so this measurement has three significant figures.
b. By rule 4, the first three zeros are not significant, but by rule 2 the zero between the sixes is; therefore, this number has four significant figures.
c. By rule 2, the two zeros between the twos are significant, so this measurement has four significant figures.
d. The four trailing zeros in the number are not significant, but the other five numbers are, so this number has five significant figures.

### Exercise 2.4.2

Give the number of significant figures in each measurement.

a. 0.000601 m
b. 65.080 kg

**Answer a**

three significant figures

**Answer b**

five significant figures

How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.21, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.

$$\begin{array}{l} 1.2 \\ \underline{4.41} \\ 5.61 \\ \uparrow \text{ limit final answer to the tenths column: } 5.6 \end{array}$$

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

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

$$\begin{array}{l} 77.2 \\ \underline{10.46} \\ 87.66 \\ \uparrow \text{ limit final answer to the tenths column and round up: } 87.7 \end{array}$$

**Example 2.4.3**

Express the final answer to the proper number of significant figures.

a. 101.2 + 18.702 = ?
b. 202.88 − 1.013 = ?

**Solution**

a. If we use a calculator to add these two numbers, we would get 119.902. However, most calculators do not understand significant figures, and we need to limit the final answer to the tenths place. Thus, we drop the 02 and report a final answer of 119.9 (rounding down).
b. A calculator would answer 201.867. However, we have to limit our final answer to the hundredths place. Because the first number being dropped is 7, which is greater than 5, we round up and report a final answer of 201.87.

**Exercise 2.4.3**

Express the answer for

$$3.445 + 90.83 - 72.4$$

to the proper number of significant figures.

**Answer**

21.9

If the operations being performed are multiplication or division, the rule is as follows: limit the answer to the number of significant figures that the data value with the *least* number of significant figures has. So if we are dividing 23 by 448, which have two and three significant figures respectively, we should limit the final reported answer to two significant figures (the lesser of two and three significant figures):

$$23 \div 448 = 0.051339286 \approx 0.051$$

The same rounding rules apply in multiplication and division as they do in addition and subtraction.

**Example 2.4.4: Significant Figures**

Express the final answer to the proper number of significant figures.

a. 76.4 × 180.4 = ?

b. 934.9 ÷ 0.00455 = ?

**Solution**

a. The first number has three significant figures, while the second number has four significant figures. Therefore, we limit our final answer to three significant figures: 76.4 × 180.4 = 13,782.56 = 13,800.

b. The first number has four significant figures, while the second number has three significant figures. Therefore, we limit our final answer to three significant figures: 934.9 ÷ 0.00455 = 205,472.5275… = 205,000.

**? Exercise 2.4.4**

Express the final answer to the proper number of significant figures.

a. 22.4 × 8.314 = ?

b. 1.381 ÷ 6.02 = ?

**Answer a**

186

**Answer b**

0.229

As you have probably realized by now, the biggest issue in determining the number of significant figures in a value is the zero. Is the zero significant or not? One way to unambiguously determine whether a zero is significant or not is to write a number in scientific notation. Scientific notation will include zeros in the coefficient of the number *only if they are significant*. Thus, the number $8.666 \times 10^6$ has four significant figures. However, the number $8.6660 \times 10^6$ has five significant figures. That last zero is significant; if it were not, it would not be written in the coefficient. So, when in doubt about expressing the number of significant figures in a quantity, use scientific notation and include the number of zeros that are truly significant.

## Summary

- Significant figures in a quantity indicate the number of known values plus one place that is estimated.
- There are rules for which numbers in a quantity are significant and which are not significant.
- In calculations involving addition and subtraction, limit significant figures based on the rightmost place that all values have in common.
- In calculations involving multiplication and division, limit significant figures to the least number of significant figures in all the data values.

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# 2.5: Converting Units

**Learning Objective**

- Convert from one unit to another unit of the same type.

In Section 2.2, we showed some examples of how to replace initial units with other units of the same type to get a numerical value that is easier to comprehend. In this section, we will formalize the process.

Consider a simple example: how many feet are there in 4 yards? Most people will almost automatically answer that there are 12 feet in 4 yards. How did you make this determination? Well, if there are 3 feet in 1 yard and there are 4 yards, then there are 4 × 3 = 12 feet in 4 yards.

This is correct, of course, but it is informal. Let us formalize it in a way that can be applied more generally. We know that 1 yard (yd) equals 3 feet (ft):

$$1\,yd = 3\,ft$$

In math, this expression is called an *equality*. The rules of algebra say that you can change (i.e., multiply or divide or add or subtract) the equality (as long as you do not divide by zero) and the new expression will still be an equality. For example, if we divide both sides by 2, we get:

$$\frac{1}{2}\,yd = \frac{3}{2}\,ft$$

We see that one-half of a yard equals 3/2, or one and a half, feet—something we also know to be true—so the above equation is still an equality. Going back to the original equality, suppose we divide both sides of the equation by 1 yard (number **and** unit):

$$\frac{1\,yd}{1\,yd} = \frac{3\,ft}{1\,yd}$$

The expression is still an equality, by the rules of algebra. The left fraction equals 1. It has the same quantity in the numerator and the denominator, so it must equal 1. The quantities in the numerator and denominator cancel, both the number *and* the unit:

$$\frac{1\,yd}{1\,yd} = \frac{3\,ft}{1\,yd}$$

When everything cancels in a fraction, the fraction reduces to 1:

$$1 = \frac{3\,ft}{1\,yd}$$

## Conversion Factors

We have an expression that equals 1.

$$\frac{3\,ft}{1\,yd} = 1$$

This is a strange way to write 1, but it makes sense: 3 ft equal 1 yd, so the quantities in the numerator and denominator are the same quantity, just expressed with different units.

The expression

$$\frac{3\,ft}{1\,yd}$$

is called a conversion factor and it is used to formally change the unit of a quantity into another unit. (The process of converting units in such a formal fashion is sometimes called *dimensional analysis* or the *factor label method*.)

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

Now let us multiply this quantity by 1. When you multiply anything by 1, you do not change the value of the quantity. Rather than multiplying by just 1, let us write 1 as:

$$\frac{3\,ft}{1\,yd}$$

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

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

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

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

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

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

How many millimeters are in 14.66 m? To answer this, we need to construct a conversion factor between millimeters and meters and apply it correctly to the original quantity. We start with the definition of a millimeter, which is:

$$1\,mm = \frac{1}{1000\,m}$$

The 1/1000 is what the prefix *milli-* means. Most people are more comfortable working without fractions, so we will rewrite this equation by bringing the 1,000 into the numerator of the other side of the equation:

$$1000\,mm = 1\,m$$

Now we construct a conversion factor by dividing one quantity into both sides. But now a question arises: which quantity do we divide by? It turns out that we have two choices, and the two choices will give us different conversion factors, both of which equal 1:

$$\frac{1000\,mm}{1000\,mm} = \frac{1\,m}{1000\,mm}$$

or

$$\frac{1000\,mm}{1\,m} = \frac{1\,m}{1\,m}$$

$$1 = \frac{1\,m}{1000\,mm}$$

or

$$\frac{1000\,mm}{1\,m} = 1$$

Which conversion factor do we use? The answer is based on *what unit you want to get rid of in your initial quantity*. The original unit of our quantity is meters, which we want to convert to millimeters. Because the original unit is assumed to be in the numerator, to get rid of it, we want the meter unit in the *denominator*; then they will cancel. Therefore, we will use the second conversion factor. Canceling units and performing the mathematics, we get:

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

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

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

### Example 2.5.1

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

**Solution**

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

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

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

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

Applying this conversion factor, we get:

$$555\,nm \times \frac{1m}{10^9 nm} = 0.000000555\,m = 5.55 \times 10^{-7}\,m$$

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

### Exercise 2.5.1

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

**Answer a**

$6.708 \times 10^{-5}$ L

**Answer b**

$5.68 \times 10^{-2}$ km

What if we have a derived unit that is the product of more than one unit, such as $m^2$? Suppose we want to convert square meters to square centimeters? The key is to remember that $m^2$ means m × m, which means we have *two* meter units in our derived unit. That means we have to include *two* conversion factors, one for each unit. For example, to convert 17.6 $m^2$ to square centimeters, we perform the conversion as follows:

$$\begin{aligned} 17.6m^2 &= 17.6(m \times m) \times \frac{100cm}{1m} \times \frac{100cm}{1m} \\ &= 176000\,cm \times cm \\ &= 1.76 \times 10^5\,cm^2 \end{aligned} \tag{2.5.1}$$

### Example 2.5.2

How many cubic centimeters are in 0.883 $m^3$?

**Solution**

With an exponent of 3, we have three length units, so by extension we need to use three conversion factors between meters and centimeters. Thus, we have:

$$0.883 m^3 \times \frac{100\,cm}{1\,m} \times \frac{100\,cm}{1\,m} \times \frac{100\,cm}{1\,m} = 883000\,cm^3 = 8.83 \times 10^5\,cm^3$$

You should demonstrate to yourself that the three meter units do indeed cancel.

**? Exercise 2.5.2**

How many cubic millimeters are present in 0.0923 $m^3$?

**Answer**

$9.23 \times 10^7$ $mm^3$

Suppose the unit you want to convert is in the denominator of a derived unit—what then? Then, in the conversion factor, the unit you want to remove must be in the *numerator*. This will cancel with the original unit in the denominator and introduce a new unit in the denominator. The following example illustrates this situation.

**✓ Example 2.5.3**

Convert 88.4 m/min to meters/second.

**Solution**

We want to change the unit in the denominator from minutes to seconds. Because there are 60 seconds in 1 minute (60 s = 1 min), we construct a conversion factor so that the unit we want to remove, minutes, is in the numerator: 1min/60s. Apply and perform the math:

$$\frac{88.4 m}{min} \times \frac{1\,min}{60\,s} = 1.47 \frac{m}{s}$$

Notice how the 88.4 automatically goes in the numerator. That's because any number can be thought of as being in the numerator of a fraction divided by 1.

**? Exercise 2.5.3**

Convert 0.203 m/min to meters/second.

**Answer**

0.00338 m/s

or

$3.38 \times 10^{-3}$ m/s

Sometimes there will be a need to convert from one unit with one numerical prefix to another unit with a different numerical prefix. How do we handle those conversions? Well, you could memorize the conversion factors that interrelate all numerical prefixes. Or you can go the easier route: first convert the quantity to the base unit—the unit with no numerical prefix—using the definition of the original prefix. Then, convert the quantity in the base unit to the desired unit using the definition of the second prefix. You can do the conversion in two separate steps or as one long algebraic step. For example, to convert 2.77 kg to milligrams:

$$2.77\,kg \times \frac{1000\,g}{1\,kg} = 2770\,g$$

(convert to the base units of grams)

$$2770\,g \times \frac{1000\,mg}{1\,g} = 2770000\,mg = 2.77 \times 10^6\,mg$$

(convert to desired unit)

Alternatively, it can be done in a single multi-step process:

$$2.77 \cancel{kg} \times \frac{1000 \cancel{g}}{1 \cancel{kg}} \times \frac{1000\,mg}{1 \cancel{g}} = 2770000\,mg$$

$$= 2.77 \times 10^6\,mg \tag{2.5.2}$$

You get the same answer either way.

**Example 2.5.4**

How many nanoseconds are in 368.09 µs?

**Solution**

You can either do this as a one-step conversion from microseconds to nanoseconds or convert to the base unit first and then to the final desired unit. We will use the second method here, showing the two steps in a single line. Using the definitions of the prefixes *micro-* and *nano-*,

$$368.0\,\mu s \times \frac{1\,s}{1000000\,\mu s} \times \frac{1000000000}{1\,s} = 3.6809 \times 10^5\,ns$$

**Exercise 2.5.4**

How many milliliters are in 607.8 kL?

**Answer**

$6.078 \times 10^8$ mL

When considering the significant figures of a final numerical answer in a conversion, there is one important case where a number does not impact the number of significant figures in a final answer: the so-called **exact number**. An exact number is a number from a defined relationship, not a measured one. For example, the prefix *kilo-* means 1,000-*exactly* 1,000, no more or no less. Thus, in constructing the conversion factor:

$$\frac{1000\,g}{1\,kg}$$

neither the 1,000 nor the 1 enter into our consideration of significant figures. The numbers in the numerator and denominator are defined exactly by what the prefix *kilo-* means. Another way of thinking about it is that these numbers can be thought of as having an infinite number of significant figures, such as:

$$\frac{1000.0000000000\ldots\,g}{1.0000000000\ldots\,kg}$$

The other numbers in the calculation will determine the number of significant figures in the final answer.

**Example 2.5.5**

A rectangular plot in a garden has the dimensions 36.7 cm by 128.8 cm. What is the area of the garden plot in square meters? Express your answer in the proper number of significant figures.

**Solution**

Area is defined as the product of the two dimensions, which we then have to convert to square meters, and express our final answer to the correct number of significant figures—which in this case will be three.

$$36.7\,cm \times 128.8\,cm \times \frac{1\,m}{100\,cm} \times \frac{1\,m}{100\,cm} = 0.472696\,m^2 = 0.473\,m^2$$

The 1 and 100 in the conversion factors do not affect the determination of significant figures because they are exact numbers, defined by the centi- prefix.

**? Exercise 2.5.5**

What is the volume of a block in cubic meters with the dimensions 2.1 cm × 34.0 cm × 118 cm?

**Answer**

$0.0084\ m^3$

**Chemistry is Everywhere: The Gimli Glider**

On July 23, 1983, an Air Canada Boeing 767 jet had to glide to an emergency landing at Gimli Industrial Park Airport in Gimli, Manitoba, because it unexpectedly ran out of fuel during flight. There was no loss of life in the course of the emergency landing, only some minor injuries associated in part with the evacuation of the craft after landing. For the remainder of its operational life (the plane was retired in 2008), the aircraft was nicknamed "the Gimli Glider."




***The Gimli Glider is the Boeing 767 that ran out of fuel and glided to safety at Gimli Airport. The aircraft ran out of fuel because of confusion over the units used to express the amount of fuel. Source: Photo courtesy of Will F., (CC BY-SA 2.5; Aero Icarus).***

The 767 took off from Montreal on its way to Ottawa, ultimately heading for Edmonton, Canada. About halfway through the flight, all the engines on the plane began to shut down because of a lack of fuel. When the final engine cut off, all electricity (which was generated by the engines) was lost; the plane became, essentially, a powerless glider. Captain Robert Pearson was an experienced glider pilot, although he had never flown a glider the size of a 767. First Officer Maurice Quintal quickly determined that the aircraft would not be able make it to Winnipeg, the next large airport. He suggested his old Royal Air Force base at Gimli Station, one of whose runways was still being used as a community airport. Between the efforts of the pilots and the flight crew, they managed to get the airplane safely on the ground (although with buckled landing gear) and all passengers off safely.

What happened? At the time, Canada was transitioning from the older English system to the metric system. The Boeing 767s were the first aircraft whose gauges were calibrated in the metric system of units (liters and kilograms) rather than the English system of units (gallons and pounds). Thus, when the fuel gauge read 22,300, the gauge meant kilograms, but the ground crew mistakenly fueled the plane with 22,300 *pounds* of fuel. This ended up being just less than half of the fuel needed to make the trip, causing the engines to quit about halfway to Ottawa. Quick thinking and extraordinary skill saved the lives of 61 passengers and 8 crew members—an incident that would not have occurred if people were watching their units.

## Key Takeaways

- Units can be converted to other units using the proper conversion factors.
- Conversion factors are constructed from equalities that relate two different units.
- Conversions can be a single step or multi-step.

- Unit conversion is a powerful mathematical technique in chemistry that must be mastered.
- Exact numbers do not affect the determination of significant figures.

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# 2.6: Other Units - 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:

$$^{\circ}\text{C} = (^{\circ}\text{F-32}) \times \tfrac{5}{9} \tag{2.6.1}$$

$$^{\circ}\text{F} = (^{\circ}\text{C} \times \tfrac{9}{5}) + 32 \tag{2.6.2}$$

**Example 2.6.1: Conversions**

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

**Solution**

a. Using Equation 2.6.1, we have

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

b. Using Equation 2.6.2, 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 2.6.1**

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 2.6.2: 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 2.6.2

What is 98.6°F on the Kelvin scale?

**Answer**

310.2 K

Figure **2.6.1** 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.

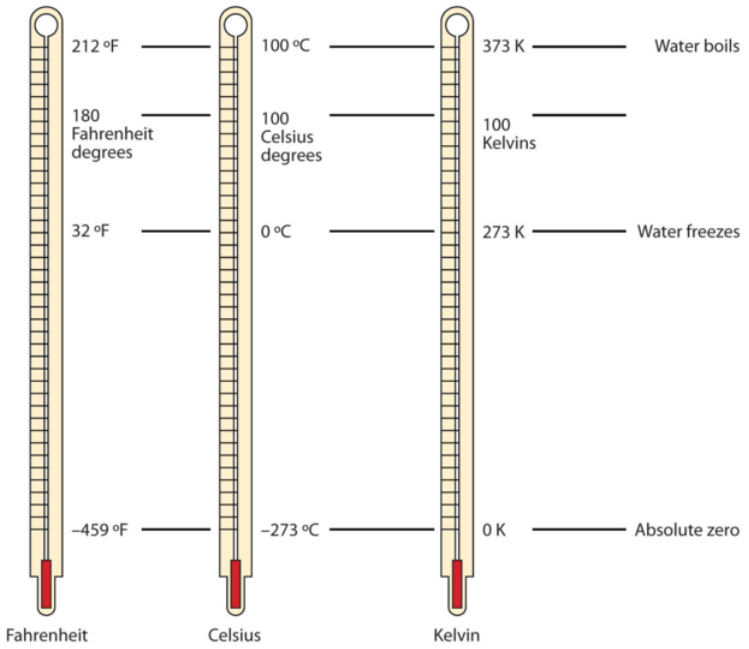



Figure 2.6.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales.

## 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 2.6.1.

Table 2.6.1: 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 2.6.1), 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 2.6.3: Mercury

What is the mass of 44.6 mL of mercury?

**Solution**

Use the density from Table **2.6.1** "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 2.6.3

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 2.6.4: 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 2.6.4**

What is the volume of 3.78 g of gold?

**Answer**

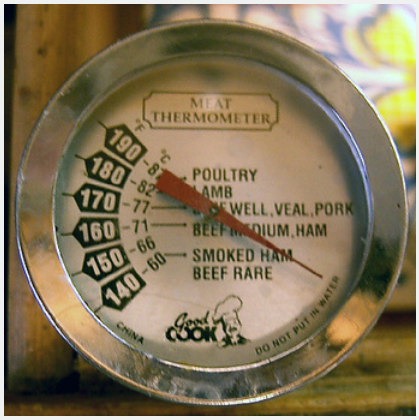
0.196 $cm^3$

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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# 2.6.1: Temperature - Random Motion of Molecules and Atoms

**Learning Objectives**

- Identify the different between temperature and heat.
- Recognize the different scales used to measure temperature

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measure of how hot or cold an object is relative to another object (its thermal energy content), whereas **heat** is the flow of thermal energy between objects with different temperatures. Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.



Figure **2.6.1.1**: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.



Figure **2.6.1.2**: Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right).

## The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist, Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as $32°F$ and the boiling point as $212°F$. The Fahrenheit scale is typically not used for scientific purposes.

## The Celsius Scale

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

## The Kelvin Scale

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

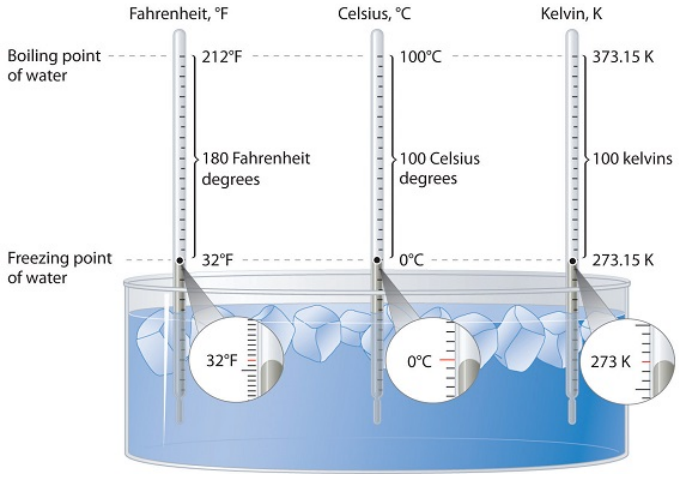



Figure **2.6.1.1**: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

## Converting Between Scales

The Kelvin is the same size as the Celsius degree, so measurements are easily converted from one to the other. The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K. The Kelvin and Celsius scales are related as follows:

$$T\,(\text{in °C}) + 273.15 = T\,(\text{in K}) \tag{3.10.1}$$

$$T\ \ (\text{in K}) - 273.15 = T\ (\text{in °C}) \tag{3.10.2}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows: where the coefficient for degrees Fahrenheit is exact. (Some calculators have a function that allows you to convert directly between °F and °C.) There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F. The relationship between the scales is as follows:

$$°C = \frac{(°F - 32)}{1.8} \tag{3.10.3}$$

$$°F = 1.8 \times (°C) + 32 \tag{3.10.4}$$

### Example 2.6.1.1: Temperature Conversions

A student is ill with a temperature of 103.5°F. What is her temperature in °C and K?

#### Solution

Converting from Fahrenheit to Celsius requires the use of Equation 3.10.3:

$$°C = \frac{(103.5°F - 32)}{1.8} \tag{2.6.1.1}$$

$$= 39.7\,°C \tag{2.6.1.2}$$

Converting from Celsius to Kelvin requires the use of Equation 3.10.1:

$$K = 39.7\,°C + 273.15 \quad (2.6.1.3)$$
$$= 312.9\,K \quad (2.6.1.4)$$

**Exercise 2.6.1.1**

Convert each temperature to °C and °F.

a. the temperature of the surface of the sun (5800 K)
b. the boiling point of gold (3080 K)
c. the boiling point of liquid nitrogen (77.36 K)

**Answer (a)**

5527 K, 9980 °F

**Answer (b)**

2807 K, 5084 °F

**Answer (c)**

-195.79 K, -320.42 °F

## Summary

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).

## Contributions & Attributions

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# 2.E: Measurements (Exercises)

## Exercises (Expressing Numbers)

1. Express these numbers in scientific notation.

    a. 56.9
    b. 563,100
    c. 0.0804
    d. 0.00000667

2. Express these numbers in scientific notation.

    a. −890,000
    b. 602,000,000,000
    c. 0.0000004099
    d. 0.000000000000011

3. Express these numbers in scientific notation.

    a. 0.00656
    b. 65,600
    c. 4,567,000
    d. 0.000005507

4. Express these numbers in scientific notation.

    a. 65
    b. −321.09
    c. 0.000077099
    d. 0.000000000218

5. Express these numbers in standard notation.

    a. $1.381 \times 10^{5}$
    b. $5.22 \times 10^{-7}$
    c. $9.998 \times 10^{4}$

6. Express these numbers in standard notation.

    a. $7.11 \times 10^{-2}$
    b. $9.18 \times 10^{2}$
    c. $3.09 \times 10^{-10}$

7. Express these numbers in standard notation.

    a. $8.09 \times 10^{0}$
    b. $3.088 \times 10^{-5}$
    c. $-4.239 \times 10^{2}$

8. Express these numbers in standard notation.

    a. $2.87 \times 10^{-8}$
    b. $1.78 \times 10^{11}$
    c. $1.381 \times 10^{-23}$

9. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.

    a. $72.44 \times 10^{3}$
    b. $9{,}943 \times 10^{-5}$
    c. $588{,}399 \times 10^{2}$

10. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.

    a. $0.000077 \times 10^{-7}$
    b. $0.000111 \times 10^{8}$
    c. $602{,}000 \times 10^{18}$

11. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.

    a. $345.1 \times 10^{2}$
    b. $0.234 \times 10^{-3}$
    c. $1{,}800 \times 10^{-2}$

12. These numbers are not written in proper scientific notation. Rewrite them so that they are in proper scientific notation.

    a. $8{,}099 \times 10^{-8}$
    b. $34.5 \times 10^{0}$
    c. $0.000332 \times 10^{4}$

13. Write these numbers in scientific notation by counting the number of places the decimal point is moved.

    a. 123,456.78
    b. 98,490
    c. 0.000000445

14. Write these numbers in scientific notation by counting the number of places the decimal point is moved.

    a. 0.000552
    b. 1,987
    c. 0.00000000887

15. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.

    a. $456 \times (7.4 \times 10^{8}) = ?$
    b. $(3.02 \times 10^{5}) \div (9.04 \times 10^{15}) = ?$
    c. $0.0044 \times 0.000833 = ?$

16. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.

    a. $98{,}000 \times 23{,}000 = ?$
    b. $98{,}000 \div 23{,}000 = ?$
    c. $(4.6 \times 10^{-5}) \times (2.09 \times 10^{3}) = ?$

17. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.

    a. $45 \times 132 \div 882 = ?$
    b. $[(6.37 \times 10^{4}) \times (8.44 \times 10^{-4})] \div (3.2209 \times 10^{15}) = ?$

18. Use your calculator to evaluate these expressions. Express the final answer in proper scientific notation.

    a. $(9.09 \times 10^{8}) \div [(6.33 \times 10^{9}) \times (4.066 \times 10^{-7})] = ?$
    b. $9{,}345 \times 34.866 \div 0.00665 = ?$

## Answers

1. a. $5.69 \times 10^{1}$
   b. $5.631 \times 10^{5}$
   c. $8.04 \times 10^{-2}$
   d. $6.67 \times 10^{-6}$

3. a. $6.56 \times 10^{-3}$
   b. $6.56 \times 10^{4}$
   c. $4.567 \times 10^{6}$
   d. $5.507 \times 10^{-6}$

5. a. 138,100
   b. 0.000000522
   c. 99,980

7. a. 8.09
   b. 0.00003088
   c. −423.9

9. a. $7.244 \times 10^{4}$
   b. $9.943 \times 10^{-2}$
   c. $5.88399 \times 10^{7}$

11. a. $3.451 \times 10^{4}$
    b. $2.34 \times 10^{-4}$
    c. $1.8 \times 10^{1}$

13. a. $1.2345678 \times 10^{5}$
    b. $9.849 \times 10^{4}$
    c. $4.45 \times 10^{-7}$

15. a. $3.3744 \times 10^{11}$
    b. $3.3407 \times 10^{-11}$
    c. $3.665 \times 10^{-6}$

17. a. $6.7346 \times 10^{0}$
    b. $1.6691 \times 10^{-14}$

## Exercises (Expressing Units)

1. Identify the unit in each quantity.

   a. 2 boxes of crayons
   b. 3.5 grams of gold

2. Identify the unit in each quantity.

   a. 32 oz of cheddar cheese
   b. 0.045 $cm^3$ of water

3. Identify the unit in each quantity.

   a. 9.58 s (the current world record in the 100 m dash)
   b. 6.14 m (the current world record in the pole vault)

4. Identify the unit in each quantity.

   a. 2 dozen eggs
   b. 2.4 km/s (the escape velocity of the moon, which is the velocity you need at the surface to escape the moon's gravity)

5. Indicate what multiplier each prefix represents.

   a. k
   b. m
   c. M

6. Indicate what multiplier each prefix represents.

   a. c
   b. G
   c. µ

7. Give the prefix that represents each multiplier.

a. 1/1,000th ×
b. 1,000 ×
c. 1,000,000,000 ×

8. Give the prefix that represents each multiplier.

a. 1/1,000,000,000th ×
b. 1/100th ×
c. 1,000,000 ×

9. Complete the following table with the missing information.

Table with Units on the right and Abbreviation on the left side. Missing units on the second and third rows on the right side of the table. Missing abbreviations on the first and fourth rows on the left side of the table.

| Unit | Abbreviation |
|---|---|
| kilosecond | |
| | mL |
| | Mg |
| centimeter | |

10. Complete the following table with the missing information.

Table with Units on the right and Abbreviation on the left side. Missing units on the third and fourth rows on the right side of the table. Missing abbreviations on the first, second, and fifth row on the left side of the table.

| Unit | Abbreviation |
|---|---|
| kilometer per second | |
| second | |
| | $cm^3$ |
| | µL |
| nanosecond | |

11. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. $3.44 \times 10^{-6}$ s
b. 3,500 L
c. 0.045 m

12. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 0.000066 m/s (Hint: you need consider only the unit in the numerator.)
b. $4.66 \times 10^{6}$ s
c. 7,654 L

13. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 43,600 mL
b. 0.0000044 m
c. 1,438 ms

14. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 0.000000345 $m^3$
b. 47,000,000 $mm^3$

c. 0.00665 L

15. Multiplicative prefixes are used for other units as well, such as computer memory. The basic unit of computer memory is the byte (b). What is the unit for one million bytes?

16. You may have heard the terms *microscale* or *nanoscale* to represent the sizes of small objects. What units of length do you think are useful at these scales? What fractions of the fundamental unit of length are these units?

17. Acceleration is defined as a change in velocity per time. Propose a unit for acceleration in terms of the fundamental SI units.

18. Density is defined as the mass of an object divided by its volume. Propose a unit of density in terms of the fundamental SI units.

## Answers

1.

a. boxes of crayons
b. grams of gold

3.

a. seconds
b. meters

5.

a. 1,000 x
b. 1/1,000 x
c. 1,000,000 x

7.

a. milli-
b. kilo-
c. giga-

9.

Table with Units on the right and Abbreviation on the left side.

| Unit | Abbreviation |
|---|---|
| kilosecond | ks |
| milliliter | mL |
| megagram | Mg |
| centimeter | cm |

11.

a. 3.44 µs
b. 3.5 kL
c. 4.5 cm

13.

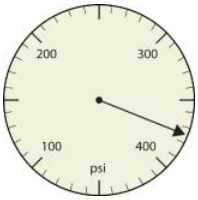
a. 43.6 mL
b. 4.4 µm
c. 1.438 s

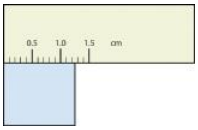
15. megabytes (Mb)

17. meters/second$^2$

## Exercises (Significant Figures)

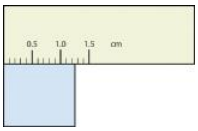
1. Express each measurement to the correct number of significant figures.



2. Express each measurement to the correct number of significant figures.



3. How many significant figures do these numbers have?

a. 23
b. 23.0
c. 0.00023
d. 0.0002302

4. How many significant figures do these numbers have?

a. $5.44 \times 10^{8}$
b. $1.008 \times 10^{-5}$
c. 43.09
d. 0.0000001381

5. How many significant figures do these numbers have?

a. 765,890
b. 765,890.0
c. $1.2000 \times 10^{5}$
d. 0.0005060

6. How many significant figures do these numbers have?

a. 0.009
b. 0.0000009
c. 65,444
d. 65,040

7. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 56.0 + 3.44 = ?


https://chem.libretexts.org/@go/page/501286

b. 0.00665 + 1.004 = ?
c. 45.99 − 32.8 = ?
d. 45.99 − 32.8 + 75.02 = ?

8. Compute and express each answer with the proper number of significant figures, rounding as necessary.

1. 1.005 + 17.88 = ?
2. 56,700 − 324 = ?
3. 405,007 − 123.3 = ?
4. 55.5 + 66.66 − 77.777 = ?

9. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 56.7 × 66.99 = ?
b. 1.0÷ 77 = ?
c. 1.00÷ 77.0 = ?
d. 6.022 × 1.89 = ?

10. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 0.000440 × 17.22 = ?
b. 203,000 ÷ 0.044 = ?
c. 67 × 85.0 × 0.0028 = ?
d. 999,999 ÷ 3,310 = ?

11. a. Write the number 87,449 in scientific notation with four significant figures.
b. Write the number 0.000066600 in scientific notation with five significant figures.

12. a. Write the number 306,000,000 in scientific notation to the proper number of significant figures.
b. Write the number 0.0000558 in scientific notation with two significant figures

13. Perform each calculation and limit each answer to three significant figures.

a. 67,883 × 0.004321 = ?
b. $(9.67 \times 10^3) \times 0.0055087$ = ?

14. Perform each calculation and limit each answer to four significant figures.

a. 18,900 × 76.33 ÷ 0.00336 = ?
b. 0.77604 ÷ 76,003 × 8.888 = ?

## Answers

1.

a. 375 psi
b. 1.30 cm

3.

a. two
b. three
c. two
d. four

5.

a. five
b. seven
c. five
d. four

7.

a. 59.4
b. 1.011
c. 13.2
d. 88.2

9.

a. $3.80 \times 10^3$
b. 0.013
c. 0.0130
d. 11.4

11. a. $\mathbf{8.745 \times 10^4}$
b. $\mathbf{6.6600 \times 10^{-5}}$

13. a. 293
b. 53.3

## Exercises (Converting Units)

1. Write the two conversion factors that exist between the two given units.

a. milliliters and liters
b. microseconds and seconds
c. kilometers and meters

2. Write the two conversion factors that exist between the two given units.

a. kilograms and grams
b. milliseconds and seconds
c. centimeters and meters

3. Perform the following conversions.

a. 5.4 km to meters
b. 0.665 m to millimeters
c. 0.665 m to kilometers

4. Perform the following conversions.

a. 90.6 mL to liters
b. 0.00066 ML to liters
c. 750 L to kiloliters

5. Perform the following conversions.

a. 17.8 µg to grams
b. $7.22 \times 10^2$ kg to grams
c. 0.00118 g to nanograms

6. Perform the following conversions.

a. 833 ns to seconds
b. 5.809 s to milliseconds
c. $2.77 \times 10^6$ s to megaseconds

7. Perform the following conversions.

a. 9.44 $m^2$ to square centimeters
b. $3.44 \times 10^8$ $mm^3$ to cubic meters

8. Perform the following conversions.

a. 0.00444 $cm^3$ to cubic meters
b. $8.11 \times 10^2$ $m^2$ to square nanometers

9. Why would it be inappropriate to convert square centimeters to cubic meters?

10. Why would it be inappropriate to convert from cubic meters to cubic seconds?

11. Perform the following conversions.

a. 45.0 m/min to meters/second
b. 0.000444 m/s to micrometers/second
c. 60.0 km/h to kilometers/second

12. Perform the following conversions.

a. $3.4 \times 10^2$ cm/s to centimeters/minute
b. 26.6 mm/s to millimeters/hour
c. 13.7 kg/L to kilograms/milliliters

13. Perform the following conversions.

a. 0.674 kL to milliliters
b. $2.81 \times 10^{12}$ mm to kilometers
c. 94.5 kg to milligrams

14. Perform the following conversions.

a. $6.79 \times 10^{-6}$ kg to micrograms
b. 1.22 mL to kiloliters
c. $9.508 \times 10^{-9}$ ks to milliseconds

15. Perform the following conversions.

a. $6.77 \times 10^{14}$ ms to kiloseconds
b. 34,550,000 cm to kilometers

16. Perform the following conversions.

a. $4.701 \times 10^{15}$ mL to kiloliters
b. $8.022 \times 10^{-11}$ ks to microseconds

17. Perform the following conversions. Note that you will have to convert units in both the numerator and the denominator.

a. 88 ft/s to miles/hour (Hint: use 5,280 ft = 1 mi.)
b. 0.00667 km/h to meters/second

18. Perform the following conversions. Note that you will have to convert units in both the numerator and the denominator.

a. $3.88 \times 10^2$ mm/s to kilometers/hour
b. 1.004 kg/L to grams/milliliter

19. What is the area in square millimeters of a rectangle whose sides are 2.44 cm × 6.077 cm? Express the answer to the proper number of significant figures.

20. What is the volume in cubic centimeters of a cube with sides of 0.774 m? Express the answer to the proper number of significant figures.

21. The formula for the area of a triangle is 1/2 × base × height. What is the area of a triangle in square centimeters if its base is 1.007 m and its height is 0.665 m? Express the answer to the proper number of significant figures.

22. The formula for the area of a triangle is 1/2 × base × height. What is the area of a triangle in square meters if its base is 166 mm and its height is 930.0 mm? Express the answer to the proper number of significant figures.

## Answers

1.

a. $\frac{1000mL}{1L}$ *and* $\frac{1L}{1000mL}$

b. $\frac{1000000\mu s}{1s}$ *and* $\frac{1s}{1000000\mu s}$

c. $\frac{1000m}{1km}$ *and* $\frac{1km}{1000m}$

3.

a. 5,400 m
b. 665 mm
c. $6.65 \times 10^{-4}$ km

5.

a. $1.78 \times 10^{-5}$ g
b. $7.22 \times 10^{5}$ g
c. $1.18 \times 10^{6}$ ng

7.

a. 94,400 $cm^2$
b. 0.344 $m^3$

9. One is a unit of area, and the other is a unit of volume.

11.

a. 0.75 m/s
b. 444 µm/s
c. $1.666 \times 10^{-2}$ km/s

13.

a. 674,000 mL
b. $2.81 \times 10^{6}$ km
c. $9.45 \times 10^{7}$ mg

15.

a. $6.77 \times 10^{8}$ ks
b. 345.5 km

17.

a. $6.0 \times 10^{1}$ mi/h
b. 0.00185 m/s

19. $1.48 \times 10^{3}$ $mm^2$

21. $3.35 \times 10^{3}$ $cm^2$

## Exercises (Other Units - Temperature and Density)

1. Perform the following conversions.

a. 255°F to degrees Celsius
b. −255°F to degrees Celsius
c. 50.0°C to degrees Fahrenheit
d. −50.0°C to degrees Fahrenheit

2. Perform the following conversions.

    a. 1,065°C to degrees Fahrenheit
    b. −222°C to degrees Fahrenheit
    c. 400.0°F to degrees Celsius
    d. 200.0°F to degrees Celsius

3. Perform the following conversions.

    a. 100.0°C to kelvins
    b. −100.0°C to kelvins
    c. 100 K to degrees Celsius
    d. 300 K to degrees Celsius

4. Perform the following conversions.

    a. 1,000.0 K to degrees Celsius
    b. 50.0 K to degrees Celsius
    c. 37.0°C to kelvins
    d. −37.0°C to kelvins

5. Convert 0 K to degrees Celsius. What is the significance of the temperature in degrees Celsius?

6. Convert 0 K to degrees Fahrenheit. What is the significance of the temperature in degrees Fahrenheit?

7. The hottest temperature ever recorded on the surface of the earth was 136°F in Libya in 1922. What is the temperature in degrees Celsius and in kelvins?

8. The coldest temperature ever recorded on the surface of the earth was −128.6°F in Vostok, Antarctica, in 1983. What is the temperature in degrees Celsius and in kelvins?

9. Give at least three possible units for density.

10. What are the units when density is inverted? Give three examples.

11. A sample of iron has a volume of 48.2 $cm^3$. What is its mass?

12. A sample of air has a volume of 1,015 mL. What is its mass?

13. The volume of hydrogen used by the *Hindenburg*, the German airship that exploded in New Jersey in 1937, was $2.000 \times 10^8$ L. If hydrogen gas has a density of 0.0899 g/L, what mass of hydrogen was used by the airship?

14. The volume of an Olympic-sized swimming pool is $2.50 \times 10^9$ $cm^3$. If the pool is filled with alcohol ($d$ = 0.789 g/$cm^3$), what mass of alcohol is in the pool?

15. A typical engagement ring has 0.77 $cm^3$ of gold. What mass of gold is present?

16. A typical mercury thermometer has 0.039 mL of mercury in it. What mass of mercury is in the thermometer?

17. What is the volume of 100.0 g of lead if lead has a density of 11.34 g/$cm^3$?

18. What is the volume of 255.0 g of uranium if uranium has a density of 19.05 g/cm$^3$?

19. What is the volume in liters of 222 g of neon if neon has a density of 0.900 g/L?

20. What is the volume in liters of 20.5 g of sulfur hexafluoride if sulfur hexafluoride has a density of 6.164 g/L?

21. Which has the greater volume, 100.0 g of iron ($d$ = 7.87 g/cm$^3$) or 75.0 g of gold ($d$ = 19.3 g/cm$^3$)?

22. Which has the greater volume, 100.0 g of hydrogen gas ($d$ = 0.0000899 g/cm$^3$) or 25.0 g of argon gas ($d$ = 0.00178 g/cm$^3$)?

### Answers

1.

a. 124°C
b. −159°C
c. 122°F
d. −58°F

3.

a. 373 K
b. 173 K
c. −173°C
d. 27°C

5. −273°C. This is the lowest possible temperature in degrees Celsius.

7. 57.8°C; 331 K

9. g/mL, g/L, and kg/L (answers will vary)

11. 379 g

13. 1.80 × 10$^7$ g

15. 15 g

17. 8.818 cm$^3$

19. 247 L

21. The 100.0 g of iron has the greater volume

## Additional Exercises

1. Evaluate 0.00000000552 × 0.0000000006188 and express the answer in scientific notation. You may have to rewrite the original numbers in scientific notation first.

2. Evaluate 333,999,500,000 ÷ 0.00000000003396 and express the answer in scientific notation. You may need to rewrite the original numbers in scientific notation first.

3. Express the number 6.022 × 10$^{23}$ in standard notation.

4. Express the number 6.626 × 10$^{-34}$ in standard notation.

5. When powers of 10 are multiplied together, the powers are added together. For example, 10$^2$ × 10$^3$ = 10$^{2+3}$ = 10$^5$. With this in mind, can you evaluate (4.506 × 10$^4$) × (1.003 × 10$^2$) without entering scientific notation into your calculator?

6. When powers of 10 are divided into each other, the bottom exponent is subtracted from the top exponent. For example, $10^5/10^3 = 10^{5-3} = 10^2$. With this in mind, can you evaluate $(8.552 \times 10^6) \div (3.129 \times 10^3)$ without entering scientific notation into your calculator?

7. Consider the quantity two dozen eggs. Is the number in this quantity "two" or "two dozen"? Justify your choice.

8. Consider the quantity two dozen eggs. Is the unit in this quantity "eggs" or "dozen eggs"? Justify your choice.

9. Fill in the blank: 1 km = _____ μm.

10. Fill in the blank: 1 Ms = _____ ns.

11. Fill in the blank: 1 cL = _____ ML.

12. Fill in the blank: 1 mg = _____ kg.

13. Express 67.3 km/h in meters/second.

14. Express 0.00444 m/s in kilometers/hour.

15. Using the idea that 1.602 km = 1.000 mi, convert a speed of 60.0 mi/h into kilometers/hour.

16. Using the idea that 1.602 km = 1.000 mi, convert a speed of 60.0 km/h into miles/hour.

17. Convert 52.09 km/h into meters/second.

18. Convert 2.155 m/s into kilometers/hour.

19. Use the formulas for converting degrees Fahrenheit into degrees Celsius to determine the relative size of the Fahrenheit degree over the Celsius degree.

20. Use the formulas for converting degrees Celsius into kelvins to determine the relative size of the Celsius degree over kelvins.

21. What is the mass of 12.67 L of mercury?

22. What is the mass of 0.663 $m^3$ of air?

23. What is the volume of 2.884 kg of gold?

24. What is the volume of 40.99 kg of cork? Assume a density of 0.22 $g/cm^3$.

## Answers

1. $3.42 \times 10^{-18}$

3. 602,200,000,000,000,000,000,000

5. $4.520 \times 10^{6}$

7. The quantity is two; dozen is the unit.

9. 1,000,000,000

11. 1/100,000,000

13. 18.7 m/s

15. 96.1 km/h

17. 14.47 m/s

19. One Fahrenheit degree is nine-fifths the size of a Celsius degree.

21. $1.72 \times 10^{5}$ g

23. 149 mL

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# CHAPTER OVERVIEW

## 3: Matter and Energy


3.1: Introduction to Matter and Energy
3.2: Classifying Matter According to Its State—Solid, Liquid, and Gas
3.3: Classifying Matter According to Its Composition
3.4: Differences in Matter- Physical and Chemical Properties
3.5: Changes in Matter - Physical and Chemical Changes
3.6: Conservation of Mass - There is No New Matter
3.7: Energy
3.8: Phase Changes


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# 3.1: Introduction to Matter and Energy

**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 3.1.1**

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 3.1.1**

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 **3.1.1**).



Figure **3.1.1**: 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 **3.1.2**). In the following chapters, we will see how descriptions of physical and chemical properties are important aspects of chemistry.



Figure **3.1.2**: 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 **3.1.3**), or alcohol in a thermometer may change volume as the temperature changes. A physical change does not affect the chemical composition of matter.



Figure **3.1.2**: 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 **3.1.2** "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 3.1.2

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 3.1.2

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 **3.1.3**).



Figure **3.1.3**: 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 3.1.3**

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 3.1.3**

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 **3.1.4**). 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 **3.1.4**: 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.

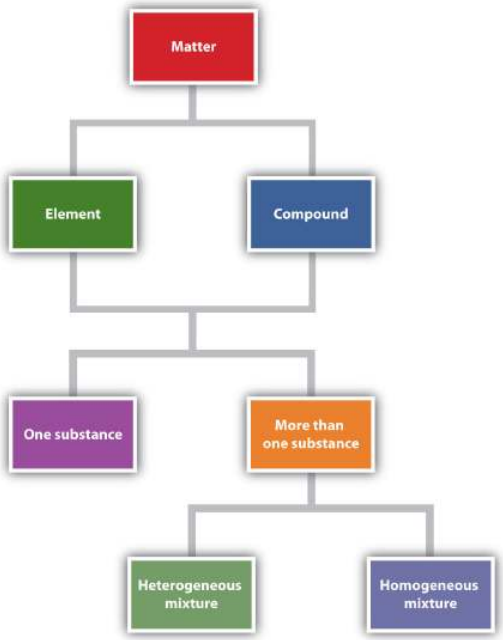



Figure **3.1.5**: Describing Matter. This flowchart shows how matter can be described. Matter forks into element & compound which fork into 1 substance & more than 1 substance. More than 1 substance is homogenous or heterogenous.

### Example **3.1.1**: 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 3.1.6: 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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# 3.2: Classifying Matter According to Its State—Solid, Liquid, and Gas

**Learning Objectives**

- To describe the solid, liquid and gas phases.

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

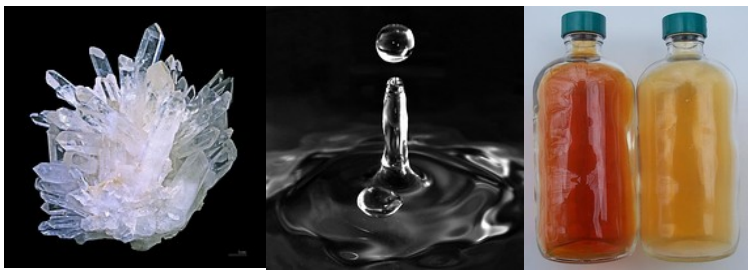


Figure **3.2.1**: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure **3.2.2** shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

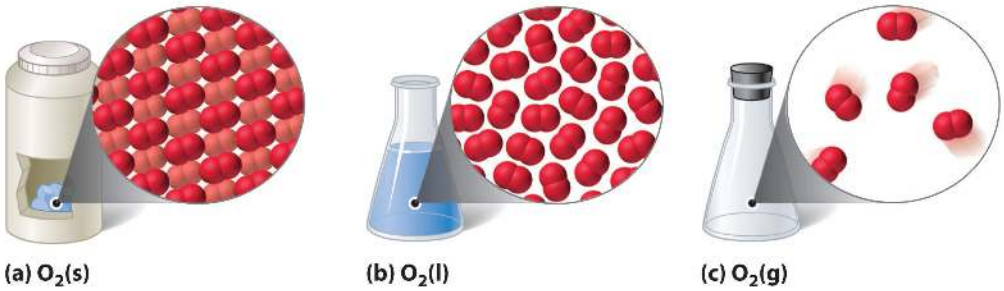
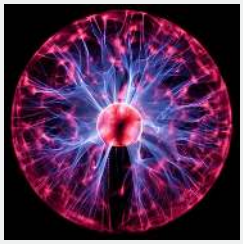



Figure **3.2.2**: A Representation of the Solid, Liquid, and Gas States. (a) Solid $O_2$ has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid $O_2$ conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous $O_2$ fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

**Plasma: A Fourth State of Matter**

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



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

## Solids

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

Solids are defined by the following characteristics:

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

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

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure **3.2.3**. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

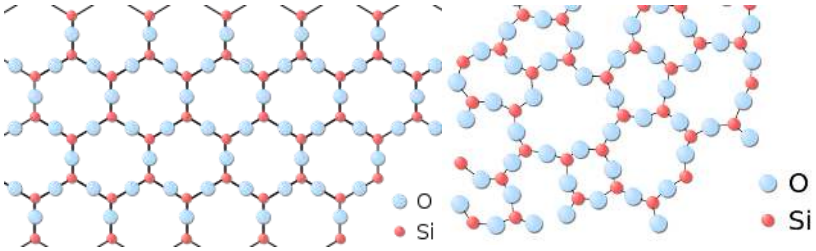



Figure **3.2.3**: (left) The periodic crystalline lattice structure of quartz $SiO_2$ in two-dimensions. (right) The random network structure of glassy $SiO_2$ in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

## Liquids

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

Liquids have the following characteristics:

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

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



***Video* 3.2.1**: *Mercury boiling to become a gas.*

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

## Gases

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

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

Table **3.2.1**: Characteristics of the Three States of Matter

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

### Example 3.2.1

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

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

#### Solution

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

**? Exercise 3.2.1**

What state or states of matter does each statement describe?

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

**Answer a:**

solid

**Answer b:**

gas

**Answer c:**

solid

## Summary

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

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# 3.3: Classifying Matter According to Its Composition

**Learning Objectives**

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

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure **3.3.1**). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

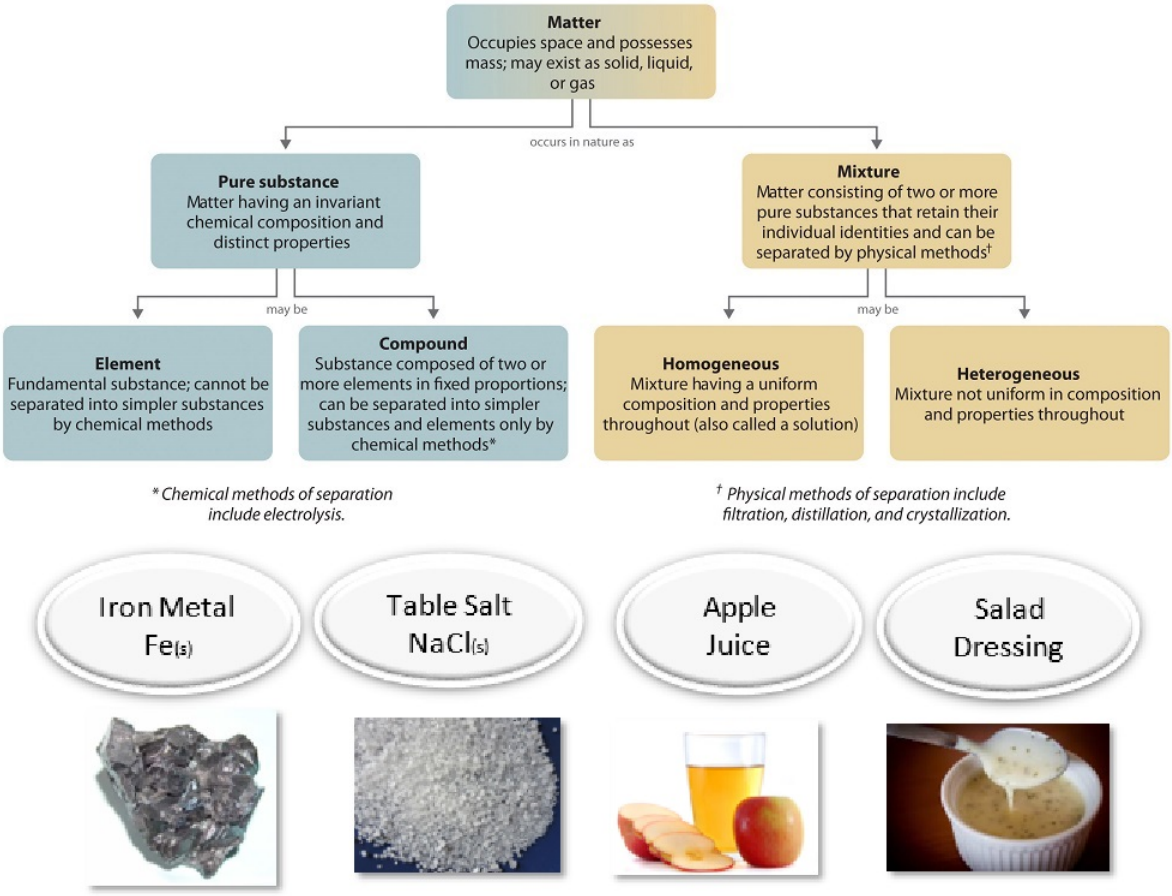



Figure **3.3.1**: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Ordinary table salt is called sodium chloride. It is considered a **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and

properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

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

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

### Phase

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

### Example 3.3.1

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 3.3.1**

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

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

**Answer a:**
homogeneous mixture (solution)

**Answer b:**
element

**Answer c:**
heterogeneous mixture

**Answer d:**
compound

**✓ Example 3.3.2**

How would a chemist categorize each example of matter?

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

Solution

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

**? Exercise 3.3.2**

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

**Answer a:**
a homogeneous mixture (solution), assuming it is filtered coffee

**Answer b:**
element

**Answer c:**
heterogeneous mixture

## Summary

Matter can be classified into two broad categories: pure substances and mixtures. A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Mixtures are physical combinations of two or more

elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

## Vocabulary

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

## Contributions & Attributions

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

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# 3.4: Differences in Matter- Physical and Chemical Properties

**Learning Objectives**

To separate physical from chemical properties.

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

## Physical Property

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

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 $g/cm^3$ at 25°C. The average densities of some common substances are in Table **3.4.1**. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table **3.4.1**: Densities of Common Substances

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

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



Figure **3.4.1**: Pencil (left) and Diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

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

## Chemical Properties

**Chemical properties** of matter describe its potential to undergo some chemical change or reaction by virtue of its composition. The elements, electrons, and bonds that are present give the matter potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, after studying chemistry for some time, you should be able to look at the formula of a compound and state some chemical property. For example, hydrogen has the potential to ignite and explode given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with an acid. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.



Figure **3.4.2**: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure **3.4.2**). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table **3.4.2**: Contrasting Physical and Chemical Properties

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

### Example 3.4.1

Which of the following is a chemical property of iron?

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

#### Solution

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

### Exercise 3.4.1A

Which of the following is a physical property of matter?

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

**Answer**

c

### Exercise 3.4.1B

Which of the following is a chemical property?

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

**Answer**

a

## Summary

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

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# 3.5: Changes in Matter - Physical and Chemical Changes

**Learning Objectives**

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

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

## Physical Change

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

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

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



Figure **3.5.1**: Ice melting is a physical change. When liquid water ($H_2O$) freezes into a solid state (ice), it appears changed; however, this change is only physical, as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

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

## Chemical Change

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

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



Figure **3.5.2**: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos )

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

Observations that help to indicate chemical change include:

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

### Example 3.5.1

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

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

#### Solution

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

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

**? Exercise 3.5.1**

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

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

**Answer a:**

physical change

**Answer b:**

chemical change

**Answer c:**

physical change

**Answer d:**

chemical change

## Separating Mixtures Through Physical Changes

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

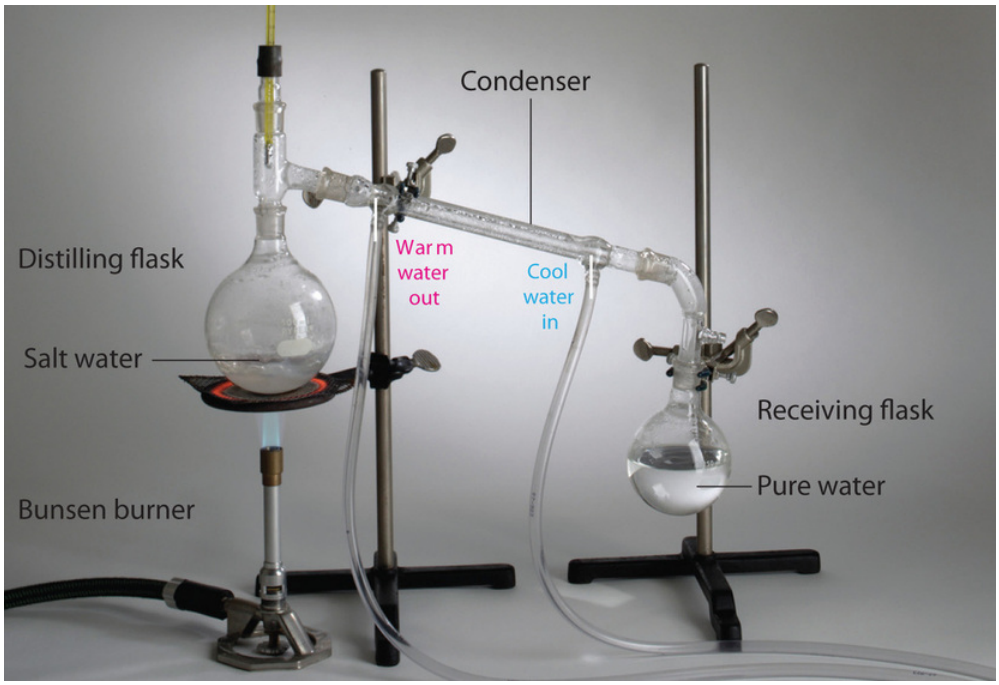



Figure **3.5.3***: The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.*

Parts of a distillation setup: Bunsen burner, salt water in distilling flask, condenser with cool water in and warm water out, pure water in receiving flask

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

Another example for using physical properties to separate mixtures is filtration (Figure **3.5.4**). Filtration is any mechanical, physical or biological operation that separates solids from fluids (liquids or gases) by adding a medium through which only the fluid can pass. The fluid that passes through is called the filtrate. There are many different methods of filtration; all aim to attain the separation of substances. Separation is achieved by some form of interaction between the substance or objects to be removed and the filter. The substance that is to pass through the filter must be a fluid, i.e. a liquid or gas. Methods of filtration vary depending on the location of the targeted material, i.e. whether it is dissolved in the fluid phase or suspended as a solid.

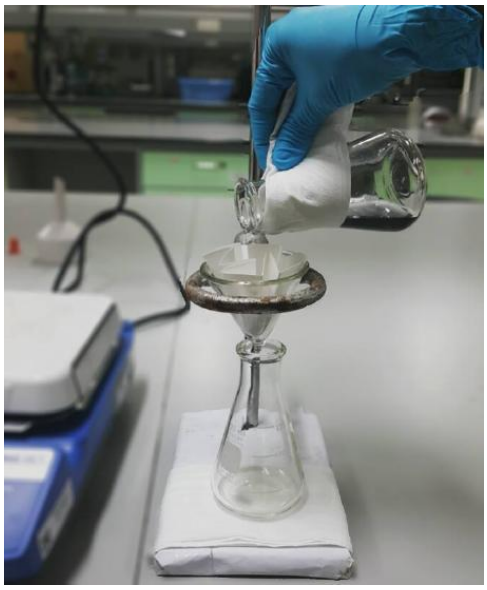


Figure **3.5.4***: Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).*

## Summary

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

## Contributions & Attributions

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- Boundless (www.boundless.com)

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# 3.6: Conservation of Mass - There is No New Matter

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Look at Figure **3.6.1** below. It shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose you had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose you had been able to measure the oxygen used by the fire and the gases produced by the fire. What would you find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

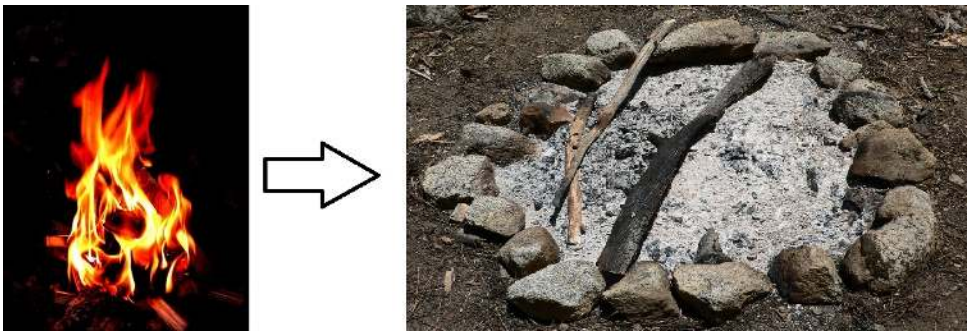


Figure **3.6.1**: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). (CC BY-SA 2.5; Einar Helland Berger for fire and Walter Siegmund for ash).

## Law of Conservation of Mass

The law of conservation of mass was created in 1789 by a French chemist, Antoine Lavoisier. The **law of conservation of mass** states that matter cannot be created or destroyed in a chemical reaction. For example, when wood burns, the mass of the soot, ashes, and gases equals the original mass of the charcoal and the oxygen when it first reacted. So the mass of the product equals the mass of the reactant. A reactant is the chemical reaction of two or more elements to make a new substance, and a product is the substance that is formed as the result of a chemical reaction (Video **3.6.1**). Matter and its corresponding mass may not be able to be created or destroyed, but can change forms to other substances like liquids, gases, and solids.



Video **3.6.1**: This is a nice little demonstration showing the Conservation of Mass in action.

If you witness a 300 kg tree burn to the ground, there are only ashes left after the burn, and all of them together weigh 10 kg. It may make you wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke, so the only thing left that you can see is the 10 kg of ash. If you know the law of conservation of mass, then you know that the other 290 kg has to go somewhere, because it has to equal the mass of the tree before it burnt down.

### Example 3.6.1

If heating 10.0 grams of calcium carbonate ($CaCO_3$) produces 4.4 g of carbon dioxide ($CO_2$) and 5.6 g of calcium oxide (CaO), show that these observations are in agreement with the law of conservation of mass.

#### Solution

$$\text{Mass of the reactants} = \text{Mass of the products}$$
$$10.0\,\text{g of } CaCO_3 = 4.4\,\text{g of } CO_2 + 5.6\,\text{g of CaO}$$
$$10.0\,\text{g of reactant} = 10.0\,\text{g of products}$$

Because the mass of the reactant is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

### Exercise 3.6.1

Potassium hydroxide (**KOH**) readily reacts with carbon dioxide ($CO_2$) to produce potassium carbonate ($K_2CO_3$) and water ( $H_2O$). How many grams of potassium carbonate are produced if 224.4 g of **KOH** reacts with 88.0 g of $CO_2$? The reaction also produces 36.0 g of water.

**Answer**

276.4 g of potassium carbonate

The Law is also applicable to both chemical and physical changes. For example, if you have an ice cube that melts into a liquid and you heat that liquid up, it becomes a gas. It will appear to have disappeared, but is still there.

## Summary

- Burning and other changes in matter do not destroy matter.
- The mass of matter is always the same before and after the changes occur.
- The law of conservation of mass states that matter cannot be created or destroyed.

## Contributions & Attributions

- Binod Shrestha (University of Lorraine)

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# 3.7: Energy

**Learning Objectives**

- Define heat and work.
- Distinguish between kinetic energy and potential energy.
- State the law of conservation of matter and energy.

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we are awake or asleep.

## Ability to Do Work or Produce Heat

When we speak of *using* energy, we are really referring to *transferring* energy from one place to another. When you use energy to throw a ball, you transfer energy from your body to the ball, and this causes the ball to fly through the air. When you use energy to warm your house, you transfer energy from the furnace to the air in your home, and this causes the temperature in your house to rise. Although energy is used in many kinds of different situations, all of these uses rely on energy being transferred in one of two ways. Energy can be transferred as *heat* or as *work*.

When scientists speak of ***heat***, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature, as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot—the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents (Figure **3.7.1**).



Figure **3.7.1**: Energy is transferred as heat from the hot stove element to the cooler pot until the pot and its contents become just as hot as the element. The energy that is transferred into the pot as heat is then used to cook the food.

Heat is only one way in which energy can be transferred. Energy can also be transferred as **work**. The scientific definition of work is *force (any push or pull) applied over a distance*. When you push an object and cause it to move, you do work, and you transfer some of *your* energy to the object. At this point, it's important to warn you of a common misconception. Sometimes we think that the amount of work done can be measured by the amount of effort put in. This may be true in everyday life, but it is not true in science. By definition, scientific work requires that force be applied *over a distance*. It does not matter how hard you push or how hard you pull. If you have not moved the object, you haven't done any work.

So far, we've talked about the two ways in which energy can be transferred from one place, or object, to another. Energy can be transferred as heat, and energy can be transferred as work. But the question still remains—*what IS energy*?

## Kinetic Energy

Machines use energy, our bodies use energy, energy comes from the sun, energy comes from volcanoes, energy causes forest fires, and energy helps us to grow food. With all of these seemingly different types of energy, it's hard to believe that there are really only two different *forms* of energy: kinetic energy and potential energy. **Kinetic energy** is energy associated with motion. When an object is moving, it has kinetic energy. When the object stops moving, it has no kinetic energy. While all moving objects have kinetic energy, not all moving objects have the same amount of kinetic energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has.

Kinetic energy is very common, and it's easy to spot examples of it in the world around you. Sometimes we even try to capture kinetic energy and use it to power things like our home appliances. If you are from California, you might have driven through the Tehachapi Pass near Mojave or the Montezuma Hills in Solano County and seen the windmills lining the slopes of the mountains (Figure **3.7.2**). These are two of the larger wind farms in North America. As wind rushes along the hills, the kinetic energy of the moving air particles turns the windmills, trapping the wind's kinetic energy so that people can use it in their houses and offices.



Figure **3.7.2**: A wind farm in Solano County harnesses the kinetic energy of the wind. (CC BY-SA 3.0 Unported; BDS2006 at Wikipedia)

## Potential Energy

**Potential energy** is *stored* energy. It is energy that remains available until we choose to use it. Think of a battery in a flashlight. If left on, the flashlight battery will run out of energy within a couple of hours, and the flashlight will die. If, however, you only use the flashlight when you need it, and turn it off when you don't, the battery will last for days or even months. The battery contains a certain amount of energy, and it will power the flashlight for a certain amount of time, but because the battery stores *potential* energy, you can choose to use the energy all at once, or you can save it and only use a small amount at a time.

Any stored energy is potential energy. There are a lot of different ways in which energy can be stored, and this can make potential energy very difficult to recognize. In general, an object has potential energy because of its *position relative to another object*. For example, when a rock is held above the earth, it has potential energy because of its position relative to the ground. This is *potential energy* because the energy is *stored* for as long as the rock is held in the air. Once the rock is dropped, though, the stored energy is released as kinetic energy as the rock falls.

## Chemical Energy

There are other common examples of potential energy. A ball at the top of a hill stores potential energy until it is allowed to roll to the bottom. When two magnets are held next to one another, they store potential energy too. For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called **chemical potential energy**. Chemical potential energy is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position?

As you learned earlier, the world, and all of the chemicals in it are made up of atoms and molecules. These store potential energy that is dependent on their positions relative to one another. Of course, you can't see atoms and molecules. Nevertheless, scientists do know a lot about the ways in which atoms and molecules interact, and this allows them to figure out how much potential energy

is stored in a specific quantity (like a cup or a gallon) of a particular chemical. *Different chemicals have different amounts of potential energy* because they are made up of different atoms, and those atoms have different positions relative to one another.

> Since different chemicals have different amounts of potential energy, scientists will sometimes say that potential energy depends not only on ***position***, but also on ***composition***. Composition affects potential energy because it determines which molecules and atoms end up next to one another. For example, the total potential energy in a cup of pure water is different than the total potential energy in a cup of apple juice, because the cup of water and the cup of apple juice are *composed* of different amounts of different chemicals.

At this point, you may wonder just how useful chemical potential energy is. If you want to release the potential energy stored in an object held above the ground, you just drop it. But how do you get potential energy out of chemicals? It's actually not difficult. Use the fact that different chemicals have *different amounts of potential energy*. If you start with chemicals that have a lot of potential energy and allow them to react and form chemicals with less potential energy, all the extra energy that was in the chemicals at the beginning, but not at the end, is released.

## Units of Energy

Energy is measured in one of two common units: the calorie and the joule. The joule (**J**) is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by $1°C$. For example, raising the temperature of $100\ g$ of water from $20°C$ to $22°C$ would require $100 \times 2 = 200\ cal$.

Calories contained within food are actually kilocalories (**kcal**). In other words, if a certain snack contains 85 food calories, it actually contains $85\ kcal$ or $85,000\ cal$. In order to make the distinction, the dietary calorie is written with a capital C.

$$1\ \text{kilocalorie} = 1\ \text{Calorie} = 1000\ \text{calories}$$

To say that the snack "contains" 85 Calories means that $85\ kcal$ of energy are released when that snack is processed by your body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1\ J = 0.2390\ cal \text{ or } 1\ cal = 4.184\ J$$

We can calculate the amount of heat released in kilojoules when a 400 Calorie hamburger is digested.

$$400\ Cal = 400\ kcal \times \frac{4.184\ kJ}{1\ kcal} = 1.67 \times 10^3\ kJ$$

## Summary

- Any time we use energy, we transfer energy from one object to another.
- Energy can be transferred in one of two ways: as heat, or as work.
- Heat is the term given to energy that is transferred from a hot object to a cooler object due to the difference in their temperatures.
- Work is the term given to energy that is transferred as a result of a force applied over a distance.
- Energy comes in two fundamentally different forms: kinetic energy and potential energy.
- Kinetic energy is the energy of motion.
- Potential energy is stored energy that depends on the position of an object relative to another object.
- Chemical potential energy is a special type of potential energy that depends on the positions of different atoms and molecules relative to one another.
- Chemical potential energy can also be thought of according to its dependence on chemical composition.
- Energy can be converted from one form to another. The total amount of mass and energy in the universe is conserved.

## Contributions & Attributions

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# 3.8: Phase Changes

**Learning Objectives**

- Determine the heat associated with a phase change.

Matter can exist in one of several different states, including a gas, liquid, or solid state. The amount of energy in molecules of matter determines the **state of matter**.

- A **gas** is a state of matter in which atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide.
- A **liquid** is a state of matter in which atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another.
- A **solid** is a state of matter in which atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another.

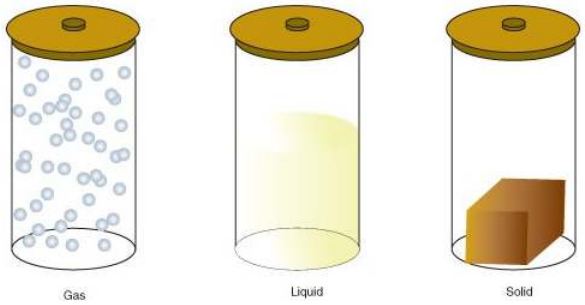



Figure **3.8.1**: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

The following are the changes of state:

Changes of State

| Phase to Phase transition | Transition Name |
|---|---|
| Solid → Liquid | Melting or fusion |
| Liquid → Gas | Vaporization |
| Liquid → Solid | Freezing |
| Gas → Liquid | Condensation |
| Solid → Gas | Sublimation |

- If heat is added to a substance, such as in melting, vaporization, and sublimation, the process is **endothermic**. In this instance, heat is increasing the speed of the molecules causing them move faster (examples: solid to liquid; liquid to gas; solid to gas).
- If heat is removed from a substance, such as in freezing and condensation, then the process is **exothermic**. In this instance, heat is decreasing the speed of the molecules causing them move slower (examples: liquid to solid; gas to liquid). These changes **release heat** to the surroundings.
- The amount of heat needed to change a sample from solid to liquid would be the same to reverse from liquid to solid. The only difference is the direction of heat transfer.

**Example 3.8.1**

Label each of the following processes as endothermic or exothermic.

a. water boiling
b. 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.

**Solution b**

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.

**3.8.1 Exercise**

Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

**Answer**

a. exothermic

b. endothermic

A phase change is a physical process in which a substance goes from one phase to another. Usually the change occurs when adding or removing heat at a particular temperature, known as the melting point or the boiling point of the substance. The melting point is the temperature at which the substance goes from a solid to a liquid (or from a liquid to a solid). The boiling point is the temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid). The nature of the phase change depends on the direction of the heat transfer. Heat going *into* a substance changes it from a solid to a liquid or a liquid to a gas. Removing heat *from* a substance changes a gas to a liquid or a liquid to a solid.

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water ($H_2O$) as an example. On the Celsius scale, $H_2O$ has a melting point of 0°C and a boiling point of 100°C. At 0°C, both the solid and liquid phases of $H_2O$ can coexist. However, if heat is added, some of the solid $H_2O$ will melt and turn into liquid $H_2O$. If heat is removed, the opposite happens: some of the liquid $H_2O$ turns into solid $H_2O$. A similar process can occur at 100°C: adding heat increases the amount of gaseous $H_2O$, while removing heat increases the amount of liquid $H_2O$ (Figure 3.8.1).

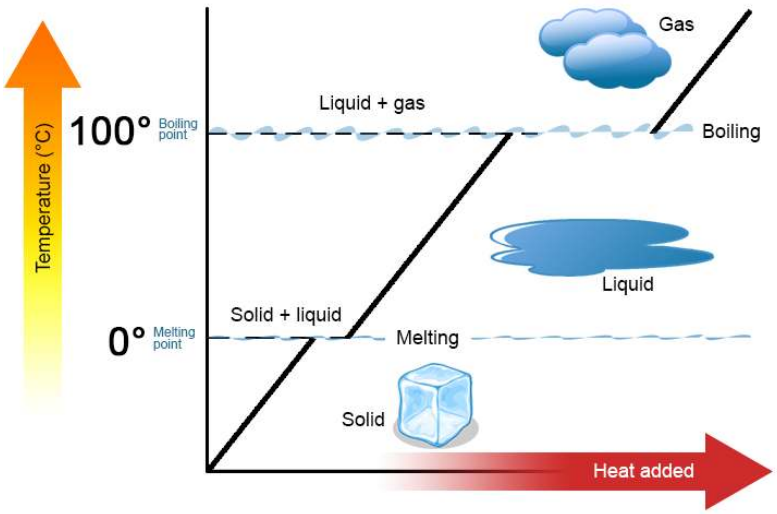



Figure **3.8.2**: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0 °C, the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).

> Water is a good substance to use as an example because many people are already familiar with it. Other substances have melting points and boiling points as well.

Second, as shown in Figure **3.8.1**, **the temperature of a substance does not change** as the substance goes **from one phase to another**. In other words, phase changes are isothermal (isothermal means "constant temperature"). Again, consider $H_2O$ as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of $H_2O$ can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

For each phase change of a substance, there is a characteristic quantity of heat needed to perform the phase change per gram (or per mole) of material. The heat of fusion ($\Delta H_{\text{fus}}$) is the amount of heat per gram (or per mole) required for a phase change that occurs at the melting point. The heat of vaporization ($\Delta H_{\text{vap}}$) is the amount of heat per gram (or per mole) required for a phase change that occurs at the boiling point. If you know the total number of grams or moles of material, you can use the $\Delta H_{\text{fus}}$ or the $\Delta H_{\text{vap}}$ to determine the total heat being transferred for melting or solidification using these expressions:

$$\text{heat} = n \times \Delta H_{fus}$$

where $n$ is the number of moles and $\Delta H_{fus}$ is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{fus}$$

where $m$ is the mass in grams and $\Delta H_{fus}$ is expressed in energy/gram.

For the boiling or condensation, use these expressions:

$$\text{heat} = n \times \Delta H_{vap}$$

where $n$ is the number of moles) and $\Delta H_{vap}$ is expressed in energy/mole or

$$\text{heat} = m \times \Delta H_{vap}$$

where $m$ is the mass in grams and $\Delta H_{vap}$ is expressed in energy/gram.

Remember that a phase change depends on the direction of the heat transfer. If heat transfers in, solids become liquids, and liquids become solids at the melting and boiling points, respectively. If heat transfers out, liquids solidify, and gases condense into liquids. At these points, there are no changes in temperature as reflected in the above equations.

### Example 3.8.2

How much heat is necessary to melt 55.8 g of ice (solid $H_2O$) at 0°C? The heat of fusion of $H_2O$ is 79.9 cal/g.

**Solution**

We can use the relationship between heat and the heat of fusion (Equation **3.8.1**) to determine how many cal of heat are needed to melt this ice:

$$\text{heat} = \text{m} \times \Delta H_{\text{fus}}$$

$$\text{heat} = (55.8 \cancel{\text{g}}) \left( \frac{79.9 \text{ cal}}{\cancel{\text{g}}} \right) = 4,460 \text{ cal}$$

### Exercise 3.8.2

How much heat is necessary to vaporize 685 g of $H_2O$ at 100°C? The heat of vaporization of $H_2O$ is 540 cal/g.

**Answer**

$$\text{heat} = \text{m} \times \Delta H_{\text{vap}}$$

$$\text{heat} = (685 \cancel{\text{g}}) \left( \frac{540 \text{ cal}}{\cancel{\text{g}}} \right) = 370,000 \text{ cal}$$

Table **3.8.1** lists the heats of fusion and vaporization for some common substances. Note the units on these quantities; when you use these values in problem solving, make sure that the other variables in your calculation are expressed in units consistent with the units in the specific heats or the heats of fusion and vaporization.

Table **3.8.1**: Heats of Fusion and Vaporization for Selected Substances

| Substance | $\Delta H_{\text{fus}}$ (cal/g) | $\Delta H_{\text{vap}}$ (cal/g) |
|---|---|---|
| aluminum (Al) | 94.0 | 2,602 |
| gold (Au) | 15.3 | 409 |
| iron (Fe) | 63.2 | 1,504 |
| water ($H_2O$) | 79.9 | 540 |
| sodium chloride (NaCl) | 123.5 | 691 |
| ethanol ($C_2H_5OH$) | 45.2 | 200.3 |
| benzene ($C_6H_6$) | 30.4 | 94.1 |

### Sublimation

There is also a phase change where a solid goes directly to a gas:

\[\text{solid} \rightarrow \text{gas} \nonumber\]

This phase change is called *sublimation*. Each substance has a characteristic heat of sublimation associated with this process. For example, the heat of sublimation ($\Delta H_{\text{sub}}$) of $H_2O$ is 620 cal/g.

We encounter sublimation in several ways. You may already be familiar with dry ice, which is simply solid carbon dioxide ($CO_2$). At −78.5°C (−109°F), solid carbon dioxide sublimes, changing directly from the solid phase to the gas phase:

$$\text{CO}_2(\text{s}) \xrightarrow{-78.5^\circ\text{C}} \text{CO}_2(\text{g})$$

Solid carbon dioxide is called dry ice because it does not pass through the liquid phase. Instead, it does directly to the gas phase. (Carbon dioxide *can* exist as liquid but only under high pressure.) Dry ice has many practical uses, including the long-term preservation of medical samples.

Even at temperatures below 0°C, solid $H_2O$ will slowly sublime. For example, a thin layer of snow or frost on the ground may slowly disappear as the solid $H_2O$ sublimes, even though the outside temperature may be below the freezing point of water. Similarly, ice cubes in a freezer may get smaller over time. Although frozen, the solid water slowly sublimes, redepositing on the colder cooling elements of the freezer, which necessitates periodic defrosting (frost-free freezers minimize this redeposition). Lowering the temperature in a freezer will reduce the need to defrost as often.

Under similar circumstances, water will also sublime from frozen foods (e.g., meats or vegetables), giving them an unattractive, mottled appearance called freezer burn. It is not really a "burn," and the food has not necessarily gone bad, although it looks unappetizing. Freezer burn can be minimized by lowering a freezer's temperature and by wrapping foods tightly so water does not have any space to sublime into.

## Key Takeaway

- There is an energy change associated with any phase change.

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# CHAPTER OVERVIEW

## 4: Atoms, Elements, and Molecules


4.1: Cutting Aluminum until you get Atoms
4.2: Indivisible - The Atomic Theory
4.3: The Nuclear Atom
4.4: The Properties of Protons, Neutrons, and Electrons
4.5: The Elements
4.6: Looking for Patterns - The Periodic Table
4.7: Ions - Losing and Gaining Electrons
4.8: Isotopes - When the Number of Neutrons Varies
4.9: Atomic Mass - The Average Mass of an Element's Atoms
4.10: Compounds Display Constant Composition
4.11: Chemical Formulas - How to Represent Compounds
4.12: A Molecular View of Elements and Compounds


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# 4.1: Cutting Aluminum until you get Atoms

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 by Democritus and other ancient Greek philosophers (Figure **4.1.1**).

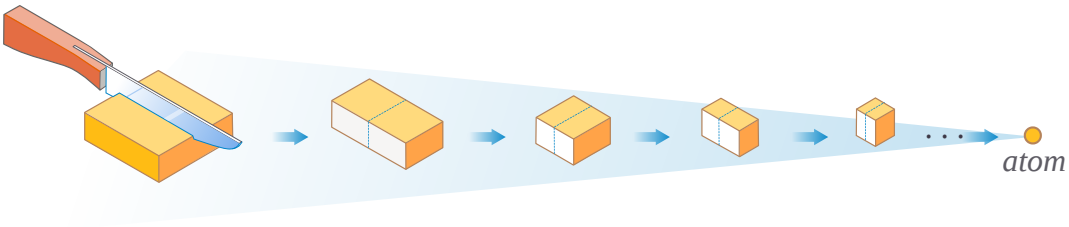



Figure **4.1.1**: Democritus argued that matter, like an aluminum block, cannot be repeatedly cut in half perpetually. At some point, a limit is reached before the substance can no longer be called "aluminum"; this is the atomic limit. (CC BY-NC; Ümit Kaya via LibreTexts)

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

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# 4.2: 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 **4.2.1**: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

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

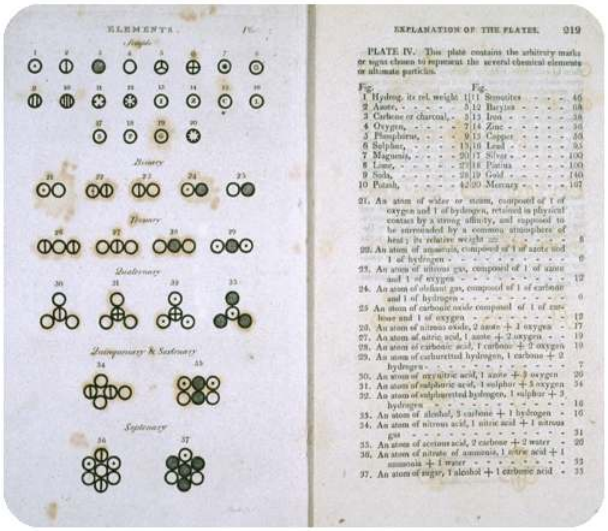
The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure **4.2.1**), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only $5.4 \times 10^{-10}$ 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. IV.

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.

Figure **4.2.2**: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video **4.2.1**, control of individual atoms can be use used create animations.



Video **4.2.1**: 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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# 4.3: The Nuclear Atom

**Learning Objectives**

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

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



Figure **4.3.1**: J. J. Thomson (center) concluded experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the model nuclear atom.

## Electrons and Plums

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

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

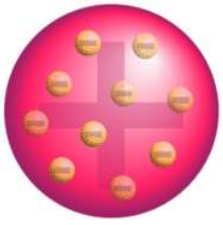


Figure **4.3.2**: The "plum pudding" model.

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

## Atoms and Gold

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

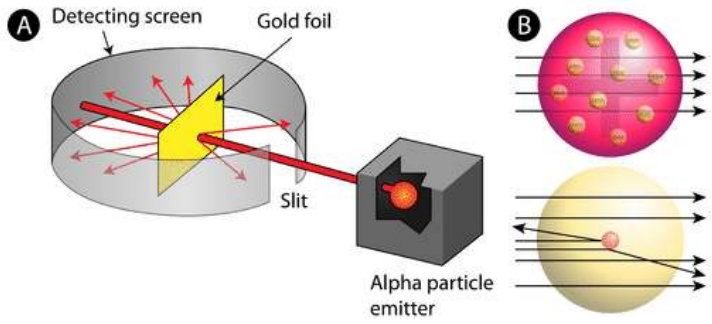



Figure **4.3.3**: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top), all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed not deflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

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

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

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

## Atomic Nucleus

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

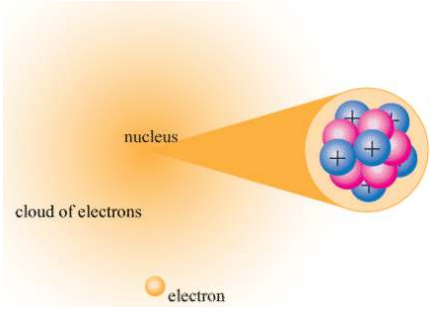



Figure **4.3.4**: The nuclear atom.

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

### Holding it all Together

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

## Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

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# 4.4: 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 4.4.1: 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 $(\text{amu})$, which is about $1.67 \times 10^{-27}$ 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 $1.67 \times 10^{-27}$ kilograms.) A neutron also has about the same diameter as a proton, or $1.7 \times 10^{-15}$ 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 4.4.1: Properties of Subatomic Particles

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

Table 4.4.1 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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# 4.5: The Elements

**Learning Objectives**

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

An element is a substance that cannot be broken down into simpler chemical substances. There are about 90 naturally occurring elements known on Earth. Using technology, scientists have been able to create nearly 30 additional elements that do not occur in nature. Today, chemistry recognizes 118 elements—some of which were created an atom at a time. Figure **4.5.1** shows some of the chemical elements.



Figure **4.5.1**: Samples of Elements. Gold is a yellowish solid, iron is a silvery solid, while mercury is a silvery liquid at room temperature. © Thinkstock

## Abundance

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

Table **4.5.1**: Elemental Composition of Earth

| Earth's Crust | | Earth (overall) | |
|---|---|---|---|
| **Element** | **Percentage** | **Element** | **Percentage** |
| oxygen | 46.1 | iron | 34.6 |
| silicon | 28.2 | oxygen | 29.5 |
| aluminum | 8.23 | silicon | 15.2 |
| iron | 5.53 | magnesium | 12.7 |
| calcium | 4.15 | nickel | 2.4 |
| sodium | 2.36 | sulfur | 1.9 |
| magnesium | 2.33 | all others | 3.7 |
| potassium | 2.09 | | |
| titanium | 0.565 | | |
| hydrogen | 0.14 | | |
| phosphorus | 0.105 | | |
| all others | 0.174 | | |

Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 14–17.

On the planet Earth, however, the situation is rather different. Oxygen makes up 46.1% of the mass of Earth's crust (the relatively thin layer of rock forming Earth's surface), mostly in combination with other elements, while silicon makes up 28.2%. Hydrogen, the most abundant element in the universe, makes up only 0.14% of Earth's crust. Table **4.5.1** lists the relative abundances of elements on Earth as a whole and in Earth's crust. Table **4.5.2** lists the relative abundances of elements in the human body. If you compare Table **4.5.1** and Table **4.5.2**, you will find disparities between the percentage of each element in the human body and on Earth. Oxygen has the highest percentage in both cases, but carbon, the element with the second highest percentage in the body, is relatively rare on Earth and does not even appear as a separate entry in Table **4.5.1**; carbon is part of the 0.174% representing "other" elements. How does the human body concentrate so many apparently rare elements?

Table **4.5.2**: Elemental Composition of a Human Body

| Element | Percentage by Mass |
|---|---|
| oxygen | 61 |
| carbon | 23 |
| hydrogen | 10 |
| nitrogen | 2.6 |
| calcium | 1.4 |
| phosphorus | 1.1 |
| sulfur | 0.20 |
| potassium | 0.20 |
| sodium | 0.14 |
| chlorine | 0.12 |
| magnesium | 0.027 |
| silicon | 0.026 |
| iron | 0.006 |
| fluorine | 0.0037 |
| zinc | 0.0033 |
| all others | 0.174 |

Source: D. R. Lide, ed. *CRC Handbook of Chemistry and Physics*, 89th ed. (Boca Raton, FL: CRC Press, 2008–9), 7–24.

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

### LOOKING CLOSER: PHOSPHOROUS, THE CHEMICAL BOTTLENECK

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

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

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

## Names and Symbols

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

> *Today, new elements are usually named after famous scientists.*

The names of the elements can be cumbersome to write in full, especially when combined to form the names of compounds. Therefore, each element name is abbreviated as a one- or two-letter chemical symbol. By convention, the first letter of a chemical symbol is a capital letter, while the second letter (if there is one) is a lowercase letter. The first letter of the symbol is usually the first letter of the element's name, while the second letter is some other letter from the name. Some elements have symbols that derive from earlier, mostly Latin names, so the symbols may not contain any letters from the English name. Table **4.5.3** lists the names and symbols of some of the most familiar elements.

Table **4.5.3**: Element Names and Symbols

| | | | |
|---|---|---|---|
| aluminum | Al | magnesium | Mg |
| argon | Ar | manganese | Mn |
| arsenic | As | mercury | Hg* |
| barium | Ba | neon | Ne |
| bismuth | Bi | nickel | Ni |
| boron | B | nitrogen | N |
| bromine | Br | oxygen | O |
| calcium | Ca | phosphorus | P |
| carbon | C | platinum | Pt |
| chlorine | Cl | potassium | K* |
| chromium | Cr | silicon | Si |
| copper | Cu* | silver | Ag* |
| fluorine | F | sodium | Na* |
| gold | Au* | strontium | Sr |
| helium | He | sulfur | S |
| hydrogen | H | tin | Sn* |
| iron | Fe | tungsten | W† |
| iodine | I | uranium | U |
| lead | Pb* | zinc | Zn |
| lithium | Li | zirconium | Zr |

*The symbol comes from the Latin name of element.

†The symbol for tungsten comes from its German name—*wolfram*.

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

### Example 4.5.1

Write the chemical symbol for each element without consulting Table **4.5.3** "Element Names and Symbols".

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

**Answer a**

Br

**Answer b**

B

**Answer c**

C

**Answer d**

Ca

**Answer e**

Au

## Exercise 4.5.1

Write the chemical symbol for each element without consulting Table 4.5.3.

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

**Answer a**

Mn

**Answer b**

Mg

**Answer c**

Ne

**Answer d**

N

**Answer e**

Ag

## Example 4.5.2

What element is represented by each chemical symbol?

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

**Answer a**

sodium

**Answer b**

mercury

**Answer c**

phosphorus

**Answer d**

potassium

**Answer e**

iodine

**? Exercise 4.5.2**

What element is represented by each chemical symbol?

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

**Answer a**

lead

**Answer b**

tin

**Answer c**

uranium

**Answer d**

oxygen

**Answer e**

fluorine

## Concept Review Exercises

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

## Answers

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

## Key Takeaways

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

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# 4.6: 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 **4.6.1**. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.

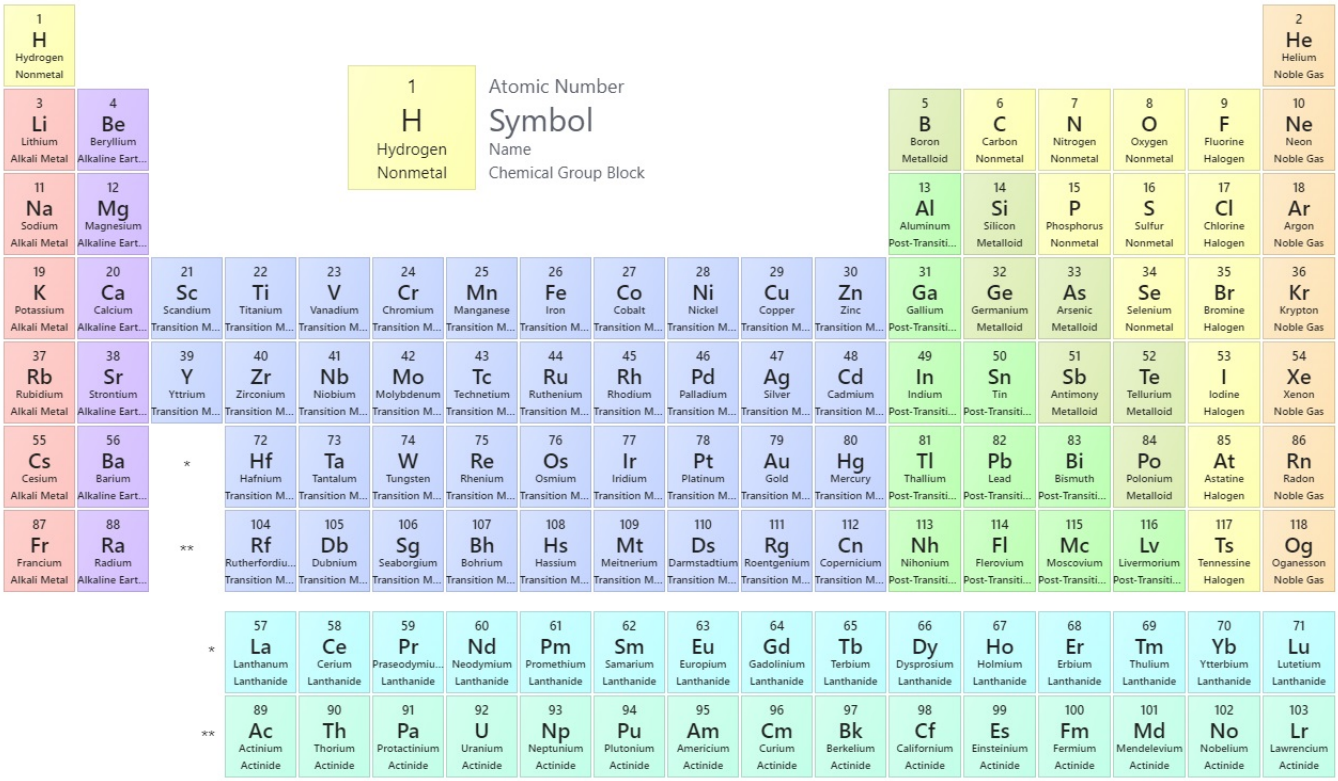



Figure **4.6.1**: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found Periodic Table of Elements. (Public Domain; PubChem via NIH)

## 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 **4.6.2**. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure **4.6.2**, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

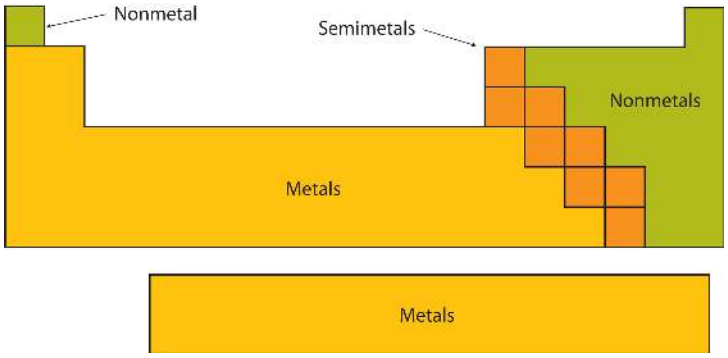



Figure **4.6.2**: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

**Example 4.6.1**

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

a. Se
b. Mg
c. Ge

Solution

a. In Figure **4.6.1**, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

**Exercise 4.6.1**

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

**Answer**

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure **4.6.3**. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

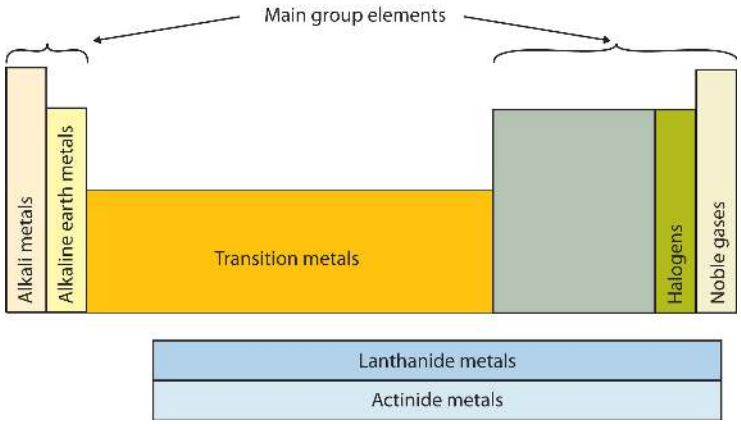



Figure **4.6.3**: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

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 4.6.2: Groups**

Provide the family or group name of each element.

a. Li
b. Ar
c. Cl

Solution

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

## ? Exercise 4.6.2: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

**Answer a:**

Fluorine is a halogen (Group 17).

**Answer b:**

Calcium is a alkaline earth metal (Group 2).

**Answer c:**

Krypton is a noble gas (Group 18).

## ✔ Example 4.6.3: Classification of Elements

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

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

Solution

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

## ? Exercise 4.6.3: Classification of Elements

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

a. F
b. U
c. Cu

**Answer a:**

Fluorine is a nonmetal.

**Answer b:**

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

**Answer c:**

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

## Summary

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

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

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# 4.7: 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 4.7.1 is a graphical depiction of this process.

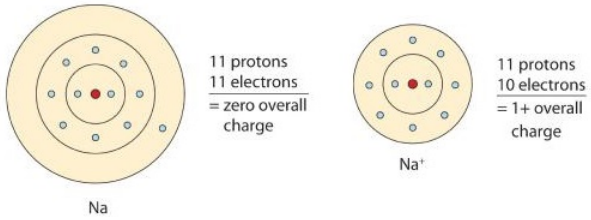



Figure 4.7.1: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

## 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 4.7.2 is a graphical depiction of this process.

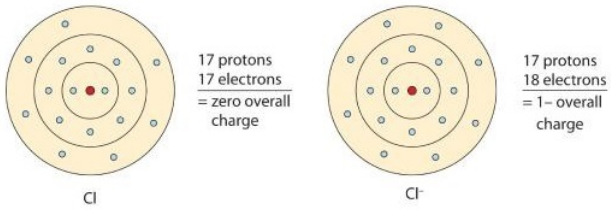



Figure **4.7.2**: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge. Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a -1 overall 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 **4.7.3** shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a ion with multiple charges. The barium cation is written $Ba^{2+}$, not $Ba^{+2}$.

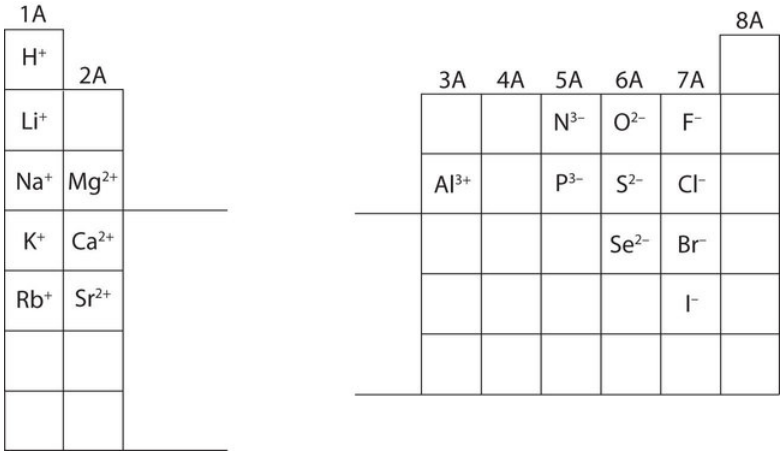



Figure **4.7.3**: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

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# 4.8: Isotopes - When the Number of Neutrons Varies

**Learning Objectives**

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

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

## An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure **4.8.1**. Most hydrogen atoms have just one proton, one electron, and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

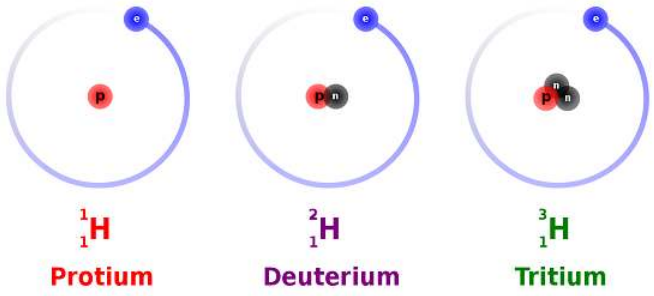



Figure **4.8.1**: The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh via Wikipedia).

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

**Example 4.8.1: Lithium Isotopes**

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

### Solution

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

a.

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

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

b.

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

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

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

## Stability of Isotopes

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

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

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

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

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




Nuclear symbol for helium-4: The element symbol is He, the mass number to the top left is 4, and the atomic number to the bottom left is 2

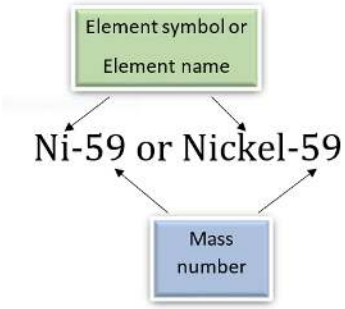
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

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

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

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

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




### Example 4.8.2: Potassium-40

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

**Solution**

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

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

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

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

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

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

### Example 4.8.3: Zinc-65

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

Solution

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

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

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

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

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

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

### Exercise 4.8.3

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

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

**Answer a:**

27 protons, 27 electrons, 33 neutrons

**Answer b:**

11 protons, 11 electrons, 13 neutrons

**Answer c:**

20 protons, 20 electrons, 25 neutrons

**Answer d:**

38 protons, 38 electrons, 52 neutrons

## Summary

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

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# 4.9: Atomic Mass - The Average Mass of an Element's Atoms

**Learning Objectives**

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

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

## Calculating Atomic Mass

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

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

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

**Example 4.9.1: Boron Isotopes**

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

### Solution

Boron has two isotopes. We will use the equation:

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

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

Substitute these into the equation, and we get:

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

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

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

**Example 4.9.2: Neon Isotopes**

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

### Solution

Neon has three isotopes. We will use the equation:

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

- Isotope 1: $\%_1 = 0.9092$ (write all percentages as decimals), $\text{mass}_1 = 19.99$

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

Substitute these into the equation, and we get:

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

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

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

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that the atomic mass of boron (symbol **B**) is 10.8, which is what we calculated in Example **4.9.1**, and the atomic mass of neon (symbol **Ne**) is 20.8, which is what we calculated in Example **4.9.2**. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

### ? Exercise 4.9.1

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

**Answer**

35.45 amu

## Summary

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

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# 4.10: 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).

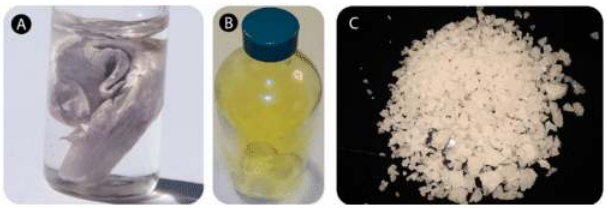



Figure **4.10.1**: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine.

## 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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# 4.11: 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:

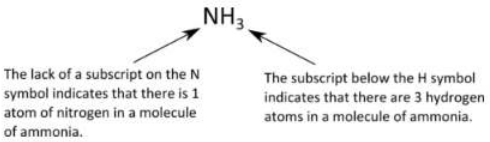



Figure 4.11.1: The molecular formula for ammonia. NH3. There is one atom of nitrogen and 3 atoms of hydrogen in a molecule of ammonia.

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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# 4.12: A Molecular View of Elements and Compounds

**Learning Objectives**

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

## Atomic Elements

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

## Molecular Elements

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

Table **4.12.1**: Elements That Exist as Diatomic Molecules

| | | | | | | |
|---|---|---|---|---|---|---|
| Hydrogen, H | Oxygen | Nitrogen | Fluorine | Chlorine | Bromine | Iodine |

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules—a molecule with only two atoms (Table **4.12.1**). As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is $H_2$, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: $O_2$, $N_2$, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, $S_8$, while phosphorus exists as a four-atom molecule, $P_4$ (Figure **4.12.1**).

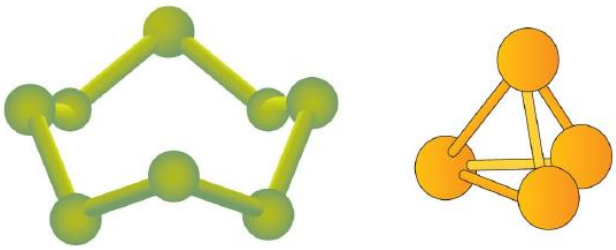


Figure **4.12.1**: Molecular Art of $S_8$ and $P_4$ Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an $S_8$ molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure **4.12.1** 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 4.12.1

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

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

**Solution**

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

### Exercise 4.12.1

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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# CHAPTER OVERVIEW

## 5: Electronic Structure

Atoms act the way that they do because of their structure. We already know that atoms are composed of protons, neutrons, and electrons. Protons and neutrons are located in the nucleus, and electrons orbit around the nucleus. But knowing structural details is key to understanding why atoms react the way they do. Virtually everything known about atoms ultimately comes from light. Before the composition of atoms (especially electrons) can be understood, the properties of light need to be understood.


5.1: Prelude to Electronic Structure
5.2: Light
5.3: Quantum Numbers for Electrons
5.4: Organization of Electrons in Atoms
5.5: Electronic Structure and the Periodic Table
5.6: Periodic Trends
5.E: Electronic Structure (Exercises)


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# 5.1: Prelude to Electronic Structure

Normal light microscopes can magnify objects up to about 1,500 times. Electron microscopes can magnify objects up to 1,000,000 times. Why can electron microscopes magnify images so much? A microscope's resolution depends on the wavelength of light used. The smaller the wavelength, the more a microscope can magnify. Light is a wave, and as such, it has a wavelength associated with it. The wavelength of visible light, which is detected by the eyes, varies from about 700 nm to about 400 nm.

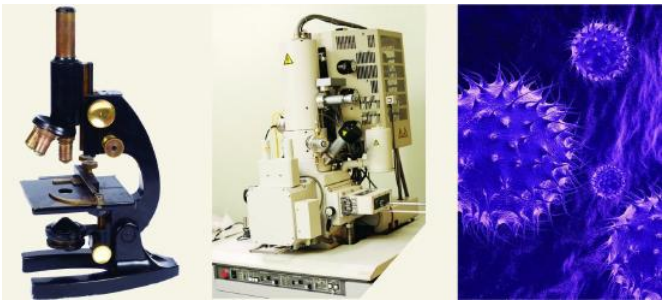


Figure **5.1.1** Microscopes © Thinkstock (a) A simple light microscope can magnify up to 1,500 times. (b) An electron microscope can magnify up to 1,000,000 times. (c) Flu viruses imaged by an electron microscope. The virus is about 100 nm in diameter.

A startling conclusions of modern science is that electrons also act as waves. However, the wavelength of electrons is much, much shorter—about 0.5 to 1 nm. This allows electron microscopes to magnify 600–700 times more than light microscopes. This allows us to see even smaller features in a world that is invisible to the naked eye.

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# 5.2: Light

**Learning Objectives**

- Describe light with its frequency and wavelength.
- Describe light as a particle of energy.

What is known as light is more properly called *electromagnetic radiation*. We know from experiments that light acts as a wave. As such, it can be described as having a frequency and a wavelength. The **wavelength** of light is the distance between corresponding points in two adjacent light cycles, and the **frequency** of light is the number of cycles of light that pass a given point in one second. Wavelength is typically represented by λ, the lowercase Greek letter *lambda*, while frequency is represented by ν, the lowercase Greek letter *nu* (although it looks like a Roman "vee," it is actually the Greek equivalent of the letter "en"). Wavelength has units of length (meters, centimeters, etc.), while frequency has units of *per second*, written as $s^{-1}$ and sometimes called a *hertz* (Hz). Figure 5.2.1 - Characteristics of Light Waves, shows how these two characteristics are defined.

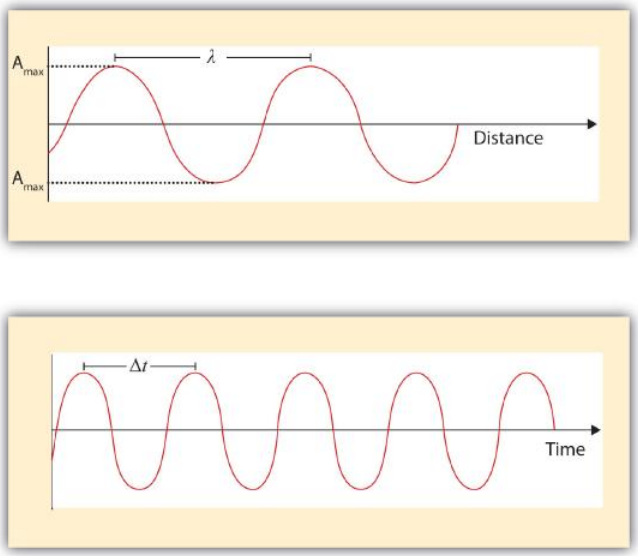



Figure 5.2.1: Characteristics of Light Waves. Light acts as a wave and can be described by a wavelength λ and a frequency ν.

One property of waves is that their speed is equal to their wavelength times their frequency. That means we have

$$speed = \lambda\nu$$

For light, however, speed is actually a universal constant when light is traveling through a vacuum (or, to a very good approximation, air). The measured speed of light ($c$) in a vacuum is $2.9979 \times 10^8$ m/s, or about $3.00 \times 10^8$ m/s. Thus, we have

$$c = \lambda\nu$$

Because the speed of light is a constant, the wavelength and the frequency of light are related to each other: as one increases, the other decreases and vice versa. We can use this equation to calculate what one property of light has to be when given the other property.

**Example 5.2.1**

What is the frequency of light if its wavelength is $5.55 \times 10^{-7}$ m?

**Solution**

We use the equation that relates the wavelength and frequency of light with its speed. We have

$$3.00 \times 10^8 m/s = \left(5.55 \times 10^{-7} m\right)\nu$$

We divide both sides of the equation by $5.55 \times 10^{-7}$ m and get

$$\nu = 5.41 \times 10^{14} s^{-1}$$

Note how the m units cancel, leaving s in the denominator. A unit in a denominator is indicated by a −1 power, $s^{-1}$, and read as "per second."

**? Exercise 5.2.1**

What is the wavelength of light if its frequency is $1.55 \times 10^{10}$ $s^{-1}$?

***Answer***

0.0194 m, or 19.4 mm

Light also behaves like a package of energy. It turns out that for light, the energy of the "package" of energy is proportional to its frequency. (For most waves, energy is proportional to wave amplitude, or the height of the wave.) The mathematical equation that relates the energy ($E$) of light to its frequency is

$$E = h\nu$$

where ν is the frequency of the light, and $h$ is a constant called Planck's constant. Its value is $6.626 \times 10^{-34}$ J·s—a very small number that is another fundamental constant of our universe, like the speed of light. The units on Planck's constant may look unusual, but these units are required so that the algebra works out.

**✓ Example 5.2.2**

What is the energy of light if its frequency is $1.55 \times 10^{10}$ $s^{-1}$?

**Solution**

Using the formula for the energy of light, we have

$E = (6.626 \times 10^{-34}$ J·s$)(1.55 \times 10^{10}$ $s^{-1})$

Seconds are in the numerator and the denominator, so they cancel, leaving us with joules, the unit of energy. So

$E = 1.03 \times 10^{-23}$ J

This is an extremely small amount of energy—but this is for only one light wave.

**? Exercise 5.2.2**

What is the frequency of a light wave if its energy is $4.156 \times 10^{-20}$ J?

***Answer***

$6.27 \times 10^{13}$ $s^{-1}$

Because a light wave behaves like a little particle of energy, light waves have a particle-type name: the photon. It is not uncommon to hear light described as photons.

Wavelengths, frequencies, and energies of light span a wide range; the entire range of possible values for light is called the **electromagnetic spectrum**. We are mostly familiar with visible light, which is light having a wavelength range between about 400 nm and 700 nm. Light can have much longer and much shorter wavelengths than this, with corresponding variations in frequency and energy. Figure **5.2.2** shows the entire electromagnetic spectrum and how certain regions of the spectrum are labeled. You may already be familiar with some of these regions; they are all light—with different frequencies, wavelengths, and energies.

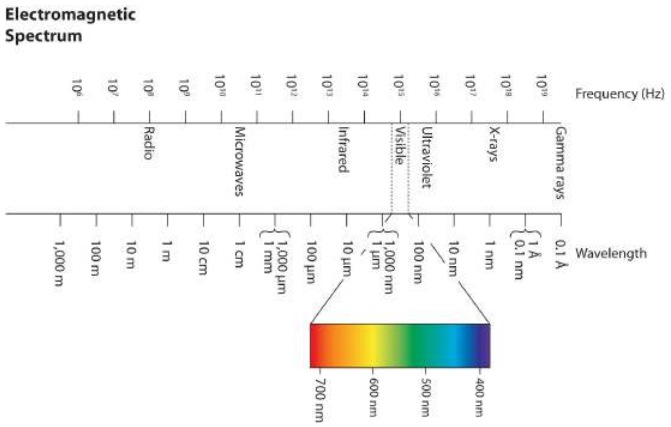



Figure **5.2.2**: The Electromagnetic Spectrum. The electromagnetic spectrum, with its various regions labeled. The borders of each region are approximate.

## Summary

Light acts like a wave, with a frequency and a wavelength. The frequency and wavelength of light are related by the speed of light, a constant. Light acts like a particle of energy, whose value is related to the frequency of light.

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# 5.3: Quantum Numbers for Electrons

**Learning Objectives**

- Explain what spectra are.
- Learn the quantum numbers that are assigned to electrons.

There are two fundamental ways of generating light: either heat an object up to be so hot that it glows, or pass an electrical current through a sample of matter (usually a gas). Incandescent lights and fluorescent lights generate light via these two methods, respectively.

A hot object gives off a continuum of light. We notice this when the visible portion of the electromagnetic spectrum is passed through a prism: the prism separates light into its constituent colors, and all colors are present in a continuous rainbow (Figure **5.3.1*a*** - Prisms and Light). This image is known as a **continuous spectrum**. However, when electricity is passed through a gas and light is emitted and this light is passed though a prism, we see only certain lines of light in the image (Figure **5.3.1*b***). This image is called a line spectrum. It turns out that every element has its own unique, characteristic line spectrum.




Figure **5.3.1**: Prisms and Light. (a) A glowing object gives off a full rainbow of colors, which are noticed only when light is passed through a prism to make a continuous spectrum. (b) However, when electricity is passed through a gas, only certain colors of light are emitted. Here are the colors of light in the line spectrum of Hg.

Why does the light emitted from an electrically excited gas have only certain colors, while light given off by hot objects has a continuous spectrum? For a long time, it was not well explained. Particularly simple was the spectrum of hydrogen gas, which could be described easily by an equation; no other element has a spectrum that is so predictable (Figure **5.3.2** - Hydrogen Spectrum). Late nineteenth century scientists found that the positions of the lines obeyed a pattern given by the equation

$$\frac{1}{\lambda} = (109,700 cm^{-1})\left(\frac{1}{4} - \frac{1}{n^2}\right)$$

where $n$ = 3, 4, 5, 6, etc. However, they could not explain why this was so.

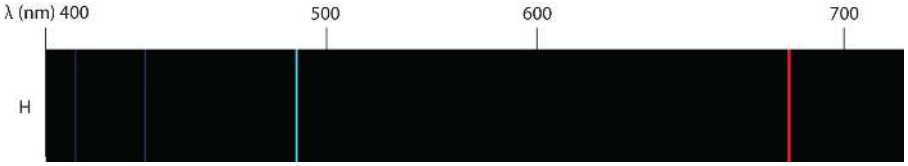



Figure **5.3.2**: Hydrogen Spectrum. The spectrum of hydrogen was particularly simple and could be predicted by a simple mathematical expression.

In 1913, the Danish scientist Niels Bohr suggested a reason why the hydrogen atom spectrum looked this way. He suggested that the electron in a hydrogen atom could not have any random energy, having *only* certain fixed values of energy that were indexed by the number $n$ (the same $n$ in the equation above and now called a **quantum number**). Quantities that have certain specific values are **quantized** values. Bohr suggested that the energy of the electron in hydrogen was quantized because it was in a specific orbit. Because the energies of the electron can have only certain values, the changes in energies can have only certain values (somewhat similar to a staircase—not only are the stair steps set at specific heights, but the height between steps is fixed). Finally, Bohr suggested that the energy of light emitted from electrified hydrogen gas was equal to the energy difference of the electron's energy states:

$$E_{light} = h\nu = \Delta E_{electron}$$

This means that only certain frequencies (and thus, certain wavelengths) of light are emitted. Figure **5.3.3** - Bohr's Model of the Hydrogen Atom, shows a model of the hydrogen atom based on Bohr's ideas.

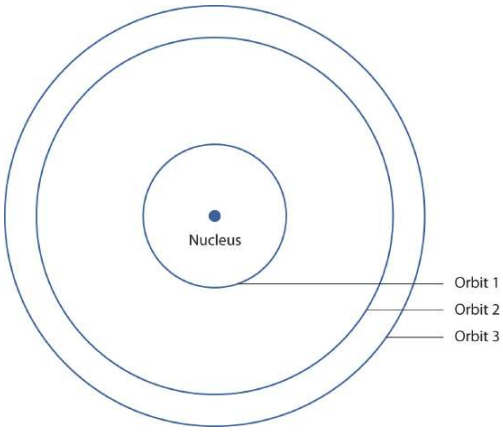



Figure **5.3.3**: Bohr's Model of the Hydrogen Atom. Bohr's description of the hydrogen atom had specific orbits for the electron, which had quantized energies.

Bohr's ideas were useful, but were applicable only to the hydrogen atom. However, later researchers generalized Bohr's ideas into a new theory called **quantum mechanics**, which explains the behavior of electrons as if they were acting as a wave, not as particles. Quantum mechanics predicts two major things: quantized energies for electrons of all atoms (not just hydrogen) and an organization of electrons within atoms. Electrons are no longer thought of as being randomly distributed around a nucleus or restricted to certain orbits (in that regard, Bohr was wrong). Instead, electrons are collected into groups and subgroups that explain much about the chemical behavior of the atom.

In the quantum-mechanical model of an atom, the state of an electron is described by four quantum numbers, not just the one predicted by Bohr. The first quantum number is called the **principal quantum number(n).** The principal quantum number largely determines the energy of an electron. Electrons in the same atom that have the same principal quantum number are said to occupy an electron **shell** of the atom. The principal quantum number can be any nonzero positive integer: 1, 2, 3, 4,….

Within a shell, there may be multiple possible values of the next quantum number, the **angular momentum quantum number (ℓ)**. The ℓ quantum number has a minor effect on the energy of the electron but also affects the spatial distribution of the electron in three-dimensional space—that is, the shape of an electron's distribution in space. The value of the ℓ quantum number can be any integer between 0 and $n - 1$: $\ell = 0, 1, 2, \ldots, n - 1$.

Thus, for a given value of $n$, there are different possible values of ℓ:

Table with two columns and four rows. The first (left) column is labeled "If n equals" and the second (right) column is labeled "ℓ can be". Underneath the columns in the rows are different values. All of the rows have numerical values.

| If *n* equals | ℓ can be |
|---|---|
| 1 | 0 |
| 2 | 0 or 1 |
| 3 | 0, 1, or 2 |
| 4 | 0, 1, 2, or 3 |

and so forth. Electrons within a shell that have the same value of ℓ are said to occupy a **subshell** in the atom. Commonly, instead of referring to the numerical value of ℓ, a letter represents the value of ℓ (to help distinguish it from the principal quantum number):

Table with two columns and four rows. The first (left) column is labeled "If ℓ equals" and the second (right) column is labeled "The letter is". The rows underneath the left column are numbers and the rows underneath the right column are letters.

| If ℓ equals | The letter is |
|---|---|
| 0 | *s* |
| 1 | *p* |
| 2 | *d* |
| 3 | *f* |

The next quantum number is called the **magnetic quantum number ($m_\ell$)**. For any value of ℓ, there are 2ℓ + 1 possible values of $m_\ell$, ranging from −ℓ to ℓ:

$$-\ell < m\ell < \ell$$

or

$$|m_\ell| < \ell$$

The following explicitly lists the possible values of $m_\ell$ for the possible values of ℓ:

Table with two columns and four rows. The first (left) column is labeled "If ℓ equals" and the second (right) column is labeled "The mℓ values can be". The rows underneath the left column are positive numbers and the rows underneath the right column are negative numbers.

| If ℓ equals | The $m_\ell$ values can be |
|---|---|
| 0 | 0 |
| 1 | −1, 0, or 1 |
| 2 | −2, −1, 0, 1, or 2 |
| 3 | −3, −2, −1, 0, 1, 2, or 3 |

The particular value of $m_\ell$ dictates the orientation of an electron's distribution in space. When ℓ is zero, $m_\ell$ can be only zero, so there is only one possible orientation. When ℓ is 1, there are three possible orientations for an electron's distribution. When ℓ is 2, there are five possible orientations of electron distribution. This goes on and on for other values of ℓ, but we need not consider any higher values of ℓ here. Each value of $m_\ell$ designates a certain orbital. Thus, there is only one orbital when ℓ is zero, three orbitals when ℓ is 1, five orbitals when ℓ is 2, and so forth. The $m_\ell$ quantum number has no effect on the energy of an electron unless the electrons are subjected to a magnetic field—hence its name.

The ℓ quantum number dictates the general shape of electron distribution in space (Figure **5.3.4** - Electron Orbitals). Any s orbital is spherically symmetric (Figure **5.3.4*a*** - Electron Orbitals), and there is only one orbital in any s subshell. Any p orbital has a two-lobed, dumbbell-like shape (Figure **5.3.4*b*** - Electron Orbitals); because there are three of them, we normally represent them as pointing along the x-, y-, and z-axes of Cartesian space. The d orbitals are four-lobed rosettes (Figure **5.3.4*c*** - Electron Orbitals) and they are oriented differently in space (the one labeled Math input error has two lobes and a torus instead of four lobes, but it is equivalent to the other orbitals). When there is more than one possible value of $m_\ell$, each orbital is labeled with one of the possible values. It should be noted that the diagrams in Figure **5.3.4** are estimates of the electron distribution in space, not surfaces electrons are fixed on.

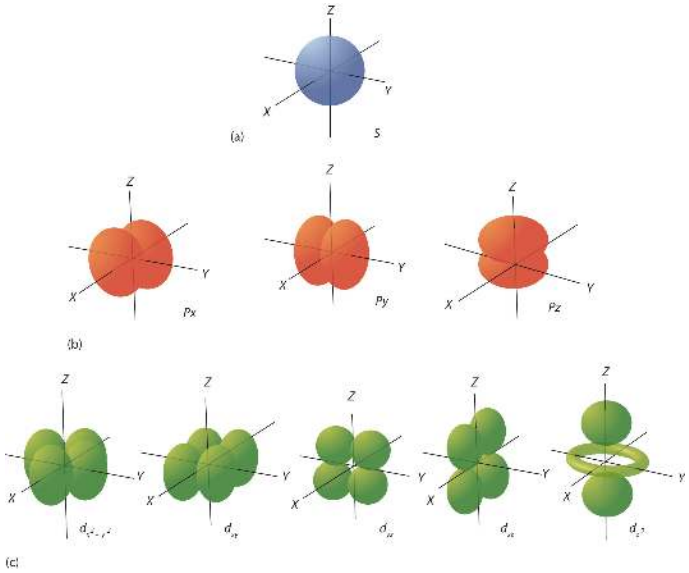


Figure **5.3.4**: Electron Orbitals. (a) The lone s orbital is spherical in distribution. (b) The three p orbitals are shaped like dumbbells, and each one points in a different direction. (c) The five d orbitals are rosette in shape, except for the dz2 orbital, which is a "dumbbell + torus" combination. They are all oriented in different directions.

The final quantum number is the **spin quantum number** ($m_S$). Electrons and other subatomic particles behave as if they are spinning (we cannot tell if they really are, but they behave as if they are). Electrons themselves have two possible spin states, and because of mathematics, they are assigned the quantum numbers +1/2 and −1/2. These are the only two possible choices for the spin quantum number of an electron.

### ✔ Example 5.3.1

Of the set of quantum numbers $\{n, \ell, m\ell, m_S\}$, which are possible and which are not allowed?

a. {3, 2, 1, +1/2}
b. {2, 2, 0, −1/2}
c. {3, −1, 0, +1/2}

**Solution**

a. The principal quantum number $n$ must be an integer, as it is here. The quantum number $\ell$ must be less than $n$, which it is. The $m\ell$ quantum number must be between $-\ell$ and $\ell$, which it is. The spin quantum number is +1/2, which is allowed. Because this set of quantum numbers follows all restrictions, it is possible.
b. The quantum number $n$ is an integer, but the quantum number $\ell$ must be less than $n$, which it is not. Thus, this is not an allowed set of quantum numbers.
c. The principal quantum number $n$ is an integer, but $\ell$ is not allowed to be negative. Therefore, this is not an allowed set of quantum numbers.

### ? Exercise 5.3.1

Of the set of quantum numbers $\{n, \ell, m\ell, m_S\}$, which are possible and which are not allowed?

a. {4, 2, −2, 1}
b. {3, 1, 0, −1/2}

***Answers***

a. Spin must be either +1/2 or −1/2, so this set of quantum number is not allowed.
b. allowed

## Chemistry is Everywhere: Neon Lights

A neon light is basically an electrified tube with a small amount of gas in it. Electricity excites electrons in the gas atoms, which then give off light as the electrons go back into a lower energy state. However, many so-called "neon" lights do not contain neon!

Although we know now that a gas discharge gives off only certain colors of light, without a prism or other component to separate the individual light colors, we see a composite of all the colors emitted. It is not unusual for a certain color to predominate. True neon lights, with neon gas in them, have a reddish-orange light due to the large amount of red-, orange-, and yellow-colored light emitted. However, if krypton is used instead of neon, a whitish light is emitted, while using argon yields a blue-purple light. A light filled with nitrogen gas glows purple, as does a helium lamp. Other gases—and mixtures of gases—emit other colors of light. Ironically, despite its importance in the development of modern electronic theory, hydrogen lamps emit little visible light and are rarely used for illumination purposes.




The Human/Need/Desire statue consists of purple "NEED", orange "HUMAN", green "HUMAN", pink "DESIRE", red "HOPE", and blue "DREAM".

***Neon lights. The different colors of these "neon" lights are caused by gases other than neon in the discharge tubes. Human/Need/Desire. Neon sculpture by Bruce Nauman (1983), who has been characterized as a conceptual artist. (CC BY-SA 3.0; eschipul).***

## Key Takeaways

Electrons in atoms have quantized energies. The state of electrons in atoms is described by four quantum numbers.

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# 5.4: Organization of Electrons in Atoms

**Learning Objectives**

- Learn how electrons are organized within atoms.
- Represent the organization of electrons by an electron configuration.

Now that you know that electrons have quantum numbers, how are they arranged in atoms? The key to understanding electronic arrangement is summarized in the **Pauli exclusion principle:** no two electrons in an atom can have the same set of four quantum numbers. This dramatically limits the number of electrons that can exist in a shell or a subshell.

Electrons are typically organized around an atom by starting at the lowest possible quantum numbers first, which are the shells-subshells with lower energies. Consider H, an atom with a single electron only. Under normal conditions, the single electron would go into the n = 1 shell, which has only a single s subshell with one orbital (because $m_\ell$ can equal only 0). The convention is to label the shell-subshell combination with the number of the shell and the letter that represents the subshell. Thus, the electron goes in the 1s shell-subshell combination. It is usually not necessary to specify the $m_\ell$ or $m_s$ quantum numbers, but for the H atom, the electron has $m_\ell = 0$ (the only possible value) and an $m_s$ of either +1/2 or −1/2.

The He atom has two electrons. The second electron can also go into the 1s shell-subshell combination, but only if its spin quantum number is different from the first electron's spin quantum number. Thus, the sets of quantum numbers for the two electrons are {1, 0, 0, +1/2} and {1, 0, 0, −1/2}. Notice that the overall set is different for the two electrons, as required by the Pauli exclusion principle.

The next atom is Li, with three electrons. However, now the Pauli exclusion principle implies that we cannot put that electron in the 1s shell-subshell because no matter how we try, this third electron would have the same set of four quantum numbers as one of the first two electrons. So this third electron must be assigned to a different shell-subshell combination. However, the n = 1 shell doesn't have another subshell; it is restricted to having just $\ell = 0$, or an s subshell. Therefore, this third electron has to be assigned to the n = 2 shell, which has an s ($\ell = 0$) subshell and a p ($\ell = 1$) subshell. Again, we usually start with the lowest quantum number, so this third electron is assigned to the 2s shell-subshell combination of quantum numbers.

The Pauli exclusion principle has the net effect of limiting the number of electrons that can be assigned a shell-subshell combination of quantum numbers. For example, in any s subshell, no matter what the shell number, there can be a maximum of only two electrons. Once the s subshell is filled up, any additional electrons must go to another subshell in the shell (if it exists), or to a higher-numbered shell. A similar analysis shows that a p subshell can hold a maximum of six electrons. A d subshell can hold a maximum of 10 electrons, while an f subshell can have a maximum of 14 electrons. By limiting subshells to these maxima, we can distribute the available electrons to their shells and subshells.

**Example 5.4.1: Carbon Atoms**

How would the six electrons for C be assigned to the n and $\ell$ quantum numbers?

**Solution**

The first two electrons go into the 1s shell-subshell combination. Two additional electrons can go into the 2s shell-subshell, but now this subshell is filled with the maximum number of electrons. The n = 2 shell also has a p subshell, so the remaining two electrons can go into the 2p subshell. The 2p subshell is not completely full because it can hold a maximum of six electrons.

**Exercise 5.4.1: Sodium Atoms**

How would the 11 electrons for Na be assigned to the n and $\ell$ quantum numbers?

**Answer**

two 1s electrons, two 2s electrons, six 2p electrons, and one 3s electron

Now that we see how electrons are partitioned among the shells and subshells, we need a more concise way of communicating this partitioning. Chemists use an **electron configuration** to represent the organization of electrons in shells and subshells in an atom. An electron configuration simply lists the shell and subshell labels, with a right superscript giving the number of electrons in that subshell. The shells and subshells are listed in the order of filling.

For example, an H atom has a single electron in the 1s subshell. Its electron configuration is

$$\text{H: } 1s^1$$

The He atom has two electrons in the 1s subshell. Its electron configuration is

$$\text{He: } 1s^2$$

The three electrons for Li are arranged in the 1s subshell (two electrons) and the 2s subshell (one electron). The electron configuration of Li is

$$\text{Li: } 1s^2 2s^1$$

The Be atom has four electrons, two in the 1s subshell and two in the 2s subshell. Its electron configuration is

$$\text{Be: } 1s^2 2s^2$$

Now that the 2s subshell is filled, electrons in larger atoms must go into the 2p subshell, which can hold a maximum of six electrons. The next six elements progressively fill up the 2p subshell:

- B: $1s^2 2s^2 2p^1$
- C: $1s^2 2s^2 2p^2$
- N: $1s^2 2s^2 2p^3$
- O: $1s^2 2s^2 2p^4$
- F: $1s^2 2s^2 2p^5$
- Ne: $1s^2 2s^2 2p^6$

Now that the 2p subshell is filled (all possible subshells in the n = 2 shell), the next electron for the next-larger atom must go into the n = 3 shell, s subshell.

### Example 5.4.2: Sodium

What is the electron configuration for Na, which has 11 electrons?

**Solution**

The first two electrons occupy the 1s subshell. The next two occupy the 2s subshell, while the next six electrons occupy the 2p subshell. This gives us 10 electrons so far, with 1 electron left. This last electron goes into the n = 3 shell, s subshell. Thus, the electron configuration of Na is $1s^2 2s^2 2p^6 3s^1$.

### Exercise 5.4.2: Magnesium

What is the electron configuration for Mg, which has 12 electrons?

**Answer**

$1s^2 2s^2 2p^6 3s^2$

For larger atoms, the electron arrangement becomes more complicated. This is because after the 3p subshell is filled, filling the 4s subshell first actually leads to a lesser overall energy than filling the 3d subshell. Recall that while the principal quantum number largely dictates the energy of an electron, the angular momentum quantum number also has an impact on energy; by the time we get to the 3d and 4s subshells, we see overlap in the filling of the shells. Thus, after the 3p subshell is completely filled (which occurs for Ar), the next electron for K occupies the 4s subshell, not the 3d subshell:

$$\text{K: } 1s^2 2s^2 2p^6 3s^2 3p^6 4s^1,$$

not

$$\text{K: } 1s^2 2s^2 2p^6 3s^2 3p^6 3d^1$$

For larger and larger atoms, the order of filling the shells and subshells seems to become even more complicated. There are some useful ways to remember the order, like that shown in Figure **5.4.1**. If you follow the arrows in order, they pass through the subshells in the order that they are filled with electrons in larger atoms. Initially, the order is the same as the expected shell-subshell order, but for larger atoms, there is some shifting around of the principal quantum numbers. However, Figure **5.4.1** gives a valid ordering of filling subshells with electrons for most atoms.

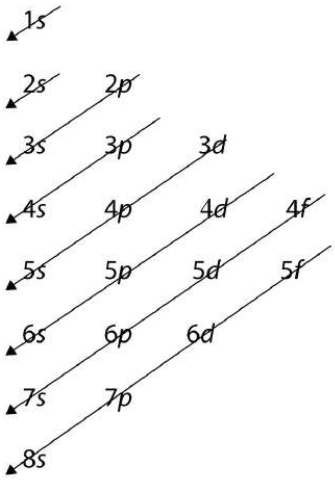



Figure **5.4.1** Electron Shell Filling Order. Starting with the top arrow, follow each arrow. The subshells you reach along each arrow give the ordering of filling of subshells in larger atoms. The n = 5 and higher shells have more subshells, but only those subshells that are needed to accommodate the electrons of the known elements are given.

The electron shell filling order goes as follows: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, then 8s.

### ✔ Example **5.4.3**: Tin

What is the predicted electron configuration for Sn, which has 50 electrons?

**Solution**

We will follow the chart in Figure 8.4., until we can accommodate 50 electrons in the subshells in the proper order: Sn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^2$

Verify by adding the superscripts, which indicate the number of electrons: 2 + 2 + 6 + 2 + 6 + 2 + 10 + 6 + 2 + 10 + 2 = 50, so we have placed all 50 electrons in subshells in the proper order.

### ? Exercise **5.4.3**: Barium

What is the electron configuration for Ba, which has 56 electrons?

**Answer**

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

As the previous example demonstrated, electron configurations can get fairly long. An **abbreviated electron configuration** uses one of the elements from the last column of the periodic table, which contains the noble gases, to represent the core of electrons up to that element. Then, the remaining electrons are listed explicitly. For example, the abbreviated electron configuration for Li, which has three electrons, would be

$$\text{Li: } [\text{He}]2s^1$$

where [He] represents the two-electron core that is equivalent to He's electron configuration. The square brackets represent the electron configuration of a noble gas. This is not much of an abbreviation. However, consider the abbreviated electron configuration for W, which has 74 electrons:

$$\text{W: [Xe]}6s^2 4f^{14} 5d^4$$

This is a significant simplification over an explicit listing of all 74 electrons. So for larger elements, the abbreviated electron configuration can be a very useful shorthand.

### Example 5.4.4: Phosphorus

What is the abbreviated electron configuration for P, which has 15 electrons?

**Solution**

With 15 electrons, the electron configuration of P is

$$\text{P: } 1s^2 2s^2 2p^6 3s^2 3p^3$$

The first immediate noble gas is Ne, which has an electron configuration of $1s^2 2s^2 2p^6$. Using the electron configuration of Ne to represent the first 10 electrons, the abbreviated electron configuration of P is

$$\text{P: [Ne]}3s^2 3p^3$$

### Exercise 5.4.4: Rubidium

What is the abbreviated electron configuration for Rb, which has 37 electrons?

**Answer**

$\text{[Kr]}5s^1$

There are some exceptions to the rigorous filling of subshells by electrons. In many cases, an electron goes from a higher-numbered shell to a lower-numbered, but later-filled, subshell to fill the later-filled subshell. One example is Ag. With 47 electrons, its electron configuration is predicted to be

$$\text{Ag: [Kr]}5s^2 4d^9$$

However, experiments have shown that the electron configuration is actually

$$\text{Ag: [Kr]}5s^1 4d^{10}$$

This, then, qualifies as an exception to our expectations. At this point, you do not need to memorize the exceptions; but if you come across one, understand that it is an exception to the normal rules of filling subshells with electrons, which can happen.

## Summary

The Pauli exclusion principle limits the number of electrons in the subshells and shells. Electrons in larger atoms fill shells and subshells in a regular pattern that can be followed. Electron configurations are a shorthand method of indicating what subshells electrons occupy in atoms. Abbreviated electron configurations are a simpler way of representing electron configurations for larger atoms. Exceptions to the strict filling of subshells with electrons do occur.

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# 5.5: Electronic Structure and the Periodic Table

**Learning Objectives**

- Relate the electron configurations of the elements to the shape of the periodic table.
- Determine the expected electron configuration of an element by its place on the periodic table.

In Chapter 3, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure **5.5.1**. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.

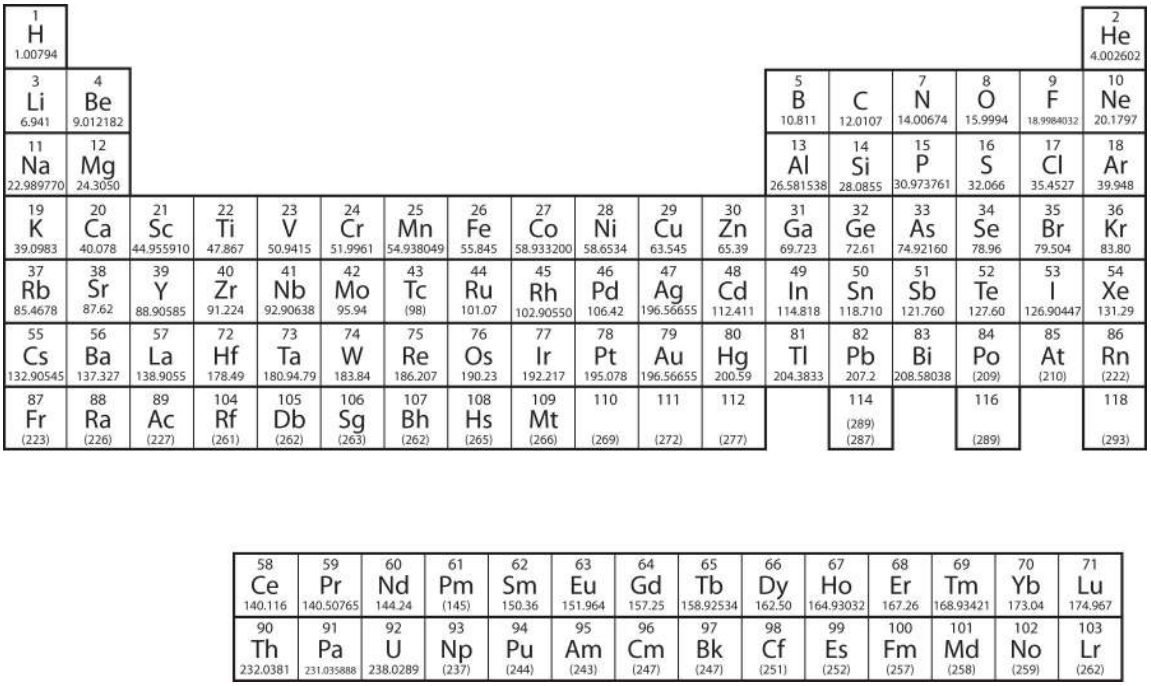



Figure **5.5.1**: The Periodic Table

Why does the periodic table have the structure it does? The answer is rather simple, if you understand electron configurations, the shape of the periodic table mimics the filling of the subshells with electrons.

> *The shape of the periodic table mimics the filling of the subshells with electrons.*

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (Figure **5.5.2**)

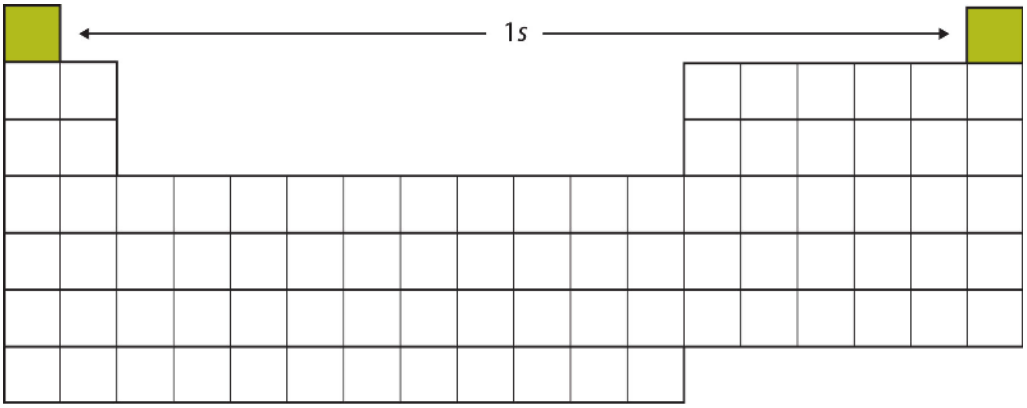



Figure 5.5.2: The 1s Subshell. H and He represent the filling of the 1s subshell.

The next two electrons, for Li and Be, would go into the 2*s* subshell. Figure 5.5.3 shows that these two elements are adjacent on the periodic table.

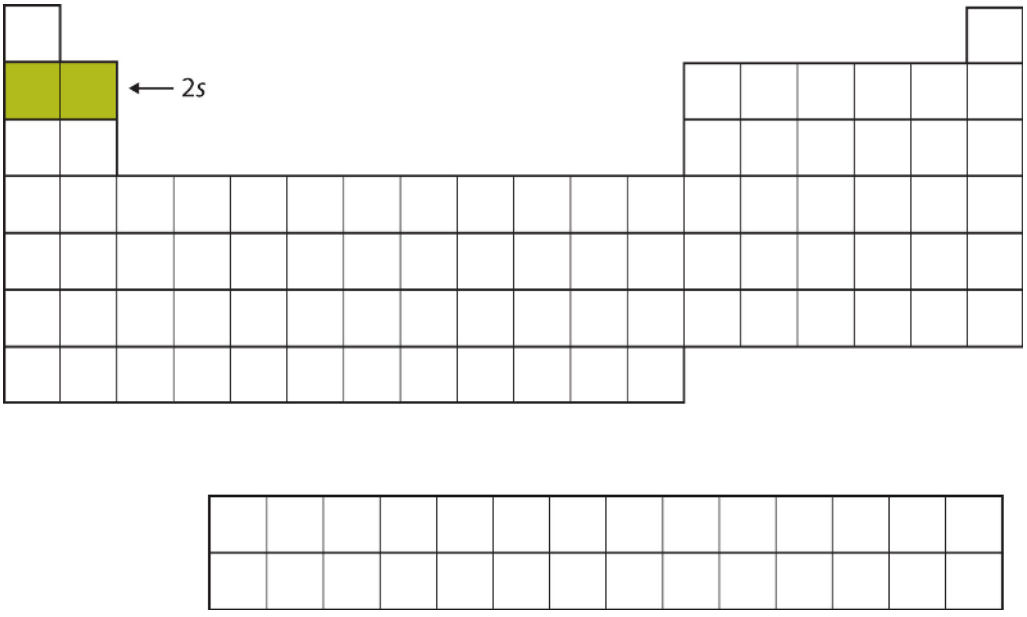



Figure 5.5.3: The 2s Subshell. In Li and Be, the 2s subshell is being filled.

For the next six elements, the 2*p* subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure 5.5.4).

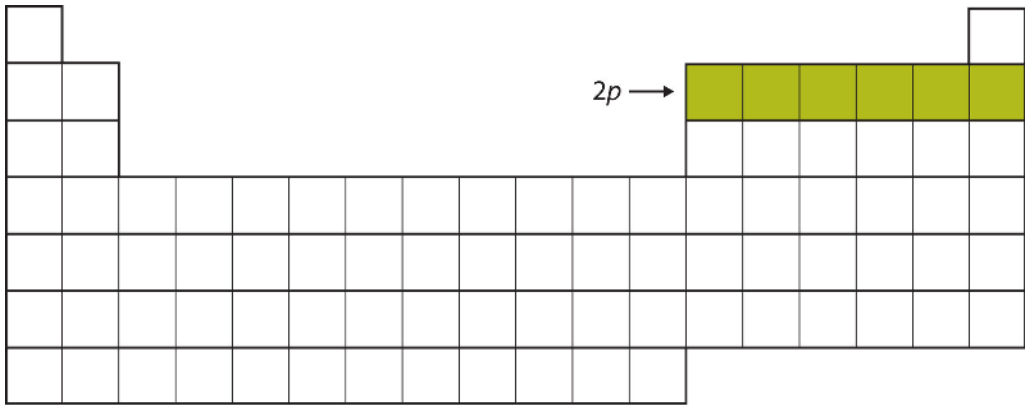



Figure 5.5.4: The 2p Subshell. For B through Ne, the 2p subshell is being occupied.

The next subshell to be filled is the 3*s* subshell. The elements coinciding this subshell being filled, Na and Mg, are back on the left side of the periodic table (Figure 5.5.5).

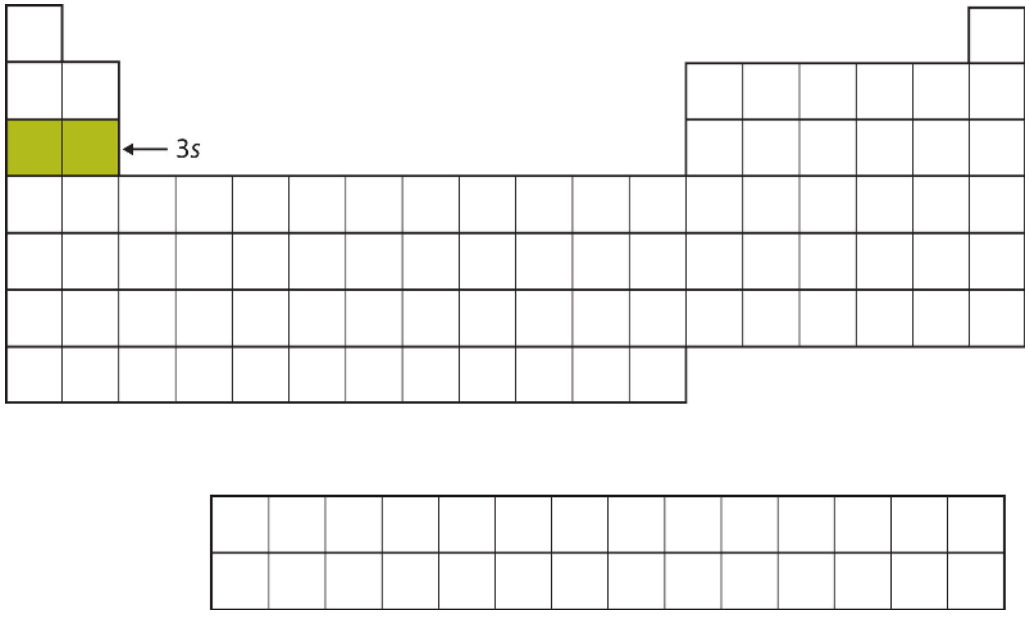



Figure 5.5.5: The 3s Subshell. Now the 3s subshell is being occupied.

Next, the 3*p* subshell is filled with the next six elements (Figure 5.5.6).

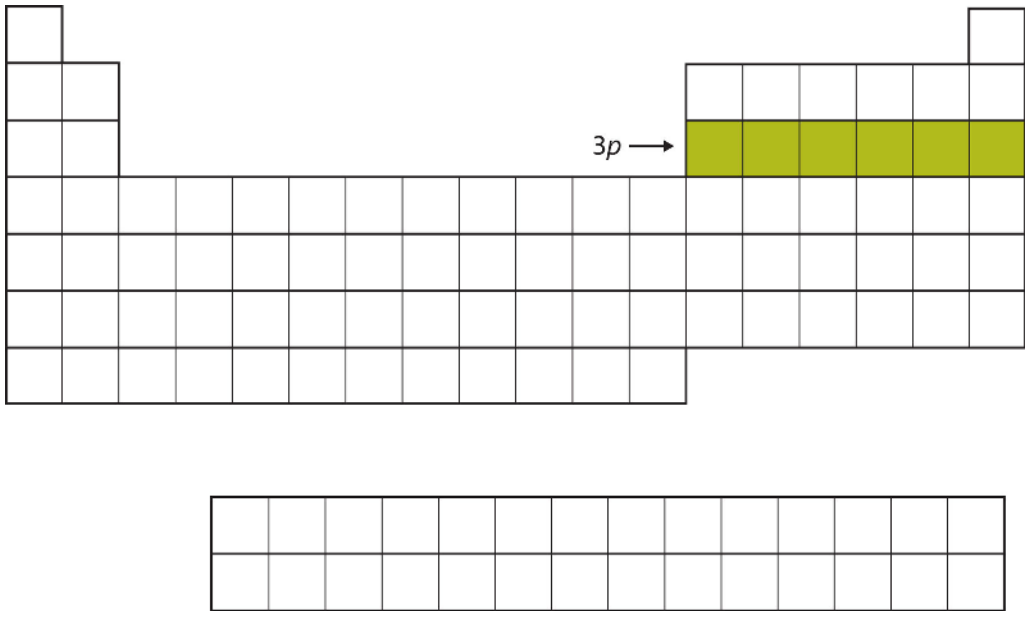



Figure8.5.6: The 3p Subshell. Next, the 3p subshell is filled with electrons.

Instead of filling the 3*d* subshell next, electrons go into the 4*s* subshell (Figure **5.5.7**).

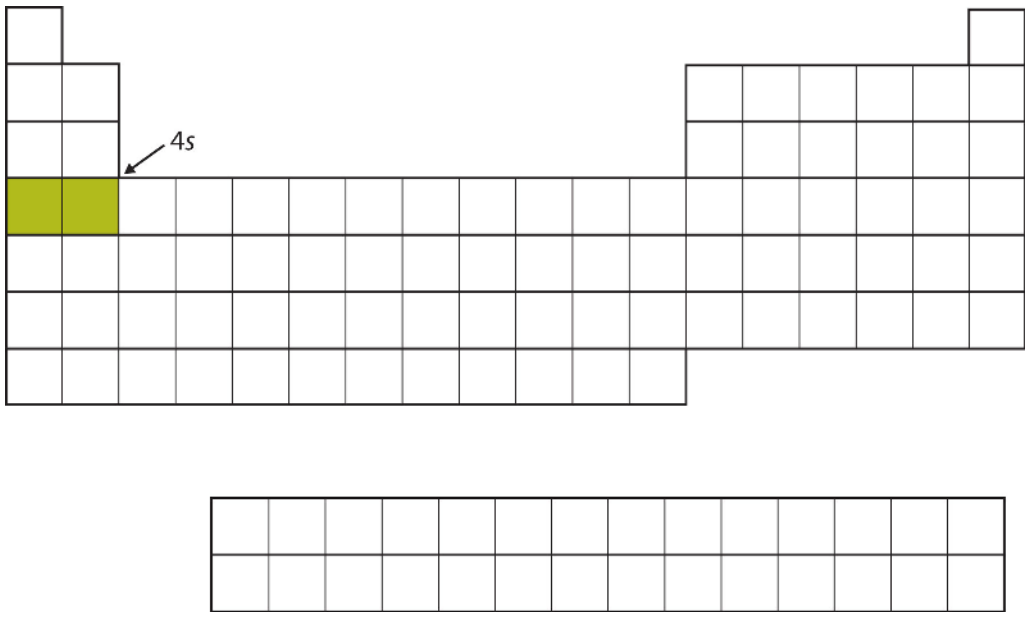



Figure **5.5.7**: The 4s Subshell. The 4s subshell is filled before the 3d subshell. This is reflected in the structure of the periodic table.

After the 4*s* subshell is filled, the 3*d* subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure **5.5.8**).

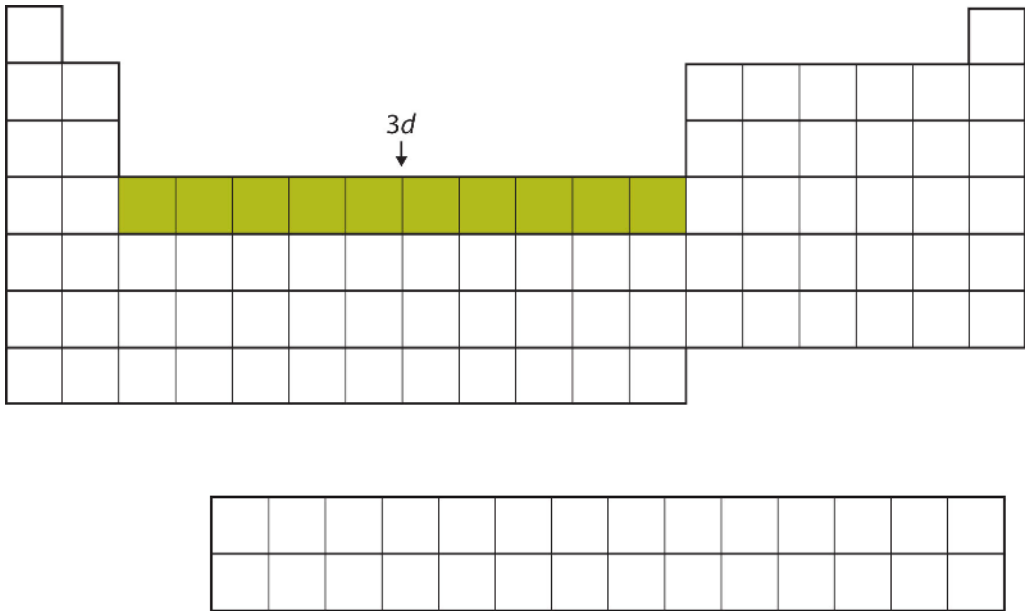



Figure 5.5.8: The 3d Subshell. The 3d subshell is filled in the middle section of the periodic table.

And so forth. As we go across the columns of the periodic table, the overall shape of the table outlines how the electrons are occupying the shells and subshells.

The first two columns on the left side of the periodic table are where the *s* subshells are being occupied. Because of this, the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** is the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure 5.5.9 shows the blocks of the periodic table.

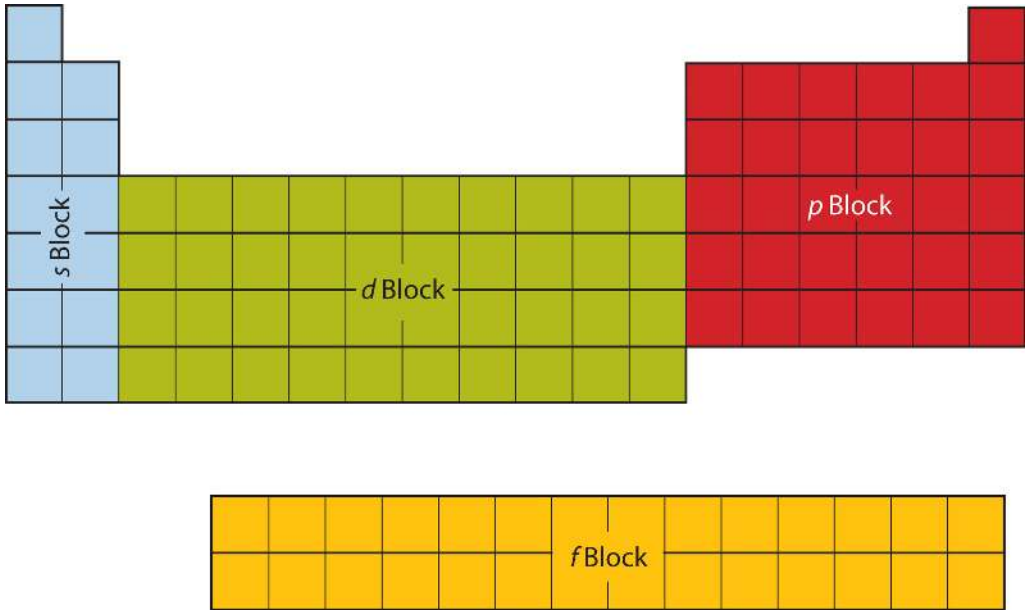



Figure 5.5.9: Blocks on the Periodic Table. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

Empty periodic table, separated into 4 different colored blocks. S block is blue, d block is green, p block is red and f block is orange.

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called *core electrons*.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column, the

valence shell's electron configuration is the same. For example, take the elements in the first column of the periodic table: H, Li, Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

Table with two columns and 6 rows. The first column on the left has various elements in the rows underneath. The second column on the right has the different corresponding electron configurations for the specified element in the rows underneath.

| | |
|---|---|
| H: | $1s^{\mathbf{1}}$ |
| Li: | $1s^2 2s^{\mathbf{1}}$ |
| Na: | $[Ne]3s^{\mathbf{1}}$ |
| K: | $[Ar]4s^{\mathbf{1}}$ |
| Rb: | $[Kr]5s^{\mathbf{1}}$ |
| Cs: | $[Xe]6s^{\mathbf{1}}$ |

They all have a similar electron configuration in their valence shells: a single *s* electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar chemistry—*and they do*. The organization of electrons in atoms explains not only the shape of the periodic table, but also the fact that elements in the same column of the periodic table have similar chemistry.

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the *s* and *p* blocks. In the *d* and *f* blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure **5.5.10**. It is in the fourth column of the *p* block. This means that its electron configuration should end in a $p^4$ electron configuration. Indeed, the electron configuration of Se is $[Ar]4s^2 3d^{10} 4p^4$, as expected.

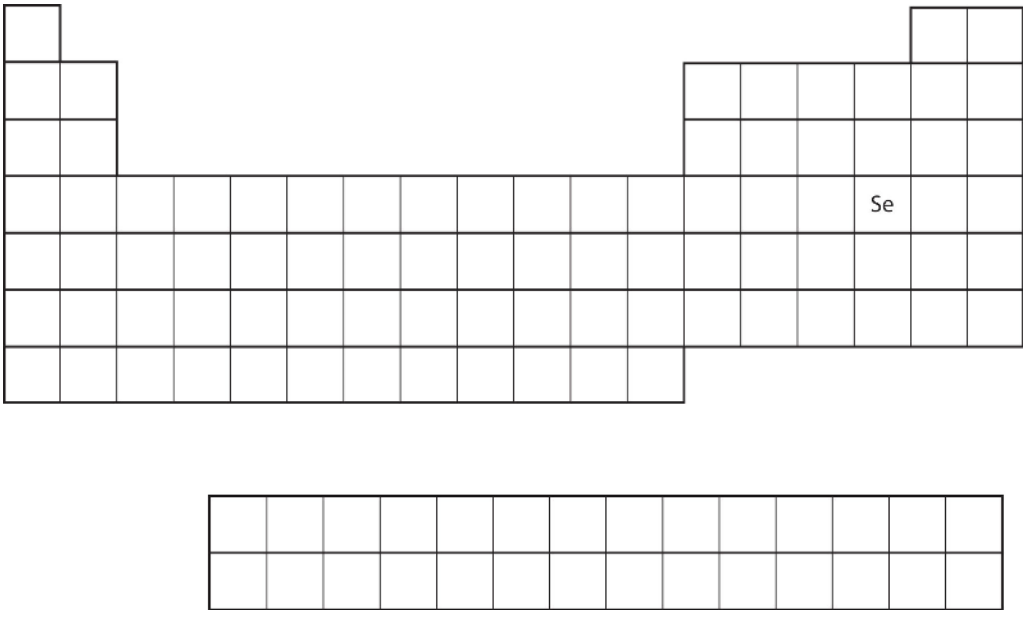



Figure **5.5.10**: Selenium on the Periodic Table

## Example 5.5.1

From the element's position on the periodic table, predict the valence shell electron configuration for each atom (Figure 5.5.11).

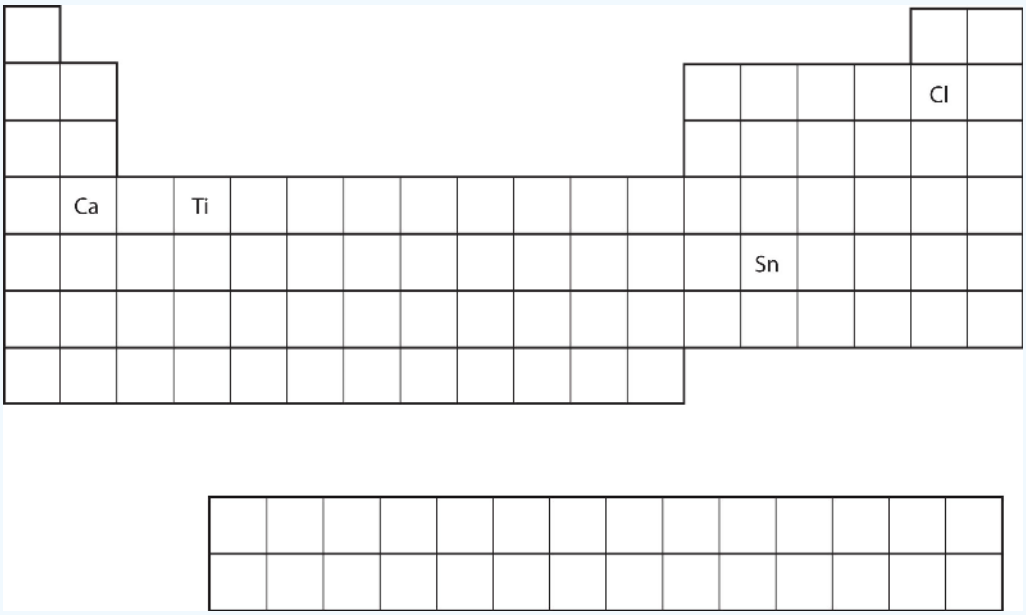



Figure 5.5.11: Various Elements on the Periodic Table

1. Ca
2. Sn

### Solution

1. Ca is located in the second column of the *s* block. We would expect that its electron configuration should end with $s^2$. Calcium's electron configuration is $[Ar]4s^2$.
2. Sn is located in the second column of the *p* block, so we expect that its electron configuration would end in $p^2$. Tin's electron configuration is $[Kr]5s^24d^{10}5p^2$.

## Exercise 5.5.1

From the element's position on the periodic table, predict the valence shell electron configuration for each atom. Figure 5.5.11.

a. Ti
b. Cl

**Answer a**

$[Ar]4s^23d^2$

**Answer b**

$[Ne]3s^23p^5$

## Food and Drink Application: Artificial Colors

The color of objects comes from a different mechanism than the colors of neon and other discharge lights. Although colored lights produce their colors, objects are colored because they preferentially reflect a certain color from the white light that shines on them. A red tomato, for example, is bright red because it reflects red light while absorbing all the other colors of the rainbow.

Many foods, such as tomatoes, are highly colored; in fact, the common statement "you eat with your eyes first" is an implicit recognition that the visual appeal of food is just as important as its taste. But what about processed foods?

Many processed foods have food colorings added to them. There are two types of food colorings: natural and artificial. Natural food colorings include caramelized sugar for brown; annatto, turmeric, and saffron for various shades of orange or yellow; betanin from beets for purple; and even carmine, a deep red dye that is extracted from the cochineal, a small insect that is a parasite on cacti in Central and South America. (That's right—you may be eating bug juice!)

Some colorings are artificial. In the United States, the Food and Drug Administration currently approves only seven compounds as artificial colorings in food, beverages, and cosmetics:

- FD&C Blue #1: Brilliant Blue FCF
- FD&C Blue #2: Indigotine
- FD&C Green #3: Fast Green FCF
- RD&C Red #3: Erythrosine
- FD&C Red #40: Allura Red AC
- FD&C Yellow #5: Tartrazine
- FD&C Yellow #6: Sunset Yellow FCF

Lower-numbered colors are no longer on the market or have been removed for various reasons. Typically, these artificial colorings are large molecules that absorb certain colors of light very strongly, making them useful even at very low concentrations in foods and cosmetics. Even at such low amounts, some critics claim that a small portion of the population (especially children) is sensitive to artificial colorings and urge that their use be curtailed or halted. However, formal studies of artificial colorings and their effects on behavior have been inconclusive or contradictory. Many people continue to enjoy processed foods with artificial coloring (like those shown in the accompanying figure).



Artificial food colorings. Artificial food colorings are found in a variety of food products, such as processed foods, candies, and egg dyes. Even pet foods have artificial food coloring in them. (CC by 2.0, Matthew Bland via Flickr.)

## Summary

The arrangement of electrons in atoms is responsible for the shape of the periodic table. Electron configurations can be predicted by the position of an atom on the periodic table.

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# 5.6: Periodic Trends

**Learning Objective**

- Be able to state how certain properties of atoms vary based on their relative position on the periodic table.

One reason the periodic table is so useful is that its structure allows us to qualitatively determine how some properties of the elements vary according to their position on the periodic table. The variations of properties according to positions on the periodic table are called **periodic trends**. There is no other tool in science that allows us to judge relative properties of a class of objects like this, which makes the periodic table a very useful tool. Many periodic trends are general. There may be a few points where an opposite trend is seen, but there is an overall trend when considered across a whole row, or down an entire column of the periodic table.

The first periodic trend we will consider is atomic radius. The **atomic radius** is an indication of the size of an atom. Although the concept of a definite radius of an atom is a bit fuzzy, atoms behave as if they have a certain radius. Such radii can be estimated from various experimental techniques, such as the x-ray crystallography of crystals.

As you go down a column of the periodic table, the atomic radii increase. This is because the valence electron shell is getting larger and there is a larger principal quantum number, so the valence shell lies physically farther away from the nucleus. This trend can be summarized as follows:

\[as\downarrow PT,atomic\; radius \uparrow\nonumber \]

where PT stands for periodic table. Going across a row on the periodic table, left to right, the trend is different. This is because although the valence shell maintains the same principal quantum number, the number of protons—and hence the nuclear charge—is increasing as you go across the row. The increasing positive charge casts a tighter grip on the valence electrons, so as you go across the periodic table, the atomic radii decrease. Again, we can summarize this trend as follows:

$$as \rightarrow PT, atomic\ radius \downarrow$$

Figure **5.6.1** shows spheres representing the atoms of the *s* and *p* blocks from the periodic table to scale, showing the two trends for the atomic radius.

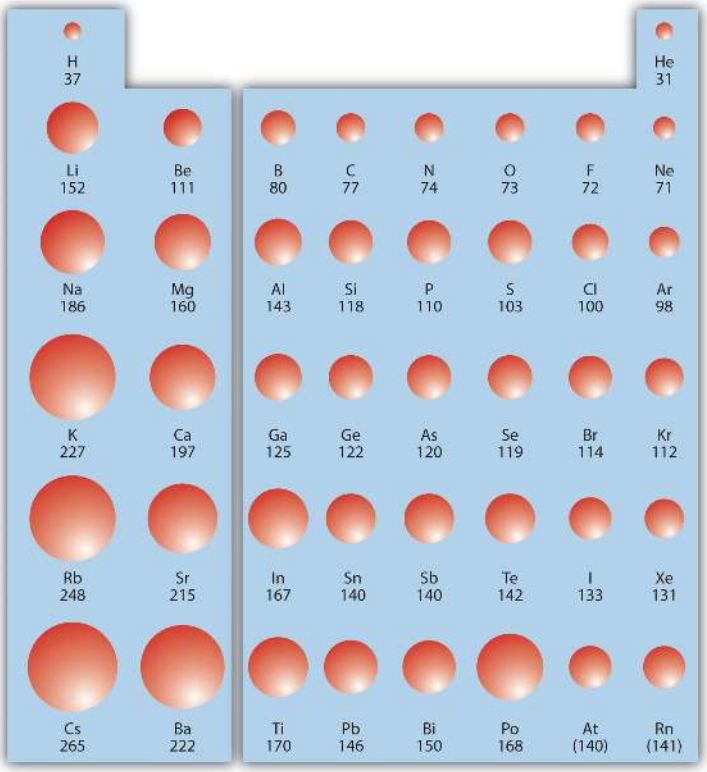



Figure 5.6.1: Atomic Radii Trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

### Example 5.6.1: Atomic Radii

Referring only to a periodic table and not to Figure 5.6.1, which atom is larger in each pair?

a. Si or S
b. S or Te

**Solution**

a. Si is to the left of S on the periodic table, so it is larger because as you go across the row, the atoms get smaller.
b. S is above Te on the periodic table, so Te is larger because as you go down the column, the atoms get larger.

### Exercise 5.6.1: Atomic Radii

Referring only to a periodic table and not to Figure 5.6.1, which atom is smaller, Ca or Br?

***Answer***

Br

**Ionization energy (IE)** is the amount of energy required to remove an electron from an atom in the gas phase:

$$A(g) \rightarrow A^{+}(g) + e^{-} \quad \Delta H \equiv IE$$

IE is usually expressed in kJ/mol of atoms. It is always positive because the removal of an electron always requires that energy be put in (i.e., it is endothermic). IE also shows periodic trends. As you go down the periodic table, it becomes easier to remove an electron from an atom (i.e., IE decreases) because the valence electron is farther away from the nucleus. Thus,

$$as \downarrow PT,\ IE \downarrow$$

However, as you go across the periodic table and the electrons get drawn closer in, it takes more energy to remove an electron. As a result, IE increases:

$$as \rightarrow PT,\ IE \uparrow$$

Figure 5.6.2 shows values of IE versus position on the periodic table. Again, the trend isn't absolute, but the general trends going across and down the periodic table should be obvious.

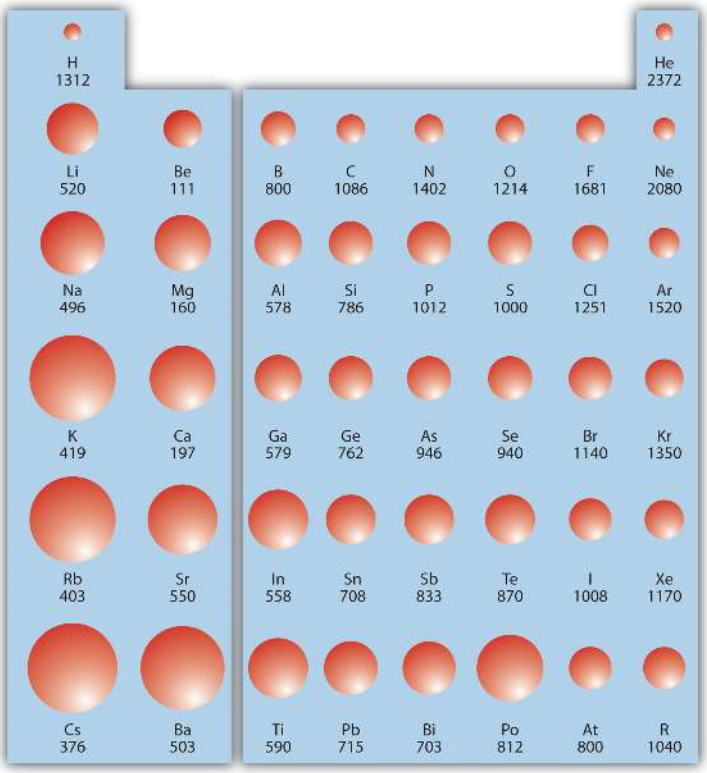



Figure 5.6.2: Ionization Energy on the Periodic Table. Values are in kJ/mol.

IE also shows an interesting trend within a given atom. This is because more than one IE can be defined by removing successive electrons (if the atom has them to begin with):

- First Ionization Energy ($IE_1$):

$$A(g) \rightarrow A^+(g) + e^-$$

- Second Ionization Energy ($IE_2$):

$$A^+(g) \rightarrow A^{2+}(g) + e^-$$

- Third Ionization Energy ($IE_3$):

$$A^{2+}(g) \rightarrow A^{3+}(g) + e^-$$

and so forth.

Each successive IE is larger than the previous because an electron is being removed from an atom with a progressively larger positive charge. However, IE takes a large jump when a successive ionization goes down into a new shell. For example, the following are the first three IEs for Mg, whose electron configuration is $1s^2 2s^2 2p^6 3s^2$:

First Ionization Energy ($IE_1$) = 738 kJ/mol:

$$Mg(g) \rightarrow Mg^+(g) + e^-$$

Second Ionization Energy ($IE_2$) = 1,450 kJ/mol:

$$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-}$$

Third Ionization Energy ($IE_3$) = 7,734 kJ/mol:

$$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^{-}$$

The second IE is twice the first, which is not a surprise: the first IE involves removing an electron from a neutral atom, while the second one involves removing an electron from a positive ion. The third IE, however, is over *five times* the previous one. Why is it so much larger? Because the first two electrons are removed from the 3*s* subshell, but the third electron has to be removed from the $n$ = 2 shell (specifically, the 2*p* subshell, which is lower in energy than the $n$ = 3 shell). Thus, it takes much more energy than just overcoming a larger ionic charge would suggest. Trends like this one demonstrate that electrons within atoms are organized in groups.

### Example 5.6.2: Ionization Energies

Which atom in each pair has the larger first ionization energy?

a. Ca or Sr
b. K or $K^+$

**Solution**

a. Because Sr is below Ca on the periodic table, it is easier to remove an electron from it; thus, Ca has the higher IE.
b. Because $K^+$ has a positive charge, it will be harder to remove another electron from it, so its IE is larger than that of K. Indeed, it will be significantly larger because the next electron in $K^+$ to be removed comes from another shell.

### Exercise 5.6.2: Ionization Energies

Which atom has the lower ionization energy, C or F?

**Answer**

C

The opposite of IE is described by **electron affinity (EA),** which is the energy change when a gas-phase atom accepts an electron:

$$A(g) + e^{-} \rightarrow A^{-}(g) \quad \Delta H \equiv EA$$

EA is also usually expressed in kJ/mol. EA also demonstrates some periodic trends, although they are less obvious than the other periodic trends discussed previously. Generally, as you go across the periodic table, EA increases its magnitude:

$$as \rightarrow PT,\ EA \uparrow$$

There is not a definitive trend as you go down the periodic table; sometimes EA increases, sometimes it decreases. Figure **5.6.3** shows EA values versus position on the periodic table for the *s*- and *p*-block elements. The trend isn't absolute, especially considering the large positive EA values for the second column. However, the general trend going across the periodic table should be obvious.

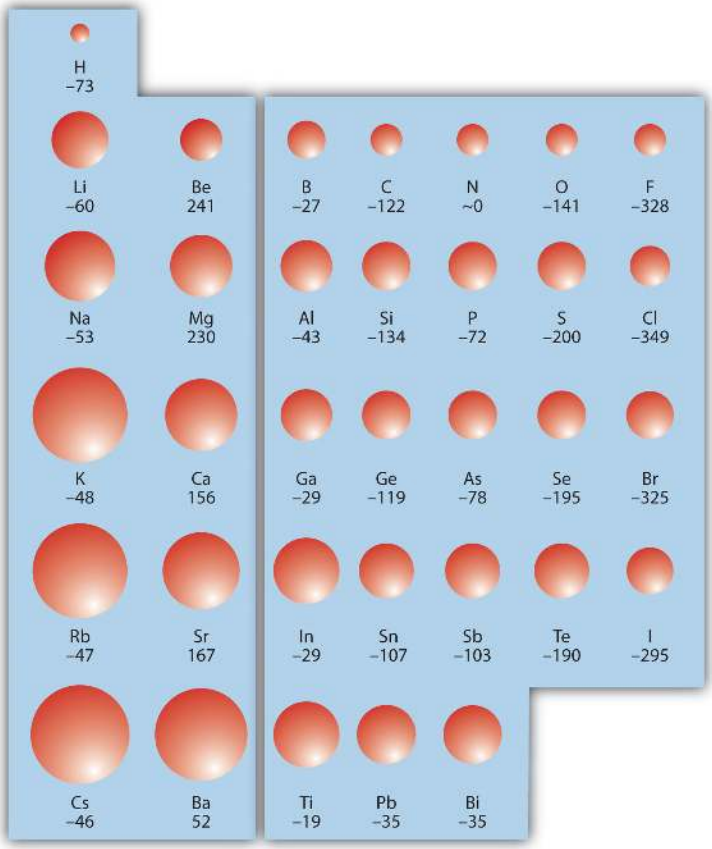



Figure **5.6.3**: Electron Affinity on the Periodic Table. Values are in kJ/mol.

### Example 5.6.3: Electron Affinities

Predict which atom in each pair will have the highest magnitude of Electron Affinity.

a. C or F
b. Na or S

**Solution**

a. C and F are in the same row on the periodic table, but F is farther to the right. Therefore, F should have the larger magnitude of EA.
b. Na and S are in the same row on the periodic table, but S is farther to the right. Therefore, S should have the larger magnitude of EA.

### Exercise 5.6.3: Electron Affinities

Predict which atom will have the highest magnitude of Electron Affinity: As or Br.

**Answer**

Br

## Summary

Certain properties—notably atomic radius, ionization energies, and electron affinities—can be qualitatively understood by the positions of the elements on the periodic table.

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# 5.E: Electronic Structure (Exercises)

## Exercises (Light)

1. Describe the characteristics of a light wave.

2. What is a characteristic of a particle of light?

3. What is the frequency of light if its wavelength is $7.33 \times 10^{-5}\ \text{m}$ ?

4. What is the frequency of light if its wavelength is $1.226\ \text{m}$ ?

5. What is the frequency of light if its wavelength is $733\ \text{nm}$ ?

6. What is the frequency of light if its wavelength is $8.528\ \text{cm}$ ?

7. What is the wavelength of light if its frequency is $8.19 \times 10^{14}\ \text{s}^{-1}$ ?

8. What is the wavelength of light if its frequency is $3.66 \times 10^{6}\ \text{s}^{-1}$ ?

9. What is the wavelength of light if its frequency is $1.009 \times 10^{6}\ \text{Hz}$ ?

10. What is the wavelength of light if its frequency is $3.79 \times 10^{-3}\ \text{Hz}$ ?

11. What is the energy of a photon if its frequency is $5.55 \times 10^{13}\ \text{s}^{-1}$ ?

12. What is the energy of a photon if its frequency is $2.06 \times 10^{18}\ \text{s}^{-1}$ ?

13. What is the energy of a photon if its wavelength is $5.88 \times 10^{-4}\ \text{m}$ ?

14. What is the energy of a photon if its wavelength is $1.888 \times 10^{2}\ \text{m}$ ?

### Answers

1. Light has a wavelength and a frequency.

3. $4.09 \times 10^{12}\ \text{s}^{-1}$

5. $4.09 \times 10^{14}\ \text{s}^{-1}$

7. $3.66 \times 10^{-7}\ \text{m}$

9. $297\ \text{m}$

11. $3.68 \times 10^{-20}\ \text{J}$

13. $3.38 \times 10^{-22}\ \text{J}$

## Exercises (Quantum Numbers for Electrons)

1. Differentiate between a continuous spectrum and a line spectrum.

2. Under what circumstances is a continuous spectrum formed? Under what circumstances is a line spectrum formed?

3. What is the wavelength of light from the hydrogen atom spectrum when $n = 3$ ?

4. What is the wavelength of light from the hydrogen atom spectrum when $n = 5$ ?

5. What are the restrictions on the principal quantum number?

6. What are the restrictions on the angular momentum quantum number?

7. What are the restrictions on the magnetic quantum number?

8. What are the restrictions on the spin quantum number?

9. What are the possible values for $\ell$ when $n = 5$ ?

10. What are the possible values for $\ell$ when $n = 1$ ?

11. What are the possible values for $m\ell$ when $\ell = 3$ ?

12. What are the possible values for $m\ell$ when $\ell = 6$ ?

13. Describe the shape of an $s$ orbital.

14. Describe the shape of a $p$ orbital.

15. Which of these sets of quantum numbers is allowed? If it is not, explain why.

a. $\{4, 1, -2, +1/2\}$
b. $\{2, 0, 0, -1/2\}$

16. Which of these sets of quantum numbers is allowed? If it is not, explain why.

a. $\{5, 2, -1, -1/2\}$
b. $\{3, -1, -1, -1/2\}$

## Answers

1. A continuous spectrum is a range of light frequencies or wavelengths; a line spectrum shows only certain frequencies or wavelengths.

3. $6.56 \times 10^{-7}$ m, or 656 nm

5. The principal quantum number is restricted to being a positive whole number.

7. The absolute value of $m\ell$ must be less than or equal to $\ell$ : $|m\ell| \leq \ell$.

9. $l$ can be $0, 1, 2, 3$, or 4 .

11. $m\ell$ can be $-3, -2, -1, 0, 1, 2$, or 3 .

13. An $s$ orbital is spherical in shape.

15. a. Because $|m\ell|$ must be less than $\ell$, this set of quantum numbers is not allowed.
    b. allowed

## Exercises (Organization of Electrons in Atoms)

1. Give two possible sets of four quantum numbers for the electron in an $\mathbf{H}$ atom.

2. Give the possible sets of four quantum numbers for the electrons in a Li atom.

3. How many subshells are completely filled with electrons for Na? How many subshells are unfilled?

4. How many subshells are completely filled with electrons for Mg? How many subshells are unfilled?

5. What is the maximum number of electrons in the entire $n = 2$ shell?

6. What is the maximum number of electrons in the entire $n = 4$ shell?

7. Write the complete electron configuration for each atom.
    a. Si, 14 electrons
    b. Sc, 21 electrons

8. Write the complete electron configuration for each atom.
    a. $\mathbf{Br, 35}$ electrons
    b. Be, 4 electrons

9. Write the complete electron configuration for each atom.
    a. Cd,48 electrons
    b. Mg, 12 electrons

10. Write the complete electron configuration for each atom.
    a. Cs, 55 electrons
    b. Ar, 18 electrons

11. Write the abbreviated electron configuration for each atom in Exercise 7.

12. Write the abbreviated electron configuration for each atom in Exercise 8.

13. Write the abbreviated electron configuration for each atom in Exercise 9.

14. Write the abbreviated electron configuration for each atom in Exercise 10.

### Answers

1. $\{1, 0, 0, 1/2\}$ and $[1, 0, 0, -1/2\}$

3. Three subshells $(1s, 2s, 2p)$ are completely filled, and one shell $(3s)$ is partially filled.

5. 8 electrons

7. a. $1s^2 2s^2 2p^6 3s^2 3p^2$

b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$

9. a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$

b. $1s^2 2s^2 2p^6 3s^2$

11. a. $[Ne] 3s^2 3p^2$

b. $[Ar] 4\ s^2 3d^1$

13. a. $[Kr] 5s^2 4d^{10}$

b. $[Ne] 3\ s^2$

## Exercises (Periodic Table)

1. Where on the periodic table are $s$ subshells being occupied by electrons?

2. Where on the periodic table are $d$ subshells being occupied by electrons?

3. In what block is Ra found?

4. In what block is Br found?

5. What are the valence shell electron configurations of the elements in the second column of the periodic table?

6. What are the valence shell electron configurations of the elements in the nextto-last column of the periodic table?

7. What are the valence shell electron configurations of the elements in the first column of the $p$ block?

8. What are the valence shell electron configurations of the elements in the last column of the $p$ block?

9. From the element's position on the periodic table, predict the electron configuration of each atom.
   a. Sr
   b. S

10. From the element's position on the periodic table, predict the electron configuration of each atom.
   a. Fe
   b. Ba

11. From the element's position on the periodic table, predict the electron configuration of each atom.
   a. \(C\)
   b. Ar

12. From the element's position on the periodic table, predict the electron configuration of each atom.
   a. Cl
   b. K

13. From the element's position on the periodic table, predict the electron configuration of each atom.
   a. Ge
   b. C

14. From the element's position on the periodic table, predict the electron configuration of each atom.
    a. **Mg**
    b. \(\mathrm{CI}\)

### Answers

1. the first two columns

3. the $s$ block

5. $ns^2$

7. $ns^2np^1$

9. a. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$
   b. $1s^22s^22p^63s^23p^4$

11. a. $1s^22s^22p^63s^23p^64s^23d^3$
    b. $1s^22s^22p^63s^23p^6$

13. a. $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$
    b. $1s^22s^22p^2$

## Exercises (Periodic Trends)

1. Write a chemical equation with an IE energy change.

2. Write a chemical equation with an EA energy change.

3. State the trends in atomic radii as you go across and down the periodic table.

4. State the trends in IE as you go across and down the periodic table.

5. Which atom of each pair is larger?
    a. **Na** or **Cs**
    b. **N** or **Bi**

6. Which atom of each pair is larger?
    a. **C** or **Ge**
    b. Be or Ba

7. Which atom of each pair is larger?
    a. **K** or **Cl**
    b. Ba or Bi

8. Which atom of each pair is larger?
    a. Si or S
    b. **H** or **He**

9. Which atom has the higher IE?
    a. **Na** or **S**
    b. **Ge** or **Br**

10. Which atom has the higher IE?
    a. C or Ne
    b. Rb or I

11. Which atom has the higher IE?
    a. Li or Cs
    b. Se or O

12. Which atom has the higher IE?
    a. Al or Ga
    b. F or I

13. A third-row element has the following successive IEs: $738; 1,450; 7,734$; and $10,550\ \text{kJ/mol}$. Identify the element.

14. A third-row element has the following successive IEs: $1{,}012; 1,903; 2,912; 4,940$; 6,270 ; and $21,300\ \text{kJ/mol}$. Identify the element.

15. For which successive IE is there a large jump in IE for Ca ?

16. For which successive IE is there a large jump in IE for Al?

17. Which atom has the greater magnitude of EA?
    a. Cor F
    b. Al or Cl

18. Which atom has the greater magnitude of EA?
    a. K or Br
    b. Mg or S

## Answers

1. $\text{Na(g)} \rightarrow \text{Na}^{+}\text{(g)} + \text{e}^{-} \Delta H = \text{IE}$ (answers will vary)

3. As you go across, atomic radii decrease; as you go down, atomic radii increase.

5. a. Cs
   b. Bi

7. a. K
   b. Ba

9. a. S
   b. Br

11. a. Li
    b. O

13. Mg

15. The third IE shows a large jump in Ca.

17. a. $\mathrm{F}$
    b. $\mathrm{Cl}$

## Additional Exercises

1. What is the frequency of light if its wavelength is $1.00\ \mathrm{m}$ ?

2. What is the wavelength of light if its frequency is $1.00\ \mathrm{s}^{-1}$ ?

3. What is the energy of a photon if its wavelength is 1.00 meter?

4. What is the energy of a photon if its frequency is $1.00\ \mathrm{s}^{-1}$ ?

5. If visible light is defined by the wavelength limits of $400\ \mathrm{nm}$ and $700\ \mathrm{nm}$, what is the energy range for visible light photons?

6. Domestic microwave ovens use microwaves that have a wavelength of $122\ \mathrm{mm}$. What is the energy of one photon of this microwave?

7. Use the equation for the wavelengths of the lines of light in the $\mathrm{H}$ atom spectrum to calculate the wavelength of light emitted when $n$ is 7 and 8 .

8. Use the equation for the wavelengths of the lines of light in the $\mathrm{H}$ atom spectrum to calculate the wavelengths of light emitted when $n$ is 5 and 6 .

9. Make a table of all the possible values of the four quantum numbers when the principal quantum number $n = 5$.

10. Make a table of all the possible values of $m\ell$ and $m_s$ when $\ell = 4$. What is the lowest value of the principal quantum number for this to occur?

11. a. Predict the electron configurations of Sc through Zn.
    b. From a source of actual electron configurations, determine how many exceptions there are from your predictions in part a.

12. a. Predict the electron configurations of Ga through Kr.
    b. From a source of actual electron configurations, determine how many exceptions there are from your predictions in part a.

13. Recently, Russian chemists reported experimental evidence of element 117. Use the periodic table to predict its valence shell electron configuration.

14. Bi (atomic number 83) is used in some stomach discomfort relievers. Using its place on the periodic table, predict its valence shell electron configuration.

15. Which atom has a higher ionization energy (IE), O or P?

16. Which atom has a higher IE, $\mathrm{F}$ or As?

17. Which atom has a smaller radius, $\mathrm{As}$ or $\mathrm{Cl}$ ?

18. Which atom has a smaller radius, $\mathrm{K}$ or $\mathrm{F}$ ?

19. How many IEs does an $\mathrm{H}$ atom have? Write the chemical reactions for the successive ionizations.

20. How many IEs does a Be atom have? Write the chemical reactions for the successive ionizations.

21. Based on what you know of electrical charges, do you expect $\mathbf{Na^+}$ to be larger or smaller than $\mathbf{Na}$ ?

22. Based on what you know of electrical charges, do you expect $\mathbf{Cl^-}$ to be larger or smaller than $\mathbf{Cl}$ ?

## Answers

1. $\mathbf{3.00 \times 10^8\ s^{-1}}$

3. $\mathbf{1.99 \times 10^{-22}\ J}$

5. $\mathbf{4.97 \times 10^{-19}\ J}$ to $\mathbf{2.84 \times 10^{-19}\ J}$

7. $\mathbf{3.97 \times 10^{-7}\ m}$ and $\mathbf{3.89 \times 10^{-7}\ m}$, respectively

9.

| n | ℓ | mℓ | ms |
|---|---|---|---|
| 5 | 0 | 0 | 1/2 or -1/2 |
| 5 | 1 | -1, 0, -1 | 1/2 or -1/2 |
| 5 | 2 | −2, −1, 0, 1, 2 | 1/2 or -1/2 |
| 5 | 3 | −3, −2, −1, 0, 1, 2, 3 | 1/2 or -1/2 |
| 5 | 4 | −4, −3, −2, −1, 0, 1, 2, 3, 4 | 1/2 or -1/2 |

11. a. The electron configurations are predicted to end in $\mathbf{3d^1, 3d^2, 3d^3, 3d^4, 3d^5}$, $\mathbf{3d^6, 3d^7, 3d^8, 3d^9}$, and $\mathbf{3d^{10}}$.
b. $\mathbf{Cr}$ and $\mathbf{Cu}$ are exceptions.

13. Element 117's valence shell electron configuration should be $\mathbf{7s^2 7p^5}$.

15. $\mathbf{O}$

17. $\mathbf{Cl}$

19. $\mathbf{H}$ has only one IE: $\mathbf{H \rightarrow H^+ + e^-}$

21. smaller

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# CHAPTER OVERVIEW

## 6: Chemical Bonding


6.1: Introduction to Chemical Bonding

6.2: Lewis Structures of Ionic Compounds- Electrons Transferred

6.3: Covalent Lewis Structures- Electrons Shared

6.4: Writing Lewis Structures for Covalent Compounds

6.5: Resonance - Equivalent Lewis Structures for the Same Molecule

6.6: Other Aspects of Covalent Bonds

6.7: Molecular Shapes


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# 6.1: Introduction to Chemical Bonding

**Learning Objective**

- Draw a Lewis electron dot diagram for an atom or a monatomic ion.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram, or a Lewis diagram, or a Lewis structure) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (The order in which the positions are used does not matter.) For example, the Lewis electron dot diagram for hydrogen is simply

$$\mathbf{H}\cdot$$

Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:

$$\dot{\mathbf{H}} \quad \textit{or} \quad \cdot\mathbf{H} \quad \textit{or} \quad \underset{\cdot}{\mathbf{H}}$$

The electron dot diagram for helium, with two valence electrons, is as follows:

$$\mathbf{He:}$$

By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1*s* subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:

$$\mathbf{Li}\cdot$$

Beryllium has two valence electrons in its 2*s* shell, so its electron dot diagram is like that of helium:

$$\mathbf{Be:}$$

The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:

$$\dot{\mathbf{B}}\mathbf{e:}$$

Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2*s* subshell and two in the 2*p* subshell. As usual, we will draw two dots together on one side, to represent the 2*s* electrons. However, conventionally, we draw the dots for the two *p* electrons on different sides. As such, the electron dot diagram for carbon is as follows:

$$\cdot\dot{\mathbf{C}}\mathbf{:}$$

With N, which has three *p* electrons, we put a single dot on each of the three remaining sides:

$$\cdot\dot{\underset{\cdot}{\mathbf{N}}}\mathbf{:}$$

For oxygen, which has four *p* electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that a side has no more than two electrons.

$$\cdot\ddot{\underset{\cdot}{\mathbf{O}}}\mathbf{:}$$

Fluorine and neon have seven and eight dots, respectively:





With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the $n = 3$ shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

### Example 6.1.1

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

**Solution**

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3*s* electrons:



2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the $n = 4$ shell, there are six electrons. Its electron dot diagram is as follows:



### Exercise 6.1.1

What is the Lewis electron dot diagram for each element?

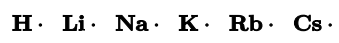
a. phosphorus
b. argon

**Answer**

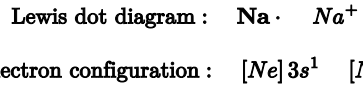




Elements in the same column of the periodic table have similar Lewis electron dot diagrams because they have the same valence shell electron configuration. The electron dot diagrams for the first column of elements are as follows:



Monatomic ions are atoms that have either lost (for cations) or gained (for anions) electrons. Electron dot diagrams for ions are the same as for atoms, except that some electrons have been removed for cations, while some electrons have been added for anions. Thus, in comparing the electron configurations and electron dot diagrams for the Na atom and the $Na^+$ ion, we note that the Na atom has a single valence electron in its Lewis diagram, while the $Na^+$ ion has lost that one valence electron:



Technically, the valence shell of the $Na^+$ ion is now the $n = 2$ shell, which has eight electrons in it. So why do we not put eight dots around $Na^+$? Conventionally, when we show electron dot diagrams for ions, we show the original valence shell of the atom, which in this case is the $n = 3$ shell and empty in the $Na^+$ ion.

In making cations, electrons are first lost from the *highest numbered shell*, not necessarily the last subshell filled. For example, in going from the neutral Fe atom to the $Fe^{2+}$ ion, the Fe atom loses its two 4*s* electrons first, not its 3*d* electrons, despite the fact that the 3*d* subshell is the last subshell being filled. Thus, we have

$$\textbf{Lewis dot diagram :} \quad \textbf{Fe} : \quad Fe^{2+}$$

$$\textbf{Electron configuration :} \quad [Ar]\,4s^2 3d^6 \quad [Ar]\,3d^6$$

Anions have extra electrons when compared to the original atom. Here is a comparison of the Cl atom with the $Cl^-$ ion:

$$\textbf{Lewis dot diagram :} \quad :\ddot{\underset{\cdot\,\cdot}{\text{Cl}}}\cdot \quad :\ddot{\underset{\cdot\,\cdot}{\text{Cl}}}:^-$$

$$\textbf{Electron configuration :} \quad [Ne]\,3s^2 3p^5 \quad [Ne]\,3s^2 3p^6$$

## Example 6.1.2

What is the Lewis electron dot diagram for each ion?

a. $Ca^{2+}$
b. $O^{2-}$

### Solution

a. Having lost its two original valence electrons, the Lewis electron dot diagram is simply: $Ca^{2+}$
b. The $O^{2-}$ ion has gained two electrons in its valence shell, so its Lewis electron dot diagram is as follows:

$$:\ddot{\underset{\cdot\,\cdot}{\text{O}}}:^{2-}$$

## Exercise 6.1.2

The electronic configuration of the thallium ion is $6s^2 4f^{14} 5d^{10}$. What is the Lewis electron dot diagram for the $Tl^+$ ion?

**Answer**

$$\textbf{Tl:}^+$$

## Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.
- Lewis electron dot diagrams for ions have less (for cations) or more (for anions) dots than the corresponding atom.

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# 6.2: LEWIS STRUCTURES OF IONIC COMPOUNDS- ELECTRONS TRANSFERRED

## LEARNING OBJECTIVES

- State the octet rule.
- Define *ionic bond*.
- Draw Lewis structures for ionic compounds.

In Section 4.7, we demonstrated that ions are formed by losing electrons to make cations, or by gaining electrons to form anions. The astute reader may have noticed something: many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The **octet rule** explains the favorable trend of atoms having eight electrons in their valence shell. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the $Na^+$ ion. We *could* remove another electron by adding even more energy to the ion, to make the $Na^{2+}$ ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the $Na^+$ ion has a complete octet in its new valence shell, the $n = 2$ shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Na}\cdot \qquad \cdot\ddot{\mathbf{Cl}}\mathbf{:}$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5$$

For the Na atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Na}\cdot \rightsquigarrow \cdot\ddot{\mathbf{Cl}}\mathbf{:}$$

resulting in two ions—the $Na^+$ ion and the $Cl^-$ ion:

$$\mathbf{Na}^+ \qquad \mathbf{:}\ddot{\mathbf{Cl}}\mathbf{:}^-$$

$$[Ne] \qquad [Ne]\,3s^2 3p^6$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the $Na^+$ and $Cl^-$ ions:

$$\mathbf{Na}^+ + \mathbf{:}\ddot{\mathbf{Cl}}\mathbf{:}^- \rightarrow Na^+Cl^- \;\; or \;\; NaCl$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:




The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

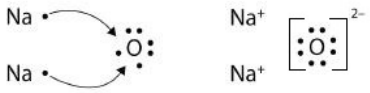
$$Mg^{2+} + \left[:\ddot{O}:\right]^{2-} \quad Mg^{2+}O^{2-} \text{ or } MgO$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$Na\cdot \rightsquigarrow \cdot\ddot{O}:$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:



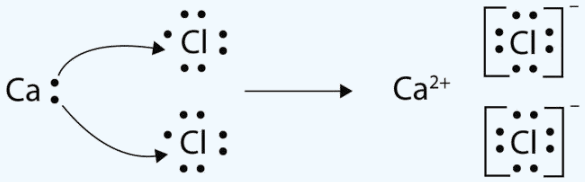

These three ions attract each other to give an overall neutral-charged ionic compound, which we write as $Na_2O$. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

### EXAMPLE 6.2.1: SYNTHESIS OF CALCIUM CHLORIDE FROM ELEMENTS

With arrows, illustrate the transfer of electrons to form calcium chloride from *Ca* atoms and *Cl* atoms.

#### Solution

A *Ca* atom has two valence electrons, while a *Cl* atom has seven electrons. A *Cl* atom needs only one more to complete its octet, while *Ca* atoms have two electrons to lose. Thus we need two *Cl* atoms to accept the two electrons from one *Ca* atom. The transfer process looks as follows:

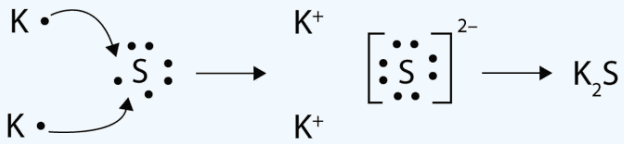



The oppositely charged ions attract each other to make $CaCl_2$.

### EXERCISE 6.2.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from *K* atoms and *S* atoms.

**Answer**




## SUMMARY

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

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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**.

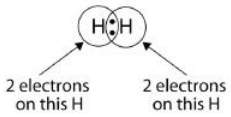
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \quad \cdot\mathbf{H}$$

The two H atoms can share their electrons:

$$\mathbf{H:H}$$

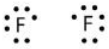
We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:




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

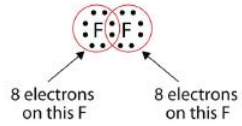
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:




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:



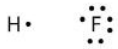

We can also write this using a dash to represent the shared electron pair:



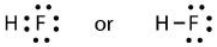

There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F

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:**

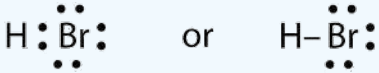
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

**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:

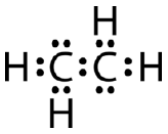


Figure **6.3.1**: Incorrect dot structure of ethene. (CK12 License)

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.

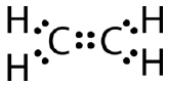


Figure **6.3.2**: Correct dot structure for ethene. (CK12 License)

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.

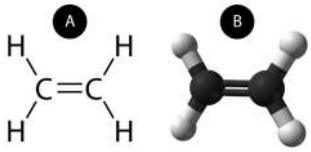



Figure **6.3.3**: (A) The structural model for $C_2H_4$ consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of $C_2H_4$.

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.

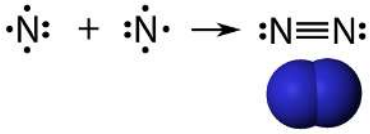


Figure **6.3.4**: Triple bond in $N_2$.

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

- 
- Anonymous by request

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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

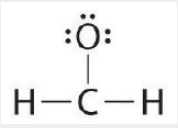
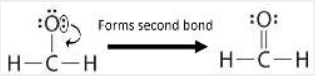
| Steps for Writing Lewis Structures | Example **6.4.1** |
|---|---|
| 1. Determine the total number of valence electrons in the molecule or ion. | Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons. |
| 2. Arrange the atoms to show specific connections. | H O H<br>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.<br>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<br>H -O- H<br>with 4 electrons left over.<br>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:<br>  |
| 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**

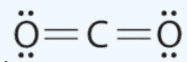
Solutions to Example 10.4.2

| 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. |  <br>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:<br>  <br>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:<br>  <br>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.<br>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:<br>  |
| 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.

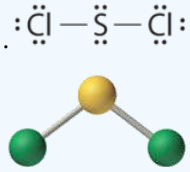
**Answer $CO_2$**





Carbon dioxide

**Answer $SCl_2$**




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.

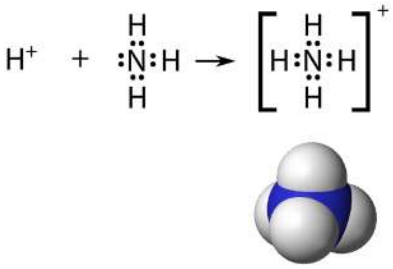


Figure **6.4.3**: The ammonium ion. (CK12 License)

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

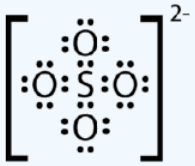
total of 8 valence electrons in the ion

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)

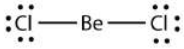
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

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:

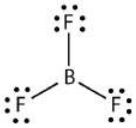



Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

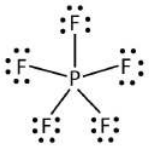
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting 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.

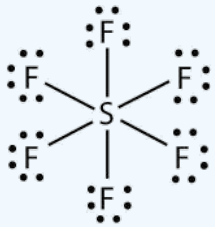
a. ClO
b. $SF_6$

**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: Resonance - Equivalent Lewis Structures for the Same Molecule

**Learning Objectives**

- Explain the concept of resonance and how it works with within molecules.

## Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone ($O_3$) molecule in Figure **6.5.1**. There are a total of 18 electrons in the structure and so the following two structures are possible.

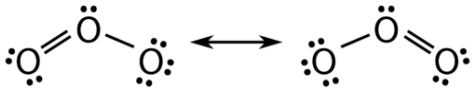


Figure **6.5.1**: Resonance forms of ozone. Note the use of the double-headed arrow.

The structure on the left (**6.5.1**) can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as $O_3$ consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the $O_3$ and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an O–O single bond and a double bond.

**Resonance** is the use of two or more Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the $O_3$ molecule, each of the covalent bonds between O atoms are best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (Figure **6.5.2**).

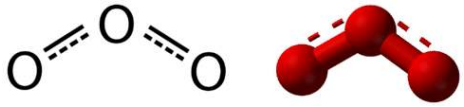


Figure **6.5.2**: "Half-bond" model of ozone molecule. This is a better description of the electronic structure of ozone than either of the resonance structures in Figure **6.5.1**.

Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, $NO_3^-$ in Figure **6.5.3**.

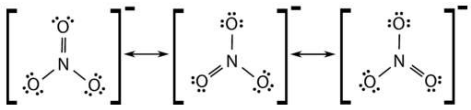


Figure **6.5.3**: Resonance structure of nitrate anion.

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

## Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

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# 6.6: Other Aspects of Covalent Bonds

**Learning Objectives**

- Describe a nonpolar bond and a polar bond.
- Use electronegativity to determine whether a bond between two elements will be nonpolar covalent, polar covalent, or ionic.
- Describe the bond energy of a covalent bond.

Consider the $H_2$ molecule:



Because the nuclei of each H atom contain protons, the electrons in the bond are attracted to the nuclei (opposite charges attract). But because the two atoms involved in the covalent bond are both H atoms, each nucleus attracts the electrons by the same amount. Thus the electron pair is equally shared by the two atoms. The equal sharing of electrons in a covalent bond is called a **nonpolar covalent bond**.

Now consider the ***HF*** molecule:



There are two different atoms involved in the covalent bond. The H atom has one proton in its nucleus that is attracting the bonding pair of electrons. However, the F atom has nine protons in its nucleus, with nine times the attraction of the H atom. The F atom attracts the electrons so much more strongly that the electrons remain closer to the F atom than to the H atom; the electrons are no longer equally balanced between the two nuclei. Instead of representing the HF molecule as



it may be more appropriate to draw the covalent bond as



with the electrons in the bond being nearer to the F atom than the H atom. Because the electrons in the bond are nearer to the F atom, this side of the molecule takes on a partial negative charge, which is represented by δ− (δ is the lowercase Greek letter delta). The other side of the molecule, the H atom, adopts a partial positive charge, which is represented by δ+:




A covalent bond between different atoms that attract the shared electrons by different amounts, and cause an imbalance of electron distribution is called a **polar covalent bond**.

Technically, any covalent bond between two different elements is polar. However, the degree of polarity is important. A covalent bond between two different elements may be so slightly unbalanced that the bond is, essentially, nonpolar. A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how strongly atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure **6.6.1**.

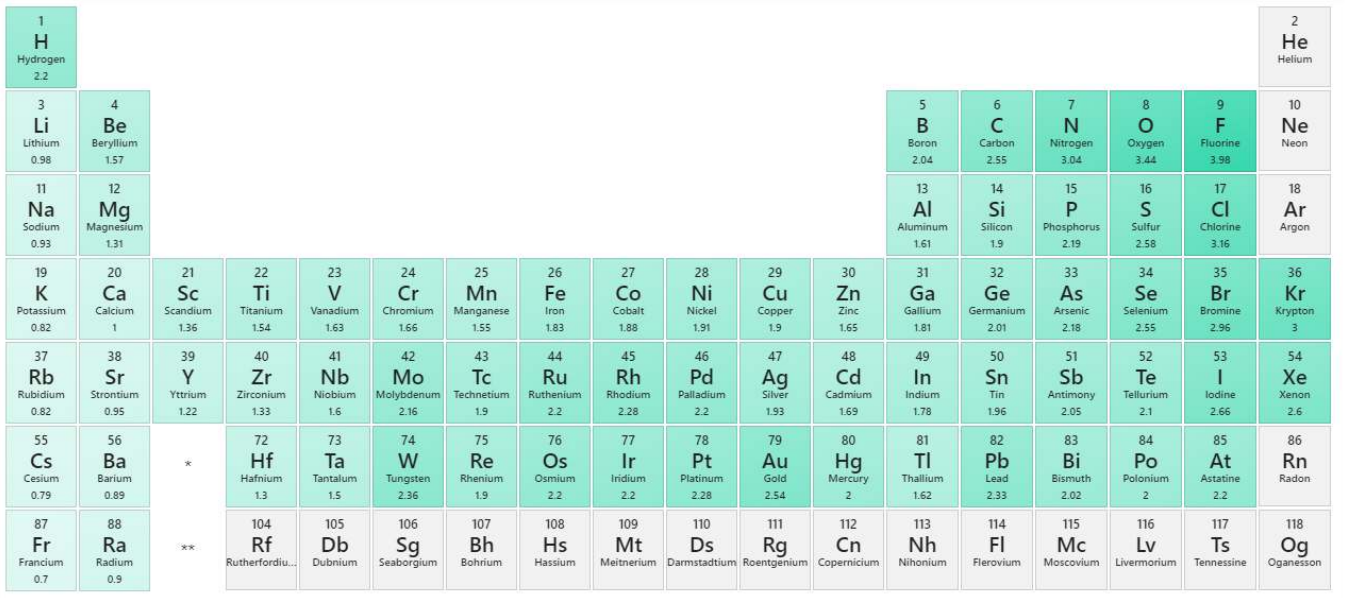



Figure **6.6.1**: Electronegativities of the Elements. Electronegativity is used to determine the polarity of covalent bonds. The more electronegative elements are in the upper right of the table (more colored), while the less electronegative are in the lower left (less colored).

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

Table with two columns and four rows. The first column on the left has different values underneath in the row. The second column on the right side has the corresponding bond type for the values underneath in the rows.

| Electronegativity Difference | Bond Type |
|---|---|
| 0 | nonpolar covalent |
| 0–0.4 | slightly polar covalent |
| 0.4–1.9 | definitely polar covalent |
| >1.9 | likely ionic |

### ✔ Example 6.6.1

What is the polarity of each bond?

a. C–H
b. O–H

#### Solution

Using Figure **6.6.1**, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 − 2.1 = 0.4. Thus we predict that this bond will be slightly polar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 − 2.1 = 1.4, so we predict that this bond will be definitely polar covalent.

### ? Exercise 6.6.1

What is the polarity of each bond?

a. Rb–F

b. P–Cl

**Answer a**

likely ionic

**Answer b**

polar covalent

The polarity of a covalent bond can have significant influence on the properties of the substance. If the overall molecule is polar, the substance may have a higher melting point and boiling point than expected; also, it may or may not be soluble in various other substances, such as water or hexane.

It should be obvious that covalent bonds are stable because molecules exist. However, they can be broken if enough energy is supplied to a molecule. For most covalent bonds between any two given atoms, a certain amount of energy must be supplied. Although the exact amount of energy depends on the molecule, the approximate amount of energy to be supplied is similar if the atoms in the bond are the same. The approximate amount of energy needed to break a covalent bond is called the **bond energy** of the covalent bond. Table **6.6.1**, lists the bond energies of some covalent bonds.

Table **6.6.1**: Bond Energies of Covalent Bonds

| Bond | Energy (kJ/mol) | Bond | Energy (kJ/mol) |
|---|---|---|---|
| C–C | 348 | N–N | 163 |
| C=C | 611 | N=N | 418 |
| C≡C | 837 | N≡N | 946 |
| C–O | 351 | N–H | 389 |
| C=O | 799 | O–O | 146 |
| C–Cl | 328 | O=O | 498 |
| C–H | 414 | O–H | 463 |
| F–F | 159 | S–H | 339 |
| H–Cl | 431 | S=O | 523 |
| H–F | 569 | Si–H | 293 |
| H–H | 436 | Si–O | 368 |

A few trends are obvious from Table **6.6.1**. For bonds that involve the same two elements, a double bond is stronger than a single bond, and a triple bond is stronger than a double bond. The energies of multiple bonds are not exact multiples of the single bond energy; for carbon-carbon bonds, the energy increases somewhat less than double or triple the C–C bond energy, while for nitrogen-nitrogen bonds the bond energy increases at a rate greater than the multiple of the N–N single bond energy. The bond energies in Table **6.6.1** are average values; the exact value of the covalent bond energy will vary slightly among molecules with these bonds, but should be close to these values.

To be broken, covalent bonds always require energy; that is, covalent bond breaking is always an *endothermic* process. Thus the $\Delta H$ for this process is positive:

$$\text{Molecule–O–H} \rightarrow \text{Molecule–O} + \text{H}$$

$$\text{with } \Delta H \approx +463 \text{ kJ/mol}$$

However, when making a covalent bond, energy is always given off; covalent bond making is always an *exothermic* process. Thus $\Delta H$ for this process is negative:

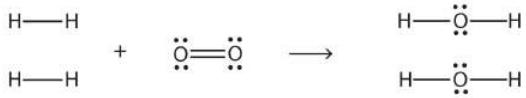
$$\text{Molecule–S} + \text{H} \rightarrow \text{Molecule–S–H}$$

with $\Delta H \approx -339$ kJ/mol

Bond energies can be used to estimate the energy change of a chemical reaction. When bonds are broken in the reactants, the energy change for this process is endothermic. When bonds are formed in the products, the energy change for this process is exothermic. We combine the positive energy change with the negative energy change to estimate the overall energy change of the reaction. For example, in

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

we can draw Lewis electron dot diagrams for each substance to see what bonds are broken and what bonds are formed:



(The lone electron pairs on the O atoms are omitted for clarity.) We are breaking two H–H bonds and one O–O double bond and forming four O–H single bonds. The energy required for breaking the bonds is as follows:

| | |
|---|---|
| 2 H–H bonds: | 2(+436 kJ/mol) |
| 1 O=O bond: | +498 kJ/mol |
| Total: | +1,370 kJ/mol |

The energy given off when the four O–H bonds are made is as follows:

| | |
|---|---|
| 4 O–H bonds: | 4(−463 kJ/mol) |
| Total: | −1,852 kJ/mol |

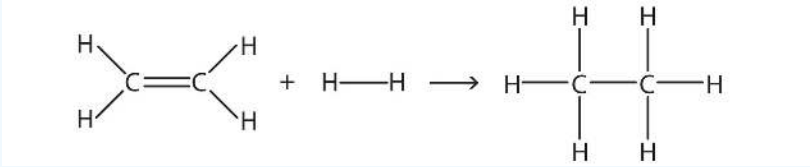
Combining these two numbers:

| | |
|---|---|
| | +1,370 kJ/mol + (−1,852 kJ/mol) |
| Net Change: | −482 kJ/mol ≈ $\Delta H$ |

The actual $\Delta H$ is −572 kJ/mol; we are off by about 16%. Although not ideal, a 16% difference is reasonable because we used estimated, not exact, bond energies.

**Example 6.6.1**

Estimate the energy change of this reaction.



Ethene reacts with hydrogen gas to create ethane. A c-c double bond is broken and two C-H bonds are formed.

**Solution**

Here, we are breaking a C–C double bond and an H–H single bond and making a C–C single bond and two C–H single bonds. Bond breaking is endothermic, while bond making is exothermic. For the bond breaking:

| | |
|---|---|
| 1 C=C: | +611 kJ/mol |
| 1 H–H: | +436 kJ/mol |
| Total: | +1,047 kJ/mol |

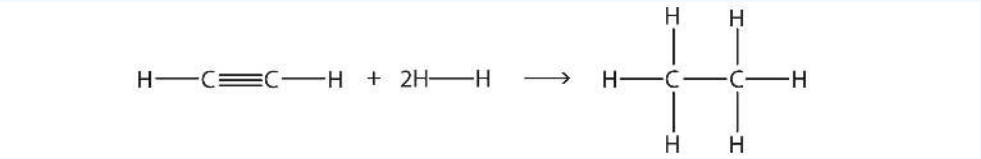
For the bond making:

| | |
|---|---|
| 1 C–C: | −348 kJ/mol |
| 2 C–H: | 2(−414 kJ/mol) |
| Total | −1,176 kJ/mol |

Overall, the energy change is +1,047 + (−1,176) = −129 kJ/mol.

**? Exercise 6.6.1**

Estimate the energy change of this reaction.



## Summary

- Covalent bonds can be nonpolar or polar, depending on the electronegativities of the atoms involved.
- Covalent bonds can be broken if energy is added to a molecule.
- The formation of covalent bonds is accompanied by energy given off.
- Covalent bond energies can be used to estimate the enthalpy changes of chemical reactions.

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# 6.7: Molecular Shapes

**Learning Objective**

- Determine the shape of simple molecules.

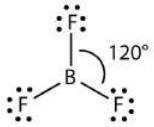
Molecules have shapes. There is an abundance of experimental evidence to that effect, from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basis of molecular shapes is called **valence shell electron pair repulsion (VSEPR).** VSEPR says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from each other as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

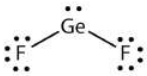
Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include $BeH_2$ and $CO_2$:




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 $BF_3$:




Some substances have a trigonal planar electron group distribution, but have atoms bonded to only two of the three electron groups. An example is $GeF_2$:




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 **6.7.1**: 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.

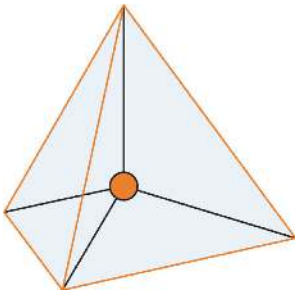
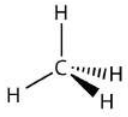
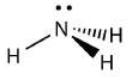


Figure **6.7.1**: Tetrahedral Geometry. Four electron groups orient themselves in the shape of a tetrahedron.

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.



$NH_3$ is an example of a molecule whose central atom has four electron groups, but only three of them are bonded to surrounding atoms.



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.

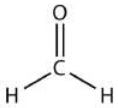


Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom, but only one electron group bonded to another atom is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. $CH_2O$ has the following Lewis electron dot diagram.



The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape:



(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

### Example 6.7.1

What is the approximate shape of each molecule?

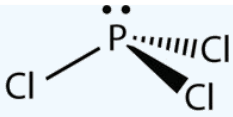
a. $\mathbf{PCl_3}$
b. **NOF**

#### Solution

The first step is to draw the Lewis electron dot diagram of the molecule.

***a:***

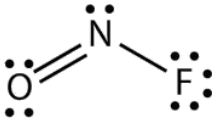
For $PCl_3$, the electron dot diagram is as follows:




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.

*b:*

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 6.7.1**

What is the approximate molecular shape of $CH_2Cl_2$?

**Answer**

Tetrahedral

Table **6.7.1** summarizes the shapes of molecules based on their number of electron groups and surrounding atoms.

Table **6.7.1**: Summary of Molecular Shapes

| Number of Electron Groups on Central Atom | Number of Surrounding Atoms | Molecular Shape |
|---|---|---|
| any | 1 | linear |
| 2 | 2 | linear |
| 3 | 3 | trigonal planar |
| 3 | 2 | bent |
| 4 | 4 | tetrahedral |
| 4 | 3 | trigonal pyramidal |
| 4 | 2 | bent |

## 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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# CHAPTER OVERVIEW

## 7: Molecules and Nomenclature


7.1: Introduction to Molecules and Nomenclature
7.2: Nomenclature- Naming Compounds
7.3: Naming Ionic Compounds
7.4: Naming Molecular Compounds
7.5: Naming Acids
7.6: Nomenclature Summary


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# 7.1: Introduction to Molecules and Nomenclature

**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 7.1.1: 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<br>1(3+) = 1(3-)<br>+3 = -3 | total charge of cations = total charge of anions<br>2(1+) = 1(2-)<br>+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 7.1.2: 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 7.1.2

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 7.1.3: 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 7.1.3

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 **7.1.1** lists the most common polyatomic ions.

Table **7.1.1**: 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 7.1.4: 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 7.1.5

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 7.1.5

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 7.1.6

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 7.1.6

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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# 7.2: Nomenclature- Naming Compounds

Nomenclature is the process of naming chemical compounds so that they can be easily identified as separate chemicals. The primary function of chemical nomenclature is to ensure that a spoken or written chemical name leaves no ambiguity concerning which chemical compound the name refers to—each chemical name should refer to a single substance. A less important aim is to ensure that each substance has a single name, although a limited number of alternative names is acceptable in some cases. Preferably, the name also conveys some information about the structure or chemistry of a compound. A common name will often suffice to identify a chemical compound in a particular set of circumstances. To be more generally applicable, the name should indicate at least the chemical formula. To be more specific still, the three-dimensional arrangement of the atoms may need to be specified.

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# 7.3: 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 **7.3.1**: 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 **7.3.1** lists the elements that use the common system, along with their respective cation names.

Table **7.3.2**: 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 **7.3.2** lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

### Example 7.3.1

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 7.3.1

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 7.3.2

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 7.3.2

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 **7.3.1** for the compound $BaCl_2$. The word *ion* is dropped from both parts.

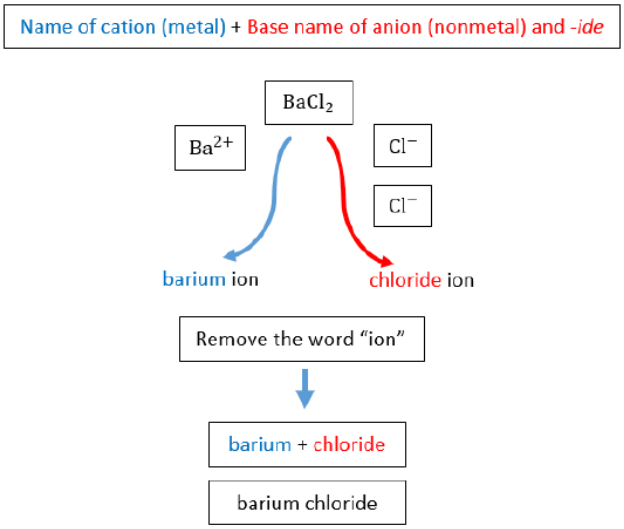



Figure **7.3.1**: Naming $BaCl_2$
Naming formula: Name of metal cation + base
name of nonmetal anion + suffix -ide. BaCl2 is named as barium chloride.

Subscripts in the formula do not affect the name.

### ✔ Example 7.3.3: 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 7.3.3

Name each ionic compound.

a. AgI
b. MgO
c. $Ca_3P_2$

**Answer a:**

silver iodide

**Answer b:**

magnesium oxide

**Answer c:**

calcium phosphide

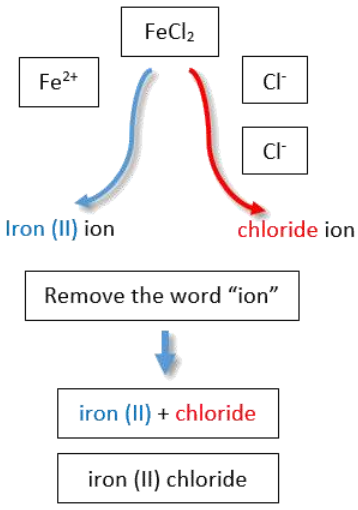
## Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

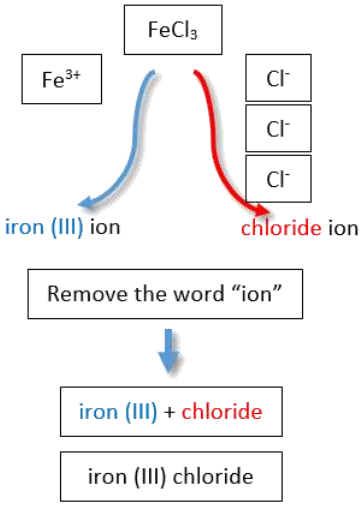
If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron

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 **7.3.2**).

Table **7.3.3**: 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***

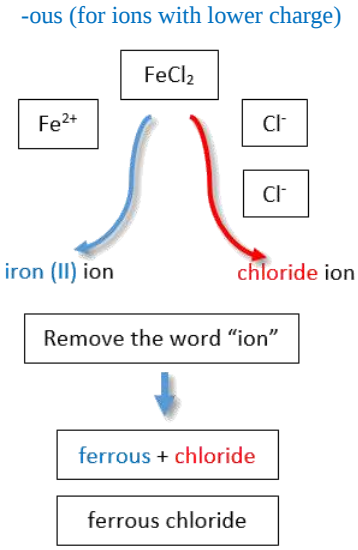


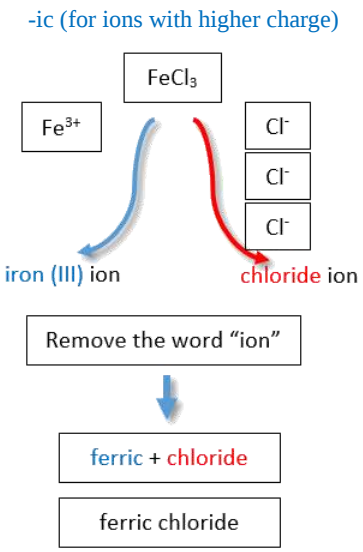




If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure **7.3.3**) .

Table **7.3.4**: 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 7.3.4:**

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.<br>• Oxide always has a 2− charge, so with three oxide ions, we have a total negative charge of 6−.<br>• This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+.<br>• 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.<br>• Chloride always has a 1− charge, so with two chloride ions, we have a total negative charge of 2−.<br>• This means that the one iron ion must have a 2+ charge.<br>• Therefore, the proper name for this ionic compound is iron(II) chloride. | iron(II) chloride |

**Exercise 7.3.4**

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 **7.3.6**.

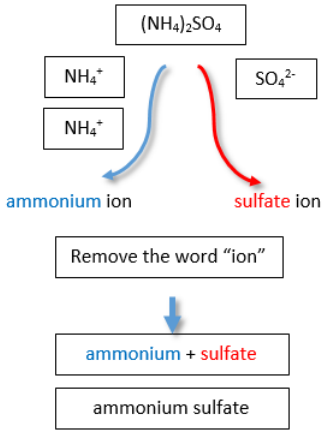



Figure **7.3.2**: Naming Ionic Compounds with Polyatomic Ions

## Example 7.3.5: 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.<br>Two ammonium ions need to balance the charge on a single sulfide ion.<br>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.<br>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.<br>To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed .<br>The compound's name is iron(II) phosphate. | iron(II) phosphate |

## Exercise 7.3.5A

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 **7.3.1** is a synopsis of how to name simple ionic compounds.

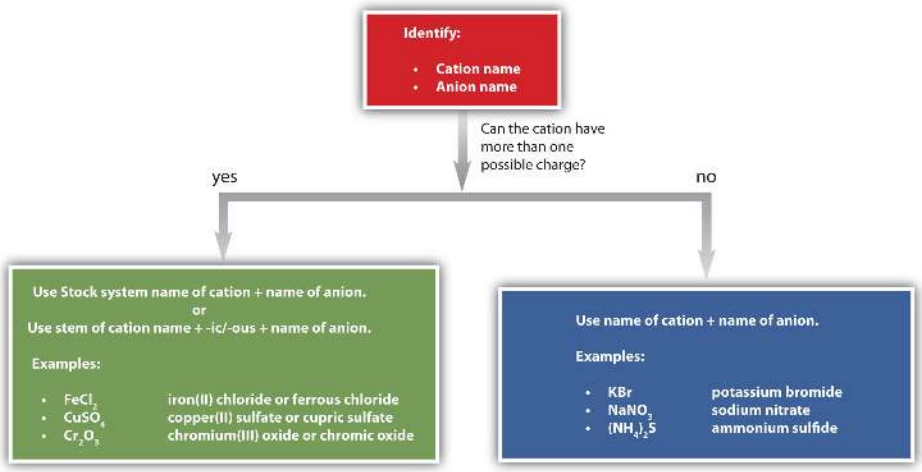



Figure **7.3.3**: A Guide to Naming Simple Ionic Compounds.

## ? Exercise 7.3.5*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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# 7.4: 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.

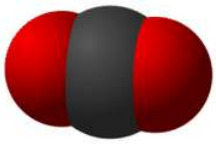


Figure 7.4.1: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

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 $C_8H_{18}$.



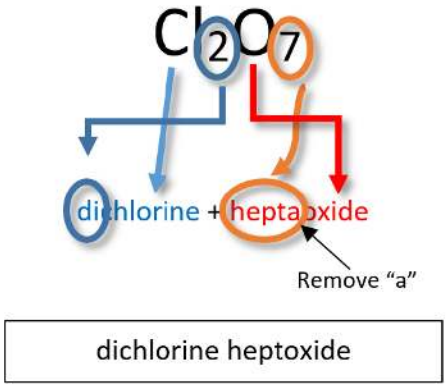
Figure 7.4.2: Nitrogen dioxide ($NO_2$) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

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 7.4.1 lists these numerical prefixes.

Table 7.4.1: 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 **7.4.2**.

Table **7.4.2**

| 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 7.4.1**

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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# 7.5: 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 7.5.1: (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:

$$HCl \rightarrow H^+ + Cl^-$$

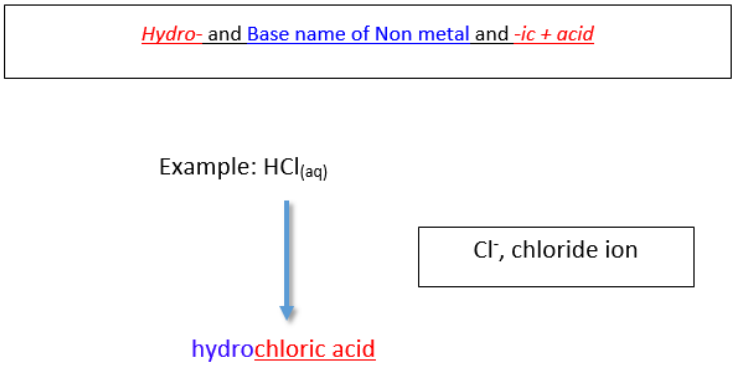
Since acids produce $H^+$ cations upon dissolving in water, the $H$ of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the $H$) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed 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*.



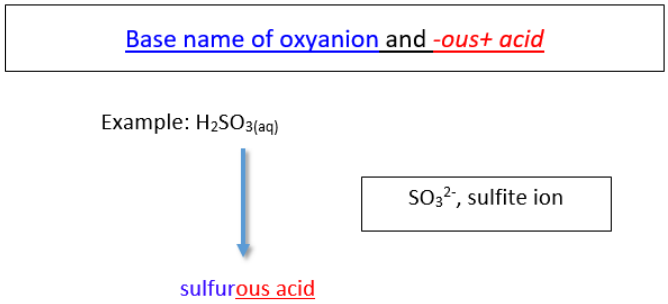

Formula for naming acids: Hydro- and Base name of nonmetal and -ic + acid. Example: HCl is hydrochloric acid.

## 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.

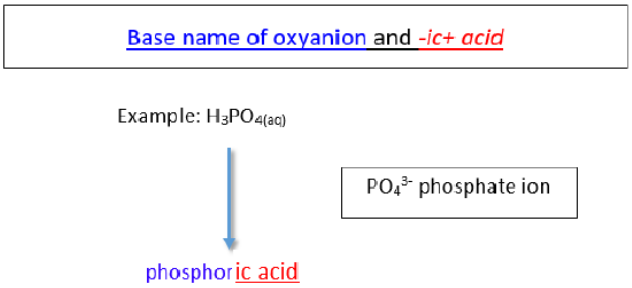
The name of the acid is the root of the anion followed by the suffix ***-ous***. There is no prefix.




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.



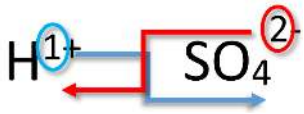

Formula for naming oxyanions with -ate ending: Base name of oxyanion and -ic + acid. Example: H3PO4 is phosphoric acid.

**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 $\mathbf{H^+}$ 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 7.5.2:*** *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.

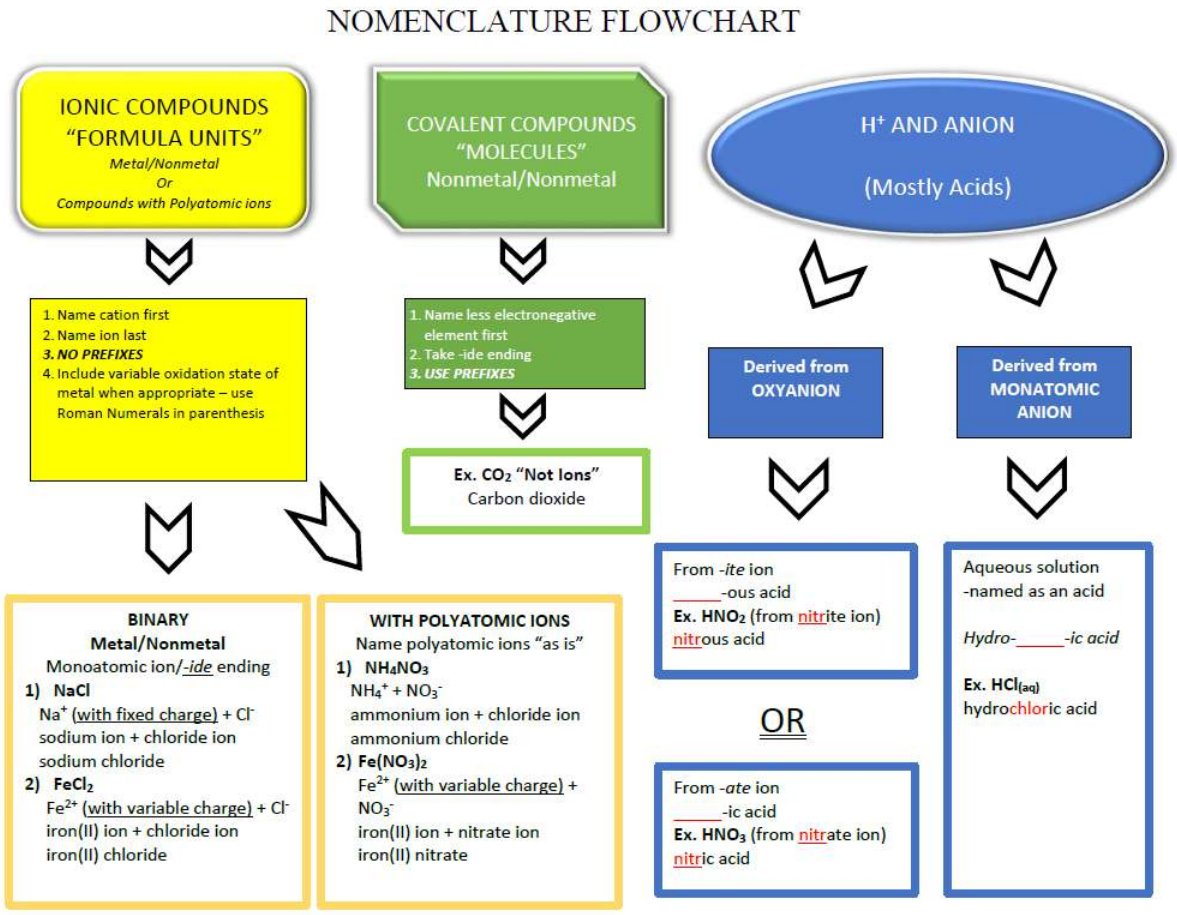
- The name of the acid is based on the anion attached to the hydrogen.

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# 7.6: Nomenclature Summary




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# CHAPTER OVERVIEW

## 8: Chemical Reactions


8.1: Introduction to Chemical Reactions

8.2: The Chemical Equation

8.3: How to Write Balanced Chemical Equations

8.4: Aqueous Solutions and Solubility - Compounds Dissolved in Water

8.5: Precipitation Reactions

8.6: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

8.7: Some Types of Chemical Reactions

8.8: Oxidation-Reduction (Redox) Reactions


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# 8.1: Introduction to Chemical Reactions

**Learning Objectives**

- Identify the evidence for chemical reactions.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

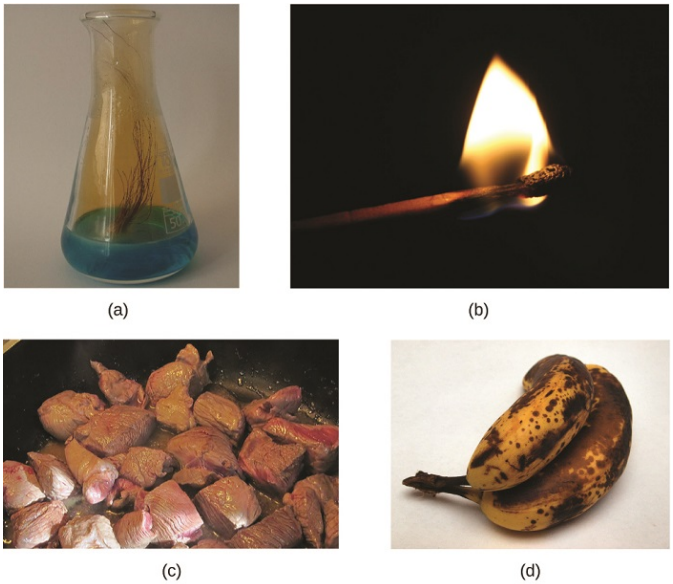



Figure **8.1.1**: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (Credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo.)

To identify a chemical reaction, we look for a **chemical change**. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include: reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure **8.1.1**).



***Video*8.1.1*:** *Evidence of a Chemical Reaction*

### Example 8.1.1: Evidence of a Chemical Reaction

Which of the following is a chemical reaction?

a. Freezing liquid mercury.
b. Adding yellow to blue to make green.
c. Cutting a piece of paper into two pieces.
d. Dropping a sliced orange into a vat of sodium dydroxide.
e. Filling a balloon with natural air.

**Solution**

A, B, C, & E involve only physical changes. A sliced orange has acid (citric acid) that can react with sodium hydroxide, so the answer is D.

### Exercise 8.1.1

Which of the following is a chemical reaction?

a. Painting a wall blue.
b. A bicycle rusting.
c. Ice cream melting.
d. Scratching a key across a desk.
e. Making a sand castle.

**Answer**

B

### Example 8.1.2: Evidence of a Chemical Reaction

Which of the following is not a chemical reaction?

a. Shattering glass with a baseball.
b. Corroding metal.
c. Fireworks exploding.
d. Lighting a match.
e. Baking a cake.

**Solution**

Shattering glass with a baseball results in glass broken into many pieces but no chemical change happens, so the answer is A.

**? Exercise 8.1.2**

Which of the following is NOT a chemical reaction?

a. Frying an egg.
b. Slicing carrots.
c. A Macbook falling out of a window.
d. Creating ATP in the human body.
e. Dropping a fizzy tablet into a glass of water.

**Answer**

B and C

## Summary

Chemical reactions can be identified via a wide range of different observable factors including change in color, energy change (temperature change or light produced), gas production, formation of precipitate and change in properties.

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# 8.2: The Chemical Equation

**Learning Objectives**

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, $(s)$, $(l)$, $(g)$, $(aq)$, and $\rightarrow$ appropriately when writing a chemical reaction.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

## Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that—what is produced—or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

$$\underbrace{\text{Flour}+\text{Butter}+\text{Sugar}+\text{Vanilla}+\text{Baking Soda}+\text{Eggs}+\text{Chocolate Chips}}_{\text{Ingredients = Reactants}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}}$$

## Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $SO_2 + O_2$, are reactants and sulfur trioxide, $SO_3$, is the product.

$$\underbrace{2\,SO_2\,(g)+O_2\,(g)}_{\text{Reactants}} \rightarrow \underbrace{2\,SO_3\,(g)}_{\text{Products}}$$

In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\text{Reactants} \rightarrow \text{Products}$$

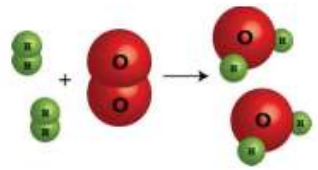
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table **8.2.1** shows a listing of symbols used in chemical equations.

Table **8.2.1**: Symbols Used in Chemical Equations

| Symbol | Description | Symbol | Description |
|---|---|---|---|
| $+$ | used to separate multiple reactants or products | $(s)$ | reactant or product in the solid state |
| $\rightarrow$ | yield sign; separates reactants from products | $(l)$ | reactant or product in the liquid state |
| $\rightleftharpoons$ | replaces the yield sign for reversible reactions that reach equilibrium | $(g)$ | reactant or product in the gas state |
| $\xrightarrow{\text{Pt}}$ | formula written above the arrow is used as a catalyst in the reaction | $(aq)$ | reactant or product in an aqueous solution (dissolved in water) |
| $\xrightarrow{\Delta}$ | triangle indicates that the reaction is being heated | | |

Chemists have a choice of methods for describing a chemical reaction.

1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:
*"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."*

3. They could write the equation in chemical shorthand.

$$2\text{H}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{H}_2\text{O}\,(g)$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:

$$\text{Ca}(\text{NO}_3)_2\,(aq) + 2\text{NaOH}\,(aq) \rightarrow \text{Ca}(\text{OH})_2\,(s) + 2\text{NaNO}_3\,(aq)$$

How much easier is that to read? Let's try it in reverse. Look at the following reaction in shorthand and write the word equation for the reaction:

$$\text{Cu}\,(s) + \text{AgNO}_3\,(aq) \rightarrow \text{Cu}(\text{NO}_3)_2\,(aq) + \text{Ag}\,(s)$$

The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver."

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).

3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table **8.2.1**: *Diatomic Elements*

| **Element Name** | Hydrogen | Nitrogen | Oxygen | Fluorine | Chlorine | Bromine | Iodine |
|---|---|---|---|---|---|---|---|
| **Formula** | $H_2$ | $N_2$ | $O_2$ | $F_2$ | $Cl_2$ | $Br_2$ | $I_2$ |

## Example 8.2.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. $\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$

b. Gaseous propane, $\mathrm{C_3H_8}$, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.

c. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

### Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane ($\mathrm{C_3H_8}$) and oxygen ($\mathrm{O_2}$)

Product: carbon dioxide ($\mathrm{CO_2}$) and water ($\mathrm{H_2O}$)

$$\mathrm{C_3H_8}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathrm{HF}\,(g) + \mathrm{K_2CO_3}\,(aq) \rightarrow \mathrm{KF}\,(aq) + \mathrm{H_2O}\,(l) + \mathrm{CO_2}\,(g)$$

## Exercise 8.2.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.

b. $\mathrm{HCl}\,(aq) + \mathrm{LiOH}\,(aq) \rightarrow \mathrm{LiCl}\,(aq) + \mathrm{H_2O}\,(l)$

c. Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

**Answer a**

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

**Answer b**

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

**Answer c**

$Cu(s) + O_2(g) \rightarrow CuO(s)$

## Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

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# 8.3: How to Write Balanced Chemical Equations

**Learning Objectives**

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

## Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products; and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

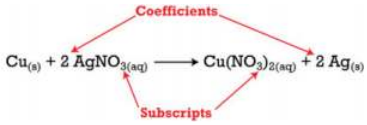
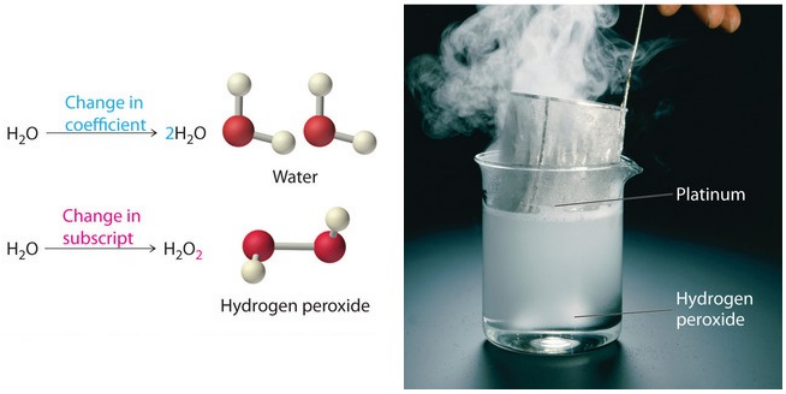



Figure **8.3.1**: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water ($H_2O$) and hydrogen peroxide ($H_2O_2$) are chemically distinct substances. $H_2O_2$ decomposes to $H_2O$ and $O_2$ gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

## Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure **8.3.1**.




Original molecule $\mathbf{H_2O}$: if the coefficient 2 is added in front, that makes 2 water molecules; but if the subscript 2 is added to make $\mathbf{H_2O_2}$, that's hydrogen peroxide.

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

## Steps in Balancing a Chemical Equation

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

## Example 8.3.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane ($C_7H_{16}$).

$$C_7H_{16}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

### Solution

Solutions to Example 8.3.1

| Steps | 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.** | a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 $CO_2$ molecules, each of which contains 1 carbon atom, on the right side: $$C_7H_{16}(l) + O_2(g) \longrightarrow \underline{7}CO_2(g) + H_2O(g)$$ • 7 carbon atoms on both reactant and product sides<br>b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 $H_2O$ molecules, each of which contains 2 hydrogen atoms, on the right side: $$C_7H_{16}(l) + O_2(g) \longrightarrow 7\,CO_2(g) + \underline{8}H_2O(g)$$ • 16 hydrogen atoms on both reactant and product sides |
| 3. **Balance polyatomic ions as a unit.** | There are no polyatomic ions to be considered in this reaction. |
| 4. **Balance the remaining atoms.** | The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, $O_2$, on the reactant side: $$C_7H_{16}(l) + \underline{11}O_2(g) \longrightarrow 7\,CO_2(g) + 8\,H_2O(g)$$ • 22 oxygen atoms on both reactant and product sides |
| 5. **Check your work.** | The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced. |

### Example 8.3.2: Combustion of Isooctane

Combustion of Isooctane ($C_8H_{18}$)

$$C_8H_{18}\,(l) + O_2\,(g) \longrightarrow CO_2\,(g) + H_2O\,(g)$$

**Solution**

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Solutions to Example 7.4.2

| Steps | Example |
|---|---|
| 1. **Identify the most complex substance.** | The most complex substance is the one with the largest number of different atoms, which is $C_8H_{18}$. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance. |
| 2. **Adjust the coefficients.** | a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 $CO_2$ molecules in the products:<br>$$C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$$<br>• 8 carbon atoms on both reactant and product sides<br>b. 18 hydrogen atoms in isooctane means that there must be 9 $H_2O$ molecules in the products:<br>$$C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$$<br>• 18 hydrogen atoms on both reactant and product sides |
| 3. **Balance polyatomic ions as a unit.** | There are no polyatomic ions to be considered in this reaction. |
| 4. **Balance the remaining atoms.** | The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, $O_2$, but because there are 2 oxygen atoms per $O_2$ molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms:<br>$$C_8H_{18}(l) + \underline{\frac{25}{2}}O_2\ (g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(g)$$<br>• 25 oxygen atoms on both reactant and product sides<br>The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**:<br>$$\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$$ |
| 5. **Check your work.** | The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side.<br>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 8.3.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

**Solution**

Solutions to Example 7.4.3

| Steps | Example |
|---|---|
| 1. **Identify the most complex substance.** | The most complex substance is lead (II) chloride.<br>$Pb(NO_3)_2\,(aq) + NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)$ |
| 2. **Adjust the coefficients.** | There are twice as many chloride ions in the reactants as there are in the products. Place a 2 in front of the NaCl in order to balance the chloride ions.<br>$Pb(NO_3)_2\,(aq) + \underline{2}\,NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)$<br>• 1 Pb atom on both reactant and product sides<br>• 2 Na atoms on reactant side, 1 Na atom on product side<br>• 2 Cl atoms on both reactant and product sides |
| 3. **Balance polyatomic ions as a unit.** | The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is:<br>$Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow \underline{2}\,NaNO_3\,(aq) + PbCl_2\,(s)$<br>• 1 Pb atom on both reactant and product sides<br>• 2 Na atoms on both reactant and product sides<br>• 2 Cl atoms on both reactant and product sides<br>• 2 $NO_3^-$ atoms on both reactant and product sides |
| 4. **Balance the remaining atoms.** | There is no need to balance the remaining atoms because they are already balanced. |
| 5. **Check your work.** | $Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow 2\,NaNO_3\,(aq) + PbCl_2\,(s)$<br>• 1 Pb atom on both reactant and product sides<br>• 2 Na atoms on both reactant and product sides<br>• 2 Cl atoms on both reactant and product sides<br>• 2 $NO_3^-$ atoms on both reactant and product sides |

## ? Exercise 8.3.1

Is each chemical equation balanced?

a. $2\,Hg(\ell) + O_2\,(g) \rightarrow Hg_2O_2\,(s)$
b. $C_2H_4\,(g) + 2\,O_2(g) \longrightarrow 2\,CO_2\,(g) + 2\,H_2O\,(g)$
c. $Mg(NO_3)_2\,(s) + 2\,Li\,(s) \rightarrow Mg(s) + 2\,LiNO_3\,(s)$

**Answer a**

yes

**Answer b**

no

**Answer c**

yes

## Exercise 8.3.2

Balance the following chemical equations.

a. $\mathbf{N_2\,(g) + O_2\,(g) \longrightarrow NO_2\,(g)}$
b. $\mathbf{Pb(NO_3)_2\,(aq) + FeCl_3\,(aq) \longrightarrow Fe(NO_3)_3\,(aq) + PbCl_2\,(s)}$
c. $\mathbf{C_6H_{14}\,(l) + O_2(g) \longrightarrow CO_2\,(g) + H_2O\,(g)}$

**Answer a**

$N_2\,(g) + 2O_2\,(g) \rightarrow 2NO_2\,(g)$

**Answer b**

$3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)$

**Answer c**

$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

## Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

## Vocabulary

- **Chemical reaction** - The process in which one or more substances are changed into one or more new substances.
- **Reactants** - The starting materials in a reaction.
- **Products** - Materials present at the end of a reaction.
- **Balanced chemical equation** - A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- **Subscripts** - Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- **Coefficient** - A small whole number that appears in front of a formula in a balanced chemical equation.

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# 8.4: Aqueous Solutions and Solubility - Compounds Dissolved in Water

**Learning Objectives**

- Define and give examples of electrolytes.

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called electrolytes. Substances that do not yield ions when dissolved are called nonelectrolytes. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 8.4.1).

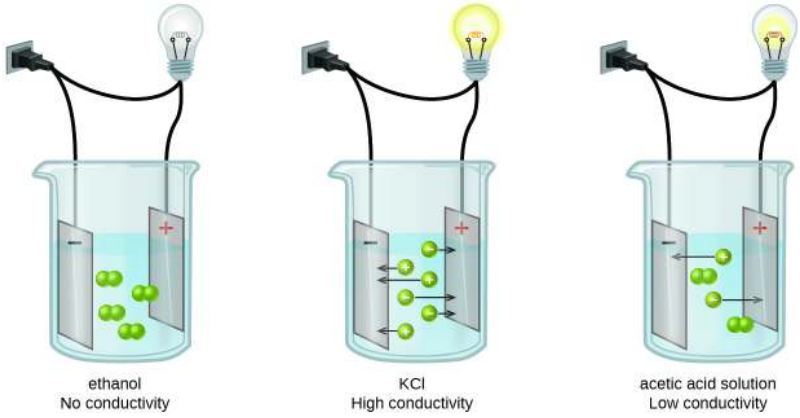



Figure 8.4.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte. This diagram shows three separate beakers. Each has a wire plugged into a wall outlet. In each case, the wire leads from the wall to the beaker and is split resulting in two ends. One end leads to a light bulb and continues on to a rectangle labeled with a plus sign. The other end leads to a rectangle labeled with a minus sign. The rectangles are in a solution. In the first beaker, labeled “Ethanol No Conductivity,” four pairs of linked small green spheres suspended in the solution between the rectangles. In the second beaker, labeled “K C l Strong Conductivity,” six individual green spheres, three labeled plus and three labeled minus are suspended in the solution. Each of the six spheres has an arrow extending from it pointing to the rectangle labeled with the opposite sign. In the third beaker, labeled “Acetic acid solution Weak conductivity,” two pairs of joined green spheres and two individual spheres, one labeled plus and one labeled minus are shown suspended between the two rectangles. The plus labeled sphere has an arrow pointing to the rectangle labeled minus and the minus labeled sphere has an arrow pointing to the rectangle labeled plus.

## Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in Figure 8.4.2. The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water.

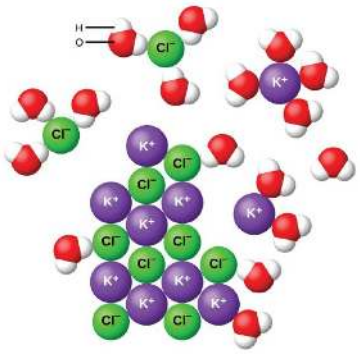


Figure 8.4.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the $K^+$ and $Cl^-$ ions. Water molecules in front of and behind the ions are not shown. The diagram shows eight purple spheres labeled K superscript plus and eight green spheres labeled C l superscript minus mixed and touching near the center of the diagram. Outside of this cluster of spheres are seventeen clusters of three spheres, which include one red and two white spheres. A red sphere in one of these clusters is labeled O. A white sphere is labeled H. Two of the green C l superscript minus spheres are surrounded by three of the red and white clusters, with the red spheres closer to the green spheres than the white spheres. One of the K superscript plus purple spheres is surrounded by four of the red and white clusters. The white spheres of these clusters are closest to the purple spheres.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. The water molecules penetrate between individual $K^+$ and $Cl^-$ ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure 8.4.2 shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system, as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

## Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all. It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (Tables 8.4.1 and 8.4.2).

Table 8.4.1: Solubility Rules for Soluble Substances

| Soluble in Water | Important Exceptions (Insoluble) |
|---|---|
| All Group IA and $NH_4^+$ salts | none |
| All nitrates, chlorates, perchlorates and acetates | none |
| All sulfates | $CaSO_4$, $BaSO_4$, $SrSO_4$, $PbSO_4$ |
| All chlorides, bromides, and iodides | AgX, $Hg_2X_2$, $PbX_2$ (X= Cl, Br, or I) |

Table 8.4.2: Solubility Rules for Sparingly Soluble Substances

| Sparingly Soluble in Water | Important Exceptions (Soluble) |
|---|---|
| All carbonates and phosphates | Group IA and $NH_4^+$ salts |
| All hydroxides | Group IA and $NH_4^+$ salts; $Ba^{2+}$, $Sr^{2+}$, $Ca^{2+}$ sparingly soluble |
| All sulfides | Group IA, IIA and $NH_4^+$ salts; MgS, CaS, BaS sparingly soluble |
| All oxalates | Group IA and $NH_4^+$ salts |
| Special note: The following electrolytes are of only moderate solubility in water: $CH_3COOAg$, $Ag_2SO_4$, $KClO_4$. They will precipitate only if rather concentrated solutions are used. | |

As an example on how to use the solubility rules, predict if a precipitate will form when solutions of cesium bromide and lead (II) nitrate are mixed.

$$Cs^+(aq) + Br^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow ?$$

The potential precipitates from a double-replacement reaction are cesium nitrate and lead (II) bromide. According to the solubility rules table, cesium nitrate is soluble because all compounds containing the nitrate ion, as well as all compounds containing the alkali metal ions, are soluble. Most compounds containing the bromide ion are soluble, but lead (II) is an exception. Therefore, the cesium and nitrate ions are spectator ions and the lead (II) bromide is a precipitate. The balanced net ionic reaction is:

$$Pb^{2+}(aq) + 2Br^-(aq) \rightarrow PbBr_2(s)$$

### ✔ Example 8.4.1: Solubility

Classify each compound as soluble or insoluble

a. $Zn(NO_3)_2$
b. $PbBr_2$
c. $Sr_3(PO_4)_2$

**Solution**

a. All nitrates are soluble in water, so $Zn(NO_3)_2$ is soluble.
b. All bromides are soluble in water, except those combined with $Pb^{2+}$, so $PbBr_2$ is insoluble.
c. All phosphates are insoluble, so $Sr_3(PO_4)_2$ is insoluble.

### ? Exercise 8.4.1: Solubility

Classify each compound as soluble or insoluble.

a. $Mg(OH)_2$
b. KBr
c. $Pb(NO_3)_2$

**Answer a**

insoluble

**Answer b**

soluble

**Answer c**

soluble

## Summary

Substances that dissolve in water to yield ions are called electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water. Solubility rules allow prediction of what products will be insoluble in water.

## Contributions & Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
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# 8.5: Precipitation Reactions

**Learning Objectives**

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate, a reddish precipitate of silver dichromate is produced.

$$\mathrm{AgNO_3\,(aq) + K_2Cr_2O_7\,(aq) \rightarrow Ag_2Cr_2O_7\,(s) + KNO_3\,(aq)} \tag{8.5.1}$$

This unbalanced equation has the general form of an exchange reaction:

$$\mathrm{AC + BD \rightarrow \underset{insoluble}{AD} + BC} \tag{8.5.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called **double-displacement reactions**. Precipitation reactions are used to isolate metals that have been extracted from their ores, and to recover precious metals for recycling.



Video: Mixing potassium dichromate and silver nitrate together to initiate a precipitation reaction (Equation 8.5.1).

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of aqueous $\mathit{NaCl}$ solution is mixed with 500 mL of aqueous $\mathit{KBr}$ solution, the final solution has a volume of 1.00 L and contains $\mathrm{Na^+\,(aq)}$, $\mathrm{Cl^-\,(aq)}$, $\mathrm{K^+\,(aq)}$, and $\mathrm{Br^-\,(aq)}$. As you will see in (Figure 8.5.1), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

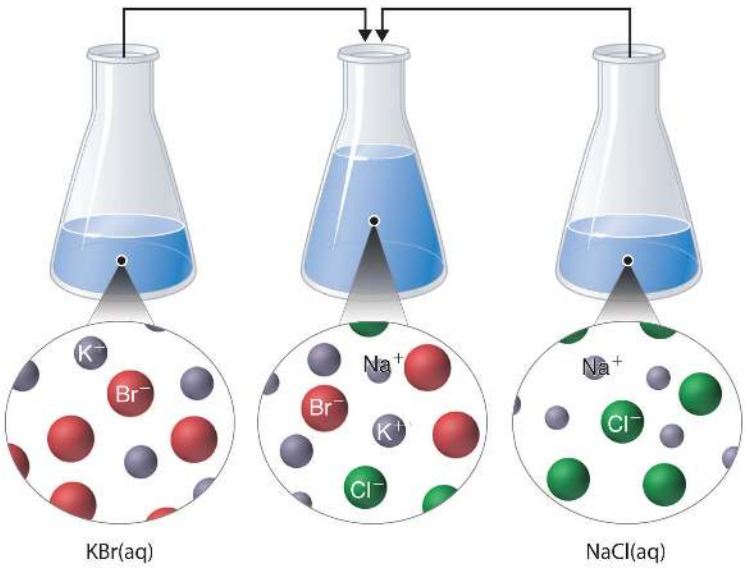


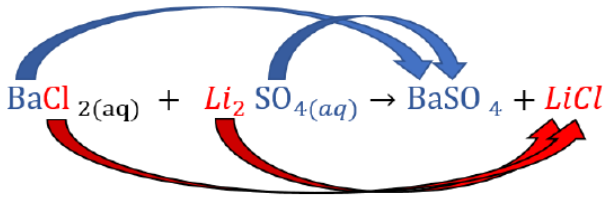

Figure 8.5.1: The Effect of Mixing Aqueous **KBr** and **NaCl** Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

## Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurred.

Predict what will happen when aqueous solutions of barium chloride and lithium sulfate are mixed.

**Change the partners of the anions and cations on the reactant side to form new compounds (products):**




Chemical equation of the reactants barium chloride and lithium sulfate forming the products barium sulfate and lithium chloride.

Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains $Ba^{2+}$, $Cl^-$, $Li^+$, and $SO_4^{2-}$ ions. The only possible exchange reaction is to form **LiCl** and **$BaSO_4$**.

### Correct the formulas of the products based on the charges of the ions.

No need to correct the formula as both compounds already have their charges balanced.

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4 + LiCl$$

### Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + LiCl\,(aq)$$

Table 7.5.1 from the previous section shows that **LiCl** is soluble in water, but **$BaSO_4$** is not soluble in water.

### Balance the equation:

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + 2\,LiCl\,(aq)$$

Although soluble barium salts are toxic, **$BaSO_4$** is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a "barium milkshake" or a "barium enema"—a suspension of very fine **$BaSO_4$** particles in water.



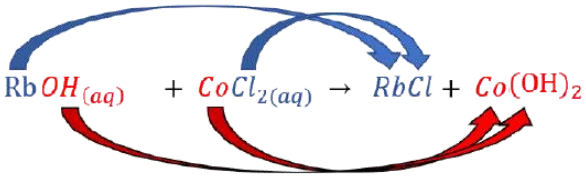
Figure 8.5.2: An x-ray of the digestive organs of a patient who has swallowed a “barium milkshake.” A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. (Public Domain; Glitzy queen00 via Wikipedia).

## Example 8.5.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

### Solution

Solutions to Example 7.6.1

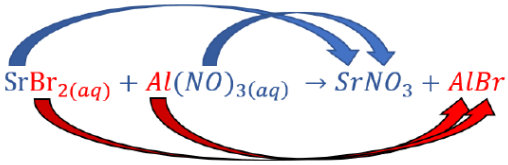
| Steps | Example |
|---|---|
| Change the partners of the anions and cations on the reactant side to form new compounds (products). |  Chemical equation of the reactants rubidium hydroxide and cobalt(II) chloride forming the products rubidium chloride and cobalt hydroxide. |
| Correct the formulas of the products based on the charges of the ions. | $\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl + Co(OH)_2}$ |
| Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate. | $\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl\,(aq) + Co(OH)_2\,(s)}$ |
| Balance the equation. | Coefficients already balanced. $\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl\,(aq) + Co(OH)_2\,(s)}$ |

## Example 8.5.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

### Solution

Solutions for Example 7.6.2

| Steps | Example |
|---|---|
| Change the partners of the anions and cations on the reactant side to form new compounds (products). |  Chemical equation of the reactants strontium bromide and aluminum nitrate forming the products strontium nitrate and aluminum bromide. |
| Correct the formulas of the products based on the charges of the ions. | $\mathbf{SrBr_2\,(aq) + Al(NO_3)_3\,(aq) \rightarrow Sr(NO_3)_2 + AlBr_3}$ |
| Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate. | $\mathbf{SrBr_2\,(aq) + Al(NO_3)_3\,(aq) \rightarrow Sr(NO_3)_2\,(aq) + AlBr_3\,(aq)}$ According to Table 7.5.1 from the previous section, both $\mathbf{AlBr_3}$ (rule 4) and $\mathbf{Sr(NO_3)_2}$ (rule 2) are soluble. |
| If all possible products are soluble, then no net reaction will occur. | $\mathbf{SrBr_2\,(aq) + Al(NO_3)_3\,(aq) \rightarrow}$ **NO REACTION** |

## ? Exercise 8.5.2

Using the information in Table 7.5.1 from the previous section, predict what will happen in each case involving strong electrolytes.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

**Answer a**

$Fe(OH)_2$ precipitate is formed.

**Answer b**

$Hg_3(PO_4)_2$ precipitate is formed.

**Answer c**

No Reaction.

**Answer d**

$CaCO_3$ is precipitate formed.

## Summary

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

## Contributions & Attributions

- Modified by Joshua Halpern (Howard University)

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# 8.6: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

A typical precipitation reaction occurs when an aqueous solution of barium chloride is mixed with one containing sodium sulfate. The **complete chemical equation** can be written to describe what happens, and such an equation is useful in making chemical calculations.

$$\underbrace{BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow BaSO_4\,(s) + 2\,NaCl\,(aq)}_{\text{Complete Chemical Equation}} \tag{8.6.1}$$

However, Equation 8.6.1 does not really represent the microscopic particles (that is, the ions) present in the solution. Below is the complete ionic equation:

$$\underbrace{Ba^{2+}\,(aq) + \overbrace{2Cl^-\,(aq)}^{\text{spectator}} + \overbrace{2Na^+\,(aq)}^{\text{spectator}} + SO_4{}^{2-}\,(aq) \longrightarrow BaSO_4\,(s) + \overbrace{2Na^+\,(aq)}^{\text{spectator}} + \overbrace{Cl^-\,(aq)}^{\text{spectator}}}_{\text{Complete Ionic Equation}} \tag{8.6.2}$$

Equation 8.6.2 is rather cumbersome and includes so many different ions that it may be confusing. In any case, we are often interested in the independent behavior of ions, not the specific compound from which they came. A precipitate of $BaSO_4\,(s)$ will form when *any* solution containing $Ba^{2+}\,(aq)$ is mixed with *any* solution containing $SO_4{}^{2-}\,(aq)$ (provided concentrations are not extremely small). This happens independently of the $Cl^-\,(aq)$ and $Na^+\,(aq)$ ions in Equation 8.6.2. These ions are called **spectator ions** because they do not participate in the reaction. When we want to emphasize the independent behavior of ions, a **net ionic equation** is written, omitting the spectator ions. For precipitation of $BaSO_4$ the net ionic equation is

$$\underbrace{Ba^{2+}\,(aq) + SO_4{}^{2-}\,(aq) \longrightarrow BaSO_4\,(s)}_{\text{Net Ionic Equation}} \tag{8.6.3}$$

## Example 8.6.1

a. When a solution of $AgNO_3$ is added to a solution of $CaCl_2$, insoluble $AgCl$ precipitates. Write three equations (complete chemical equation, complete ionic equation, and net ionic equation) that describe this process.

b. Write the balanced net ionic equation to describe any reaction that occurs when the solutions of $Na_2SO_4$ and $NH_4I$ are mixed.

### Solution

Solutions to Example 7.7.1

| Equation Type | Example 8.6.1*a* | Example 8.6.1*b* |
|---|---|---|
| Complete Chemical Equation | $2\,AgNO_3\,(aq) + CaCl_2\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)$<br>The proper states and formulas of all products are written and the chemical equation is balanced. | $Na_2SO_4\,(aq) + 2\,NH_4I\,(aq) \longrightarrow 2\,NaI\,(aq) + (NH_4)_2SO_4\,(aq)$<br>Both products are aqueous so there is no net ionic equation that can be written. |
| Complete Ionic Equation | $2\,Ag^+\,(aq) + 2\,NO_3{}^-\,(aq) + Ca^{2+}\,(aq) + 2\,Cl^-\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca^{2+}\,(aq) + 2\,NO_3{}^-\,(aq)$<br>AgCl is a solid so it does not break up into ions in solution. | |
| Net Ionic Equation | $Ag^+\,(aq) + Cl^-\,(aq) \longrightarrow AgCl\,(s)$<br>All spectator ions are removed.<br>The chemical equation is written using the lowest common coefficients. | $NaI$ and $(NH_4)_2SO_4$ are both soluble.<br>No net ionic equation.<br>There is no reaction. |

The occurrence or nonoccurrence of precipitates can be used to detect the presence or absence of various species in solution. A $BaCl_2$ solution, for instance, is often used as a test for the presence of $SO_4^{2-}(aq)$ ions. There are several insoluble salts of Ba, but they all dissolve in dilute acid except for $BaSO_4$. Thus, if $BaCl_2$ solution is added to an unknown solution which has previously been acidified, the occurrence of a white precipitate is proof of the presence of the $SO_4^{2-}$ ion.



Figure 8.6.1: The three common silver halide precipitates: AgI, AgBr and AgCl (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

$AgNO_3$ solutions are often used in a similar way to test for halide ions. If $AgNO_3$ solution is added to an acidified unknown solution, a white precipitate indicates the presence of $Cl^-$ ions, a cream-colored precipitate indicates the presence of $Br^-$ ions, and a yellow precipitate indicates the presence of $I^-$ ions (Figure 8.6.1). Further tests can then be made to see whether perhaps a mixture of these ions is present. When $AgNO_3$ is added to tap water, a white precipitate is almost always formed. The $Cl^-$ ions in tap water usually come from the $Cl_2$ which is added to municipal water supplies to kill microorganisms.

**? Exercise 8.6.1**

Write balanced net ionic equations to describe any reaction that occurs when the following solutions are mixed.

a. $K_2CO_3 + SrCl_2$
b. $FeSO_4 + Ba(NO_3)_2$

**Answer a**

$$Sr^{2+}(aq) + CO_3^{2-}(aq) \longrightarrow SrCO_3(s)$$

**Answer b**

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Ba(SO_4)(s)$$

Precipitates are also used for quantitative analysis of solutions, that is, to determine the amount of solute or the mass of solute in a given solution. For this purpose it is often convenient to use the first of the three types of equations described above. Then the rules of stoichiometry may be applied.

## Contributions & Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

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# 8.7: Some Types of Chemical Reactions

**Learning Objectives**

- To classify a given chemical reaction into a variety of types.

Although there are untold millions of possible chemical reactions, most can be classified into a small number of general reaction types. Classifying reactions has two purposes: it helps us to recognize similarities among them, and it enables us to predict the products of certain reactions. A particular reaction may fall into more than one of the categories that we will define in this book.

## Combination (composition) Reactions

A combination (composition) reaction is a chemical reaction that makes a single substance from two or more reactants. There may be more than one molecule of product in the balanced chemical equation, but there is only one substance produced. For example, the equation

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

is a combination reaction that produces $Fe_2O_3$ from its constituent elements — $Fe$ and $O_2$. Combination reactions do not have to combine elements, however. The chemical equation

$$Fe_2O_3 + 3\,SO_3 \rightarrow Fe_2(SO_4)_3$$

shows a combination reaction in which $Fe_2O_3$ combines with three molecules of $SO_3$ to make $Fe_2(SO_4)_3$.

**Example 8.7.1**

Which equations are combination reactions?

a. $Co\,(s) + Cl_2\,(g) \longrightarrow CoCl_2\,(s)$
b. $CO\,(g) + Cl_2\,(g) \longrightarrow COCl_2\,(g)$
c. $N_2H_4(\ell) + O_2\,(g) \longrightarrow N_2\,(g) + 2\,H_2O(\ell)$

**Solution**

a. This is a combination reaction.
b. This is a combination reaction. (The compound $COCl_2$ is called phosgene and, in the past, was used as a gassing agent in chemical warfare.)
c. This is not a combination reaction.

**Exercise 8.7.1**

Which equations are combination reactions?

a. $P_4\,(s) + 6\,Cl_2\,(g) \longrightarrow 4\,PCl_3\,(g)$
b. $SO_3(\ell) + H_2O(\ell) \longrightarrow H_2SO_4(\ell)$
c. $NaOH\,(s) + HCl\,(g) \longrightarrow NaCl\,(s) + H_2O(\ell)$

**Answer**

a and b are combination reactions

## Decomposition Reactions

A decomposition reaction is the reverse of a combination reaction. In a decomposition reaction, a single substance is converted into two or more products. There may be more than one molecule of the reactant, but there is only one substance initially. For example, the equation

$$2\,NaHCO_3\,(s) \rightarrow Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(\ell)$$

is a decomposition reaction that occurs when $NaHCO_3$ is exposed to heat. Another example is the decomposition of $KClO_3$:

$$2\,KClO_3\,(s) \rightarrow 2\,KCl\,(s) + 3\,O_2\,(g)$$

This reaction was once commonly used to generate small amounts of oxygen in the chemistry lab.

> The decomposition reaction of $NaHCO_3$ is the reaction that occurs when baking soda is poured on a small kitchen fire. The intent is that the $H_2O$ and $CO_2$ produced by the decomposition will smother the flames.

## Combustion Reactions

A combustion reaction occurs when a substance combines with molecular oxygen to make oxygen-containing compounds of other elements in the reaction. Many combustion reactions occur with a hydrocarbon, a compound made up solely of carbon and hydrogen. The products of the combustion of hydrocarbons are carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amounts of heat energy. An example is propane ($C_3H_8$), a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$C_3H_8\,(g) + 5\,O_2\,(g) \longrightarrow 3\,CO_2\,(g) + 4\,H_2O\,(g)$$

Another example is the burning of acetylene ($C_2H_2$) in welding torches:

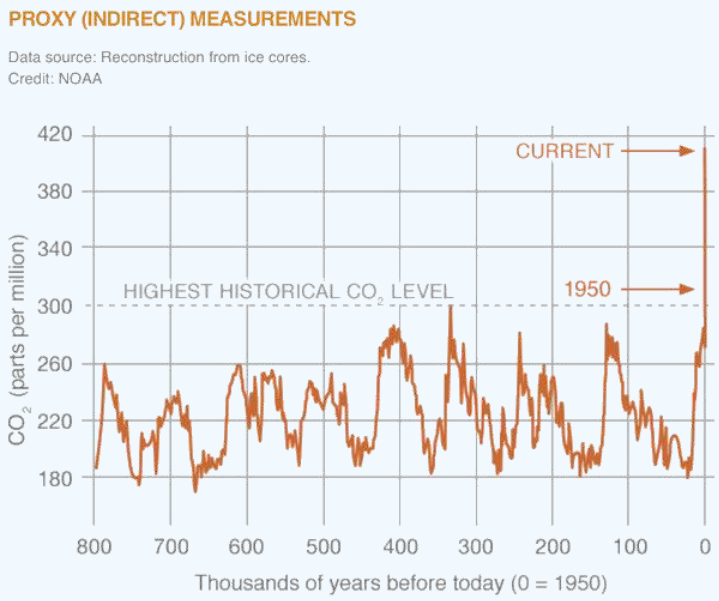
$$2\,C_2H_2 + 5\,O_2 \rightarrow 4\,CO_2 + 2\,H_2O$$

Oxygen (in its elemental form) is a crucial reactant in combustion reactions.

> *Energy in the form of heat is usually given off as a product in a combustion reaction as well.*

**Example 8.7.2**

Carbon dioxide ($CO_2$) is an important heat-trapping (greenhouse) gas, which is released through human activities such as deforestation and burning fossil fuels, as well as natural processes such as respiration and volcanic eruptions. The graph above shows $CO_2$ levels during the last three glacial cycles, as reconstructed from ice cores.



Carbon dioxide is the primary greenhouse gas emitted through human activities. In 2015, $CO_2$ accounted for about 82.2% of all U.S. greenhouse gas emissions from human activities. Carbon dioxide is naturally present in the atmosphere as part of the Earth's carbon cycle (the natural circulation of carbon among the atmosphere, oceans, soil, plants, and animals). Human activities are altering the carbon cycle–both by adding more $CO_2$ to the atmosphere and by influencing the ability of natural sinks, like forests, to remove $CO_2$ from the atmosphere. While $CO_2$ emissions come from a variety of natural sources, human-related emissions are responsible for the increase that has occurred in the atmosphere since the industrial revolution.

The main human activity that emits $CO_2$ is the combustion of fossil fuels (coal, natural gas, and oil) for energy and transportation, although certain industrial processes and land-use changes also emit $CO_2$. An example of how $CO_2$ can be generated is the **combustion of octane** ($C_8H_{10}$), a component of gasoline.

Write the balanced equation to represent the combustion of octane.

### Solution

$$2\,C_8H_{18}\,(l) + 25\,O_2\,(g) \longrightarrow 16\,CO_2\,(g) + 18\,H_2O\,(g)$$

The balanced reaction argues that for every two molecules of octane that are burned, 16 molecules of $CO_2$ are generated.

## Example 8.7.2

Identify each type of reaction.

a. $2\,K\,(s) + S\,(s) + 2\,O_2\,(g) \longrightarrow K_2SO_4\,(s)$
b. $(NH_4)_2Cr_2O_7\,(s) \longrightarrow N_2\,(g) + Cr_2O_3\,(s) + 4\,H_2O(\ell)$
c. $CH_4\,(g) + 2\,O_2\,(g) \longrightarrow CO_2\,(g) + 2\,H_2O(\ell)$

### Solution

a. Multiple reactants are combining to make a single product, so this reaction is a combination reaction.
b. A single substance reacts to make several products, so we have a decomposition reaction.
c. Oxygen reacts with a compound to make carbon dioxide (an oxide of carbon) and water (an oxide of hydrogen). This is a combustion reaction.

## Exercise 8.7.2

Identify each type of reaction.

a. $C_2H_5OH + 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$
b. $2\,Mg\,(s) + O_2\,(g) \longrightarrow 2\,MgO\,(s)$
c. $CaCO_3\,(s) \longrightarrow CaO\,(s) + CO_2\,(g)$

**Answer a**

combustion

**Answer b**

combination (also combustion)

**Answer c**

decomposition

## Key Takeaway

There are several recognizable types of chemical reactions: combination, decomposition, and combustion reactions are examples.

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https://chem.libretexts.org/@go/page/501335

# 8.8: Oxidation-Reduction (Redox) Reactions

**Learning Objectives**

- To identify a chemical reaction as an oxidation-reduction reaction.

When zinc metal is submerged into a quantity of aqueous $\mathbf{HCl}$, the following reaction occurs (Figure 8.8.1):

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow H_2\,(g) + ZnCl_2\,(aq)} \tag{8.8.1}$$

This is one example of what is sometimes called a *single replacement reaction* because $\mathbf{Zn}$ replaces $\mathbf{H}$ in combination with $\mathbf{Cl}$.



Figure 8.8.1: Zinc Metal plus Hydrochloric Acid. It is fairly obvious that zinc metal reacts with aqueous hydrochloric acid! The bubbles are hydrogen gas (right side of Equation 8.8.1).

Because some of the substances in this reaction are aqueous, we can separate them into ions:

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) + 2\,Cl^-\,(aq) \longrightarrow H_2\,(g) + Zn^{2+}\,(aq) + 2\,Cl^-\,(aq)}$$

Viewed this way, the net reaction seems to be a charge transfer between zinc and hydrogen atoms. (There is no net change experienced by the chloride ion.) In fact, electrons are being transferred from the zinc atoms to the hydrogen atoms (which ultimately make a molecule of diatomic hydrogen), changing the charges on both elements.

To understand electron-transfer reactions like the one between zinc metal and hydrogen ions, chemists separate them into two parts: one part focuses on the loss of electrons, and one part focuses on the gain of electrons. The loss of electrons is called oxidation. The gain of electrons is called reduction. Because any loss of electrons by one substance must be accompanied by a gain in electrons by something else, oxidation and reduction always occur together. As such, electron-transfer reactions are also called oxidation-reduction reactions, or simply **redox reactions**. The atom that loses electrons is **oxidized**, and the atom that gains electrons is **reduced**. Also, because we can think of the species being oxidized as causing the reduction, the species being oxidized is called the **reducing agent**, and the species being reduced is called the **oxidizing agent**.

> *Because batteries are used as sources of electricity (that is, of electrons), all batteries are based on redox reactions.*

Although the two reactions occur together, it can be helpful to write the oxidation and reduction reactions separately as half reactions. In half reactions, we include only the reactant being oxidized or reduced, the corresponding product species, any other species needed to balance the half reaction, and the electrons being transferred. Electrons that are lost are written as products; electrons that are gained are written as reactants. For example, in our earlier equation, now written without the chloride ions,

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) \longrightarrow Zn^{2+}\,(aq) + H_2\,(g)}$$

zinc atoms are oxidized to $Zn^{2+}$. The **half reaction for the oxidation** reaction, omitting phase labels, is as follows:

$$\mathrm{Zn \longrightarrow Zn^{2+} + 2\,e^-}$$

This half reaction is balanced in terms of the number of zinc atoms, and it also shows the two electrons that are needed as products to account for the zinc atom losing two negative charges to become a 2+ ion. With half reactions, there is one more item to balance:

the overall charge on each side of the reaction. If you check each side of this reaction, you will note that both sides have a zero net charge.

Hydrogen is reduced in the reaction. The balanced **reduction half reaction** is as follows:

$$2\,H^+ + 2\,e^- \longrightarrow H_2$$

There are two hydrogen atoms on each side, and the two electrons written as reactants serve to neutralize the 2+ charge on the reactant hydrogen ions. Again, the overall charge on both sides is zero.

The overall reaction is simply the combination of the two half reactions and is shown by adding them together.

$$\begin{aligned} &Zn \rightarrow Zn^{2+} + \cancel{2e^-} \\ &\underline{2H^+ + \cancel{2e^-} \rightarrow H_2} \\ &Zn + 2H^+ \rightarrow Zn^{2+} + H_2 \end{aligned}$$

Because we have two electrons on each side of the equation, they can be canceled. This is the key criterion for a balanced redox reaction: the electrons have to cancel exactly. If we check the charge on both sides of the equation, we see they are the same—2+. (In reality, this positive charge is balanced by the negative charges of the chloride ions, which are not included in this reaction because chlorine does not participate in the charge transfer.)

Redox reactions are often balanced by balancing each individual half reaction and then combining the two balanced half reactions. **Sometimes a half reaction must have all of its coefficients multiplied by some integer for all the electrons to cancel**. The following example demonstrates this process.

### Example 8.8.1: Reducing Silver Ions

Write and balance the redox reaction that has silver ions and aluminum metal as reactants and silver metal and aluminum ions as products. Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

#### Solution

We start by using symbols of the elements and ions to represent the reaction:

$$Ag^+ + Al \longrightarrow Ag + Al^{3+}$$

The equation looks balanced as it is written. However, when we compare the overall charges on each side of the equation, we find a charge of +1 on the left but a charge of +3 on the right. This equation is not properly balanced. To balance it, let us write the two half reactions. Silver ions are reduced, and it takes one electron to change $Ag^+$ to $Ag$:

**Reduction half-reaction**:

$$Ag^+ + e^- \longrightarrow Ag$$

Aluminum is oxidized, losing three electrons to change from $Al$ to $Al^{3+}$:

**Oxidation half-reaction**:

$$Al \longrightarrow Al^{3+} + 3\,e^-$$

To combine these two half reactions and cancel out all the electrons, we need to multiply the silver reduction reaction by 3:

$$\begin{aligned} &3\left(Ag^+ + e^- \rightarrow Ag\right) \\ &\underline{\qquad Al \rightarrow Al^{3+} + 3e^-} \\ &3Ag^+ + Al \rightarrow 3Ag + Al^{3+} \end{aligned}$$

Now the equation is balanced, not only in terms of elements but also in terms of charge.

- The **substance oxidized** is the reactant that had undergone oxidation: $Al$
- The **substance reduced** is the reactant that had undergone reduction: $Ag^+$
- The **reducing agent** is the same as the substance oxidized: $Al$

- The **oxidizing agent** is the same as the substance reduced: $\mathbf{Ag^+}$

## Exercise 8.8.1

Write and balance the redox reaction that has calcium ions and potassium metal as reactants and calcium metal and potassium ions as products. Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

**Answer**

Reduction: $\mathbf{Ca^{2+} + 2\,e^- \longrightarrow Ca}$

Oxidation: \(2 \times (\ce{K -> K^{+} + e^{-}})\)

Combined: $\mathbf{Ca^{2+} + 2\,K \longrightarrow Ca + 2\,K^+}$

- The **substance oxidized** is the reactant that had undergone oxidation: $\mathbf{K}$
- The **substance reduced** is the reactant that had undergone reduction: $\mathbf{Ca^{2+}}$
- The **reducing agent** is the same as the substance oxidized: $\mathbf{K}$
- The **oxidizing agent** is the same as the substance reduced: $\mathbf{Ca^{2+}}$

Potassium has been used as a reducing agent to obtain various metals in their elemental form.

## To Your Health: Redox Reactions and Pacemaker Batteries

All batteries use redox reactions to supply electricity because electricity is basically a stream of electrons being transferred from one substance to another. Pacemakers—surgically implanted devices for regulating a person's heartbeat—are powered by tiny batteries, so the proper operation of a pacemaker depends on a redox reaction.

Pacemakers used to be powered by NiCad batteries, in which nickel and cadmium (hence the name of the battery) react with water according to this redox reaction:

$$\mathbf{Cd\,(s) + 2\,NiOOH\,(s) + 2\,H_2O(\ell) \longrightarrow Cd(OH)_2\,(s) + 2\,Ni(OH)_2\,(s)}$$

The cadmium is oxidized, while the nickel atoms in $\mathbf{NiOOH}$ are reduced. Except for the water, all the substances in this reaction are solids, allowing NiCad batteries to be recharged hundreds of times before they stop operating. Unfortunately, NiCad batteries are fairly heavy batteries to be carrying around in a pacemaker. Today, the lighter lithium/iodine battery is used instead. The iodine is dissolved in a solid polymer support, and the overall redox reaction is as follows:

$$\mathbf{2\,Li\,(s) + I_2\,(s) \longrightarrow 2\,LiI\,(s)}$$

Lithium is oxidized, and iodine is reduced. Although the lithium/iodine battery cannot be recharged, one of its advantages is that it lasts up to 10 years. Thus, a person with a pacemaker does not have to worry about periodic recharging; about once per decade a person requires minor surgery to replace the pacemaker/battery unit. Lithium/iodine batteries are also used to power calculators and watches.



Figure **8.8.1**: A small button battery like this is used to power a watch, pacemaker, or calculator. (CC BY-SA; Gerhard H Wrodnigg via Wikipedia)

Oxidation and reduction can also be defined in terms of changes in composition. The original meaning of **oxidation** was "**adding oxygen**," so when oxygen is added to a molecule, the molecule is being oxidized. The reverse is true for **reduction**: if a molecule

**loses oxygen** atoms, the molecule is being reduced. For example, the acetaldehyde ($CH_3CHO$) molecule takes on an oxygen atom to become acetic acid ($CH_3COOH$).

$$2\,CH_3CHO + O_2 \longrightarrow 2\,CH_3COOH$$

Thus, acetaldehyde is being oxidized.

Similarly, **reduction and oxidation can be defined in terms of the gain or loss of hydrogen atoms**. If a molecule adds hydrogen atoms, it is being reduced. If a molecule loses hydrogen atoms, the molecule is being oxidized. For example, in the conversion of acetaldehyde into ethanol ($CH_3CH_2OH$), hydrogen atoms are added to acetaldehyde, so the acetaldehyde is being reduced:

$$CH_3CHO + H_2 \longrightarrow CH_3CH_2OH$$

Table shows how each process affects the change in oxygen and change in hydrogen.

| Process | Change in oxygen (some reactions) | Change in hydrogen (some reactions) |
|---|---|---|
| Oxidation | gain | lose |
| Reduction | lose | gain |

### Example 8.8.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $N_2 \longrightarrow NH_3$
b. $CH_3CH_2OHCH_3 \longrightarrow CH_3COCH_3$
c. $HCHO \longrightarrow HCOOH$

**Solution**

a. Hydrogen is being added to the original reactant molecule, so reduction is occurring.
b. Hydrogen is being removed from the original reactant molecule, so oxidation is occurring.
c. Oxygen is being added to the original reactant molecule, so oxidation is occurring.

### Exercise 8.8.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $CH_4 \longrightarrow CO_2 + H_2O$
b. $NO_2 \longrightarrow N_2$
c. $CH_2{=}CH_2 \longrightarrow CH_3CH_3$

**Answer a:**

Oxygen is being added. Oxidation is occurring.

**Answer b:**

Oxygen is being removed. Reduction is occurring.

**Answer a:**

Hydrogen is being added. Reduction is occurring.

## Key Takeaway

Chemical reactions in which electrons are transferred are called oxidation-reduction, or redox, reactions. Oxidation is the loss of electrons. Reduction is the gain of electrons. Oxidation and reduction always occur together, even though they can be written as separate chemical equations.

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# CHAPTER OVERVIEW

## 9: Stoichiometry and the Mole


9.1: Introduction to Stoichiometry and the Mole

9.2: The Mole

9.3: The Mole in Chemical Reactions

9.4: Mole-Mass and Mass-Mass Calculations

9.5: Limiting Reactant and Theoretical Yield

9.6: Limiting Reactant, Theoretical Yield, and Percent Yield from Initial Masses of Reactants


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# 9.1: Introduction to Stoichiometry and the Mole

**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,

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

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:

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

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 9.1.1

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 9.1.1

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 9.1.2**

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 9.1.2**

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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# 9.2: 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 9.2.1**

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 9.2.1**

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

**Answer**

$2.78 \times 10^{22}$ 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 $6.022 \times 10^{23}$ 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 $1.50 \times 10^{25}$ 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 $6.022 \times 10^{23}$ units: $6.022 \times 10^{23}$ atoms of H, $6.022 \times 10^{23}$ molecules of $H_2$ and $H_2O$, $6.022 \times 10^{23}$ units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

### Example 9.2.2: 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 9.2.2

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 9.2.3

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 9.2.3

What is the mass of 33.7 mol of $H_2O$?

**Answer**

607 g

### Example 9.2.4

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 9.2.4

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 9.2.5

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 9.2.5**

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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# 9.3: 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 9.3.1**

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 9.3.1**

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 9.3.2

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 9.3.2

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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# 9.4: 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 9.4.1**

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 9.4.1

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 9.4.2

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 9.4.2**

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):

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

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

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

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

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

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

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

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of $SO_3$, which is what we are looking for, as our final answer.

**✓ Example 9.4.3**

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,

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

**? Exercise 9.4.3**

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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## 9.5: LIMITING REACTANT AND THEORETICAL YIELD

**LEARNING OBJECTIVES**

- Identify the limiting reactant (limiting reagent) in a given chemical reaction.
- Calculate how much product will be produced from the limiting reactant.
- Calculate how much reactant(s) remains when the reaction is complete.

In all examples discussed thus far, the reactants were assumed to be present in stoichiometric quantities. Consequently, none of the reactants were left over at the end of the reaction. This is often desirable—as in the case of a space shuttle—where excess oxygen or hydrogen is not only extra freight to be hauled into orbit, but also an explosion hazard. More often, however, reactants are present in mole ratios that are not the same as the ratio of the coefficients in the balanced chemical equation. As a result, one or more of them will not be used up completely, but will be left over when the reaction is completed. In this situation, the amount of product that can be obtained is limited by the amount of only one of the reactants. The reactant that restricts the amount of product obtained is called the limiting reactant. The reactant that remains after a reaction has gone to completion is in excess.

Consider a nonchemical example. Assume you have invited some friends for dinner and want to bake brownies for dessert. You find two boxes of brownie mix in your pantry and see that each package requires two eggs. The balanced equation for brownie preparation is:

$$\textbf{1 box mix} + \textbf{2 eggs} \rightarrow \textbf{1 batch brownies} \tag{9.5.1}$$

If you have a dozen eggs, which ingredient will determine the number of batches of brownies that you can prepare? Because each box of brownie mix requires two eggs and you have two boxes, you need four eggs. Twelve eggs is eight more eggs than you need. Although the ratio of eggs to boxes in is 2:1, the ratio in your possession is 6:1. Hence the eggs are the ingredient (reactant) present in excess, and the brownie mix is the limiting reactant. Even if you had a refrigerator full of eggs, you could make only two batches of brownies.

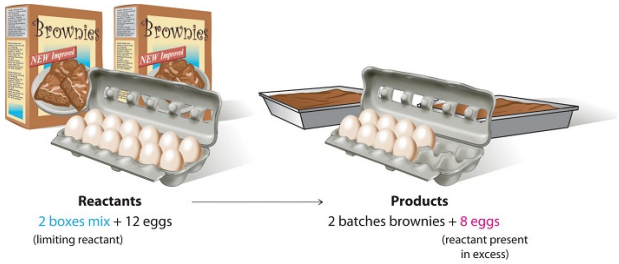
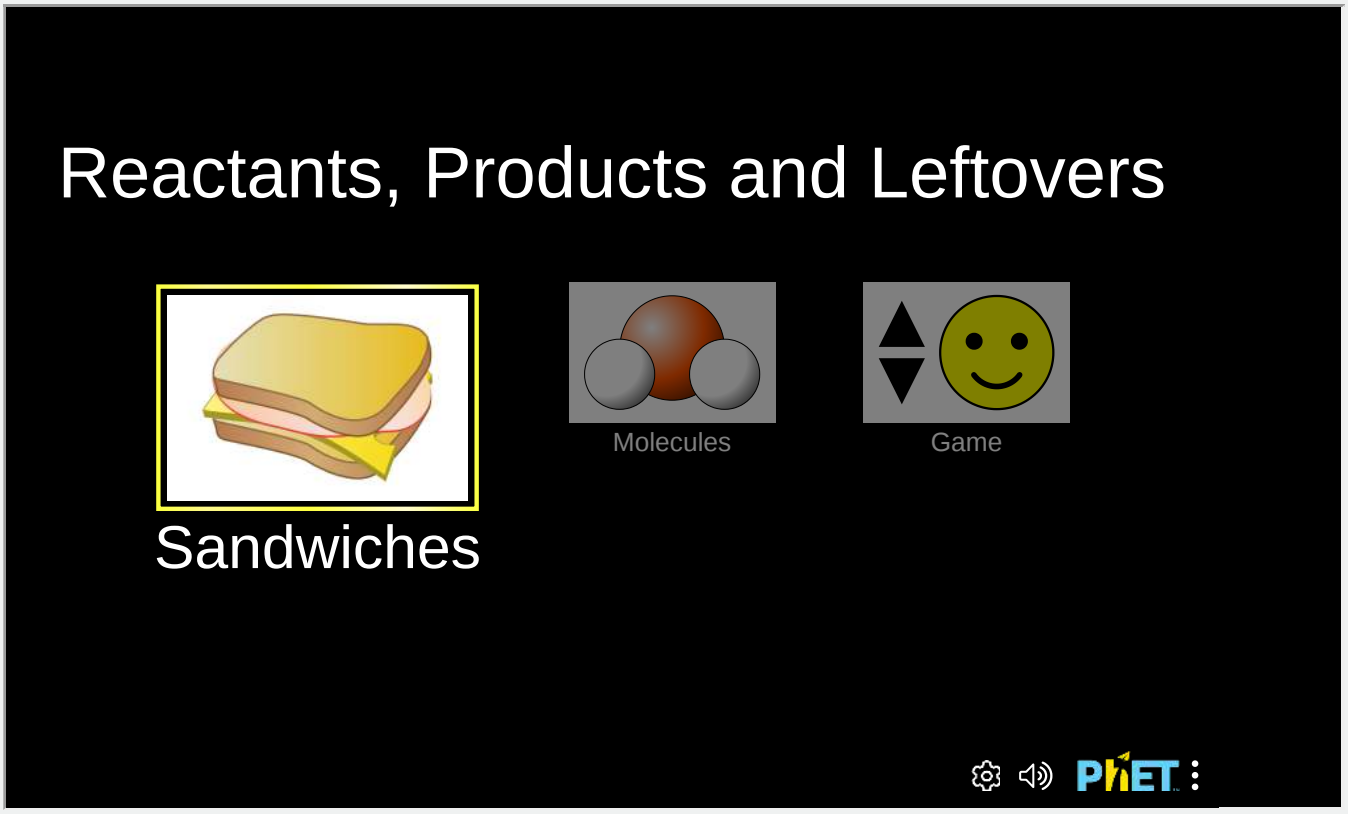



Figure **9.5.1**: The Concept of a Limiting Reactant in the Preparation of Brownies. For a chemist, the **balanced** chemical equation is the recipe that must be followed. 2 boxes of brownie mix and 12 eggs results in 2 batches of brownies and 8 eggs; in this case the 8 eggs are reactant present in excess

**PHET SIMULATION: REACTANTS, PRODUCTS AND LEFTOVERS**

View this interactive simulation illustrating the concepts of limiting and excess reactants.




Consider this concept now with regard to a chemical process, the reaction of hydrogen with chlorine to yield hydrogen chloride:

$$H_2 + Cl_2(g) \rightarrow 2\,HCl(g)$$

The balanced equation shows that hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation

requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation. For example, imagine combining 3 moles of $H_2$ and 2 moles of $Cl_2$. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all the provided chlorine (2 mol) will consume 2 mol of the 3 mol of hydrogen provided, leaving 1 mol of hydrogen non-reacted.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant. For the example, in the previous paragraph, complete reaction of the hydrogen would yield:

$$\text{mol HCl produced} = 3\ \text{mol } H_2 \times \frac{2\ \text{mol HCl}}{1\ \text{mol } H_2} = 6\ \text{mol HCl}$$

Complete reaction of the provided chlorine would produce:

$$\text{mol HCl produced} = 2\ \text{mol } Cl_2 \times \frac{2\ \text{mol HCl}}{1\ \text{mol } Cl_2} = 4\ \text{mol HCl}$$

The chlorine will be completely consumed once 4 moles of HCl have been produced. Since enough hydrogen was provided to yield 6 moles of HCl, there will be non-reacted hydrogen remaining once this reaction is complete. Chlorine, therefore, is the limiting reactant and hydrogen is the excess reactant (Figure **9.5.2**).

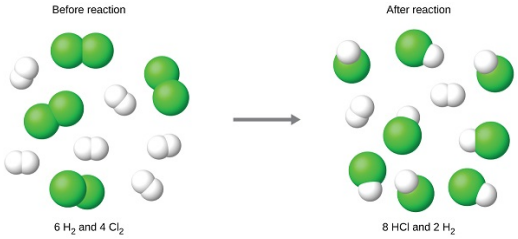



Figure **9.5.2**: When $H_2$ and $Cl_2$ are combined in nonstoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant. The figure shows a space-filling molecular models reacting. There is a reaction arrow pointing to the right in the middle. To the left of the reaction arrow there are three molecules each consisting of two green spheres bonded together. There are also five molecules each consisting of two smaller, white spheres bonded together. Above these molecules is the label, "Before reaction," and below these molecules is the label, "6 H subscript 2 and 4 C l subscript 2." To the right of the reaction arrow, there are eight molecules each consisting of one green sphere bonded to a smaller white sphere. There are also two molecules each consisting of two white spheres bonded together. Above these molecules is the label, "After reaction," and below these molecules is the label, "8 H C l and 2 H subscript 2."

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reactant; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reactant and which is in excess.

### HOW TO IDENTIFY THE LIMITING REACTANT (LIMITING REAGENT)

There are two ways to determine the limiting reactant. One method is to find and compare the mole ratio of the reactants used in the reaction (Approach 1). Another way is to calculate the grams of products produced from the given quantities of reactants; the reactant that produces the **smallest** amount of product is the limiting reactant (Approach 2). This section will focus more on the second method.

**Approach 1 (The "Reactant Mole Ratio Method"):** Find the limiting reactant by looking at the number of moles of each reactant.

1. Determine the balanced chemical equation for the chemical reaction.
2. Convert all given information into moles (most likely, through the use of molar mass as a conversion factor).
3. Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.
4. Use the amount of limiting reactant to calculate the amount of product produced.
5. If necessary, calculate how much is left in excess of the non-limiting (excess) reactant.

**Approach 2 (The "The Product Method"):** Find the limiting reactant by calculating and comparing the amount of product that each reactant will produce.

1. Balance the chemical equation for the chemical reaction.
2. Convert the given information into moles.
3. Use stoichiometry for each individual reactant to find the mass of product produced.
4. The reactant that produces a lesser amount of product is the limiting reactant.
5. The reactant that produces a larger amount of product is the excess reactant.
6. To find the amount of remaining excess reactant, subtract the mass of excess reactant consumed from the total mass of excess reactant given.

The key to recognizing which reactant is the limiting reactant is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reactant. What we need to do is determine an amount of one product (either moles or mass) assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reactant. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

### EXAMPLE 9.5.1: IDENTIFYING THE LIMITING REACTANT

As an example, consider the balanced equation

$$4\,C_2H_3Br_3 + 11\,O_2 \rightarrow 8\,CO_2 + 6\,H_2O + 6\,Br_2$$

What is the limiting reactant if 76.4 grams of $C_2H_3Br_3$ reacted with 49.1 grams of $O_2$?

**Solution**

Using Approach 1:

**Step 1: *Balance the chemical equation.***

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol $O_2$ to 6 mol $H_2O$ to 6 mol Br

***Step 2: Convert all given information into moles.***

$$76.4\ \cancel{\text{g } C_2H_3Br_3} \times \frac{1\ \text{mol } C_2H_3Br_3}{266.72\ \cancel{\text{g } C_2H_3Br_3}} = 0.286\ \text{mol } C_2H_3Br_3$$

$$49.1\ \cancel{\text{g } O_2} \times \frac{1\ \text{mol } O_2}{32.00\ \cancel{\text{g } O_2}} = 1.53\ \text{mol } O_2$$

***Step 3: Calculate the mole ratio from the given information. Compare the calculated ratio to the actual ratio.***

Assuming that all of the oxygen is used up,

$$1.53\ \cancel{\text{mol } O_2} \times \frac{4\ \text{mol } C_2H_3Br_3}{11\ \cancel{\text{mol } O_2}} = 0.556\ \text{mol } C_2H_3Br_3 \text{ are required.}$$

**Because 0.556 moles of $C_2H_3Br_3$ required > 0.286 moles of $C_2H_3Br_3$ available, $C_2H_3Br_3$ is the limiting reactant.**

Using Approach 2:

**Step 1:** ***Balance the chemical equation.***

The equation is already balanced with the relationship

4 mol $C_2H_3Br_3$ to 11 mol $O_2$ to 6 mol $H_2O$ to 6 mol $Br_2$

***Step 2 and Step 3: Convert mass to moles and stoichiometry.***

$$76.4\ \text{g}\ C_2H_3Br_3 \times \frac{1\ \text{mol}\ C_2H_3Br_3}{266.72\ \text{g}\ C_2H_3Br_3} \times \frac{8\ \text{mol}\ CO_2}{4\ \text{mol}\ C_2H_3Br_3} \times \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2} = 25.2\ \text{g}\ CO_2$$

$$49.1\ \text{g}\ O_2 \times \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2} \times \frac{8\ \text{mol}\ CO_2}{11\ \text{mol}\ O_2} \times \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2} = 49.1\ \text{g}\ CO_2$$

***Step 4: The reactant that produces a smaller amount of product is the limiting reactant.***

Therefore, by either method, $C_2H_3Br_3$ is the limiting reactant.

## EXAMPLE 9.5.2: IDENTIFYING THE LIMITING REACTANT AND THE MASS OF EXCESS REACTANT

For example, in the reaction of magnesium metal and oxygen, calculate the mass of magnesium oxide that can be produced if 2.40 g ***Mg*** reacts with 10.0 g ***$O_2$***. Also determine the amount of excess reactant. **MgO** is the only product in the reaction.

**Solution**

Following Approach 1:

***Step 1: Balance the chemical equation.***

$$2\ \text{Mg (s)} + O_2\ \text{(g)} \rightarrow 2\ \text{MgO (s)}$$

The balanced equation provides the relationship of 2 mol Mg to 1 mol $O_2$ to 2 mol MgO

***Step 2 and Step 3: Convert mass to moles and stoichiometry.***

$$2.40\ \text{g}\ Mg \times \frac{1\ \text{mol}\ Mg}{24.31\ \text{g}\ Mg} \times \frac{2\ \text{mol}\ MgO}{2\ \text{mol}\ Mg} \times \frac{40.31\ \text{g}\ MgO}{1\ \text{mol}\ MgO} = 3.98\ \text{g}\ MgO$$

$$10.0\ \text{g}\ O_2 \times \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2} \times \frac{2\ \text{mol}\ MgO}{1\ \text{mol}\ O_2} \times \frac{40.31\ \text{g}\ MgO}{1\ \text{mol}\ MgO} = 25.2\ \text{g}\ MgO$$

***Step 4: The reactant that produces a smaller amount of product is the limiting reactant.***

Mg produces less MgO than does $O_2$ (3.98 g MgO vs. 25.2 g MgO), therefore Mg is the limiting reactant in this reaction.

***Step 5: The reactant that produces a larger amount of product is the excess reactant.***

$O_2$ produces more amount of MgO than Mg (25.2g MgO vs. 3.98 MgO), therefore $O_2$ is the excess reactant in this reaction.

***Step 6: Find the amount of remaining excess reactant by subtracting the mass of the excess reactant consumed from the total mass of excess reactant given.***

Mass of excess reactant calculated using the limiting reactant:

$$2.40\ \text{g}\ Mg \times \frac{1\ \text{mol}\ Mg}{24.31\ \text{g}\ Mg} \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ Mg} \times \frac{32.00\ \text{g}\ O_2}{1\ \text{mol}\ O_2} = 1.58\ \text{g}\ O_2$$

**OR**

Mass of excess reactant calculated using the mass of the product:

$$3.98\ \text{g}\ MgO \times \frac{1\ \text{mol}\ MgO}{40.31\ \text{g}\ MgO} \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ MgO} \times \frac{32.0\ \text{g}\ O_2}{1\ \text{mol}\ O_2} = 1.58\ \text{g}\ O_2$$

Mass of total excess reactant given – mass of excess reactant consumed in the reaction:

10.0g $O_2$ - (available) 1.58g $O_2$ (used) = 8.42g $O_2$ (excess)

**Therefore, $O_2$ is in excess.**

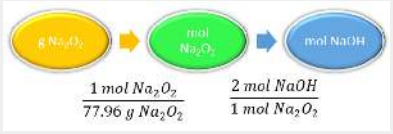
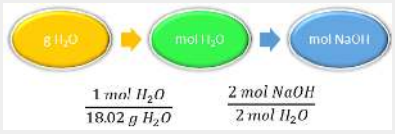
## EXAMPLE 9.5.3: LIMITING REACTANT

What is the limiting reactant if 78.0 grams of $Na_2O_2$ were reacted with 29.4 grams of $H_2O$? The unbalanced chemical equation is

$$Na_2O_2\,(s) + H_2O\,(l) \longrightarrow NaOH\,(aq) + H_2O_2\,(l)$$

**Solution**

Solutions to Example 8.4.3

| Steps for Problem Solving- The Product Method | Example 9.5.1 |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: 78.0 grams of $Na_2O_2$<br>29.4 g $H_2O$<br>Find: limiting reactant |
| List other known quantities. | 1 mol $Na_2O_2$= 77.96 g/mol<br>1 mol $H_2O$ = 18.02 g/mol<br>Since the amount of product in grams is not required, only the molar mass of the reactants is needed. |
| Balance the equation. | $Na_2O_2$ (s) + $\underline{2}H_2O$ (l) → $\underline{2}$NaOH (aq) + $H_2O_2$ (l)<br>The balanced equation provides the relationship of 1 mol $Na_2O_2$ to 2 mol $H_2O$ 2mol NaOH to 1 mol $H_2O_2$ |
| Prepare a concept map and use the proper conversion factor. |  <br>  <br>Because the question only asks for the limiting reactant, we can perform two **mass-mole** calculations and determine which amount is less. |
| Cancel units and calculate. | $78.0\ g\ Na_2O_2 \times \frac{1\ mol\ Na_2O_2}{77.96\ g\ Na_2O_2} \times \frac{2\ mol\ NaOH}{1\ mol\ Na_2O_2} \times \frac{40\ g\ NaOH}{1\ mol\ NaOH} = 2.00\ mol\ NaOH$<br>$29.4\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} \times \frac{2\ mol\ NaOH}{2\ mol\ Na_2O_2} \times \frac{40\ g\ NaOH}{1\ mol\ NaOH} = 1.63\ mol\ NaOH$<br>Therefore, $H_2O$ is the *limiting reactant*. |

Think about your result.

## ✔ EXAMPLE 9.5.4: LIMITING REACTANT AND MASS OF EXCESS REACTANT

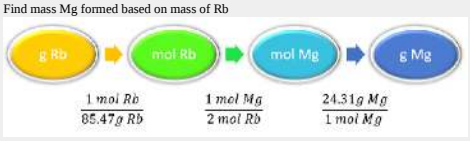
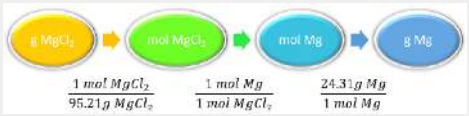
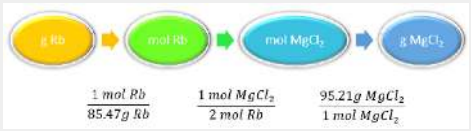
A 5.00 g quantity of **Rb** is combined with 3.44 g of **$MgCl_2$** according to this chemical reaction:

$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of **Mg** is formed, and what mass of remaining reactant is left over?

## Solution

Solutions to Example 8.4.10

| Steps for Problem Solving- The Product Method | Example 9.5.2 |
|---|---|
| Steps for Problem Solving | A 5.00 g quantity of Rb is combined with 3.44 g of $MgCl_2$ according to this chemical reaction:<br>$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$<br>What mass of Mg is formed, and what mass of remaining reactant is left over? |
| Identify the "given" information and what the problem is asking you to "find." | Given: 5.00g Rb, 2.44g MgCl2<br>Find: mass of Mg formed, mass of remaining reactant |
| List other known quantities. | • molar mass: Rb = 85.47 g/mol<br>• molar mass: $MgCl_2$ = 95.21 g/mol<br>• molar mass: Mg = 24.31 g/mol |
| Prepare concept maps and use the proper conversion factor. | Find mass Mg formed based on mass of Rb<br>  <br>Find mass of Mg formed based on mass of $MgCl_2$<br>  <br>Conversion factors: 1 mole MgCl2 to 95.21 grams MgCl2, 1 mole Mg to 1 mole MgCl2, 24.31 grams Mg to 1 mole Mg<br>Use limiting reactant to determine amount of excess reactant consumed<br>  <br>Conversion factors: 1 mole Rb to 85.47 grams Rb, 1 mole MgCl2 to 2 moles Rb, 95.21 grams MgCl2 to 1 mole MgCl2 |
| Cancel units and calculate. | Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less.<br>$$5.00\ g\ Rb \times \frac{1\ mol\ Rb}{85.47\ g\ Rb} \times \frac{1\ mol\ Mg}{2\ mol\ Rb} \times \frac{24.31\ g\ Mg}{1\ mol\ Mg} = 0.711\ g\ Mg$$<br>$$3.44\ g\ MgCl_2 \times \frac{1\ mol\ MgCl_2}{95.21\ g\ MgCl_2} \times \frac{1\ mol\ Mg}{1\ mol\ MgCl_2} \times \frac{24.31\ g\ Mg}{1\ mol\ Mg} = 0.878\ g\ Mg$$<br>The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reactant. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of $MgCl_2$ reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount.<br>$$5.00\ g\ Rb \times \frac{1\ mol\ Rb}{85.47\ g\ Rb} \times \frac{1\ mol\ MgCl_2}{2\ mol\ Rb} \times \frac{95.21\ g\ MgCl_2}{1\ mol\ MgCl_2} = 2.78\ g\ MgCl_2\ reacted$$<br>Because we started with 3.44 g of $MgCl_2$, we have<br>3.44 g $MgCl_2$ − 2.78 g $MgCl_2$ reacted = 0.66 g $MgCl_2$ left |
| Think about your result. | It usually is not possible to determine the limiting reactant using just the initial masses, as the reagents have different molar masses and coefficients. |

### ? EXERCISE 9.5.1

Given the initial amounts listed, what is the limiting reactant, and what is the mass of the leftover reactant?

$$\underset{MgO(s)}{22.7\ g} + \underset{H_2S}{17.9\ g} \rightarrow MgS(s) + H_2O(l)$$

***Answer***

$H_2S$ is the limiting reagent; 1.5 g of MgO are left over.

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# 9.6: LIMITING REACTANT, THEORETICAL YIELD, AND PERCENT YIELD FROM INITIAL MASSES OF REACTANTS

## LEARNING OBJECTIVES

- Calculate percentage or actual yields from known amounts of reactants.

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

## PERCENT YIELD

Chemical reactions in the real world don't always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than predicted. Besides spills and other experimental errors, there are usually losses due to an incomplete reaction, undesirable side reactions, etc. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. This is called the **theoretical yield**, the maximum amount of product that can be formed from the given amounts of reactants. The **actual yield** is the amount of product that is actually formed when the reaction is carried out in the laboratory. The **percent yield** is the ratio of the actual yield to the theoretical yield, expressed as a percentage.

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and unnecessary expense.

Typically, percent yields are understandably less than $100\%$ because of the reasons indicated earlier. However, percent yields greater than $100\%$ are possible if the measured product of the reaction contains impurities that cause its mass to be greater than it actually would be if the product was pure. When a chemist synthesizes a desired chemical, he or she is always careful to purify the products of the reaction. Example **9.6.1** illustrates the steps for determining percent yield.

## EXAMPLE 9.6.1: DECOMPOSITION OF POTASSIUM CHLORATE

Potassium chlorate decomposes upon slight heating in the presence of a catalyst, according to the reaction below:

$$2KClO_3\,(s) \rightarrow 2KCl\,(s) + 3O_2\,(g)$$

In a certain experiment, $40.0\ g\ KClO_3$ is heated until it completely decomposes. The experiment is performed and the oxygen gas is collected and its mass is found to be $14.9\ g$.

a. What is the theoretical yield of oxygen gas?
b. What is the percent yield for the reaction?

### Solution

a. Calculation of theoretical yield

First, we will calculate the theoretical yield based on the stoichiometry.

#### *Step 1: Identify the "given" information and what the problem is asking you to "find".*

Given: Mass of $KClO_3 = 40.0\ g$

Mass of $O_2$ collected = 14.9g

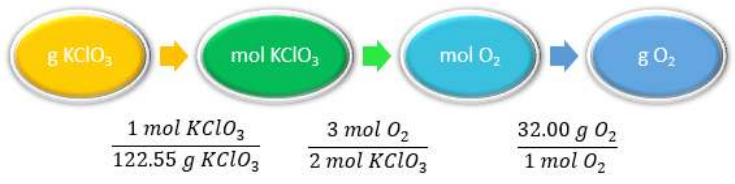
Find: Theoretical yield, g $O_2$

#### *Step 2: List other known quantities and plan the problem.*

1 mol $KClO_3$ = 122.55 g/mol

1 mol $O_2$ = 32.00 g/mol

**Step 3: Apply stoichiometry to convert from the mass of a reactant to the mass of a product:**




***Step 4: Solve.***

$$40.0\ \cancel{\text{g KClO}_3} \times \frac{1\ \cancel{\text{mol KClO}_3}}{122.55\ \cancel{\text{g KClO}_3}} \times \frac{3\ \cancel{\text{mol O}_2}}{2\ \cancel{\text{mol KClO}_3}} \times \frac{32.00\ \text{g O}_2}{1\ \cancel{\text{mol O}_2}} = 15.7\ \text{g O}_2$$

The theoretical yield of $O_2$ is **15.7 g**, 15.67 g unrounded.

***Step 5: Think about your result.***

The mass of oxygen gas must be less than the **40.0 g** of potassium chlorate that was decomposed.

b. Calculation of percent yield

Now we will use the actual yield and the theoretical yield to calculate the percent yield.

***Step 1: Identify the "given" information and what the problem is asking you to "find".***

Given: Theoretical yield =15.67 g, use the un-rounded number for the calculation.

Actual yield = 14.9g

Find: Percent yield, % Yield

***Step 2: List other known quantities and plan the problem.***

***No other quantities needed.***

***Step 3: Use the percent yield equation below.***

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

***Step 4: Solve.***

$$\text{Percent Yield} = \frac{14.9\ \text{g}}{15.67\ \text{g}} \times 100\% = 94.9\%$$

***Step 5: Think about your result.***

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under **100%**.

## EXAMPLE 9.6.2: OXIDATION OF ZINC

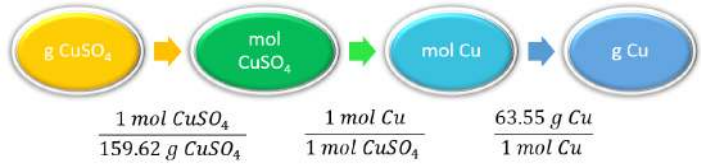
Upon reaction of 1.274 g of copper sulfate with excess zinc metal, 0.392 g copper metal was obtained according to the equation:

$$\text{CuSO}_4(aq) + \text{Zn}(s) \rightarrow \text{Cu}(s) + \text{ZnSO}_4(aq)$$

What is the percent yield?

## Solution

Solutions to Example 8.6.2

| Steps for Problem Solving-The Product Method | Example 9.6.1 |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: 1.274 g CuSO4<br>Actual yield = 0.392 g Cu<br><br>Find: Percent yield |
| List other known quantities. | 1 mol CuSO4= 159.62 g/mol<br>1 mol Cu = 63.55 g/mol<br>Since the amount of product in grams is not required, only the molar mass of the reactants is needed. |
| Balance the equation. | The chemical equation is already balanced.<br>The balanced equation provides the relationship of 1 mol CuSO4 to 1 mol Zn to 1 mol Cu to 1 mol ZnSO4. |
| Prepare a concept map and use the proper conversion factor. |  <br>The provided information identifies copper sulfate as the limiting reactant, and so the theoretical yield (g Cu) is found by performing **mass-mass** calculation based on the initial amount of $CuSO_4$. |
| Cancel units and calculate. | Double subscripts: use braces to clarify<br><br>Using this theoretical yield and the provided value for actual yield, the percent yield is calculated to be:<br>$\text{percent yield} = \left(\frac{\text{actual yield}}{\text{theoretical yield}}\right) \times 100$<br>$= \left(\frac{0.392\,g\,\text{Cu}}{0.5072\,g\,\text{Cu}}\right) \times 100$<br>$= 77.3\%$ |
| Think about your result. | Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%. |

## ? EXERCISE 9.6.1

What is the percent yield of a reaction that produces 12.5 g of the Freon $CF_2Cl_2$ from 32.9 g of $CCl_4$ and excess HF?

$$CCl_4 + 2\,HF \rightarrow CF_2Cl_2 + 2\,HCl$$

***Answer***

48.3%

## SUMMARY

Theoretical yield is calculated based on the stoichiometry of the chemical equation. The actual yield is experimentally determined. The percent yield is determined by calculating the ratio of actual yield to theoretical yield.

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# CHAPTER OVERVIEW

## 10: Solutions


10.1: Introduction to Solutions

10.2: Definitions

10.3: Quantitative Units of Concentration

10.4: Dilutions and Concentrations

10.5: Concentrations as Conversion Factors


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# 10.1: Introduction to Solutions

More than 70% of the earth's surface is covered by a very important solution—seawater. It is likely that without seawater, no life would exist on Earth.

At its simplest, seawater is mostly $H_2O$. However, about 3.5% of seawater is dissolved solids—mostly NaCl—but other ions as well. Table **10.1.1** lists the percentage by mass of the various ions in seawater. Because it is highly likely that life on Earth originated in the oceans, it should not be surprising that many bodily fluids resemble seawater, especially blood. Table **10.1.1** also lists the percentage by mass of ions in a typical sample of blood.

Table **10.1.1**: Percentage by Mass of Ions in Seawater and Blood

| Ion | Percentage in Seawater | Percentage in Blood |
|---|---|---|
| $Na^+$ | 2.36 | 0.322 |
| $Cl^-$ | 1.94 | 0.366 |
| $Mg^{2+}$ | 0.13 | 0.002 |
| $SO_4^{2-}$ | 0.09 | - |
| $K^+$ | 0.04 | 0.016 |
| $Ca^{2+}$ | 0.04 | 0.0096 |
| $HCO_3^-$ | 0.002 | 0.165 |
| $HPO_4^{2-}$, $H_2PO_4^-$ | - | 0.01 |

Most ions are more abundant in seawater than they are in blood, with some notable exceptions. There is far more hydrogen carbonate ion ($HCO_3^-$) in blood than in seawater; indeed, it is the third most common ion in blood. This difference is significant because the $HCO_3^-$ ion and some related species [$CO_3^{2-}$, $CO_2(aq)$] have an important role in controlling the acid-base properties of blood. Although there is a negligible amount of the two hydrogen phosphate ions ($HPO_4^{2-}$ and $H_2PO_4^-$) in seawater, there is a small amount in blood, where these ions affect acid-base properties. Another notable difference is that blood has a negligible amount of the sulfate ion ($SO_4^{2-}$), but this ion is present in seawater. Gold is present in seawater—but only a tiny amount. A current estimate of the amount of gold is about 1 part per every $1 \times 10^{13}$ parts of seawater, which makes the extraction of gold from seawater unfeasible. However, it does mean that there are about $1.4 \times 10^{14}$ g of gold in the world's oceans!



Figure **10.1.1**: Gold in the Ocean. There are approximately $1.4 \times 10^{14}$ g of gold in the oceans, but extracting it effectively is beyond current technologies. (CC SA-BY; Evgeni Dinev, Flickr.)

A solution is a *homogeneous mixture*: a mixture of two or more substances that are so intimately mixed that the mixture behaves in many ways like a single substance. Many chemical reactions occur when the reactants are dissolved in solution. In this chapter, we will introduce concepts that are applicable to solutions and the chemical reactions that occur in them.

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# 10.2: Definitions

**Learning Objective**

- Learn some terminology involving solutions.
- Recognize which terminology is qualitative and which terminology is quantitative.
- Explain why certain substances dissolve in other substances.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. "Major" and "minor" indicate which component has the greater or lesser presence by mass or moles, respectively. Sometimes this becomes confusing, especially when considering substances with very different molar masses. We will confine the discussion here to solutions for which the major component and the minor component are obvious.

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water; soda water is a solution of gaseous $CO_2$ in liquid water, while air is a solution of a gaseous solute ($O_2$) in a gaseous solvent ($N_2$). In all cases, however, the overall phase of the solution is the same phase as the solvent.

**Example 10.2.1: Sugar Water**

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

**Solution**

Either by mass or by moles, the obvious minor component is sucrose, so it is the solute. Water—the majority component—is the solvent. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

**Exercise 10.2.1**

A solution is made by dissolving 3.33 g of HCl(g) in 40.0 g of liquid methyl alcohol ($CH_3OH$). Identify the solvent and solute in the resulting solution.

**Answer**

solute: HCl(g); solvent: $CH_3OH$

One important concept of solutions is in defining how much solute is dissolved in a given amount of solvent. This concept is called **concentration**. Various words are used to describe the relative amounts of solute. **Dilute** describes a solution that has very little solute, while **concentrated** describes a solution that has a lot of solute. One problem is that these terms are qualitative; they describe more or less but not exactly how much.

In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table **10.2.1** lists the solubilities of some simple ionic compounds. These solubilities vary widely: NaCl can dissolve up to 31.6 g per 100 g of $H_2O$, while AgCl can dissolve only 0.00019 g per 100 g of $H_2O$.

Table **10.2.1**: Solubilities of Some Ionic Compounds

| Solute | Solubility (g per 100 g of $H_2O$ at 25°C) |
|---|---|
| AgCl | 0.00019 |
| $CaCO_3$ | 0.0006 |
| KBr | 70.7 |
| NaCl | 36.1 |
| $NaNO_3$ | 94.6 |

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of $H_2O$ may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of $H_2O$ is also saturated but rather concentrated. Ideally, we need more precise ways of specifying the amount of solute in a solution. We will introduce such ways in Section 11.3.

In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.

It should be obvious that some solutes dissolve in certain solvents but not others. NaCl, for example, dissolves in water but not in vegetable oil. Beeswax dissolves in liquid hexane but not water. What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding (as described in Chapter 10). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane but not in polar water. This concept leads to the general rule that "like dissolves like" for predicting whether a solute is soluble in a given solvent. However, this is a general rule, not an absolute statement, so it must be applied with care.

### ✔ Example 10.2.2: Polar and Nonpolar Solvents

Would $I_2$ be more soluble in $CCl_4$ or $H_2O$? Explain your answer.

**Solution**

$I_2$ is nonpolar. Of the two solvents, $CCl_4$ is nonpolar and $H_2O$ is polar, so $I_2$ would be expected to be more soluble in $CCl_4$.

### ? Exercise 10.2.2

Would $C_3H_7OH$ be more soluble in $CCl_4$ or $H_2O$? Explain your answer.

**Answer**

$H_2O$ because both experience hydrogen bonding

## Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Concentration is the expression of the amount of solute in a given amount of solvent and can be described by several qualitative terms.
- Solubility is a specific amount of solute that can dissolve in a given amount of solvent.

- "Like dissolves like" is a useful rule for deciding if a solute will be soluble in a solvent.

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# 10.3: Quantitative Units of Concentration

**Learning Objective**

- Learn to determine specific concentrations with several common units.

Rather than qualitative terms (Section 11.2 - Definitions), we need quantitative ways to express the amount of solute in a solution; that is, we need specific units of concentration. In this section, we will introduce several common and useful units of concentration.

**Molarity** (M) is defined as the number of moles of solute divided by the number of liters of solution:

$$molarity = \frac{moles\ of\ solute}{liters\ of\ solution}$$

which can be simplified as

$$M = \frac{mol}{L},\ or\ mol/L$$

As with any mathematical equation, if you know any two quantities, you can calculate the third, unknown, quantity.

For example, suppose you have 0.500 L of solution that has 0.24 mol of NaOH dissolved in it. The concentration of the solution can be calculated as follows:

$$molarity = \frac{0.24\ mol\ NaOH}{0.500 L} = 0.48\ M\ NaOH$$

The concentration of the solution is 0.48 M, which is spoken as "zero point forty-eight molarity" or "zero point forty-eight molar." If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L? First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ \cancel{g HCl} \times \frac{1\ mol\ Hcl}{36.5\ \cancel{g HCl}} = 0.614\ M\ HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M = \frac{0.614\ mol\ HCl}{1.56 L} = 0.394\ M$$

**Example 10.3.1**

What is the molarity of a solution made when 32.7 g of NaOH are dissolved to make 445 mL of solution?

Solution

To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$445\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}} = 0.445\ L$$

Now we convert the amount of solute to moles, using the molar mass of NaOH, which is 40.0 g/mol:

$$32.7\ \cancel{g NaOH} \times \frac{1\ mol\ NaOH}{40.0\ \cancel{g NaOH}} = 0.818\ mol\ NaOH$$

Now we can use the definition of molarity to determine the molar concentration:

$$M = \frac{0.818\ mol\ NaOH}{0.445 L} = 1.84\ M\ NaOH$$

**Exercise 10.3.1**

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

**Answer**

1.57 M

The definition of molarity can be used to determine the amount of solute or the volume of solution, if the other information is given. Example 4 illustrates this situation.

**Example 10.3.2**

How many moles of solute are present in 0.108 L of a 0.887 M NaCl solution?

Solution

We know the volume and the molarity; we can use the definition of molarity to mathematically solve for the amount in moles. Substituting the quantities into the definition of molarity:

$$0.887\,M = \frac{mol\ NaCl}{0.108 L}$$

We multiply the 0.108 L over to the other side of the equation and multiply the units together; "molarity × liters" equals moles, according to the definition of molarity. So

mol NaCl = (0.887 M)(0.108 L) = 0.0958 mol

**Exercise 10.3.2**

How many moles of solute are present in 225 mL of a 1.44 M $CaCl_2$ solution?

**Answer**

0.324 mol

If you need to determine volume, remember the rule that the unknown quantity must be by itself and in the numerator to determine the correct answer. Thus, rearrangement of the definition of molarity is required.

**Example 10.3.3**

What volume of a 2.33 M $NaNO_3$ solution is needed to obtain 0.222 mol of solute?

Solution

Using the definition of molarity, we have

$$2.33\,M = \frac{0.222 mol}{L}$$

To solve for the number of liters, we bring the 2.33 M over to the right into the denominator, and the number of liters over to the left in the numerator. We now have

$$L = \frac{0.222 mol}{2.33\,M}$$

Dividing, **the volume is 0.0953 L = 95.3 mL.**

**? Exercise 10.3.3**

What volume of a 0.570 M $K_2SO_4$ solution is needed to obtain 0.872 mol of solute?

**Answer**

1.53 L

A similar unit of concentration is **molality** (*m*), which is defined as the number of moles of solute per kilogram of solvent, not per liter of solution:

$$molality = \frac{moles\ solute}{kilograms\ solvent}$$

Mathematical manipulation of molality is the same as with molarity.

Another way to specify an amount is **percentage composition by mass** (or *mass percentage*, % m/m). It is defined as follows:

$$\%m/m = \frac{mass\ of\ solute}{mass\ of\ entire\ sample} \times 100\%$$

It is not uncommon to see this unit used on commercial products (see Figure **10.3.1** - Concentration in Commercial Applications).




Figure **10.3.1** Concentration in Commercial Applications © Thinkstock. *The percentage of urea in this package is 5% m/m, meaning that there are 5 g of urea per 100 g of product.*

**✓ Example 10.3.4**

What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

Solution

Using the definition of mass percentage, we have

$$\%m/m = \frac{87.9\,g\,Fe}{113\,g\,sample} \times 100\% = 77.8\%\,Fe$$

**Exercise 10.3.4**

What is the mass percentage of $H_2O_2$ in a solution with 1.67 g of $H_2O_2$ in a 55.5 g sample?

**Answer**

3.01%

Related concentration units are **parts per thousand (ppth), parts per million (ppm)** and **parts per billion (ppb).** Parts per thousand is defined as follows:

$$ppth = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1000$$

There are similar definitions for parts per million and parts per billion:

$$ppm = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000$$

and

$$ppb = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000,000$$

Each unit is used for progressively lower and lower concentrations. The two masses must be expressed in the same unit of mass, so conversions may be necessary.

**Example 10.3.5**

If there are 0.6 g of Pb present in 277 g of solution, what is the Pb concentration in parts per thousand?

Solution

Use the definition of parts per thousand to determine the concentration. Substituting

$$\frac{0.6g\ Pb}{277g\ solution} \times 1000 = 2.17\ ppth$$

**Exercise 10.3.5**

If there are 0.551 mg of As in 348 g of solution, what is the As concentration in ppm?

**Answer**

1.58 ppm

As with molarity and molality, algebraic rearrangements may be necessary to answer certain questions.

**Example 10.3.6**

The concentration of $Cl^-$ ion in a sample of $H_2O$ is 15.0 ppm. What mass of $Cl^-$ ion is present in 240.0 mL of $H_2O$, which has a density of 1.00 g/mL?

Solution

First, use the density of $H_2O$ to determine the mass of the sample:

$$240.0\ \cancel{mL} \times \frac{1.00\ g}{\cancel{mL}} = 240.0\ g$$

Now we can use the definition of ppm:

$$15.0\,ppm = \frac{mass\ of\ solute}{240.0\ g\ solution} \times 1,000,000$$

Rearranging to solve for the mass of solute,

$$mass\ solute = \frac{(15.0\,ppm)(240.0\ g\ solution)}{1,000,000} = 0.0036g = 3.6\,mg$$

**Exercise 10.3.6**

The concentration of $Fe^{3+}$ ion in a sample of $H_2O$ is 335.0 ppm. What mass of $Fe^{3+}$ ion is present in 3,450 mL of $H_2O$, which has a density of 1.00 g/mL?

**Answer**

1.16 g

For ionic solutions, we need to differentiate between the concentration of the salt versus the concentration of each individual ion. Because the ions in ionic compounds go their own way when a compound is dissolved in a solution, the resulting concentration of the ion may be different from the concentration of the complete salt. For example, if 1 M NaCl were prepared, the solution could also be described as a solution of 1 M $Na^+(aq)$ and 1 M $Cl^-(aq)$ because there is one $Na^+$ ion and one $Cl^-$ ion per formula unit of the salt. However, if the solution were 1 M $CaCl_2$, there are two $Cl^-(aq)$ ions for every formula unit dissolved, so the concentration of $Cl^-(aq)$ would be 2 M, not 1 M.

In addition, the total ion concentration is the sum of the individual ion concentrations. Thus for the 1 M NaCl, the total ion concentration is 2 M; for the 1 M $CaCl_2$, the total ion concentration is 3 M.

**Key Takeaway**

- Quantitative units of concentration include molarity, molality, mass percentage, parts per thousand (ppth), parts per million (ppm), and parts per billion (ppb).

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# 10.4: Dilutions and Concentrations

**Learning Objective**

- Learn how to dilute and concentrate solutions.

Often, a worker will need to change the concentration of a solution by changing the amount of solvent. **Dilution** is the addition of solvent, which decreases the concentration of the solute in the solution. **Concentration** is the removal of solvent, which increases the concentration of the solute in the solution. (Do not confuse the two uses of the word *concentration* here!)

In both dilution and concentration, the amount of solute stays the same. This gives us a way to calculate what the new solution volume must be for the desired concentration of solute. From the definition of molarity,

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

we can solve for the number of moles of solute:

moles of solute = (molarity)(liters of solution)

A simpler way of writing this is to use $M$ to represent molarity and $V$ to represent volume. So the equation becomes

moles of solute = $MV$

Because this quantity does not change before and after the change in concentration, the product $MV$ must be the same before and after the concentration change. Using numbers to represent the initial and final conditions, we have

$$M_1V_1 = M_2V_2$$

as the **dilution equation**. The volumes must be expressed in the same units. Note that this equation gives only the initial and final conditions, not the amount of the change. The amount of change is determined by subtraction.

**Example 10.4.1**

If 25.0 mL of a 2.19 M solution are diluted to 72.8 mL, what is the final concentration?

**Solution**

It does not matter which set of conditions is labeled 1 or 2, as long as the conditions are paired together properly. Using the dilution equation, we have

$$(2.19\text{ M})(25.0\text{ mL}) = M_2(72.8\text{ mL})$$

Solving for the second concentration (noting that the milliliter units cancel),

$$M_2 = 0.752\text{ M}$$

The concentration of the solution has decreased. In going from 25.0 mL to 72.8 mL, 72.8 − 25.0 = 47.8 mL of solvent must be added.

**Exercise 10.4.1**

A 0.885 M solution of KBr whose initial volume is 76.5 mL has more water added until its concentration is 0.500 M. What is the new volume of the solution?

**Answer**

135.4 mL

Concentrating solutions involves removing solvent. Usually this is done by evaporating or boiling, assuming that the heat of boiling does not affect the solute. The dilution equation is used in these circumstances as well.

## Chemistry is Everywhere: Preparing IV Solutions

In a hospital emergency room, a physician orders an intravenous (IV) delivery of 100 mL of 0.5% KCl for a patient suffering from hypokalemia (low potassium levels). Does an aide run to a supply cabinet and take out an IV bag containing this concentration of KCl?

Not likely. It is more probable that the aide must make the proper solution from an IV bag of sterile solution and a more concentrated, sterile solution, called a *stock solution*, of KCl. The aide is expected to use a syringe to draw up some stock solution and inject it into the waiting IV bag and dilute it to the proper concentration. Thus the aide must perform a dilution calculation.



Figure **10.4.1** Preparing IV Solution © Thinkstock. Medical personnel commonly must perform dilutions for IV solutions.

If the stock solution is 10.0% KCl and the final volume and concentration need to be 100 mL and 0.50%, respectively, then it is easy to calculate how much stock solution to use:

$$(10\%)V_1 = (0.50\%)(100\text{ mL})V_1 = 5\text{ mL}$$

Of course, the addition of the stock solution affects the total volume of the diluted solution, but the final concentration is likely close enough even for medical purposes.

Medical and pharmaceutical personnel are constantly dealing with dosages that require concentration measurements and dilutions. It is an important responsibility: calculating the *wrong* dose can be useless, harmful, or even fatal!

## Summary

- Molarity and volume are used to determine dilutions and concentrations of solutions.

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# 10.5: Concentrations as Conversion Factors

**Learning Objective**

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108\ \cancel{L\ NaCl} \times \frac{0.887\ mol\ NaCl}{\cancel{L\ NaCl}} = 0.0958\ mol\ NaCl$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

**Example 10.5.1**

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

**Solution**

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88\ \cancel{mol\ CuSO_4} \times \frac{1\ L}{2.35\ \cancel{mol}} = 2.08\ L\ of\ solution$$

**Exercise 10.5.1**

Using concentration as a conversion factor, how many liters of 0.0444 M $CH_2O$ are needed to obtain 0.0773 mol of $CH_2O$?

**Answer:**

1.74 L

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

**Example 10.5.2**

What mass of solute is present in 0.765 L of 1.93 M NaOH?

**Solution**

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765\ \cancel{L} \times \frac{1.93\ \cancel{mol\ NaOH}}{\cancel{L\ solution}} \times \frac{40.0\ g\ NaOH}{1\ \cancel{mol\ NaOH}} = 59.1\ g\ NaOH$$

**Exercise 10.5.2**

What mass of solute is present in 1.08 L of 0.0578 M $H_2SO_4$?

**Answer**

6.12 g

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:

$$2\,AgNO_3\,(aq) + CaCl_2\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)$$

If we wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, we first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution:

$$1.25\ \cancel{mol\ AgNO_3} \times \frac{1\ \cancel{mol\ CaCl_2}}{2\ \cancel{mol\ AgNO_3}} \times \frac{1L\ solution}{0.555\ \cancel{mol\ CaCl_2}} = 1.13\,L\,CaCl_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

**Example 10.5.3**

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\,Ag\,(s) + Al(NO_3)_3\,(aq) \longrightarrow 3\,AgNO_3 + Al\,(s)$$

**Solution**

Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66\ \cancel{g\ Ag} \times \frac{1\ \cancel{mol\ Ag}}{107.97\ \cancel{g\ Ag}} \times \frac{1\ \cancel{mol\ Al(NO_3)_3}}{3\ \cancel{mol\ Ag}} \times \frac{1L\ solution}{0.0995\ \cancel{mol\ Al(NO_3)_3}} = 0.114\,L$$

The strikeouts show how the units cancel.

**Exercise 10.5.3**

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$H_2C_2O_4\,(s) + 2\,NaOH\,(aq) \longrightarrow Na_2C_2O_4\,(aq) + 2\,H_2O(\ell)$$

**Answer:**

0.777 L

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

**Example 10.5.4**

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of $FeCl_3$. The student carefully adds 0.1074 M $Na_2C_2O_4$ until all the $Fe^{3+}(aq)$ has precipitated as $Fe_2(C_2O_4)_3(s)$. Using a precisely measured tube called a burette, the student finds that 9.04 mL of the $Na_2C_2O_4$ solution was added to completely precipitate the $Fe^{3+}(aq)$. What was the concentration of the $FeCl_3$ in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:

$$2\,FeCl_3\,(aq) + 3\,Na_2C_2O_4\,(aq) \longrightarrow Fe_2(C_2O_4)_3\,(s) + 6\,NaCl\,(aq)$$

### Solution

First we need to determine the number of moles of $Na_2C_2O_4$ that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04\,\cancel{mL} \times \frac{1\,\cancel{L}}{1000\,\cancel{mL}} \times \frac{0.1074 mol\,Na_2C_2O_4}{\cancel{L}} = 0.000971\,mol\,Na_2C_2O_4$$

Now we will use the balanced chemical equation to determine the number of moles of $Fe^{3+}$(aq) that were present in the initial aliquot:

$$0.000971\,\cancel{mol\,Na_2C_2O_4} \times \frac{2 mol\,FeCl_3}{3\,\cancel{mol Na_2C_2O_4}} = 0.000647 mol\,FeCl_3$$

Then we determine the concentration of $FeCl_3$ in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{mol}{L} = \frac{0.000647 mol\,FeCl_3}{0.01000 L} = 0.0647 M\,FeCl_3$$

## Exercise 10.5.4

A student titrates 25.00 mL of $H_3PO_4$ with 0.0987 M KOH. She uses 54.06 mL to complete the chemical reaction. What is the concentration of $H_3PO_4$?

$$H_3PO_4\,(aq) + 3\,KOH\,(aq) \longrightarrow K_3PO_4\,(aq) + 3\,H_2O$$

**Answer:**

0.0711 M



Figure **10.5.1** Titration © Thinkstock. When a student performs a titration, a measured amount of one solution is added to another reactant.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

## Example 10.5.5

$H_2O_2$ is used to determine the amount of Mn according to this balanced chemical equation:

$$2\,MnO_4^-(aq) + 5\,H_2O_2(aq) + 6\,H^+(aq) \longrightarrow 2\,Mn_2^+(aq) + 5\,O_2(g) + 8\,H_2O(\ell)$$

What mass of 3.00% m/m $H_2O_2$ solution is needed to react with 0.355 mol of $MnO_4^-$(aq)?

### Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of $H_2O_2$ and then convert to find the mass of $H_2O_2$. Knowing that the $H_2O_2$ solution is 3.00% by mass, we can determine the mass of solution needed:

$$0.355\ \cancel{mol\ MnO_4^-} \times \frac{5\ \cancel{mol\ H_2O_2}}{2\ \cancel{mol\ MnO_4^-}} \times \frac{34.02\ \cancel{g\ H_2O_2}}{\cancel{mol\ H_2O_2}} \times \frac{100 g\ solution}{3\ \cancel{g\ H_2O_2}} = 1006 g\ sol$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of $H_2O_2$, and the third conversion factor comes from the definition of percentage concentration by mass.

## Exercise 10.5.5

Use the balanced chemical reaction for $MnO_4^-$ and $H_2O_2$ to determine what mass of $O_2$ is produced if 258 g of 3.00% m/m $H_2O_2$ is reacted with $MnO_4^-$.

**Answer**

7.28 g

## Summary

Know how to apply concentration units as conversion factors.

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# CHAPTER OVERVIEW

## 11: Acids and Bases

Acids and bases are common substances found in many every day items, from fruit juices and soft drinks to soap. In this chapter, we will examine the properties of acids and bases, and learn about the chemical nature of these important compounds. We will cover pH, and how to calculate the pH of a solution.


11.1: Sour Patch Kids and International Spy Movies
11.2: Acids- Properties and Examples
11.3: Bases- Properties and Examples
11.4: Molecular Definitions of Acids and Bases
11.5: Reactions of Acids and Bases
11.6: Acid–Base Titration
11.7: Strong and Weak Acids and Bases
11.8: Water - Acid and Base in One
11.9: The pH and pOH Scales - Ways to Express Acidity and Basicity
11.10: Buffers- Solutions that Resist pH Change


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# 11.1: Sour Patch Kids and International Spy Movies

Sour Patch Kids are a soft candy with a coating of invert sugar and sour sugar. The candy's slogan, "Sour. Sweet. Gone.", refers to its sour-to-sweet taste.



Figure **11.1.1**: Sour Patch Kids candies. Image courtesy of Evan-Amos (public domain).

Sour sugar is a food ingredient that is used to impart a sour flavor, made from citric or tartaric acid and sugar. It is used to coat sour candies like Sour Patch Kids. Eating large amounts of sour sugar can cause irritation of the tongue because of the acid. It can also cause irreversible dental erosion.

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# 11.2: Acids- Properties and Examples

**Learning Objectives**

- Examine properties of acids.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

## Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g) \tag{11.2.1}$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

It should not be hard for you to name several common acids (but you might find that listing bases is a little more difficult). Below is a partial list of some common acids, along with some chemical formulas:

Table **11.2.1**: Common Acids and Their Uses

| Chemist Name | Common Name | Uses |
|---|---|---|
| hydrochloric acid, HCl | muriatic acid (used in pools) and stomach acid is HCl | Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning. |
| sulfuric acid, $H_2SO_4$ | | Used in car batteries, and in the manufacture of fertilizers. |
| nitric acid, $HNO_3$ | | Used in the manufacture of fertilizers, explosives and in extraction of gold. |
| acetic acid, $HC_2H_3O_2$ | vinegar | Main ingredient in vinegar. |
| carbonic acid, $H_2CO_3$ | responsible for the "fizz" in carbonated drinks | As an ingredient in carbonated drinks. |
| citric acid, $C_6H_8O_7$ | | Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions. |
| acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$ | aspirin | The active ingredient in aspirin. |

What exactly makes an acid an acid, and what makes a base act as a base? Take a look at the formulas given in the above table and take a guess.

## Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. A colorless, highly pungent solution of hydrogen chloride (HCl) in water. Hydrochloric acid is usually prepared by treating $\mathbf{HCl}$ with water.

$$\mathbf{HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)}$$

Hydrochloric acid can therefore be used to prepare chloride salts. Hydrochloric acid is a strong acid, since it is completely dissociated in water. Hydrochloric acid is the preferred acid in titration for determining the amount of bases.

## Sulfuric Acid

Sulfuric acid is a highly corrosive strong mineral acid with the molecular formula $\mathbf{H_2SO_4}$. Sulfuric acid is a diprotic acid and has a wide range of applications including use in domestic acidic drain cleaners,[] as an electrolyte in lead-acid batteries, and in various cleaning agents. It is also a central substance in the chemical industry.



Figure **11.2.1**: Drops of concentrated sulfuric acid rapidly decompose a piece of cotton towel by dehydration. (CC BY-SA 3.0; Toxic Walker).

Because the hydration of sulfuric acid is thermodynamically favorable (and is highly exothermic) and the affinity of it for water is sufficiently strong, sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating

property, removing water ($H_2O$) from other compounds including sugar and other carbohydrates and producing carbon, heat, steam. Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas (Equation 11.2.2).

$$M + H_2SO_4 \longrightarrow M(SO_4) + H_2 \tag{11.2.2}$$

## Nitric Acid

Nitric acid ($HNO_3$) is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. Nitric acid can be made by reacting nitrogen dioxide ($NO_2\,(g)$) with water.

$$3\,NO_2\,(g) + H_2O\,(l) \longrightarrow 2\,HNO_3\,(ag) + NO\,(g)$$

Nitric acid reacts with most metals, but the details depend on the concentration of the acid and the nature of the metal. Dilute nitric acid behaves as a typical acid in its reaction with most metals (e.g., nitric acid with magnesium, manganese or zinc will liberate $H_2$ gas):

$$Mg + 2\,HNO_3 \longrightarrow Mg(NO_3)_2 + H_2$$

$$Mn + 2\,HNO_3 \longrightarrow Mn(NO_3)_2 + H_2$$

$$Zn + 2\,HNO_3 \longrightarrow Zn(NO_3)_2 + H_2$$

Nitric acid is a corrosive acid and a powerful oxidizing agent. The major hazard it poses is chemical burn, as it carries out acid hydrolysis with proteins (amide) and fats (ester) which consequently decomposes living tissue (Figure 11.2.2). Concentrated nitric acid stains human skin yellow due to its reaction with the keratin



Figure 11.2.2: Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).

## Carbonic Acid

Carbonic acid is a chemical compound with the chemical formula $H_2CO_3$ and is also a name sometimes given to solutions of carbon dioxide in water (carbonated water), because such solutions contain small amounts of $H_2CO_3\,(aq)$. Carbonic acid, which is a weak acid, forms two kinds of salts: the carbonates and the bicarbonates. In geology, carbonic acid causes limestone to dissolve, producing calcium bicarbonate—which leads to many limestone features such as stalactites and stalagmites. Carbonic acid is a polyprotic acid, specifically it is diprotic, meaning that it has two protons which may dissociate from the parent molecule.

When carbon dioxide dissolves in water, it exists in chemical equilibrium (discussed in Chapter 15), producing carbonic acid:

$$CO_2 + H_2O \rightleftharpoons H_2CO_3$$

The reaction can be pushed to favor the reactants to generate $CO_2\,(g)$ from solution, which is key to the bubbles observed in carbonated beverages (Figure 11.2.3).



Figure 11.2.3: A glass of sparkling water. (CC BY-SA 3.0; Nevit Dilmen).

## Formic Acid

Formic acid ($HCO_2H$) is the simplest carboxylic acid and is an important intermediate in chemical synthesis and occurs naturally, most notably in some ants. The word "formic" comes from the Latin word for ant, formica, referring to its early isolation by the distillation of ant bodies. Formic acid occurs widely in nature as its conjugate base formate.

## Citric Acid

Citric acid ($C_6H_8O_7$) is a weak organic tricarboxylic acid that occurs naturally in citrus fruits. The citrate ion is an intermediate in the TCA cycle (Krebs cycle), a central metabolic pathway for animals, plants and bacteria. Because it is one of the stronger edible acids, the dominant use of citric acid is used as a flavoring and preservative in food and beverages, especially soft drinks.

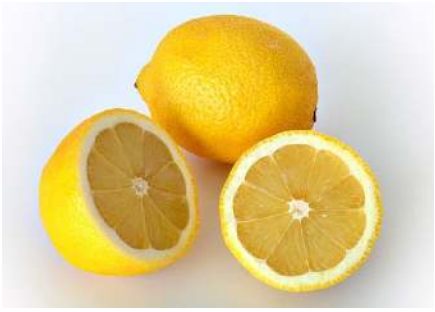


Figure 11.2.4: Lemons, oranges, limes, and other citrus fruits possess high concentrations of citric acid (CC BY-SA 2.5; André Karwath).

## Acetylsalicylic Acid

Acetylsalicylic acid (also known as aspirin) is a medication used to treat pain, fever, and inflammation. Aspirin, in the form of leaves from the willow tree, has been used for its health effects for at least 2,400 years.

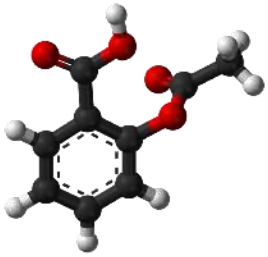


Figure 11.2.5: Ball-and-stick model of the aspirin molecule. (Public Domain; Ben Mills).

Aspirin is a white, crystalline, weakly acidic substance.

## Summary

A brief summary of key aspects of several acids commonly encountered by students was given. Acids are a distinct class of compounds because of the properties of their aqueous solutions.

## Contributions & Attributions

- 
- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

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# 11.3: Bases- Properties and Examples

**Learning Objectives**

- Examine properties of bases.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the **HCl** in the stomach and neutralize it, taking care of that unpleasant feeling.

## Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base, while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.



Figure **11.3.1**: Phenolphthalein indicator in presence of base.

**Warning!**

Tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words—don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate. Various common bases and corresponding uses are given in Table **11.3.2**.

Table **11.3.1**: Common Bases and Corresponding Uses

| Some Common Bases | Uses |
|---|---|
| sodium hydroxide, NaOH (lye or caustic soda) | Used in the manufacture of soaps and detergents, and as the main ingredient in oven and drain cleaners. |
| potassium hydroxide, KOH (lye or caustic potash) | Used in the production of liquid soaps and soft soaps. Used in alkaline batteries. |
| magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia) | Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater. |
| calcium hydroxide, $Ca(OH)_2$ (slaked lime) | Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil. |
| aluminum hydroxide | Used in water purification and as an ingredient in antacids. |
| ammonia, $NH_3$ | Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers. |

## Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula $\mathbf{NaOH}$. It is a white solid ionic compound consisting of sodium cations $\mathbf{Na^+}$ and hydroxide anions $\mathbf{OH^-}$.

Dissolution of solid sodium hydroxide in water is a highly exothermic reaction:

$$\mathbf{NaOH\,(s) \rightarrow Na^+\,(aq) + OH^-\,(aq)}$$

The resulting solution is usually colorless and odorless and feels slippery when it comes in contact with skin.



Figure 11.3.1: Sample of sodium hydroxide as pellets in a watch glass. (Public Domain; Walkerma.)

## Potassium Hydroxide

Potassium hydroxide is an inorganic compound with the formula $\mathbf{KOH}$, and is commonly called caustic potash. Along with sodium hydroxide (NaOH), this colorless solid is a prototypical strong base. It has many industrial and niche applications, most of which exploit its corrosive nature and its reactivity toward acids. Its dissolution in water is strongly exothermic.

$$\mathbf{KOH\,(s) \rightarrow K^+\,(aq) + OH^-\,(aq)}$$

Concentrated aqueous solutions are sometimes called *potassium lyes*.

## Magnesium Hydroxide

Magnesium hydroxide is the inorganic compound with the chemical formula $\mathbf{Mg(OH)_2}$. Magnesium hydroxide is a common component of antacids, such as milk of magnesia, as well as laxatives.




Figure 11.3.1: Bottle of Antacid tablets. (CC BY 2.,5; Midnightcomm).

It is a white solid with low solubility in water. Combining a solution of many magnesium salts with basic water induces precipitation of solid $\mathbf{Mg(OH)_2}$. However, a weak concentration of dissociated ions can be found in solution:

$$\mathrm{Mg(OH)_2\,(s) \rightleftharpoons Mg^{2+}\,(aq) + 2\,OH^-\,(aq)}$$

## Calcium Hydroxide

Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $\mathbf{Ca(OH)_2}$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime. Calcium hydroxide is used in many applications, including food preparation. Limewater is the common name for a saturated solution of calcium hydroxide.

Calcium hydroxide is relatively insoluble in water, but is large enough that its solutions are basic according to the following reaction:

$$\mathrm{Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2\,OH^-\,(aq)}$$

## Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula $\mathbf{NH_3}$ and is a colorless gas with a characteristic pungent smell. It is the active product of “smelling salts,” and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form.

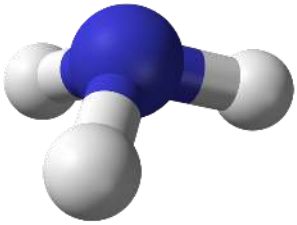


Figure 11.3.1: Ball-and-stick model of the ammonia molecule. (Public Domain; Ben Mills).

In aqueous solution, ammonia acts as a base, acquiring hydrogen ions from $\mathbf{H_2O}$ to yield ammonium and hydroxide ions:

$$\mathrm{NH_3\,(g) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)}$$

Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

## Summary

- A brief summary of properties of bases was given.
- The properties of bases mostly contrast those of acids.
- Bases have many, varied uses.

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# 11.4: Molecular Definitions of Acids and Bases

**Learning Objectives**

- Identify an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

There are three major classifications of substances known as acids or bases. The theory developed by Svante Arrhenius in 1883, the Arrhenius definition, states that an acid produces $H^+$ in solution and a base produces $OH^-$. Later, two more sophisticated and general theories were proposed. These theories are the Brønsted-Lowry and Lewis definitions of acids and bases. This section will cover the Arrhenius and Brønsted-Lowry theories; the Lewis theory is discussed elsewhere.

## The Arrhenius Theory of Acids and Bases

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. An **Arrhenius acid** is a compound that **increases the concentration of $H^+$ ions** that are present when added to water. These $H^+$ ions form the hydronium ion ($H_3O^+$) when they combine with water molecules. This process is represented in a chemical equation by adding $H_2O$ to the reactants side.

$$HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-\,(aq)$$

In this reaction, hydrochloric acid (*HCl*) dissociates completely into hydrogen ($H^+$) and chlorine ($Cl^-$) ions when dissolved in water, thereby releasing $H^+$ ions into solution. Formation of the hydronium ion equation:

$$HCl\,(aq) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)$$

An **Arrhenius base** is a compound that **increases the concentration of $OH^-$ ions** that are present when added to water. The dissociation is represented by the following equation:

$$NaOH\ \ (aq) \rightarrow Na^+\ \ (aq) + OH^-\ \ (aq)$$

In this reaction, sodium hydroxide (NaOH) disassociates into sodium ($Na^+$) and hydroxide ($OH^-$) ions when dissolved in water, thereby releasing $OH^-$ ions into solution.

> *Arrhenius acids are substances which produce hydrogen ions in solution and Arrhenius bases are substances which produce hydroxide ions in solution.*

**Limitations to the Arrhenius Theory**

The Arrhenius theory has many more limitations than the other two theories. The theory does not explain the weak base ammonia ($NH_3$), which in the presence of water, releases hydroxide ions into solution, but does not contain OH- itself. The Arrhenius definition of acid and base is also limited to aqueous (i.e., water) solutions.

## The Brønsted-Lowry Theory of Acids and Bases

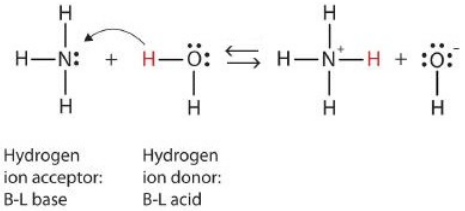
In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton ($H^+$) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, **a Brønsted-Lowry acid** is a **proton donor (PD)**, while **a Brønsted-Lowry base** is a **proton acceptor (PA)**.

> *A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.*

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$NH_3\,(aq) + H_2O\,(\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq) \tag{11.4.1}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:

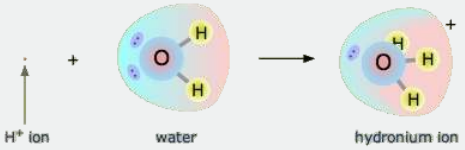



Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the $H^+$ ion attaches itself to $H_2O$ to make $H_3O^+$, which is called the *hydronium ion*. For most purposes, $H^+$ and $H_3O^+$ represent the same species, but writing $H_3O^+$ instead of $H^+$ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

### The Hydronium Ion

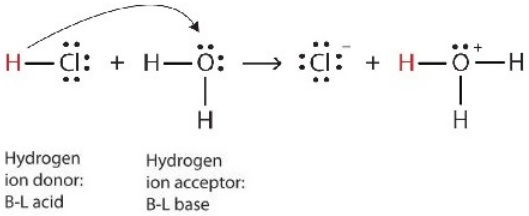
A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than $H_3O^+$. It is simpler, however, to use $H_3O^+$ to represent the hydronium ion.




With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in $H_2O$:

$$HCl\,(g) + H_2O\,(\ell) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq) \tag{11.4.2}$$

We can depict this process using Lewis electron dot diagrams:




Now we see that a hydrogen ion is transferred from the HCl molecule to the $H_2O$ molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, $H_2O$ is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense, but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, $H_2O$ is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled $H_2O$ a base in this circumstance.

- A Brønsted-Lowry acid is a proton (hydrogen ion) donor.
- A Brønsted-Lowry base is a proton (hydrogen ion) acceptor.
- All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

### Example 11.4.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule, just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

**Solution**

$C_6H_5NH_2$ and $H_2O$ are the reactants. When $C_6H_5NH_2$ accepts a proton from $H_2O$, it gains an extra H and a positive charge and leaves an $OH^-$ ion behind. The reaction is as follows:

$$\mathbf{C_6H_5NH_2\,(aq) + H_2O(\ell) \rightleftharpoons C_6H_5NH_3^+\,(aq) + OH^-\,(aq)}$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The $H_2O$ molecule, because it donates a proton, is the Brønsted-Lowry acid.

### Exercise 11.4.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$\mathbf{H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+}$$

**Answer**

Brønsted-Lowry acid: $H_2PO_4^-$; Brønsted-Lowry base: $H_2O$

### Exercise 11.4.2

Which of the following compounds is a Bronsted-Lowry base?

a. HCl
b. $HPO_4^{2-}$
c. $H_3PO_4$
d. $NH_4^+$
e. $CH_3NH_3^+$

**Answer**

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an $H^+$. This eliminates $\mathbf{HCl}$, $\mathbf{H_3PO_4}$, $\mathbf{NH_4^+}$ and $\mathbf{CH_3NH_3^+}$ because they are Bronsted-Lowry acids. They all give away protons. In the case of $\mathbf{HPO_4^{2-}}$, consider the following equation:

$$\mathbf{HPO_4^{2-}\,(aq) + H_2O\,(l) \rightarrow PO_4^{3-}\,(aq) + H_3O^+\,(aq)}$$

Here, it is clear that $HPO_4^{2-}$ is the acid since it donates a proton to water to make $H_3O^+$ and $PO_4^{3-}$. Now consider the following equation:

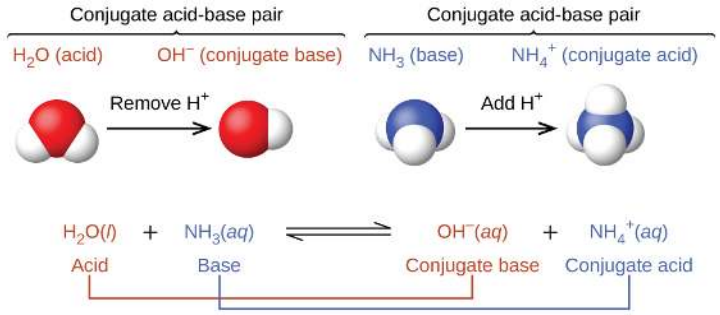
$$\mathbf{HPO_4^{2-}\,(aq) + H_2O\,(l) \rightarrow H_2PO_4^- + OH^-\,(aq)}$$

In this case, $HPO_4^{2-}$ is the base since it accepts a proton from water to form $H_2PO_4^-$ and $OH^-$. Thus, $HPO_4^{2-}$ is an acid and base together, making it amphoteric.

Since $HPO_4^{2-}$ is the only compound from the options that can act as a base, the answer is **(b) $\mathbf{HPO_4^{2-}}$.**

## Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, $\mathbf{OH^-}$, and the conjugate acid of ammonia, $\mathbf{NH_4^+}$:




This figure has three parts in two rows. In the first row, two diagrams of acid-base pairs are shown. On the left, a space filling model of H subscript 2 O is shown with a red O atom at the center and two smaller white H atoms attached in a bent shape. Above this model is the label “H subscript 2 O (acid)” in purple. An arrow points right, which is labeled “Remove H superscript plus.” To the right is another space filling model with a single red O atom to which a single smaller white H atom is attached. The label in purple above this model reads, “O H superscript negative (conjugate base).” Above both of these red and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” To the right is a space filling model with a central blue N atom to which three smaller white H atoms are attached in a triangular pyramid arrangement. A label in green above reads “N H subscript 3 (base).” An arrow labeled “Add H superscript plus” points right. To the right of the arrow is another space filling model with a blue central N atom and four smaller white H atoms in a tetrahedral arrangement. The green label above reads “N H subscript 3 superscript plus (conjugate acid).” Above both of these blue and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” The second row of the figure shows the chemical reaction, H subscript 2 O ( l ) is shown in purple, and is labeled below in purple as “acid,” plus N H subscript 3 (a q) in green, labeled below in green as “base,” followed by a double sided arrow arrow and O H superscript negative (a q) in purple, labeled in purple as “conjugate base,” plus N H subscript 4 superscript plus (a q)” in green, which is labeled in green as “conjugate acid.” The acid on the left side of the equation is connected to the conjugate base on the right with a purple line. Similarly, the base on the left is connected to the conjugate acid on the right side.

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are $\boldsymbol{NH_4^+/NH_3}$ and $\boldsymbol{H_2O/OH^-}$.

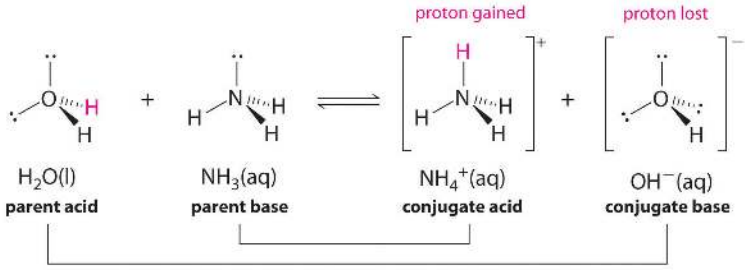



Figure **11.4.1**. The pairing of parent acids and bases with conjugate acids and bases.

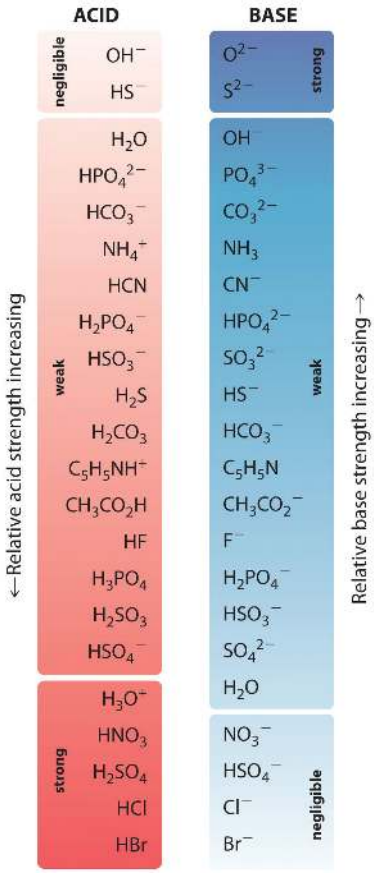



Figure 11.4.1: The Relative Strengths of Some Common Conjugate Acid-Base Pairs. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

## Example 11.4.2

Identify the conjugate acid-base pairs in this equilibrium.

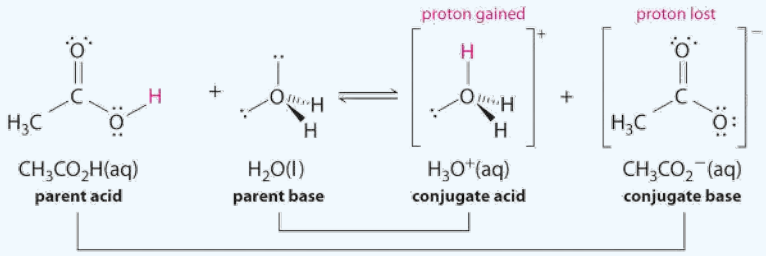
$$\mathrm{CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-}$$

### Solution

Similarly, in the reaction of acetic acid with water, acetic acid **donates** a proton to water, which acts as the base. In the reverse reaction, $\boldsymbol{H_3O^+}$ is the acid that donates a proton to the acetate ion, which acts as the base.

Once again, we have two conjugate acid-base pairs:

- the parent acid and its conjugate base ($\boldsymbol{CH_3CO_2H/CH_3CO_2^-}$) and
- the parent base and its conjugate acid ($\boldsymbol{H_3O^+/H_2O}$).



**Example 11.4.3**

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

**Solution**

One pair is $H_2O$ and $OH^-$, where $H_2O$ has one more $H^+$ and is the conjugate acid, while $OH^-$ has one less $H^+$ and is the conjugate base.

The other pair consists of $(CH_3)_3N$ and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and $(CH_3)_3N$ is the conjugate base.

**Exercise 11.4.3**

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

**Answer**

$H_2O$ (acid) and $OH^-$ (base); $NH_2^-$ (base) and $NH_3$ (acid)

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# 11.5: Reactions of Acids and Bases

**Learning Objectives**

- Write acid-base neutralization reactions.
- Write reactions of acids with metals.
- Write reactions of bases with metals.

## Neutralization Reactions

The reaction that happens when an acid, such as **HCl**, is mixed with a base, such as **NaOH**:

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)}$$

When an acid and a base are combined, water and a salt are the products. Salts are ionic compounds containing a positive ion other than $\mathrm{H^+}$ and a negative ion other than the hydroxide ion, $\mathrm{OH^-}$. Double displacement reactions of this type are called **neutralization reactions**. We can write an expanded version of this equation, with aqueous substances written in their longer form:

$$\mathrm{H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow Na^+\,(aq) + Cl^-\,(aq) + H_2O\,(l)}$$

After removing the spectator ions, we get the net ionic equation:

$$\mathrm{H^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O\,(l)}$$

When a strong acid and a strong base are combined in the proper amounts—when $[\mathrm{H^+}]$ equals $[\mathrm{OH^-}]$\)—a neutral solution results in which pH = 7. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

Salt solutions do not always have a pH of 7, however. Through a process known as **hydrolysis**, the ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. As a general concept, if a strong acid is mixed with a weak base, the resulting solution will be slightly acidic. If a strong base is mixed with a weak acid, the solution will be slightly basic.




Video: Equimolar (~0.01 M) and equivolume solutions of **HCl** and **NaOH** are combined to make salt water. https://youtu.be/TS-I9KrUjB0

**Example 11.5.1: Propionic Acid + Calcium Hydroxide**

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($\mathrm{CH_3CH_2CO_2H}$) with aqueous calcium hydroxide [$\mathrm{Ca(OH)_2}$].

**Solution**

Solutions to Example 14.5.1

| Steps | Reaction |
|---|---|
| *Write the unbalanced equation.*<br>This is a double displacement reaction, so the cations and anions swap to create the water and the salt. | $CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + H_2O(l)$ |
| *Balance the equation.*<br>Because there are two $OH^-$ ions in the formula for $Ca(OH)_2$, we need two moles of propionic acid, $CH_3CH_2CO_2H$, to provide $H^+$ ions. | $\underline{2}CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + \underline{2}H_2O(l)$ |

**? Exercise 11.5.1**

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

**Answer**

$$Ba(OH)_2\,(s) + 2\,CH_3CO_2H\,(aq) \longrightarrow Ba(CH_3CO_2)_2\,(aq) + 2\,H_2O\,(l)$$

## Acids and Bases React with Metals

Acids react with most metals to form a salt and hydrogen gas. As discussed previously, metals that are more active than acids can undergo a **single displacement reaction**. For example, zinc metal reacts with hydrochloric acid, producing zinc chloride and hydrogen gas.

$$Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)$$

Bases also react with certain metals, like zinc or aluminum, to produce hydrogen gas. For example, sodium hydroxide reacts with zinc and water to form sodium zincate and hydrogen gas.

$$Zn\,(s) + 2\,NaOH\,(aq) + 2\,H_2O\,(l) \longrightarrow Na_2Zn(OH)_4\,(aq) + H_2\,(g).$$

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# 11.6: Acid–Base Titration

**Learning Objectives**

- Understand the basics of acid-base titrations.
- Understand the use of indicators.
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because so many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

During an acid-base titration, an acid with a known concentration (a **standard solution**) is slowly added to a base with an unknown concentration (or vice versa). A few drops of indicator solution are added to the base. The indicator will signal, by color change, when the base has been neutralized (when $[H^+] = [OH^-]$). At that point—called the **equivalence point,** or **end point**—the titration is stopped. By knowing the volumes of acid and base used, and the concentration of the standard solution, calculations allow us to determine the concentration of the other solution.

It is important to accurately measure volumes when doing titrations. The instrument you would use is called a burette (or buret).



Figure **11.6.1**: Equipment for Titrations. A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed.

For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the sample? We can calculate the number of moles of HCl reacted:

$$\text{\# mol HCl} = (0.02566\ \text{L})(0.1078\ \text{M}) = 0.002766\ \text{mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:

$$\mathrm{HCl + NaOH \longrightarrow NaCl + H_2O}$$

So we can construct a conversion factor to convert to number of moles of NaOH reacted:

$$0.002766\ \cancel{mol\ HCl} \times \frac{1\ mol\ NaOH}{1\ \cancel{mol\ HCl}} = 0.002766\ mol\ NaOH$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766\ \cancel{mol\ HCl} \times \frac{40.00\ g\ NaOH}{1\ \cancel{mol\ HCl}} = 0.1106\ g\ NaOH$$

This type of calculation is performed as part of a titration.

**Example 11.6.1: Equivalence Point**

What mass of $Ca(OH)_2$ is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M $HNO_3$? The balanced chemical equation is as follows:

$$\mathrm{2\,HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2\,H_2O}$$

**Solution**

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

# moles $HNO_3$ = (0.04402 L)(0.0885 M) = 0.00390 mol $HNO_3$

Using the balanced chemical equation, we can determine the number of moles of $Ca(OH)_2$ present in the analyte:

$$0.00390\ \cancel{mol\ HNO_3} \times \frac{1\ mol\ Ca(OH)_2}{2\ \cancel{mol\ HNO_3}} = 0.00195\ mol\ Ca(OH)_2$$

Then we convert this to a mass using the molar mass of $Ca(OH)_2$:

$$0.00195\ \cancel{mol\ Ca(OH)_2} \times \frac{74.1\ g\ Ca(OH)_2}{\cancel{mol\ Ca(OH)_2}} = 0.144\ g\ Ca(OH)_2$$

**? Exercise 11.6.1**

What mass of $H_2C_2O_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is as follows:

$$H_2C_2O_4 + 2\,NaOH \longrightarrow Na_2C_2O_4 + 2\,H_2O$$

**Answer**

0.182 g

**? Exercise 11.6.2**

If 25.00 mL of HCl solution with a concentration of 0.1234 M is neutralized by 23.45 mL of NaOH, what is the concentration of the base?

**Answer**

0.1316 M NaOH

**? Exercise 11.6.3**

A 20.0 mL solution of strontium hydroxide, $Sr(OH)_2$, is placed in a flask and a drop of indicator is added. The solution turns color after 25.0 mL of a standard 0.0500 M HCl solution is added. What was the original concentration of the $Sr(OH)_2$ solution?

**Answer**

$3.12 \times 10^{-2}\ M$ $Sr(OH)_2$

## Indicator Selection for Titrations

The indicator used depends on the type of titration performed. The indicator of choice should change color when enough of one substance (acid or base) has been added to exactly use up the other substance. Only when a strong acid and a strong base are produced will the resulting solution be neutral. The three main types of acid-base titrations, and suggested indicators, are:

The three main types of acid-base titrations, suggested indicators, and explanations

| Titration between . . . | Indicator | Explanation |
|---|---|---|
| strong acid and strong base | any | |
| strong acid and weak base | methyl orange | changes color in the acidic range (3.2 - 4.4) |
| weak acid and strong base | phenolphthalein | changes color in the basic range (8.2 - 10.6) |

## Summary

A titration is the quantitative reaction of an acid and a base. Indicators are used to show that all the analyte has reacted with the titrant.

## Contributions & Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

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# 11.7: Strong and Weak Acids and Bases

**Learning Objectives**

- Define a strong and a weak acid and base.
- Recognize an acid or a base as strong or weak.
- Determine if a salt produces an acidic or a basic solution.

## Strong and Weak Acids

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider HCl(aq). When HCl is dissolved in $H_2O$, it completely dissociates into $H^+$(aq) and $Cl^-$(aq) ions; all the HCl molecules become ions:

$$HCl \xrightarrow{100\%} H^+(aq) + Cl^-(aq)$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. $HC_2H_3O_2$ is an example of a weak acid:

$$HC_2H_3O_2 \xrightarrow{\sim 6\%} H^+(aq) + C_2H_3O_2^-(aq)$$

Because this reaction does not go 100% to completion, it is more appropriate to write it as **a reversible reaction**:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

As it turns out, there are very few strong acids, which are given in Table **11.7.1**. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

> *Any acid that dissociates 100% into ions is called a strong acid. If it does not dissociate 100%, it is a weak acid.*

Table **11.7.1**: Strong Acids and Bases

| Acids | Bases |
|---|---|
| HCl | LiOH |
| HBr | NaOH |
| HI | KOH |
| $HNO_3$ | RbOH |
| $H_2SO_4$ | CsOH |
| $HClO_3$ | $Mg(OH)_2$ |
| $HClO_4$ | $Ca(OH)_2$ |
| | $Sr(OH)_2$ |
| | $Ba(OH)_2$ |

## Strong and Weak Bases

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it is a **weak base**. There are very few strong bases (Table **11.7.1**); any base not listed is a weak base. All strong bases are $OH^-$ compounds. So a base based on some other mechanism, such as $NH_3$ (which does not contain $OH^-$ ions as part of its formula), will be a weak base.

**Example 11.7.1**: Identifying Strong and Weak Acids and Bases

Identify each acid or base as strong or weak.

a. HCl
b. $Mg(OH)_2$
c. $C_5H_5N$

**Solution**

a. Because HCl is listed in Table **11.7.1**, it is a strong acid.
b. Because $Mg(OH)_2$ is listed in Table **11.7.1**, it is a strong base.
c. The nitrogen in $C_5H_5N$ would act as a proton acceptor and therefore can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

**Exercise 11.7.1**

Identify each acid or base as strong or weak.

a. **RbOH**
b. **$HNO_2$**

**Answer a**

strong base

**Answer b**

weak acid

**Example 11.7.2**: Characterizing Base Ionization

Write the balanced chemical equation for the dissociation of $Ca(OH)_2$ and indicate whether it proceeds 100% to products or not.

**Solution**

This is an ionic compound of $Ca^{2+}$ ions and $OH^-$ ions. When an ionic compound dissolves, it separates into its constituent ions:

$$Ca(OH)_2 \longrightarrow Ca^{2+}(aq) + 2\,OH^-(aq)$$

Because $Ca(OH)_2$ is listed in Table **11.7.1**, this reaction proceeds 100% to products.

**? Exercise 11.7.2**

Write the balanced chemical equation for the dissociation of hydrazoic acid ($HN_3$) and indicate whether it proceeds 100% to products or not.

**Answer a**

The reaction is as follows:

$$HN_3 \longrightarrow H^+(aq) + N_3^-(aq)$$

It does not proceed 100% to products because hydrazoic acid is not a strong acid.

## Key Takeaways

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.
- Salts of weak acids or bases can affect the acidity or basicity of their aqueous solutions.

## Contributions & Attributions

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# 11.8: Water - Acid and Base in One

**Learning Objectives**

- Describe the autoionization of water.
- Calculate the concentrations of $H_3O^+$ and $OH^-$ in aqueous solutions, knowing the other concentration.

We have already seen that $H_2O$ can act as an acid or a base:

$$\underbrace{NH_3}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons NH_4^+ + OH^-$$

where $H_2O$ acts as an acid (in red).

$$\underbrace{HCl}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \longrightarrow H_3O^+ + Cl^-$$

where $H_2O$ acts as an base (in blue).

It may not surprise you to learn, then, that within any given sample of water, some $H_2O$ molecules are acting as acids, and other $H_2O$ molecules are acting as bases. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{11.8.1}$$

This occurs only to a very small degree: only about 6 in $10^8$ $H_2O$ molecules are participating in this process, which is called the **autoionization of water**.

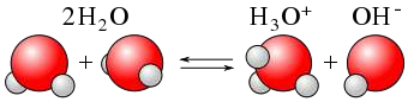



Figure 11.8.1: Autoionization of water, resulting in hydroxide and hydronium ions.

At this level, the concentration of both $H_3O^+(aq)$ and $OH^-(aq)$ in a sample of pure $H_2O$ is about $1.0 \times 10^{-7}\ M$ (at room temperature). If we use square brackets—[ ]—around a dissolved species to imply the molar concentration of that species, we have

$$[H_3O^+]=[OH^-] = 1.0 \times 10^{-7} \tag{11.8.2}$$

for *any* sample of pure water because $H_2O$ can act as both an acid and a base. The product of these two concentrations is $1.0 \times 10^{-14}$:

$$[H_3O^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

- For acids, the concentration of $H_3O^+(aq)$ (i.e., $[H_3O^+]$) is greater than $1.0 \times 10^{-7}\ M$.
- For bases the concentration of $OH^-(aq)$ (i.e., $[OH^-]$) is greater than $1.0 \times 10^{-7}\ M$.

However, the *product* of the two concentrations—$[H_3O^+][OH^-]$—is *always* equal to $1.0 \times 10^{-14}$, no matter whether the aqueous solution is an acid, a base, or neutral:

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted $K_w$:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{11.8.3}$$

This means that if you know $[H_3O^+]$ for a solution, you can calculate what $[OH^-]$ has to be for the product to equal $1.0 \times 10^{-14}$; or if you know $[OH^-]$, you can calculate $[H_3O^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of $K_w$.

**Warning: Temperature Matters**

The degree of autoionization of water (Equation 11.8.1)—and hence the value of $K_w$—changes with temperature, so Equations 11.8.2 - 11.8.3 are accurate only at room temperature.

**Example 11.8.1: Hydroxide Concentration**

What is $[OH^-]$) of an aqueous solution if $[H_3O^+]$ is $1.0 \times 10^{-4} M$?

**Solution**

Solutions to Example 14.7.1

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: $[H_3O^+] = 1.0 \times 10^{-4}\ M$<br>Find: $[OH^-]$ = ? M |
| List other known quantities. | none |
| Plan the problem. | Using the expression for $K_w$, (Equation 11.8.3), rearrange the equation algebraically to solve for $[OH^-]$.<br>$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$$ |
| Calculate. | Now substitute the known quantities into the equation and solve.<br>$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} M$$<br>It is assumed that the concentration unit is molarity, so $[OH^-]$ is $1.0 \times 10^{-10}$ M. |
| Think about your result. | The concentration of the acid is high (> 1 x $10^{-7}$ M), so $[OH^-]$ should be low. |

**Exercise 11.8.1**

What is $[OH^-]$ in a 0.00032 M solution of $H_2SO_4$?

**Hint**

Assume **both** protons ionize from the molecule...although this is not the case.

**Answer**

$3.1 \times 10^{-11}\ M$

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of $H_3O^+$ or $OH^-$ ions in the formula unit because $[H_3O^+]$ or $[OH^-]$) may not be the same as the concentration of the acid or base itself.

**Example 11.8.2: Hydronium Concentration**

What is $[H_3O^+]$ in a 0.0044 M solution of $Ca(OH)_2$?

**Solution**

Solutions to Example 14.7.2

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: $[Ca(OH)_2] = 0.0044\,M$<br>Find: $[H_3O^+]$ = ? M |
| List other known quantities. | We begin by determining $[OH^-]$. The concentration of the solute is 0.0044 M, but because $Ca(OH)_2$ is a strong base, there are two $OH^-$ ions in solution for every formula unit dissolved, so the actual $[OH^-]$ is two times this:<br>$$[OH^-] = 2 \times 0.0044\ \text{M} = 0.0088\,\text{M}\cdot$$ |
| Plan the problem. | Use the expression for $K_w$ (Equation 11.8.3) and rearrange the equation algebraically to solve for $[H_3O^+]$.<br>$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{[OH^-]}$$ |
| Calculate. | Now substitute the known quantities into the equation and solve.<br>$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{(0.0088)} = 1.1 \times 10^{-12} M$$<br>$[H_3O^+]$ has decreased significantly in this basic solution. |
| Think about your result. | The concentration of the base is high (> 1 x $10^{-7}$ M) so Extra close brace or missing open brace should be low. |

**? Exercise 11.8.2**

What is $[H_3O^+]$ of an aqueous solution if $[OH^-]$ is $1.0 \times 10^{-9}\,M$?

**Answer**

$1.0 \times 10^{-5}$ M

In any aqueous solution, the product of $[H_3O^+]$ and $[OH^-]$ equals $1.0 \times 10^{-14}$ (at room temperature).

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# 11.9: The pH and pOH Scales - Ways to Express Acidity and Basicity

**Learning Objectives**

- Define pH and pOH.
- Determine the pH of acidic and basic solutions.
- Determine the hydronium ion concentration and pOH from pH.

As we have seen, $[H_3O^+]$ and $[OH^-]$ values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

$pH$ is a logarithmic function of $[H_3O^+]$:

$$pH = -\log[H_3O^+] \tag{11.9.1}$$

$pH$ is usually (but not always) between 0 and 14. Knowing the dependence of $pH$ on $[H_3O^+]$, we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

This is known as the $pH$ scale. The pH scale is the range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use $pH$ to make a quick determination whether a given aqueous solution is acidic, basic, or neutral. Figure **11.9.1** illustrates this relationship, along with some examples of various solutions. Because hydrogen ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$), the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the ***negative* log of** $[H_3O^+]$, which will give a positive value for pH.

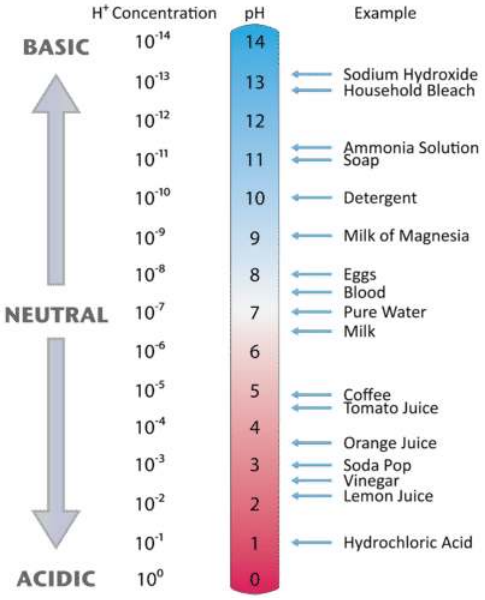



Figure **11.9.1**: The $pH$ values for several common materials.

**Example 11.9.1**

Label each solution as acidic, basic, or neutral based only on the stated $pH$.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

***Answer***

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)

b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

**? Exercise 11.9.1**

Identify each substance as acidic, basic, or neutral based only on the stated $pH$.

a. human blood with $pH$ = 7.4
b. household ammonia with $pH$ = 11.0
c. cherries with $pH$ = 3.6

**Answer a**

basic

**Answer b**

basic

**Answer c**

acidic

## Calculating pH from Hydronium Concentration

The pH of solutions can be determined by using logarithms as illustrated in the next example for stomach acid. Stomach acid is a solution of $HCl$ with a hydronium ion concentration of $1.2 \times 10^{-3}\ M$, what is the $pH$ of the solution?

$$\begin{aligned} \text{pH} &= -\log[\text{H}_3\text{O}^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) = 2.92 \end{aligned}$$

**Logarithms**

To get the **log value** on your calculator, enter the number (in this case, the hydronium ion concentration) first, then press the LOG key.

If the number is 1.0 x $10^{-5}$ (for $[H_3O^+]$ = 1.0 x $10^{-5}$ M) you should get an answer of "-5".

If you get a different answer, or an error, try pressing the LOG key before you enter the number.

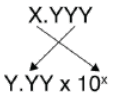
**Example 11.9.2: Converting Ph to Hydronium Concentration**

Find the pH, given the $[H_3O^+]$ of the following:

a. 1 ×$10^{-3}$ M
b. 2.5 ×$10^{-11}$ M
c. 4.7 ×$10^{-9}$ M

**Solution**

Solutions to Example 14.9.2

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>a. $[H_3O^+] = 1 \times 10^{-3}$ M<br>b. $[H_3O^+] = 2.5 \times 10^{-11}$ M<br>c. $[H_3O^+] = 4.7 \times 10^{-9}$ M<br>Find: ? pH |
| Plan the problem. | Need to use the expression for pH (Equation 11.9.1).<br>$pH = - \log [H_3O^+]$ |
| Calculate. | Now substitute the known quantity into the equation and solve.<br>a. $pH = - \log [1 \times 10^{-3}] = 3.0$ (1 decimal places since 1 has 1 significant figure)<br>b. $pH = - \log [2.5 \times 10^{-11}] = 10.60$ (2 decimal places since 2.5 has 2 significant figures)<br>c. $pH = - \log [4.7 \times 10^{-9}] = 8.30$ (2 decimal places since 4.7 has 2 significant figures)<br>The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer:<br>  |

## ? Exercise 11.9.2

Find the pH, given $[H_3O^+]$ of the following:

a. $5.8 \times 10^{-4}$ M
b. $1.0 \times 10^{-7}$

**Answer a**

3.22

**Answer b**

7.00

### Calculating Hydronium Concentration from pH

Sometimes you need to work "backwards"—you know the pH of a solution and need to find $[H_3O^+]$, or even the concentration of the acid solution. How do you do that? To convert pH into $[H_3O^+]$ we solve Equation 11.9.1 for $[H_3O^+]$. This involves taking the antilog (or inverse log) of the negative value of pH .

$$[H_3O^+] = \text{antilog}(-pH)$$

or

$$[H_3O^+] = 10^{-pH} \tag{11.9.2}$$

As mentioned above, different calculators work slightly differently—make sure you can do the following calculations using ***your*** calculator.

## Calculator Skills

We have a solution with a pH = 8.3. What is $[H_3O^+]$ ?

**With some calculators you will do things in the following order:**

1. Enter 8.3 as a negative number (use the key with both the +/- signs, not the subtraction key).
2. Use your calculator's 2nd or Shift or INV (inverse) key to type in the symbol found above the LOG key. The shifted function should be $10^x$.
3. You should get the answer $\mathbf{5.0 \times 10^{-9}}$.

**Other calculators require you to enter keys in the order they appear in the equation.**

1. Use the Shift or second function to key in the $10^x$ function.
2. Use the +/- key to type in a negative number, then type in 8.3.
3. You should get the answer $\mathbf{5.0 \times 10^{-9}}$.

If neither of these methods work, try rearranging the order in which you type in the keys. Don't give up—you must master your calculator!

## Example 11.9.3: Calculating Hydronium Concentration from pH

Find the hydronium ion concentration in a solution with a pH of 12.6. Is this solution an acid or a base? How do you know?

**Solution**

Solutions to Example 14.9.3

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: pH = 12.6<br>Find: $[H_3O^+]$ = ? M |
| Plan the problem. | Need to use the expression for $[H_3O^+]$ (Equation 11.9.2).<br>$[H_3O^+]$ = antilog (-pH) or $[H_3O^+] = 10^{-pH}$ |
| Calculate. | Now substitute the known quantity into the equation and solve.<br>$[H_3O^+]$ = antilog (12.60) = 2.5 x $10^{-13}$ M (2 significant figures since 4.7 has 12.60 2 decimal places)<br>or<br>$[H_3O^+] = 10^{-12.60} = 2.5 \times 10^{-13}$ M (2 significant figures since 4.7 has 12.60 2 decimal places)<br>The other issue that concerns us here is significant figures. Because the number(s) before the decimal point in a logarithm relate to the power on 10, the number of digits *after* the decimal point is what determines the number of significant figures in the final answer:<br>X.YYY<br>Y.YY x $10^x$ |

## Exercise 11.9.3

If moist soil has a pH of 7.84, what is $[H_3O^+]$ of the soil solution?

**Answer**

1.5 x $10^{-8}$ M

## The pOH scale

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** of a solution is the negative logarithm of the hydroxide-ion concentration.

$$\mathrm{pOH} = -\log\left[\mathrm{OH}^-\right]$$

The relation between the hydronium and hydroxide ion concentrations expressed as p-functions is easily derived from the $K_w$ expression:

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH}^-] \tag{11.9.3}$$

$$-\log K_\mathrm{w} = -\log([\mathrm{H_3O^+}][\mathrm{OH}^-]) = -\log[\mathrm{H_3O^+}] + -\log[\mathrm{OH}^-] \tag{11.9.4}$$

$$\mathrm{p}K_\mathrm{w} = \mathrm{pH} + \mathrm{pOH} \tag{11.9.5}$$

At 25 °C, the value of $K_w$ is $1.0 \times 10^{-14}$, and so:

$$14.00 = \mathrm{pH} + \mathrm{pOH} \tag{11.9.6}$$

The hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7}\ M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{11.9.7}$$

$$\mathrm{pOH} = -\log[\mathrm{OH}^-] = -\log(1.0 \times 10^{-7}) = 7.00 \tag{11.9.8}$$

And so, at this temperature, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities less than $1.0 \times 10^{-7}\ M$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7}\ M$ and hydroxide ion molarities greater than $1.0 \times 10^{-7}\ M$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

### Example 11.9.4:

Find the pOH of a solution with a pH of 4.42.

**Solution**

Solutions to Example 14.9.4

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: pH =4.42<br>Find: ? pOH |
| Plan the problem. | Need to use the expression<br>pOH = 14 - pH |
| Calculate. | Now substitute the known quantity into the equation and solve.<br>pOH=14−4.42=9.58 |
| Think about your result. | The pH is that of an acidic solution, and the resulting pOH is the difference after subtracting from 14. The answer has two significant figures because the given pH has two decimal places. |

### Exercise 11.9.4

The pH of a solution is 8.22. What is the pOH?

**Answer**

5.78

The diagram below shows all of the interrelationships between [H3O+][H3O+], [OH−][OH−], pH, and pOH.

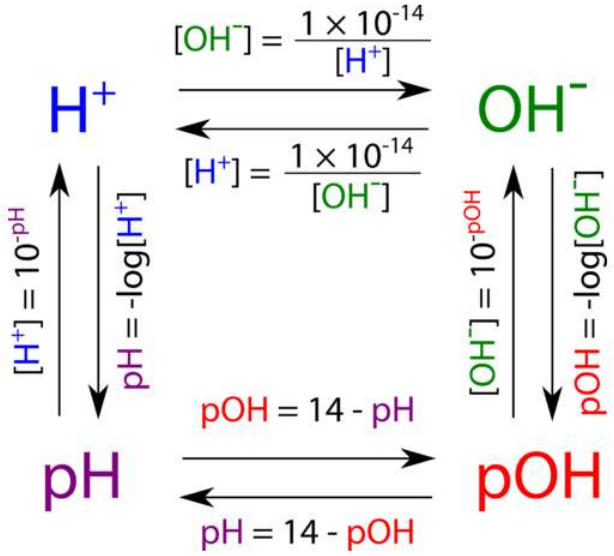



Figure **11.9.1**: Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH.

## Contributions & Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
-
- Template:OpenStax

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# 11.10: Buffers- Solutions that Resist pH Change

**Learning Objective**

- Define *buffer* and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9—a pH that is not conducive to life. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

This mechanism involves a buffer, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid, or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate ($NaC_2H_3O_2$, a salt derived from that acid). Another example of a buffer is a solution containing ammonia ($NH_3$, a weak base) and ammonium chloride ($NH_4Cl$, a salt derived from that base).

Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. If a strong base—a source of $\mathbf{OH^-(aq)}$ ions—is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$\mathbf{HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(\ell) + C_2H_3O_2^-(aq)} \tag{11.10.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

> Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid—a source of $H^+$ ions—is added to the buffer solution, the $H^+$ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$\mathbf{H^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq)} \tag{11.10.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure **11.10.1** illustrates both actions of a buffer.

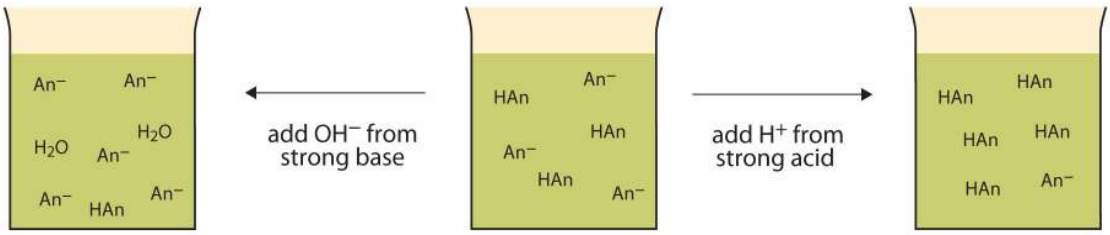



Figure **11.10.1**: The Action of Buffers. Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing $NH_3$ and $NH_4Cl$, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:

$$\mathbf{NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)} \tag{11.10.3}$$

while the **ammonium ion** ($\mathbf{NH_4^+(aq)}$) can react with any hydroxide ions introduced by strong bases:

$$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(\ell) \quad (11.10.4)$$

### Example 11.10.1: Making Buffer Solutions

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

a. $HCHO_2$ and $NaCHO_2$
b. HCl and NaCl
c. $CH_3NH_2$ and $CH_3NH_3Cl$
d. $NH_3$ and NaOH

**Solution**

a. Formic acid ($HCHO_2$) is a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid—the formate ion ($CHO_2^-$). The combination of these two solutes would make a buffer solution.
b. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
c. Methylamine ($CH_3NH_2$) is like ammonia with one of its hydrogen atoms substituted with a $CH_3$ (methyl) group. Because it is not on our list of strong bases, we can assume that it is a weak base. The compound $CH_3NH_3Cl$ is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
d. Ammonia ($NH_3$) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

### Exercise 11.10.1

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

a. $NaHCO_3$ and NaCl
b. $H_3PO_4$ and $NaH_2PO_4$
c. $NH_3$ and $(NH_4)_3PO_4$
d. NaOH and NaCl

**Answer a**

Yes.

**Answer b**

No. Need a weak acid or base and a salt of its conjugate base or acid.

**Answer c**

Yes.

**Answer d**

No. Need a weak base or acid.

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain **capacity**. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of $HCO_3^-$ and $H_2CO_3$ [$H_2CO_3$ is another way to write $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

## Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/− type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.

Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must conform to standards established by the American Association of Blood Banks.

## Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

## Contributions & Attributions

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# CHAPTER OVERVIEW

## 12: Gases


12.1: Introduction to Gases

12.2: Boyle’s Law - Pressure and Volume

12.3: Charles’s Law- Volume and Temperature

12.4: Gay-Lussac's Law- Temperature and Pressure

12.5: The Combined Gas Law- Pressure, Volume, and Temperature

12.6: Avogadro’s Law- Volume and Moles

12.7: The Ideal Gas Law- Pressure, Volume, Temperature, and Moles

12.8: Mixtures of Gases - Why Deep-Sea Divers Breathe a Mixture of Helium and Oxygen


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# 12.1: Introduction to Gases

**Learning Objectives**

- To describe the gas phase.

The gas phase is unique among the three states of matter in that there are some simple models we can use to predict the physical behavior of all gases—independent of their identities. We cannot do this for the solid and liquid states. In fact, the development of this understanding of the behavior of gases represents the historical dividing point between alchemy and modern chemistry. Initial advances in the understanding of gas behavior were made in the mid 1600s by Robert Boyle, an English scientist who founded the Royal Society (one of the world's oldest scientific organizations).

How is it that we can model all gases independent of their chemical identity? The answer is in a group of statements called the kinetic theory of gases:

- Gases are composed of tiny particles that are separated by large distances.
- Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container.
- The velocity of gas particles is related to the temperature of a gas.
- Gas particles do not experience any force of attraction or repulsion with each other.

Did you notice that none of these statements relates to the identity of the gas? This means that all gases should behave similarly. A gas that follows these statements perfectly is called an *ideal gas*. Most gases show slight deviations from these statements and are called *real gases*. However, the existence of real gases does not diminish the importance of the kinetic theory of gases.

One of the statements of the kinetic theory mentions collisions. As gas particles are constantly moving, they are also constantly colliding with each other and with the walls of their container. There are forces involved as gas particles bounce off the container walls (Figure **12.1.1**). The force generated by gas particles divided by the area of the container walls yields pressure. Pressure is a property we can measure for a gas, but we typically do not consider pressure for solids or liquids.

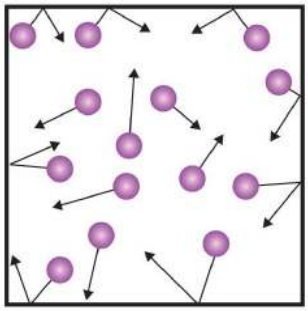


Figure **12.1.1**: Gas Pressure. Pressure is what results when gas particles rebound off the walls of their container.

The basic unit of pressure is the newton per square meter ($N/m^2$). This combined unit is redefined as a pascal (Pa). One pascal is not a very large amount of pressure. A more useful unit of pressure is the bar, which is 100,000 Pa (1 bar = 100,000 Pa). Other common units of pressure are the atmosphere (atm), which was originally defined as the average pressure of Earth's atmosphere at sea level; and mmHg (millimeters of mercury), which is the pressure generated by a column of mercury 1 mm high. The unit millimeters of mercury is also called a torr, named after the Italian scientist Evangelista Torricelli, who invented the barometer in the mid-1600s. A more precise definition of atmosphere, in terms of torr, is that there are exactly 760 torr in 1 atm. A bar equals 1.01325 atm. Given all the relationships between these pressure units, the ability to convert from one pressure unit to another is a useful skill.

**Example 12.1.1: Converting Pressures**

Write a conversion factor to determine how many atmospheres are in 1,547 mmHg.

**Solution**

Because 1 mmHg equals 1 torr, the given pressure is also equal to 1,547 torr. Because there are 760 torr in 1 atm, we can use this conversion factor to do the mathematical conversion:

$$1,547 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 2.04 \text{ atm}$$

Note how the torr units cancel algebraically.

**? Exercise 12.1.1: Converting Pressures**

Write a conversion factor to determine how many millimeters of mercury are in 9.65 atm.

**Answer**

$$9.65 \text{ atm} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 7,334 \text{ mmHg}.$$

The kinetic theory also states that there is no interaction between individual gas particles. Although we know that there are, in fact, intermolecular interactions in real gases, the kinetic theory assumes that gas particles are so far apart that the individual particles don't "feel" each other. Thus, we can treat gas particles as tiny bits of matter whose identity isn't important to certain physical properties.

## Key Takeaway

- The gas phase has certain general properties characteristic of that phase.

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# 12.2: Boyle's Law - Pressure and Volume

**Learning Objectives**

- Learn what is meant by the term *gas laws*.
- Learn and apply Boyle's Law.

When seventeenth-century scientists began studying the physical properties of gases, they noticed some simple relationships between some of the measurable properties of the gas. Take pressure ($P$) and volume ($V$), for example. Scientists noted that for a given amount of a gas (usually expressed in units of moles [$n$]), if the temperature ($T$) of the gas was kept constant, pressure and volume were related: as one increases, the other decreases. As one decreases, the other increases. This means that pressure and volume are *inversely related*.

There is more to it, however: pressure and volume of a given amount of gas at constant temperature are *numerically* related. If you take the pressure value and multiply it by the volume value, the product is a constant for a given amount of gas at a constant temperature:

$$P \times V = \text{ constant at constant n and T}$$

If either volume or pressure changes while amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labeled $P_1$ and $V_1$ and the new conditions are labeled $P_2$ and $V_2$, we have

$$P_1V_1 = \text{constant} = P_2V_2$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply

$$P_1V_1 = P_2V_2 \text{ at constant n and T}$$

This equation is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behavior of a gas. This particular gas law is called **Boyle's Law**, after the English scientist Robert Boyle, who first announced it in 1662. Figure **12.2.1** shows two representations of how Boyle's Law works.

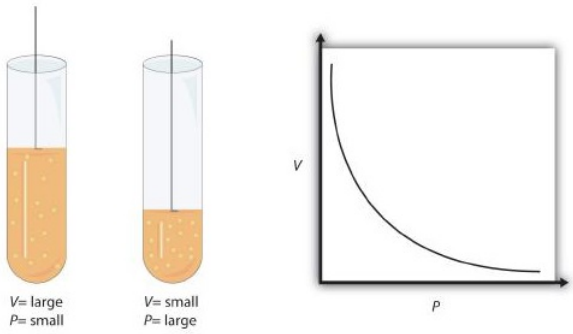



Figure **12.2.1**: Boyle's Law. A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot $P$ versus $V$ for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

Boyle's Law is an example of a second type of mathematical problem we see in chemistry—one based on a mathematical formula. Tactics for working with mathematical formulas are different from tactics for working with conversion factors. First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables. Second, in most formulas, some mathematical rearrangements (i.e., algebra) must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation. Finally, units must be consistent. For example, in Boyle's Law there are two pressure variables; they must have the same unit. There are also two volume variables; they also must have the same unit. In most cases, it won't matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

## Example 12.2.1

A sample of gas has an initial pressure of 2.44 atm and an initial volume of 4.01 L. Its pressure changes to 1.93 atm. What is the new volume if temperature and amount are kept constant?

### Solution

Solutions to Example 11.8.1

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: $P_1$ = 2.44 atm and $V_1$ = 4.01 L<br>$P_2$ = 1.93 atm<br>Find: $V_2$ = ? L |
| List other known quantities. | none |
| Plan the problem. | First, rearrange the equation algebraically to solve for $V_2$.<br>$$V_2 = \frac{P_1 \times V_1}{P_2}$$ |
| Cancel units and calculate. | Now substitute the known quantities into the equation and solve.<br>$$V_2 = \frac{2.44 \cancel{\text{ atm}} \times 4.01 \text{ L}}{1.93 \cancel{\text{ atm}}} = 5.07 \text{ L}$$ |
| Think about your result. | We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreasing (from 2.44 atm to 1.93 atm), so volume should be increasing to compensate, and it is (from 4.01 L to 5.07 L). So the answer makes sense based on Boyle's Law. |

## Exercise 12.2.1

If $P_1$ = 334 torr, $V_1$ = 37.8 mL, and $P_2$ = 102 torr, what is $V_2$?

**Answer**

124 mL

As mentioned, you can use any units for pressure and volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

## Example 12.2.2:

A sample of gas has an initial pressure of 722 torr and an initial volume of 88.8 mL. Its volume changes to 0.663 L. What is the new pressure?

### Solution

Solutions to Example 11.8.2

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given: $P_1$ = 722 torr and $V_1$ = 88.8 mL<br>$V_2$ = 0.633 L<br>Find: $P_2$ = ? torr |
| List other known quantities. | 1 L = 1000 mL to have the same units for volume. |
| Plan the problem. | 1. Perform the conversion of the second volume unit from L to mL.<br>2. Rearrange the equation algebraically to solve for $P_2$.<br>$$P_2 = \frac{P_1 \times V_1}{V_2}$$ |
| Cancel units and calculate. | 1.<br>$$0.663 \cancel{L} \times \frac{1000\,ml}{1\,\cancel{L}} = 663\,ml$$<br>2. Substitute the known quantities into the equation and solve.<br>$$P_2 = \frac{722\text{ torr} \times 88.8\ \cancel{mL}}{663\ \cancel{mL}} = 96.7\text{ torr}$$ |
| Think about your result. | When the volume increased, the pressure decreased, which is as expected for Boyle's Law. |

**? Exercise 12.2.2**

If $V_1$ = 456 mL, $P_1$ = 308 torr, and $P_2$ = 1.55 atm, what is $V_2$?

**Answer**

119 mL

## Summary

- The behavior of gases can be modeled with gas laws.
- Boyle's Law relates the pressure and volume of a gas at constant temperature and amount.

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# 12.3: Charles's Law- Volume and Temperature

**Learning Objectives**

- Learn and apply Charles's Law.

Everybody enjoys the smell and taste of freshly-baked bread. It is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which at high temperatures causes the dough to expand. The end result is an enjoyable treat, especially when covered with melted butter.

## Charles's Law

French physicist Jacques Charles (1746-1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stop of molecular motion.

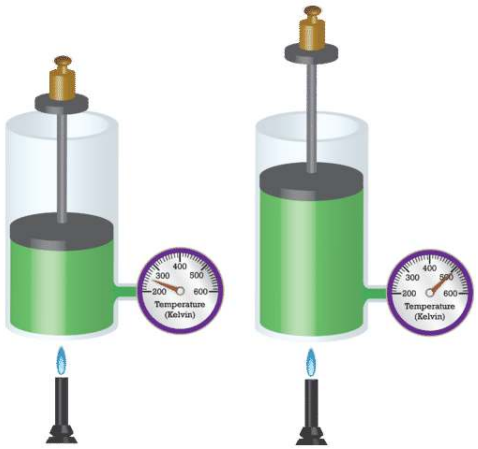



Figure **12.3.1**: As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume.

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k$$

As with Boyle's Law, $k$ is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table **12.3.1**: Temperature-Volume Data

| **Temperature (K)** | **Volume (mL)** | $\frac{V}{T} = k \left(\frac{\text{mL}}{\text{K}}\right)$ |
|---|---|---|
| 50 | 20 | 0.40 |
| 100 | 40 | 0.40 |
| 150 | 60 | 0.40 |
| 200 | 80 | 0.40 |
| 300 | 120 | 0.40 |
| 500 | 200 | 0.40 |
| 1000 | 400 | 0.40 |

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

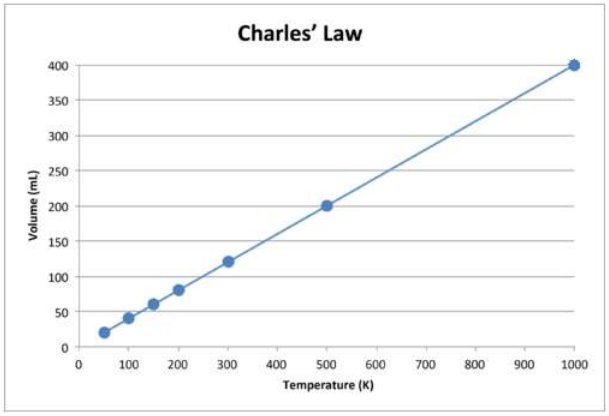



Figure 12.3.2:The volume of a gas increases as the Kelvin temperature increases.

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use $V_1$ and $T_1$ to stand for the initial volume and temperature of a gas, while $V_2$ and $T_2$ stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $\mathrm{K} = {}^\circ\mathrm{C} + 273$.

**Example 12.3.1:**

A balloon is filled to a volume of $2.20\ \mathrm{L}$ at a temperature of $22^\circ\mathrm{C}$. The balloon is then heated to a temperature of $71^\circ\mathrm{C}$. Find the new volume of the balloon.

**Solution**

Solutions to Example 11.5.1

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>$V_1 = 2.20\ \text{L}$ and<br>$T_1 = 22°\text{C} = 295\ \text{K}$<br>$T_2 = 71°\text{C} = 344\ \text{K}$<br>Find: $V_2$ = ? L |
| List other known quantities. | The temperatures have first been converted to Kelvin. |
| Plan the problem. | First, rearrange the equation algebraically to solve for $V_2$.<br>$$V_2 = \frac{V_1 \times T_2}{T_1}$$ |
| Cancel units and calculate. | Now substitute the known quantities into the equation and solve.<br>$$V_2 = \frac{2.20\ \text{L} \times 344\ \cancel{\text{K}}}{295\ \cancel{\text{K}}} = 2.57\ \text{L}$$ |
| Think about your result. | The volume increases as the temperature increases. The result has three significant figures. |

**? Exercise 12.3.1**

If $V_1$ = 3.77 L and $T_1$ = 255 K, what is $V_2$ if $T_2$ = 123 K?

**Answer**

1.82 L

**✔ Example 12.3.2:**

A sample of a gas has an initial volume of 34.8 L and an initial temperature of −67°C. What must be the temperature of the gas for its volume to be 25.0 L?

**Solution**

Solutions to Example 11.5.2

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>Given:$T_1$ = -27°C and $V_1$ = 34.8 L<br>$V_2$ = 25.0 L<br>Find: $T_2$ = ? K |
| List other known quantities. | K = -27°C + 273 |
| Plan the problem. | 1. Convert the initial temperature to Kelvin<br>2. Rearrange the equation algebraically to solve for $T_2$.<br>$$T_2 = \frac{V_2 \times T_1}{V_1}$$ |
| Cancel units and calculate. | 1. −67°C + 273 = 206 K<br>2. Substitute the known quantities into the equation and solve.<br>$$T_2 = \frac{25.0 \cancel{L} \times 206\ \text{K}}{34.8 \cancel{L}} = 148\ \text{K}$$ |
| Think about your result. | This is also equal to −125°C. As temperature decreases, volume decreases—which it does in this example. |

**? Exercise 12.3.2**

If $V_1$ = 623 mL, $T_1$ = 255°C, and $V_2$ = 277 mL, what is $T_2$?

**Answer**

235 K, or −38°C

## Summary

- Charles's Law relates the volume and temperature of a gas at constant pressure and amount.

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# 12.4: Gay-Lussac's Law- Temperature and Pressure

**Learning Objectives**

- Explain Gay-Lussac's Law.

Propane tanks are widely used with barbeque grills. But it's not fun to find out halfway through grilling that you have run out of gas. You can buy gauges that measure the pressure inside the tank to see how much is left. The gauge measures pressure and will register a higher pressure on a hot day than it will on a cold day. So you need to take the air temperature into account when you decide whether or not to refill the tank before your next cook-out.

## Gay-Lussac's Law

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. The increase in kinetic energy results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778-1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.



Figure 12.4.1: Joseph Gay-Lussac.

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} \text{ and } \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A graph of pressure vs. temperature also illustrates a direct relationship. As a gas is cooled at constant volume, its pressure continually decreases until the gas condenses to a liquid.

**Example 12.4.1**

The gas in an aerosol can is under a pressure of $3.00\ \text{atm}$ at a temperature of $25°\text{C}$. It is dangerous to dispose of an aerosol can by incineration. What would the pressure in the aerosol can be at a temperature of $845°\text{C}$?

**Solution**

Solutions to Example 11.10.1

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>$P_1 = 3.00\ \text{atm}$<br>$T_1 = 25°\text{C} = 298\ \text{K}$<br>$T_2 = 845°\text{C} = 1118\ \text{K}$<br>Find: $P_2 = ?\ \text{atm}$ |
| List other known quantities. | The temperatures have first been converted to Kelvin. |
| Plan the problem. | First, rearrange the equation algebraically to solve for $P_2$.<br>$$P_2 = \frac{P_1 \times T_2}{T_1}$$ |
| Calculate. | Now substitute the known quantities into the equation and solve.<br>$$P_2 = \frac{3.00\ \text{atm} \times 1118\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.3\ \text{atm}$$ |
| Think about your result. | The pressure increases dramatically due to a large increase in temperature. |

## Summary

- Pressure and temperature at constant volume are directly proportional.

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# 12.5: The Combined Gas Law- Pressure, Volume, and Temperature

**Learning Objectives**

- Learn and apply the Combined Gas Law.

One thing we notice about all the gas laws is that, collectively, volume and pressure are always in the numerator, and temperature is always in the denominator. This suggests that we can propose a gas law that combines pressure, volume, and temperature. This gas law is known as the **Combined Gas Law**, and its mathematical form is

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \; at \; constant \; n$$

This allows us to follow changes in all three major properties of a gas. Again, the usual warnings apply about how to solve for an unknown algebraically (isolate it on one side of the equation in the numerator), units (they must be the same for the two similar variables of each type), and units of temperature must be in Kelvin.

**Example 12.5.1:**

A sample of gas at an initial volume of 8.33 L, an initial pressure of 1.82 atm, and an initial temperature of 286 K simultaneously changes its temperature to 355 K and its volume to 5.72 L. What is the final pressure of the gas?

**Solution**

Solutions to Example 11.4.1

| **Steps for Problem Solving** | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>$V_1$ = 8.33 L, $P_1$ = 1.82 atm, and $T_1$ = 286 K<br>V2 = 5.72 L and $T_2$ = 355 K<br>Find: $P2$ = ? atm |
| List other known quantities. | none |
| Plan the problem. | First, rearrange the equation algebraically to solve for $V_2$.<br>$P_2 = \frac{P_1V_1T_2}{T_1V_2}$ |
| Calculate. | Now substitute the known quantities into the equation and solve.<br>$P_2 = \frac{(1.82\,atm)(8.33\,\cancel{L})(355\,\cancel{K})}{(286\,\cancel{K})(5.72\,\cancel{L})} = 3.22atm$ |
| Think about your result. | Ultimately, the pressure increased, which would have been difficult to predict because two properties of the gas were changing. |

**Exercise 12.5.1**

If $P_1$ = 662 torr, $V_1$ = 46.7 mL, $T_1$ = 266 K, $P_2$ = 409 torr, and $T_2$ = 371 K, what is $V_2$?

**Answer**

105 mL

As with other gas laws, if you need to determine the value of a variable in the denominator of the combined gas law, you can either cross-multiply all the terms or just take the reciprocal of the combined gas law. Remember, the variable you are solving for must be in the numerator and all by itself on one side of the equation.

## Summary

- The Combined Gas Law relates pressure, volume, and temperature of a gas.

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# 12.6: Avogadro’s Law- Volume and Moles

A flat tire is not very useful. It does not cushion the rim of the wheel and creates a very uncomfortable ride. When air is added to the tire, the pressure increases as more molecules of gas are forced into the rigid tire. How much air should be put into a tire depends on the pressure rating for that tire. Too little pressure and the tire will not hold its shape. Too much pressure and the tire could burst.

## Avogadro's Law

You have learned about Avogadro's hypothesis: equal volumes of any gas at the same temperature and pressure contain the same number of molecules. It follows that the volume of a gas is directly proportional to the number of moles of gas present in the sample. **Avogadro's Law** states that the volume of a gas is directly proportional to the number of moles (or number of particles) of gas when the temperature and pressure are held constant. The mathematical expression of Avogadro's Law is:

$$V = k \times n$$

or

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

where $n$ is the number of moles of gas and $k$ is a constant. Avogadro's Law is in evidence whenever you blow up a balloon. The volume of the balloon increases as you add moles of gas to the balloon by blowing it up.

If the container holding the gas is rigid rather than flexible, pressure can be substituted for volume in Avogadro's Law. Adding gas to a rigid container makes the pressure increase.

### Example 12.6.1

A balloon has been filled to a volume of $1.90\ \text{L}$ with $0.0920\ \text{mol}$ of helium gas. If $0.0210\ \text{mol}$ of additional helium is added to the balloon while the temperature and pressure are held constant, what is the new volume of the balloon?

**Solution**

Solutions to Example 11.11.1

| **Steps for Problem Solving** | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>$V_1 = 1.90\ \text{L}$<br>$n_1 = 0.0920\ \text{mol}$<br>Find: $V_2 = ?\ \text{L}$ |
| List other known quantities. | Note that the final number of moles has to be calculated by adding the original number of moles to the moles of added helium.<br>$n_2 = 0.0920 + 0.0210 = 0.1130\ \text{mol}$ |
| Plan the problem. | First, rearrange the equation algebraically to solve for $V_2$.<br>$V_2 = \frac{V_1 \times n_2}{n_1}$ |
| Calculate. | Now substitute the known quantities into the equation and solve.<br>$V_2 = \frac{1.90\ \text{L} \times 0.1130\ \cancel{\text{mol}}}{0.0920\ \cancel{\text{mol}}} = 2.33\ \text{L}$ |
| Think about your result. | Since a relatively small amount of additional helium was added to the balloon, its volume increases slightly. |

## Exercise 12.6.1

A 12.8 L volume of gas contains .000498 moles of oxygen gas. At constant temperature and pressure, what volume does .0000136 moles of the gas fill?

**Answer**

0.350 L

## Summary

- Calculations for relationships between volume and number of moles of a gas can be performed using Avogadro's Law.

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# 12.7: The Ideal Gas Law- Pressure, Volume, Temperature, and Moles

**Learning Objectives**

- Explain the Ideal Gas Law.

There are a number of chemical reactions that require ammonia. In order to carry out the reaction efficiently, we need to know how much ammonia we have for stoichiometric purposes. Using gas laws, we can determine the number of moles present in the tank if we know the volume, temperature, and pressure of the system.

## Ideal Gas Law

The Combined Gas Law shows that the pressure of a gas is inversely proportional to volume and directly proportional to temperature. Avogadro's Law shows that volume or pressure is directly proportional to the number of moles of gas. Putting these together leaves us with the following equation:

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2}$$

As with the other gas laws, we can also say that $\frac{(P \times V)}{(T \times n)}$ is equal to a constant. The constant can be evaluated provided that the gas being described is considered to be ideal.

The **Ideal Gas Law** is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas. If we substitute in the variable $R$ for the constant, the equation becomes:

$$\frac{P \times V}{T \times n} = R$$

The Ideal Gas Law is conveniently rearranged to look this way, with the multiplication signs omitted:

$$PV = nRT$$

The variable $R$ in the equation is called the **ideal gas constant**.

## Evaluating the Ideal Gas Constant

The value of $R$, the ideal gas constant, depends on the units chosen for pressure, temperature, and volume in the ideal gas equation. It is necessary to use Kelvin for the temperature and it is conventional to use the SI unit of liters for the volume. However, pressure is commonly measured in one of three units: $\mathbf{kPa}$, $\mathbf{atm}$, or $\mathbf{mm\ Hg}$. Therefore, $R$ can have three different values.

We will demonstrate how $R$ is calculated when the pressure is measured in $\mathbf{kPa}$. The volume of $\mathbf{1.00\ mol}$ of any gas at STP (Standard temperature, 273.15 K and pressure, 1 atm) is measured to be $\mathbf{22.414\ L}$. We can substitute $\mathbf{101.325\ kPa}$ for pressure, $\mathbf{22.414\ L}$ for volume, and $\mathbf{273.15\ K}$ for temperature into the ideal gas equation and solve for $R$.

$$\begin{aligned} R &= \frac{PV}{nT} \\ &= \frac{101.325\ \text{kPa} \times 22.414\ \text{L}}{1.000\ \text{mol} \times 273.15\ \text{K}} \\ &= 8.314\ \text{kPa} \cdot \text{L/K} \cdot \text{mol} \end{aligned}$$

This is the value of $R$ that is to be used in the ideal gas equation when the pressure is given in $\mathbf{kPa}$. The table below shows a summary of this and the other possible values of $R$. It is important to choose the correct value of $R$ to use for a given problem.

Table 12.7.1: Values of the Ideal Gas Constant

| Unit of $P$ | Unit of $V$ | Unit of $n$ | Unit of $T$ | Value and Unit of $R$ |
|---|---|---|---|---|
| kPa | L | mol | K | $8.314\ \text{J/K} \cdot \text{mol}$ |
| atm | L | mol | K | $0.08206\ \text{L} \cdot \text{atm/K} \cdot \text{mol}$ |
| mm Hg | L | mol | K | $62.36\ \text{L} \cdot \text{mm Hg/K} \cdot \text{mol}$ |

Notice that the unit for $R$ when the pressure is in kPa has been changed to $\text{J/K} \cdot \text{mol}$. A kilopascal multiplied by a liter is equal to the SI unit for energy, a joule (J).

## Example 12.7.1 Oxygen Gas

What volume is occupied by 3.76 g of oxygen gas at a pressure of 88.4 kPa and a temperature of 19°C? Assume the oxygen is ideal.

### Solution

Solutions to Example 11.5.1

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>• $P = 88.4\ \text{kPa}$<br>• $T = 19°\text{C} = 292\ \text{K}$<br>Mass $\text{O}_2 = 3.76\ \text{g}$<br>Find: $V$ = ? L |
| List other known quantities. | $\text{O}_2 = 32.00\ \text{g/mol}$<br>$R = 8.314\ \text{J/K} \cdot \text{mol}$ |
| Plan the problem. | 1. First, determine the number of moles of $\text{O}_2$ from the given mass and the molar mass.<br>2. Then, rearrange the equation algebraically to solve for V<br>$$V = \frac{nRT}{P}$$ |
| Calculate. | 1.<br>$$3.76\ \cancel{g} \times \frac{1\ \text{mol O}_2}{32.00\ \cancel{g}\ \text{O}_2} = 0.1175\ \text{mol O}_2$$<br>2. Now substitute the known quantities into the equation and solve.<br>$$V = \frac{nRT}{P}$$<br>$$= \frac{0.1175\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{J/K}} \cdot \cancel{\text{mol}} \times 292\ \cancel{\text{K}}}{88.4\ \cancel{\text{kPa}}}$$<br>$$= 3.23\ \text{L O}_2$$ |
| Think about your result. | The number of moles of oxygen is far less than one mole, so the volume should be fairly small compared to molar volume $(22.4\ \text{L/mol})$ since the pressure and temperature are reasonably close to standard. The result has three significant figures because of the values for $T$ and $P$. Since a joule $(\text{J}) = \text{kPa} \cdot \text{L}$, the units cancel out correctly, leaving a volume in liters. |

## Example 12.7.2: Argon Gas

A 4.22 mol sample of Ar has a pressure of 1.21 atm and a temperature of 34°C. What is its volume?

### Solution

Solutions to Example 11.5.2

| Steps for Problem Solving | |
|---|---|
| Identify the "given" information and what the problem is asking you to "find." | Given:<br>n = 4.22 mol<br>P = 1.21 atm<br>T = 34°C<br>Find: $V$ = ? L |
| List other known quantities. | none |
| Plan the problem. | 1. The first step is to convert temperature to Kelvin.<br>2. Then, rearrange the equation algebraically to solve for V<br>$$V = \frac{nRT}{P}$$ |
| Calculate. | 1. 34 + 273 = 307 K<br>2. Now substitute the known quantities into the equation and solve.<br>$$V = \frac{(4.22\ \cancel{mol})(0.08205 \frac{L.\ \cancel{atm}}{\cancel{mol.K}})(307\ \cancel{K})}{1.21\ \cancel{atm}}$$<br>$$= 87.9\ L$$ |
| Think about your result. | The number of moles of Ar is large so the expected volume should also be large. |

## Exercise 12.7.1

A 0.0997 mol sample of $O_2$ has a pressure of 0.692 atm and a temperature of 333 K. What is its volume?

**Answer**

3.94 L

## Exercise 12.7.2

For a 0.00554 mol sample of $H_2$, $P$ = 23.44 torr and $T$ = 557 K. What is its volume?

**Answer**

8.21 L

## Summary

- The Ideal Gas Law is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas.

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# 12.8: Mixtures of Gases - Why Deep-Sea Divers Breathe a Mixture of Helium and Oxygen

**Learning Objectives**

- Explain Dalton's Law of Partial Pressures.

The atmosphere of Venus is markedly different from that of Earth. The gases in the Venusian atmosphere are $96.5\%$ carbon dioxide and $3\%$ nitrogen. The atmospheric pressure on Venus is roughly 92 times that of Earth, so the amount of nitrogen on Venus would contribute a pressure well over $2700\ \text{mm Hg}$. And there is no oxygen present, so we couldn't breathe there. Not that we would want to go to Venus, as the surface temperature is usually over $460^\circ\text{C}$.

## Dalton's Law of Partial Pressures

Gas pressure results from collisions between gas particles and the inside walls of their container. If more gas is added to a rigid container, the gas pressure increases. The identities of the two gases do not matter. John Dalton, the English chemist who proposed the atomic theory, also studied mixtures of gases. He found that each gas in a mixture exerts a pressure independently of every other gas in the mixture. For example, our atmosphere is composed of about $78\%$ nitrogen and $21\%$ oxygen, with smaller amounts of several other gases making up the rest. Since nitrogen makes up $78\%$ of the gas particles in a given sample of air, it exerts $78\%$ of the pressure. If the overall atmospheric pressure is $1.00\ \text{atm}$, then the pressure of just the nitrogen in the air is $0.78\ \text{atm}$. The pressure of the oxygen in the air is $0.21\ \text{atm}$.

The **partial pressure** of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture. The partial pressure of nitrogen is represented by $P_{N_2}$. **Dalton's Law of Partial Pressures** states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases. Dalton's Law can be expressed with the following equation:

$$P_{\text{total}} = P_1 + P_2 + P_3 + \cdots$$

The figure below shows two gases that are in separate, equal-sized containers at the same temperature and pressure. Each exerts a different pressure, $P_1$ and $P_2$, reflective of the number of particles in the container. On the right, the two gases are combined into the same container, with no volume change. The total pressure of the gas mixture is equal to the sum of the individual pressures. If $P_1 = 300\ \text{mm Hg}$ and $P_2 = 500\ \text{mm Hg}$, then $P_{\text{total}} = 800\ \text{mm Hg}$.

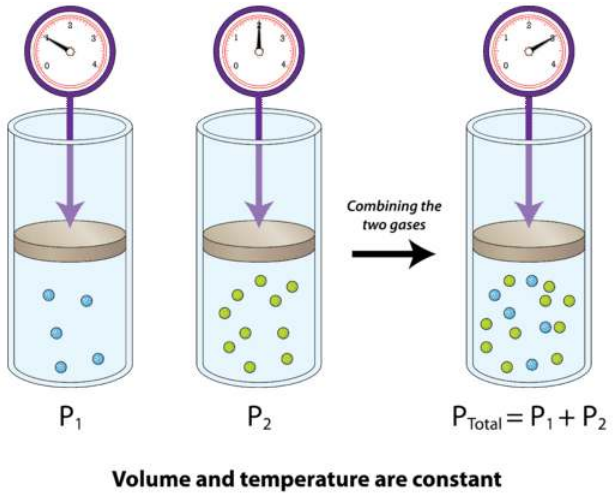



Figure 12.8.1: Dalton's Law states that the pressure of a gas mixture is equal to the partial pressures of the combining gases.

## Collecting Gases Over Water

You need to do a lab experiment where hydrogen gas is generated. In order to calculate the yield of gas, you have to know the pressure inside the tube where the gas is collected. But how can you get a barometer in there? Very simple: you don't. All you need

is the atmospheric pressure in the room. As the gas pushes out the water, it is pushing against the atmosphere, so the pressure inside is equal to the pressure outside.

## Gas Collection by Water Displacement

Gases that are produced in laboratory experiments are often collected by a technique called **water displacement** (Figure 12.8.2). A bottle is filled with water and placed upside-down in a pan of water. The reaction flask is fitted with rubber tubing, which is then fed under the bottle of water. As the gas is produced in the reaction flask, it exits through the rubber tubing and displaces the water in the bottle. When the bottle is full of the gas, it can be sealed with a lid.



Figure 12.8.2: A gas produced in a chemical reaction can be collected by water displacement.

Because the gas is collected over water, it is not pure, but is mixed with vapor from the evaporation of the water. Dalton's Law can be used to calculate the amount of the desired gas by subtracting the contribution of the water vapor.

$$P_{\text{Total}} = P_g + P_{H_2O}$$

where $P_g$ is the pressure of the desired gas, which can be solved for:

$$P_g = P_{Total} - P_{H_2O}$$

In order to solve a problem, it is necessary to know the vapor pressure of water at the temperature of the reaction (see table below). The sample problem illustrates the use of Dalton's Law when a gas is collected over water.

Table 12.8.1: Vapor Pressure of Water (mm Hg) at Selected Temperatures (°C)

| 0 | 5 | 10 | 15 | 20 | 25 | 30 | 35 | 40 | 45 | 50 | 55 | 60 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 4.58 | 6.54 | 9.21 | 12.79 | 17.54 | 23.76 | 31.82 | 42.18 | 55.32 | 71.88 | 92.51 | 118.04 | 149.38 |

### Example 14.14.1

A certain experiment generates 2.58 L of hydrogen gas, which is collected over water. The temperature is 20°C and the atmospheric pressure is 98.60 kPa. Find the volume that the dry hydrogen would occupy at STP.

**Solution**

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

**Known**

- $V_{\text{Total}} = 2.58\ \text{L}$
- $T = 20°\text{C} = 293\ \text{K}$
- $P_{\text{Total}} = 98.60\ \text{kPa} = 739.7\ \text{mm Hg}$

**Unknown**

- $V_{H_2}$ at STP $=?\ \text{L}$

The atmospheric pressure is converted from kPa to mm Hg in order to match units with the table. The sum of the pressures of the hydrogen and the water vapor is equal to the atmospheric pressure. The pressure of the hydrogen is found by subtraction. Then, the volume of the gas at STP can be calculated by using the combined gas law.

***Step 2: Solve.***

$$\begin{aligned} P_{H_2} &= P_{\text{Total}} - P_{H_2O} \\ &= 739,7 \text{ mm Hg} - 17.54 \text{ mm Hg} \\ &= 722.2 \text{ mm Hg} \end{aligned}$$

Now the combined gas law is used, solving for $V_2$, the volume of hydrogen at STP.

$$\begin{aligned} V_2 &= \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1} \\ &= \frac{722.2 \text{ mm Hg} \times 2.58 \text{ L} \times 273 \text{ K}}{760 \text{ mm Hg} \times 293 \text{ K}} \\ &= 2.28 \text{ L } H_2 \end{aligned}$$

***Step 3: Think about your result.***

If the hydrogen gas were to be collected at STP and without the presence of the water vapor, its volume would be **2.28 L**. This is less than the actual collected volume because some of that is water vapor. The conversion using STP is useful for stoichiometry purposes.

## Summary

- Dalton's Law of Partial Pressures states that the total pressure in a system is equal to the sum of the partial pressures of the gases present.
- The vapor pressure due to water in a sample can be corrected for, in order to get the true value for the pressure of the gas.

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