

GENERAL, ORGANIC, AND BIOCHEMISTRY WITH PROBLEMS, CASE STUDIES, AND ACTIVITIES



General, Organic, and Biochemistry with Problems, Case Studies, and Activities

This text is disseminated via the Open Education Resource (OER) LibreTexts Project (<https://LibreTexts.org>) and like the thousands of other texts available within this powerful platform, it is freely available for reading, printing, and "consuming."

The LibreTexts mission is to bring together students, faculty, and scholars in a collaborative effort to provide an accessible, and comprehensive platform that empowers our community to develop, curate, adapt, and adopt openly licensed resources and technologies; through these efforts we can reduce the financial burden born from traditional educational resource costs, ensuring education is more accessible for students and communities worldwide.

Most, but not all, pages in the library have licenses that may allow individuals to make changes, save, and print this book. Carefully consult the applicable license(s) before pursuing such effects. Instructors can adopt existing LibreTexts texts or Remix them to quickly build course-specific resources to meet the needs of their students. Unlike traditional textbooks, LibreTexts' web based origins allow powerful integration of advanced features and new technologies to support learning.



LibreTexts is the adaptable, user-friendly non-profit open education resource platform that educators trust for creating, customizing, and sharing accessible, interactive textbooks, adaptive homework, and ancillary materials. We collaborate with individuals and organizations to champion open education initiatives, support institutional publishing programs, drive curriculum development projects, and more.

The LibreTexts libraries are Powered by [NICE CXone Expert](#) and was supported by the Department of Education Open Textbook Pilot Project, the California Education Learning Lab, the UC Davis Office of the Provost, the UC Davis Library, the California State University Affordable Learning Solutions Program, and Merlot. This material is based upon work supported by the National Science Foundation under Grant No. 1246120, 1525057, and 1413739.

Any opinions, findings, and conclusions or recommendations expressed in this material are those of the author(s) and do not necessarily reflect the views of the National Science Foundation nor the US Department of Education.

Have questions or comments? For information about adoptions or adaptations contact info@LibreTexts.org or visit our main website at <https://LibreTexts.org>.

This text was compiled on 01/12/2026

TABLE OF CONTENTS

Licensing

1: Scientific Method, Matter, and Measurements

- 1.1: The Scientific Method
- 1.2: Units
- 1.3: Uncertainty in Measurement
- 1.4: Matter

2: Atoms and Elements

- 2.1: Electrons and electronic structure

3: Ionic and Covalent Compounds

- 3.1: Ionic and Covalent Compounds
- 3.2: Nomenclature

4: Chemical Bonding

- 4.1: Lewis structures
- 4.2: Covalent bonding and electronegativity
- 4.3: Valence Bond Theory
- 4.4: Molecular Structure and Polarity
- 4.5: Atomic Structure

5: Unit Conversions

- 5.1: Unit conversions
- 5.2: Density

6: Moles

- 6.1: Moles
- 6.2: Formula mass
- 6.3: Converting between grams and moles

7: Chemical Equations and Reactions

- 7.1: Balancing Chemical Equations
- 7.2: Energy in chemical reactions
- 7.3: Equilibrium
- 7.4: Catalysis

8: Gases, Liquids, and Solids

- 8.1: Kinetic Molecular Theory
- 8.2: Pressure
- 8.3: The Gas Laws
- 8.4: Intermolecular forces

- 8.5: Types of solutions
- 8.6: Electrolytes
- 8.7: Concentration
- 8.8: Osmosis, diffusion, and dialysis

9: Acids and Bases

- 9.1: What are acids and bases?
- 9.2: pH and pOH

10: Organic Chemistry

- 10.1: Drawing organic structures
- 10.2: Alkanes and isomers
- 10.3: Functional groups
- 10.4: Organic Reactions
- 10.5: Carbonyl substitution reactions

11: Carbohydrates

- 11.1: Carbohydrates
- 11.2: Fischer and Haworth projections

12: Lipids

- 12.1: Lipids

13: Proteins

- 13.1: Properties of proteins
- 13.2: Amino acids
- 13.3: Primary structure- peptides
- 13.4: Secondary, tertiary, and quaternary structure of proteins

14: Molecular Genetics

- 14.1: DNA structure

[Index](#)

[Glossary](#)

[Detailed Licensing](#)

Licensing

A detailed breakdown of this resource's licensing can be found in [Back Matter/Detailed Licensing](#).

CHAPTER OVERVIEW

1: Scientific Method, Matter, and Measurements

1.1: The Scientific Method

1.2: Units

1.3: Uncertainty in Measurement

1.4: Matter

1: Scientific Method, Matter, and Measurements is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

1.1: The Scientific Method

1

Harper College Chemistry Department

The Scientific Method

Chemistry is a science based on observation and experimentation. Doing chemistry involves attempting to answer questions and explain observations in terms of the laws and theories of chemistry, using procedures that are accepted by the scientific community. There is no single route to answering a question or explaining an observation, but there is an aspect common to every approach: Each uses knowledge based on experiments that can be reproduced to verify the results. Some routes involve a hypothesis, a tentative explanation of observations that acts as a guide for gathering and checking information. We test a hypothesis by experimentation, calculation, and/or comparison with the experiments of others and then refine it as needed.

Some hypotheses are attempts to explain the behavior that is summarized in laws. The laws of science summarize a vast number of experimental observations and describe or predict some facet of the natural world. If such a hypothesis turns out to be capable of explaining a large body of experimental data, it can reach the status of a theory. Scientific theories are well-substantiated, comprehensive, testable explanations of particular aspects of nature. Theories are accepted because they provide satisfactory explanations, but they can be modified if new data become available. The path of discovery that leads from question and observation to hypothesis or law to theory, combined with experimental verification of the hypothesis and any necessary modification of the theory, is called the scientific method ([Figure 1](#)).

 In this flowchart, the observation and curiosity box has an arrow pointing to a box labeled form hypothesis; make prediction. A curved arrow labeled next connects this box to a box labeled perform experiment; make more observations. Another arrow points back to the box that says form hypothesis; make prediction. This arrow is labeled results not consistent with prediction. Another arrow, labeled results are consistent with prediction points from the perform experiment box to a box labeled contributes to body of knowledge. However, an arrow also points from contributes to body of knowledge back to the form hypothesis; make prediction box. This arrow is labeled further testing does not support hypothesis. There are also two other arrows leading out from contributes to body of knowledge. One arrow is labeled much additional testing yields constant observations. This leads to the observation becomes law box. The other arrow is labeled much additional testing supports hypothesis. This arrow leads to the hypothesis becomes theory box.

The scientific method follows a process similar to the one shown in this diagram. All the key components are shown, in roughly the right order. Scientific progress is seldom neat and clean: It requires open inquiry and the reworking of questions and ideas in response to findings.

1.1: The Scientific Method is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

1.2: Units

2

Harper College Chemistry Department

Measurements describe the behavior of matter and energy in both the macroscopic and microscopic domains of chemistry. Every measurement provides three kinds of information: the size or magnitude of the measurement (a number), a standard of comparison for the measurement (a unit) and an indication of the uncertainty of the measurement. While the number and unit are explicitly represented when a quantity is written, the uncertainty is an aspect of the measurement result that is more implicitly represented and will be discussed later.

The number in the measurement can be represented in different ways, including decimal form and scientific notation (exponential notation, a review can be found [here](#)). For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms can also be written as 2.98×10^5 kg. The mass of the average mosquito is about 0.0000025 kilograms which can be written as 2.5×10^{-6} kg.

Units, such as liters, pounds and centimeters, are standards of comparison for measurements. When we buy a 2-liter bottle of a soft drink, we expect that the volume of the drink was measured. Thus, the measurement represents two times larger than the volume that everyone agrees to be 1 liter. The meat used to prepare a 0.25-pound hamburger is measured so it weighs one-fourth as much as 1 pound. Without units, a number can be meaningless, confusing or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient's seizures and states a dosage of "100" without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

We usually report the results of scientific measurements in SI units, an updated version of the metric system, using the units listed in Table 1 below. Other units can be derived from these base units. The standards for these units are fixed by international agreement and they are called the International System of Units or SI Units (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964.

Base Units of the SI System		
Property Measured	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K
electric current	ampere	A
amount of substance	mole	mol
luminous intensity	candela	cd

Sometimes we use units that are fractions or multiples of a base unit. Ice cream is sold in quarts (a familiar, non-SI base unit), pints (0.5 quart) or gallons (4 quarts). We also use fractions or multiples of units in the SI system, but these fractions or multiples are always powers of 10. Fractional or multiple SI units are named using a prefix and the name of the base unit. For example, a length of 1000 meters is also called a kilometer because the prefix *kilo* means "one thousand," which in scientific notation is 10^3 (1 kilometer = 1000 m = 10^3 m). The prefixes used and the powers to which 10 are raised are listed in Table 2.

Common Unit Prefixes			
Prefix	Symbol	Factor	Example
femto	f	10^{-15}	1 femtosecond (fs) = 1×10^{-15} s (0.000000000000001 s)
pico	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m (0.000000000001 m)
nano	n	10^{-9}	4 nanograms (ng) = 4×10^{-9} g (0.000000004 g)
micro	μ	10^{-6}	1 microliter (μ L) = 1×10^{-6} L (0.000001 L)
milli	m	10^{-3}	2 millimoles (mmol) = 2×10^{-3} mol (0.002 mol)
centi	c	10^{-2}	7 centimeters (cm) = 7×10^{-2} m (0.07 m)
deci	d	10^{-1}	1 deciliter (dL) = 1×10^{-1} L (0.1 L)
kilo	k	10^3	1 kilometer (km) = 1×10^3 m (1000 m)
mega	M	10^6	3 megahertz (MHz) = 3×10^6 Hz (3,000,000 Hz)
giga	G	10^9	8 gigayears (Gyr) = 8×10^9 yr (8,000,000,000 yr)
tera	T	10^{12}	5 terawatts (TW) = 5×10^{12} W (5,000,000,000,000 W)

Need a refresher or more practice with scientific notation? Visit this [site](#) to go over the basics of scientific notation.

SI Base Units

The initial units of the metric system, which eventually evolved into the SI system, were established in France during the French Revolution. The original standards for the meter and the kilogram were adopted there in 1799 and eventually by other countries. This section introduces four of the SI base units commonly used in chemistry. Other SI units will be introduced in subsequent chapters.

Length

The standard unit of length in both the SI and original metric systems is the meter (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. A meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

 One meter is slightly larger than a yard and one centimeter is less than half the size of one inch. 1 inch is equal to 2.54 cm. 1 m is equal to 1.094 yards which is equal to 39.36 inches.

The relative lengths of 1 m, 1 yd, 1 cm and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

Mass

Mass describes the amount of a substance. The standard unit of mass in the SI system is the kilogram (kg). A kilogram was originally defined as the mass of a liter of water (a cube of water with an edge length of exactly 0.1 meter). It is now defined by a

certain cylinder of platinum-iridium alloy, which is kept in France (Figure 2). Any object with the same mass as this cylinder is said to have a mass of 1 kilogram. One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (10^{-3} kg).

 The photo shows a small metal cylinder on a stand. The cylinder is covered with 2 glass lids, with the smaller glass lid encased within the larger glass lid.

This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland (credit: National Institutes of Standards and Technology).

Temperature

Temperature is an intensive property, a local physical property of a system that does not depend on the system size or the amount of material in the system. Temperature describes the amount of kinetic energy, motion, a system has. The SI unit of temperature is the kelvin (K). The IUPAC convention is to use kelvin (all lowercase) for the word, K (uppercase) for the unit symbol and neither the word “degree” nor the degree symbol ($^{\circ}$). The degree Celsius ($^{\circ}\text{C}$) is also allowed in the SI system with both the word “degree” and the degree symbol used for Celsius measurements. Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places. Water freezes at 273.15 K (0°C) and boils at 373.15 K (100°C) by definition and normal human body temperature is approximately 310 K (37°C). The conversion between these two units and the Fahrenheit scale will be discussed later.

Time

Time is described as how long it takes to accomplish a task. The SI base unit of time is the second (s). Small and large time intervals can be expressed with the appropriate prefixes; for example, 3 microseconds = $0.000003\text{ s} = 3 \times 10^{-6}$ and 5 megaseconds = $5,000,000\text{ s} = 5 \times 10^6\text{ s}$. Alternatively, hours, days and years can be used.

Derived SI Units

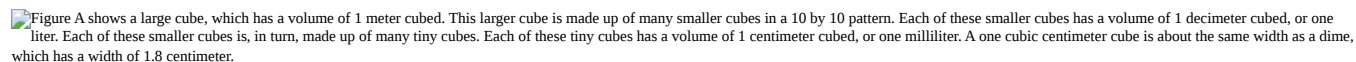
We can derive many units from the seven SI base units. For example, we can use the base unit of length to define a unit of volume, and the base units of mass and length to define a unit of density. We will talk about density later.

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (Figure 3). The standard volume is a cubic meter (m^3), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

A more commonly used unit of volume is derived from the decimeter (0.1 m or 10 cm). A cube with edge lengths of exactly one decimeter contains a volume of one cubic decimeter (dm^3). A liter (L) is the more common name for the cubic decimeter. One liter is about 1.06 quarts.

A cubic centimeter (cm^3) is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for cubic centimeter) is often used by health professionals. A cubic centimeter is also called a milliliter (mL) and is 1/1000 of a liter.

 Figure A shows a large cube, which has a volume of 1 meter cubed. This larger cube is made up of many smaller cubes in a 10 by 10 pattern. Each of these smaller cubes has a volume of 1 decimeter cubed, or one liter. Each of these smaller cubes is, in turn, made up of many tiny cubes. Each of these tiny cubes has a volume of 1 centimeter cubed, or one milliliter. A one cubic centimeter cube is about the same width as a dime, which has a width of 1.8 centimeter.

(a) The relative volumes are shown for cubes of 1 m^3 , 1 dm^3 (1 L) and 1 cm^3 (1 mL), not to scale. (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

Key Concepts and Summary

Measurements provide quantitative information that is critical in studying and practicing chemistry. Each measurement has an amount, a unit for comparison and an uncertainty. Measurements can be represented in either decimal or scientific notation. Scientists primarily use the SI (International System) or metric systems. We use base units such as meters, seconds, grams and liters. In many cases, we find it convenient to use unit prefixes that yield fractional and multiple units, such as microseconds (10^{-6} seconds) and megahertz (10^6 hertz), respectively.

[1]

1. OpenStax Chemistry: Atom First, Chemistry: Atoms First. OpenStax CNX. Nov 13, 2018. Download for free at <http://cnx.org/contents/4539ae23-1ccc-421e-9b25-843acbb6c4b0@9.17>. ↵
-

1.2: Units is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

1.3: Uncertainty in Measurement

3

Harper College Chemistry Department

By the end of this section, you will be able to:

- Define accuracy and precision
- Distinguish exact and uncertain numbers
- Correctly represent uncertainty in quantities using significant figures
- Apply proper rounding rules to computed quantities

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to know both the precision and accuracy of their results. Measurements are said to be precise if they yield very similar results when repeated in the same manner. A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value. These characterizations can be extended to other contexts, such as the results of an archery competition (Figure 1).

 Figures A through C each show targets with holes where the arrows hit. The archer in figure A was both accurate and precise as all 3 arrows are clustered in the center of the target. In figure B, the archer is precise but not accurate, as all 3 arrows are clustered together but to the upper right of the center of the target. In Figure C, the archer is neither accurate nor precise as the 3 holes are not close together and are located both to the upper right and right of the target.

Above are the results of an archery competition. (a) The results that are both accurate and precise, have arrows that are close to both the bullseye and one another. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Suppose a quality control chemist at a pharmaceutical company is tasked with checking the accuracy and precision of three different machines that are meant to dispense

10 ounces (296 mL) of cough syrup into storage bottles. She proceeds to use each machine to fill five bottles and then carefully determines the actual volume dispensed, obtaining the results in Table 1.

Volume (mL) of Cough Medicine Delivered by 10-oz (296 mL) Dispensers		
Dispenser #1	Dispenser #2	Dispenser #3
283.3	298.3	296.1
284.1	294.2	295.9
283.9	296.0	296.1
284.0	297.8	296.0
284.1	293.9	296.1

Considering these results, she will report that dispenser #1 is precise (values all close to one another, within a few tenths of a milliliter) but not accurate (none of the values are close to the target value of 296 mL, each being more than 10 mL too low). Results for dispenser #2 represent improved accuracy (each volume is less than 3 mL away from 296 mL) but worse precision (volumes vary by more than 4 mL). Finally, she can report that dispenser #3 is working well, dispensing cough syrup both accurately (all volumes within 0.1 mL of the target volume) and precisely (volumes differing from each other by no more than 0.2 mL).

Counting is the only type of measurement that is free from uncertainty, provided the number of objects being counted does not change while the counting process is underway. The result of such a counting measurement is an example of an exact number. If we count eggs in a carton, we know *exactly* how many eggs the carton contains. The numbers of defined quantities are also exact. By definition, 1 foot is exactly 12 inches, 1 inch is exactly 2.54 centimeters, and 1 gram is exactly 0.001 kilogram. Quantities derived

from measurements other than counting, however, are uncertain to varying extents due to practical limitations of the measurement process used.

Significant Figures in Measurement

The numbers of measured quantities, unlike defined or directly counted quantities, are not exact. To measure the volume of liquid in a graduated cylinder, you should make a reading at the bottom of the meniscus, the lowest point on the curved surface of the liquid.

 This diagram shows a 25 milliliter graduated cylinder filled with about 20.8 milliliters of fluid. The diagram zooms in on the meniscus, which is the curved surface of the water that is visible when the graduated cylinder is viewed from the side. You make the reading at the lowest point of the curve of the meniscus.

To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter and then make a reading (estimate) at the bottom of the meniscus.

Refer to the illustration in Figure 2. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark and so a reasonable estimate of the liquid's volume would be 21.6 mL. Thus, in the number 21.6, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain. In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

This concept holds true for all measurements, even if you do not actively make an estimation. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter weighs *about* 6.72 grams with a nominal uncertainty in the measurement of ± 0.01 gram. If we weigh the quarter on a more sensitive balance, we may find that its mass is 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram. Every measurement has some uncertainty, which depends on the device used (and the user's ability). All of the digits in a measurement, including the uncertain last digit, are called significant figures or significant digits. Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows "120," then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

Whenever you make a measurement properly, all the digits in the result are significant. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Well, for starters, all nonzero digits are significant, and it is only zeros that require some thought. We will use the terms "leading," "trailing," and "captive" for the zeros and will consider how to deal with them.

 The left diagram uses the example of 3090. The zero in the hundreds place is labeled "captive" and the zero in the ones place is labeled trailing. The right diagram uses the example 0.008020. The three zeros in the ones, tenths, and hundredths places are labeled "leading." The zero in the ten-thousandths place is labeled "captive" and the zero in the millionths place is labeled "trailing."

Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point. Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.

 The left diagram uses the example of 1267 meters. The number 1 is the first nonzero figure on the left. 1267 has 4 significant figures in total. The right diagram uses the example of 55.0 grams. The number 5 in the tens place is the first nonzero figure on the left. 55.0 has 3 significant figures. Note that the 0 is to the right of the decimal point and therefore is a significant figure.

In the above left example, all four digits are significant since all four are nonzero. On the right, the trailing zero is to the right of the decimal, hence all three digits are significant.

 The left diagram uses the example of 70.607 milliliters. The number 7 is the first nonzero figure on the left. 70.607 has 5 significant figures in total, as all figures are measured including the 2 zeros. The right diagram uses the example of 0.00832407 M L. The number 8 is the first nonzero figure on the left. 0.00832407 has 6 significant figures.

When zeros are captive (in between nonzero numbers), they are considered significant. In the the example of 70.607 mL, all digits are significant because both zeros are captive. The leading zeros in the left example are not significant. We could use exponential notation (review [here](#)) and express the number as 8.32407×10^{-3} , then the number 8.32407 contains all of the significant figures and 10^{-3} locates the decimal point.

The number of significant figures is uncertain in a number that ends with a zero to the left of the decimal point location. The zeros in the measurement 1,300 grams could be significant or they could simply indicate where the decimal point is located. The

ambiguity can be resolved with the use of exponential notation: 1.3×10^3 (two significant figures), 1.30×10^3 (three significant figures, if the tens place was measured), or 1.300×10^3 (four significant figures, if the ones place was also measured). In cases where only the decimal-formatted number is available, it is prudent to assume that all trailing zeros are not significant.

 This figure uses the example of 1300 grams. The one and the 3 are significant figures as they are clearly the result of measurement. The 2 zeros could be significant if they were measured or they could be placeholders.

When determining significant figures, be sure to pay attention to reported values and think about the measurement and significant figures in terms of what is reasonable or likely when evaluating whether the value makes sense. For example, the official January 2014 census reported the resident population of the US as 317,297,725. Do you think the US population was correctly determined to the reported nine significant figures, that is, to the exact number of people? People are constantly being born, dying, or moving into or out of the country, and assumptions are made to account for the large number of people who are not actually counted. Because of these uncertainties, it might be more reasonable to expect that we know the population to within perhaps a million or so, in which case the population should be reported as 3.17×10^8 people.

Key Concepts and Summary

Quantities can be exact or measured. Measured quantities have an associated uncertainty that is represented by the number of significant figures in the measurement. Measured values can be accurate (close to the true value) and/or precise (showing little variation when measured repeatedly).

Chemistry End of Chapter Exercises

1. Express each of the following numbers in scientific notation with correct significant figures:

- (a) 711.0
- (b) 0.239
- (c) 90743
- (d) 134.2
- (e) 0.05499
- (f) 10000.0
- (g) 0.000000738592

2. Express each of the following numbers in exponential notation with correct significant figures:

- (a) 704
- (b) 0.03344
- (c) 547.9
- (d) 22086
- (e) 1000.00
- (f) 0.0000000651
- (g) 0.007157

3. Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

- (a) the number of eggs in a basket
- (b) the mass of a dozen eggs
- (c) the number of gallons of gasoline necessary to fill an automobile gas tank
- (d) the number of cm in 2 m
- (e) the mass of a textbook
- (f) the time required to drive from San Francisco to Kansas City at an average speed of 53 mi/h

4. Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

- (a) the number of seconds in an hour
- (b) the number of pages in this book
- (c) the number of grams in your weight
- (d) the number of grams in 3 kilograms
- (e) the volume of water you drink in one day
- (f) the distance from San Francisco to Kansas City

5. How many significant figures are contained in each of the following measurements?

- (a) 38.7 g
- (b) 2×10^{18} m
- (c) 3,486,002 kg
- (d) 9.74150×10^{-4} J
- (e) 0.0613 cm^3
- (f) 17.0 kg
- (g) 0.01400 g/mL

6. How many significant figures are contained in each of the following measurements?

- (a) 53 cm
- (b) 2.05×10^8 m
- (c) 86,002 J
- (d) 9.740×10^4 m/s
- (e) 10.0613 m^3
- (f) 0.17 g/mL
- (g) 0.88400 s

7. The following quantities were reported on the labels of commercial products. Determine the number of significant figures in each.

- (a) 0.0055 g active ingredients
- (b) 12 tablets
- (c) 3% hydrogen peroxide
- (d) 5.5 ounces
- (e) 473 mL
- (f) 1.75% bismuth
- (g) 0.001% phosphoric acid
- (h) 99.80% inert ingredients

8. Round off each of the following numbers to two significant figures:

- (a) 0.436
- (b) 9.000
- (c) 27.2

- (d) 135
- (e) 1.497×10^{-3}
- (f) 0.445

9. Round off each of the following numbers to two significant figures:

- (a) 517
- (b) 86.3
- (c) 6.382×10^3
- (d) 5.0008
- (e) 22.497
- (f) 0.885

12. Consider the results of the archery contest shown in this figure.

- (a) Which archer is most precise?
- (b) Which archer is most accurate?
- (c) Who is both least precise and least accurate?

 4 targets are shown each with 4 holes indicating where the arrows hit the targets. Archer W put all 4 arrows closely around the center of the target. Archer X put all 4 arrows in a tight cluster but far to the lower right of the target. Archer Y put all 4 arrows at different corners of the target. All 4 arrows are very far from the center of the target. Archer Z put 2 arrows close to the target and 2 other arrows far outside of the target.

13. Classify the following sets of measurements as accurate, precise, both, or neither.

- (a) Checking for consistency in the weight of chocolate chip cookies: 17.27 g, 13.05 g, 19.46 g, 16.92 g
- (b) Testing the volume of a batch of 25-mL pipettes: 27.02 mL, 26.99 mL, 26.97 mL, 27.01 mL
- (c) Determining the purity of gold: 99.9999%, 99.9998%, 99.9998%, 99.9999%

Glossary

accuracy

how closely a measurement aligns with a correct value

exact number

number derived by counting or by definition

meniscus

the curve in the upper surface of a liquid close to the surface of the container, caused by surface tension

precision

how closely a measurement matches the same measurement when repeated

rounding

procedure used to ensure that calculated results properly reflect the uncertainty in the measurements used in the calculation

significant figures

(also, significant digits) all of the measured digits in a determination, including the uncertain last digit

uncertainty

estimate of amount by which measurement differs from true value

1.3: Uncertainty in Measurement is shared under a [CC BY-SA](https://creativecommons.org/licenses/by-sa/4.0/) license and was authored, remixed, and/or curated by LibreTexts.

1.4: Matter

4

Harper College Chemistry Department

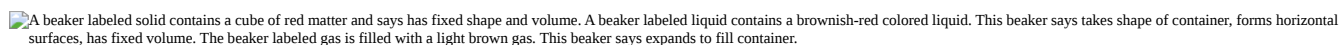
▼ Details

By the end of this section, you will be able to:

- Describe the basic properties of each physical state of matter: solid, liquid, and gas
- Define and give examples of atoms and molecules
- Classify matter as an element, compound, homogeneous mixture, or heterogeneous mixture with regard to its physical state and composition
- Distinguish between mass and weight
- Apply the law of conservation of matter

Matter is defined as anything that occupies space (volume) and has mass (amount). Matter is all around us. Solids and liquids are more obviously matter: since, we can see that they take up space and their weight tells us that they have mass. Gases are also matter; if gases did not take up space, a balloon would stay collapsed rather than inflate when filled with some amount of gas.

Solids, liquids and gases are the three states (phases) of matter commonly found on earth ([Figure 1](#)). A solid is rigid and possesses a definite shape. A liquid flows and takes the shape of a container, except that it forms a flat or slightly curved upper surface when acted upon by gravity (in zero gravity, liquids assume a spherical shape). Both liquid and solid samples have volumes that are very nearly independent of pressure. A gas takes both the shape and volume of its container. Remember volume is the amount of space occupied by a substance where shape is the outline of the substance.

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

The three most common states or phases of matter are solid, liquid, and gas.

A fourth state of matter, plasma, occurs naturally in the interiors of stars. A plasma is a gaseous state of matter that contains appreciable numbers of electrically charged particles. The presence of these charged particles imparts unique properties to plasmas that justify their classification as a state of matter distinct from gases.

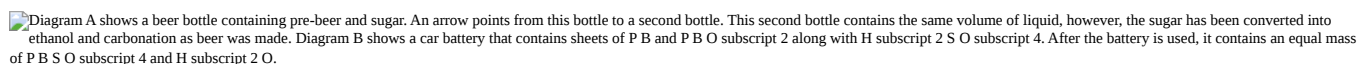
Some samples of matter appear to have properties of solids, liquids and/or gases at the same time. This can occur when the sample is composed of many small pieces. For example, we can pour sand as if it were a liquid because it is composed of many small grains of solid sand. Matter can also have properties of more than one state when it is a mixture, such as with clouds. Clouds appear to behave somewhat like gases, but they are actually mixtures of air (gas) and tiny particles of water (liquid or solid).

The mass of an object is a measure of the amount of matter in it. One way to measure an object's mass is to measure the force it takes to accelerate the object. It takes much more force to accelerate a car than a bicycle because the car has much more mass. A more common way to determine the mass of an object is to use a balance to compare its mass with a standard mass.

Although weight is related to mass, it is not the same thing. Weight refers to the force that gravity exerts on an object. This force is directly proportional to the mass of the object. The weight of an object changes as the force of gravity changes, but its mass does not. An astronaut's mass does not change just because she goes to the moon. But her weight on the moon is only one-sixth her earth-bound weight because the moon's gravity is only one-sixth that of the earth's. She may feel "weightless" during her trip when she experiences negligible external forces (gravitational or any other), although she is, of course, never "massless."

The law of conservation of matter summarizes many scientific observations about matter: It states that *there is no detectable change in the total quantity of matter present when matter converts from one type to another (a chemical change) or changes among solid, liquid, or gaseous states (a physical change)*. Brewing beer and the operation of batteries provide examples of the conservation of matter ([Figure 2](#)). During the brewing of beer, the ingredients (water, yeast, grains, malt, hops and sugar) are converted into beer (water, alcohol, carbonation and flavoring substances) with no actual loss of substance. This is most clearly seen during the bottling process, when glucose turns into ethanol and carbon dioxide, and the total mass of the substances does not change. This can also be seen in a lead-acid car battery: the original substances (lead, lead oxide and sulfuric acid), which are

capable of producing electricity, are changed into other substances (lead sulfate and water) that do not produce electricity, with no change in the actual amount of matter.

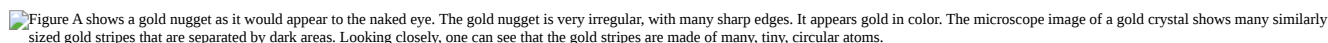
Diagram A shows a beer bottle containing pre-beer and sugar. An arrow points from this bottle to a second bottle. This second bottle contains the same volume of liquid, however, the sugar has been converted into ethanol and carbonation as beer was made. Diagram B shows a car battery that contains sheets of Pb and PbO subscript 2 along with H subscript 2 SO subscript 4. After the battery is used, it contains an equal mass of Pb SO subscript 4 and H subscript 2 O.

- (a) The mass of beer precursor materials is the same as the mass of beer produced: sugar has become alcohol and carbonation. (b) The mass of the lead, lead oxide plates and sulfuric acid that goes into the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

Although this conservation law holds true for all conversions of matter, convincing examples are few and far between because, outside of the controlled conditions in a laboratory, we seldom collect all of the material that is produced during a particular conversion. For example, when you eat, digest and assimilate food, all of the matter in the original food is preserved. But because some of the matter is incorporated into your body, and much is excreted as various types of waste, it is challenging to verify by measurement.

Atoms and Molecules

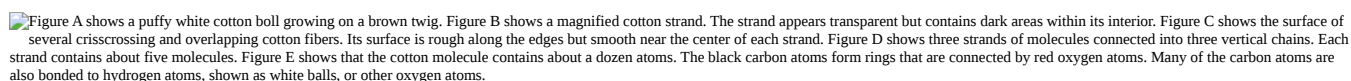
An atom is the smallest particle of an element that has the properties of that element. Consider the element gold, for example. Imagine cutting a gold nugget in half, then cutting one of the halves in half and repeating this process until a piece of gold remained that was so small that it could not be cut in half (regardless of how tiny your knife may be). This minimally sized piece of gold is an atom (from the Greek *atomos*, meaning “indivisible”) (Figure 3). This atom would no longer be gold if it were divided any further.

Figure A shows a gold nugget as it would appear to the naked eye. The gold nugget is very irregular, with many sharp edges. It appears gold in color. The microscope image of a gold crystal shows many similarly sized gold stripes that are separated by dark areas. Looking closely, one can see that the gold stripes are made of many, tiny, circular atoms.

- (a) This photograph shows a gold nugget. (b) A scanning-tunneling microscope (STM) can generate views of the surfaces of solids, such as this image of a gold crystal. Each sphere represents one gold atom (credit a: modification of work by United States Geological Survey; credit b: modification of work by “Erwinrossen”/Wikimedia Commons).

The first suggestion that matter is composed of atoms is attributed to the Greek philosophers Leucippus and Democritus, who developed their ideas in the 5th century BCE. However, it was not until the early nineteenth century that John Dalton (1766–1844), a British school teacher with a keen interest in science, supported this hypothesis with quantitative measurements. Since then, repeated experiments have confirmed many aspects of this hypothesis and it has become one of the central theories of chemistry. Other aspects of Dalton’s atomic theory are still used but with minor revisions (details of Dalton’s theory are provided in the chapter on atoms and molecules).

An atom is so small that its size is difficult to imagine. One of the smallest things we can see with our unaided eye is a single thread of a spider web: These strands are about 1/10,000 of a centimeter (0.0001 cm) in diameter. Although the cross-section of one strand is almost impossible to see without a microscope, it is huge on an atomic scale. A single carbon atom in the web has a diameter of about 0.00000015 centimeter and it would take about 7,000 carbon atoms to span the diameter of the strand. To put this in perspective, if a carbon atom were the size of a dime, the cross-section of one strand would be larger than a football field. This would require about 150 million “dimes” or carbon atoms to cover it. (Figure 4) shows increasingly close microscopic and atomic-level views of ordinary cotton.

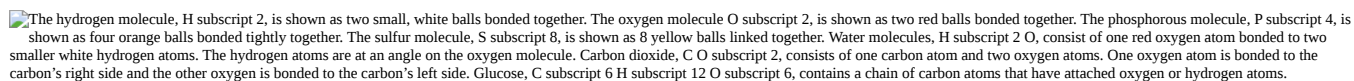
Figure A shows a puffy white cotton ball growing on a brown twig. Figure B shows a magnified cotton strand. The strand appears transparent but contains dark areas within its interior. Figure C shows the surface of several crisscrossing and overlapping cotton fibers. Its surface is rough along the edges but smooth near the center of each strand. Figure D shows three strands of molecules connected into three vertical chains. Each strand contains about five molecules. Figure E shows that the cotton molecule contains about a dozen atoms. The black carbon atoms form rings that are connected by red oxygen atoms. Many of the carbon atoms are also bonded to hydrogen atoms, shown as white balls, or other oxygen atoms.

- These images provide an increasingly closer view: (a) a cotton ball, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) a cotton fiber viewed under an electron microscope (much higher magnification than with the optical microscope) and (d and e) atomic-level models of the fiber (spheres of different colors represent atoms of different elements) (credit c: modification of work by “Featheredtar”/Wikimedia Commons).

An atom is so light that its mass is also difficult to imagine. A billion lead atoms (1,000,000,000 atoms) masses about 3×10^{-13} grams, a mass that is far too light to be measured on even the world’s most sensitive balances. It would require over 300,000,000,000,000 lead atoms (300 trillion, or 3×10^{14}) to be measured and they would mass only 0.0000001 gram.

It is rare to find collections of individual atoms. Only a few elements, such as the gases helium, neon and argon, consist of a collection of individual atoms that move about independently of one another. Other elements, such as the gas states of hydrogen, nitrogen, oxygen and chlorine, are composed of units that consist of pairs of atoms (Figure 5). One form of the element phosphorus consists of units composed of four phosphorus atoms. The element sulfur exists in various forms, one of which consists of units

composed of eight sulfur atoms. These units are called molecules. A molecule consists of two or more atoms joined by strong forces called chemical bonds. The atoms in a molecule move around as a unit, much like the cans of soda in a six-pack or a bunch of keys joined together on a single key ring. A molecule may consist of two or more identical atoms, as in the molecules found in the elements hydrogen, oxygen and sulfur or it may consist of two or more different atoms, as in the molecules found in water. Each water molecule is a unit that contains two hydrogen atoms and one oxygen atom. Each glucose molecule is a unit that contains 6 carbon atoms, 12 hydrogen atoms and 6 oxygen atoms. Like atoms, molecules are incredibly small and light. If an ordinary glass of water were enlarged to the size of the earth, the water molecules inside it would be about the size of golf balls.

The hydrogen molecule, H_2 , is shown as two small, white balls bonded together. The oxygen molecule O_2 , is shown as two red balls bonded together. The phosphorous molecule, P_4 , is shown as four orange balls bonded tightly together. The sulfur molecule, S_8 , is shown as 8 yellow balls linked together. Water molecules, H_2O , consist of one red oxygen atom bonded to two smaller white hydrogen atoms. The hydrogen atoms are at an angle on the oxygen molecule. Carbon dioxide, CO_2 , consists of one carbon atom and two oxygen atoms. One oxygen atom is bonded to the carbon's right side and the other oxygen is bonded to the carbon's left side. Glucose, $C_6H_{12}O_6$, contains a chain of carbon atoms that have attached oxygen or hydrogen atoms.

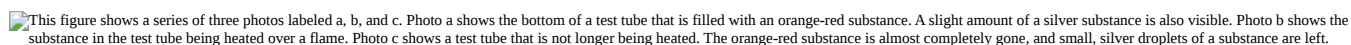
The elements hydrogen, oxygen, phosphorus and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide and glucose consist of combinations of atoms of different elements.

Classifying Matter

We can classify matter into several categories. Two broad categories are pure substances and mixtures. A pure substance has a constant composition. All specimens of a pure substance have exactly the same makeup and properties. Any sample of sucrose (table sugar) consists of 42.1% carbon, 6.5% hydrogen and 51.4% oxygen by mass. Any sample of sucrose also has the same physical properties, such as melting point, color and sweetness, regardless of the source from which it is isolated.

We can divide pure substances into two classes: elements and compounds. Pure substances that cannot be broken down into simpler substances by chemical changes are called elements. Iron, silver, gold, aluminum, sulfur, oxygen, and copper are familiar examples of the more than 100 known elements. Of these, about 90 occur naturally on the earth and two dozen or so have been created in laboratories.

Pure substances that can be broken down by chemical changes are called compounds. This breakdown may produce either elements or other compounds, or both. Mercury(II) oxide, an orange, crystalline solid, can be broken down by heat into the elements mercury and oxygen ([Figure 6](#)). When heated in the absence of air, the compound sucrose is broken down into the element carbon and the compound water. (The initial stage of this process, when the sugar is turning brown, is known as caramelization—this is what imparts the characteristic sweet and nutty flavor to caramel apples, caramelized onions, and caramel). Silver chloride is a white solid that can be broken down into its elements, silver and chlorine, by the absorption of light. This property is the basis for the use of this compound in photographic films and photochromic eyeglasses (those with lenses that darken when exposed to light).

This figure shows a series of three photos labeled a, b, and c. Photo a shows the bottom of a test tube that is filled with an orange-red substance. A slight amount of a silver substance is also visible. Photo b shows the substance in the test tube being heated over a flame. Photo c shows a test tube that is not longer being heated. The orange-red substance is almost completely gone, and small, silver droplets of a substance are left.

(a) The compound mercury(II) oxide, (b) when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas (credit: modification of work by Paul Flowers).

Many compounds break down when heated. This [site](#) shows the breakdown of mercury(II) oxide, HgO . You can also view an example of the [photochemical decomposition of silver chloride](#) ($AgCl$), the basis of early photography.

The properties of combined elements are different from those in the free, or uncombined, state. For example, white crystalline sugar (sucrose) is a compound resulting from the chemical combination of the element carbon, which is a black solid in one of its uncombined forms, and the two elements hydrogen and oxygen, which are colorless gases when uncombined. Free sodium, an element that is a soft, shiny, metallic solid, and free chlorine, an element that is a yellow-green gas, combine to form sodium chloride (table salt), a compound that is a white, crystalline solid.

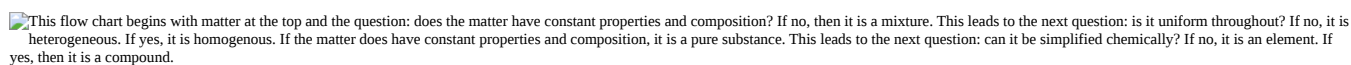
A mixture is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes, such as evaporation (you will learn more about this later). A mixture with a composition that varies from point to point is called a heterogeneous mixture. Italian dressing is an example of a heterogeneous mixture ([Figure 7a](#)). Its composition can vary because we can make it from varying amounts of oil, vinegar and herbs. It is not the same from point to point throughout the mixture—one drop may be mostly vinegar, whereas a different drop may be mostly oil or herbs because the oil and vinegar separate and the herbs settle. Other examples of heterogeneous mixtures are chocolate chip cookies (we can see the separate bits of chocolate, nuts and cookie dough) and granite (we can see the quartz, mica, feldspar and more).

Diagram A shows a glass containing a red liquid with a layer of yellow oil floating on the surface of the red liquid. A zoom in box is magnifying a portion of the red liquid that contains some of the yellow oil. The zoomed in image shows that oil is forming round droplets within the red liquid. Diagram B shows a photo of Gatorade G 2. A zoom in box is magnifying a portion of the Gatorade, which is uniformly red.

(a) Oil and vinegar salad dressing is a heterogeneous mixture because its composition is not uniform throughout. (b) A commercial sports drink is a homogeneous mixture because its composition is uniform throughout. (credit a “left”: modification of work by John Mayer; credit a “right”: modification of work by Umberto Salvagnin; credit b “left: modification of work by Jeff Bedford)

A homogeneous mixture, also called a solution, exhibits a uniform composition and appears visually the same throughout. An example of a solution is a sports drink, consisting of water, sugar, coloring, flavoring, and electrolytes mixed together uniformly (Figure 7b). Each drop of a sports drink tastes the same because each drop contains the same amounts of water, sugar and other components. Note that the composition of a sports drink can vary—it could be made with somewhat more or less sugar, flavoring or other components, and still be a sports drink. Other examples of homogeneous mixtures include air, maple syrup, gasoline and a solution of salt in water.

Although there are just over 100 elements, tens of millions of chemical compounds result from different combinations of these elements. Each compound has a specific composition and possesses definite chemical and physical properties by which we can distinguish it from all other compounds. In addition, of course, there are innumerable ways to combine elements and compounds to form different mixtures. A summary of how to distinguish between the various major classifications of matter is shown in (Figure 8).

This flow chart begins with matter at the top and the question: does the matter have constant properties and composition? If no, then it is a mixture. This leads to the next question: is it uniform throughout? If no, it is heterogeneous. If yes, it is homogenous. If the matter does have constant properties and composition, it is a pure substance. This leads to the next question: can it be simplified chemically? If no, it is an element. If yes, then it is a compound.

Depending on its properties, a given substance can be classified as a homogeneous mixture, a heterogeneous mixture, a compound or an element.

Eleven elements make up about 99% of the earth’s crust and atmosphere (Table 1). Oxygen constitutes nearly one-half and silicon about one-quarter of the total quantity of these elements. A majority of elements on earth are found in chemical combinations with other elements; about one-quarter of the elements are also found in the free state.

Table 1: Elemental Composition of Earth

Element	Symbol	Percent Mass	Element	Symbol	Percent Mass
oxygen	O	49.20	chlorine	Cl	0.19
silicon	Si	25.67	phosphorus	P	0.11
aluminum	Al	7.50	manganese	Mn	0.09
iron	Fe	4.71	carbon	C	0.08
calcium	Ca	3.39	sulfur	S	0.06
sodium	Na	2.63	barium	Ba	0.04
potassium	K	2.40	nitrogen	N	0.03
magnesium	Mg	1.93	fluorine	F	0.03
hydrogen	H	0.87	strontium	Sr	0.02
titanium	Ti	0.58	all others	—	0.47

Conversion of Temperature Units

We use the word temperature to refer to the hotness or coldness of a substance. One way we measure a change in temperature is to use the fact that most substances expand when their temperature increases and contract when their temperature decreases. The mercury or alcohol in a common glass thermometer changes its volume as the temperature changes. Because the volume of the liquid changes more than the volume of the glass, we can see the liquid expand when it gets warmer and contract when it gets cooler.

To mark a scale on a thermometer, we need a set of reference values: Two of the most commonly used are the freezing and boiling temperatures of water at a specified atmospheric pressure. On the Celsius scale, 0 °C is defined as the freezing temperature of water

and 100 °C as the boiling temperature of water. The space between the two temperatures is divided into 100 equal intervals, which we call degrees. On the Fahrenheit scale, the freezing point of water is defined as 32 °F and the boiling temperature as 212 °F. The space between these two points on a Fahrenheit thermometer is divided into 180 equal parts (degrees).

Defining the Celsius and Fahrenheit temperature scales as described in the previous paragraph results in a slightly more complex relationship between temperature values on these two scales than for different units of measure for other properties. Most measurement units for a given property are directly proportional to one another ($y = mx$). Using familiar length units as one example:

$$\text{length in feet} = \left(\frac{\text{SI}\{1\}\{\text{ft}\}}{\text{SI}\{12\}\{\text{in}\}} \right) \times \text{length in inches}$$

where y = length in feet, x = length in inches, and the proportionality constant, m , is the conversion factor. The Celsius and Fahrenheit temperature scales, however, do not share a common zero point, and so the relationship between these two scales is a linear one rather than a proportional one ($y = mx + b$). Consequently, converting a temperature from one of these scales into the other requires more than simple multiplication by a conversion factor, m , it also must take into account differences in the scales' zero points (b).

The linear equation relating Celsius and Fahrenheit temperatures is easily derived from the two temperatures used to define each scale. Representing the Celsius temperature as x and the Fahrenheit temperature as y , the slope, m , is computed to be:

$$m = \frac{\Delta y}{\Delta x} = \frac{\text{SI}\{212\}\{\text{degree F}\} - \text{SI}\{32\}\{\text{degree F}\}}{\text{SI}\{100\}\{\text{degree C}\} - \text{SI}\{0\}\{\text{degree C}\}} \\ = \frac{\text{SI}\{180\}\{\text{degree F}\}}{\text{SI}\{100\}\{\text{degree C}\}} = \frac{\text{SI}\{9\}\{\text{degree F}\}}{\text{SI}\{5\}\{\text{degree C}\}}$$

The y-intercept of the equation, b , is then calculated using either of the equivalent temperature pairs, (100 °C, 212 °F) or (0 °C, 32 °F), as:

$$b = y - mx = \text{SI}\{32\}\{\text{degree F}\} - \left(\frac{\text{SI}\{9\}\{\text{degree F}\}}{\text{SI}\{5\}\{\text{degree C}\}} \right) \times \text{SI}\{0\}\{\text{degree C}\} = \text{SI}\{32\}\{\text{degree F}\}$$

The equation relating the temperature scales is then:

$$T_{\{\text{degree F}\}} = \left(\frac{\text{SI}\{9\}\{\text{degree F}\}}{\text{SI}\{5\}\{\text{degree C}\}} \right) \times T_{\{\text{degree C}\}} + \text{SI}\{32\}\{\text{degree C}\}$$

An abbreviated form of this equation that omits the measurement units is:

$$T_{\{\text{degree F}\}} = \left(\frac{9}{5} \right) \times T_{\{\text{degree C}\}} + 32$$

Rearrangement of this equation yields the form useful for converting from Fahrenheit to Celsius:

$$T_{\{\text{degree C}\}} = \frac{5}{9} (T_{\{\text{degree F}\}} - 32)$$

As mentioned earlier in this chapter, the SI unit of temperature is the kelvin (K). Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale in which 0 (zero) K corresponds to the lowest temperature that can theoretically be achieved. The early 19th-century discovery of the relationship between a gas's volume and temperature suggested that the volume of a gas would be zero at -273.15 °C. In 1848, British physicist William Thompson, who later adopted the title of Lord Kelvin, proposed an absolute temperature scale based on this concept (further treatment of this topic is provided in this text's chapter on gases).

The freezing temperature of water on this scale is 273.15 K and its boiling temperature 373.15 K. Notice the numerical difference in these two reference temperatures is 100, the same as for the Celsius scale, and so the linear relation between these two temperature scales will exhibit a slope of $\frac{\text{SI}\{K\}}{\text{SI}\{\text{degree C}\}}$. Following the same approach, the equations for converting between the kelvin and Celsius temperature scales are derived to be:

$$T_{\{K\}} = T_{\{\text{degree C}\}} + 273.15$$

$$T_{\{\text{degree C}\}} = T_{\{K\}} - 273.15$$

The 273.15 in these equations has been determined experimentally, so it is not exact. [Figure 1](#) shows the relationship among the three temperature scales. Recall that we do not use the degree sign with temperatures on the kelvin scale.

 A thermometer is shown for the Fahrenheit, Celsius and Kelvin scales. Under the Fahrenheit scale, the boiling point of water is 212 degrees while the freezing point of water is 32 degrees. Therefore, there are 180 Fahrenheit degrees between the boiling point of water and the freezing point of water. Under the Celsius scale, the boiling point of water is 100 degrees while the freezing point of water is 0 degrees. Therefore, there are 100 Celsius degrees between the boiling point and freezing point of water. Under the kelvin scale, the boiling point of water is 373.15 K, while the freezing point of water is 273.15 K. 233.15 K is equal to negative 40 degrees Celsius, which is also equal to negative 40 degrees Fahrenheit.

The Fahrenheit, Celsius, and kelvin temperature scales are compared.

Although the kelvin (absolute) temperature scale is the official SI temperature scale, Celsius is commonly used in many scientific contexts and is the scale of choice for nonscience contexts in almost all areas of the world. Very few countries (the U.S. and its territories, the Bahamas, Belize, Cayman Islands, and Palau) still use Fahrenheit for weather, medicine, and cooking.

Conversion from Celsius

Normal body temperature has been commonly accepted as 37.0 °C (although it varies depending on time of day and method of measurement, as well as among individuals). What is this temperature on the kelvin scale and on the Fahrenheit scale?

Solution

$$\text{K} = \text{C} + 273.15 = 37.0 + 273.2 = 310.2 \text{ K}$$

$$\text{F} = \frac{9}{5} \text{C} + 32.0 = \left(\frac{9}{5} \times 37.0 \right) + 32.0 = 66.6 + 32.0 = 98.6 \text{ F}$$

Check Your Learning

Convert 80.92 °C to K and °F.

Answer:

354.07 K, 177.7 °F

Conversion from Fahrenheit

Baking a ready-made pizza calls for an oven temperature of 450 °F. If you are in Europe, and your oven thermometer uses the Celsius scale, what is the setting? What is the kelvin temperature?

Solution

$$\text{C} = \frac{5}{9} (\text{F} - 32) = \frac{5}{9} (450 - 32) = \frac{5}{9} \times 418 = 232 \text{ C} \rightarrow \text{set oven to } 230 \text{ C (two significant figures)}$$

$$\text{K} = \text{C} + 273.15 = 230 + 273 = 503 \text{ K} \rightarrow 5.0 \times 10^2 \text{ K (two significant figures)}$$

Check Your Learning

Convert 50 °F to °C and K.

Answer:

10 °C, 280 K

In case you are wondering

DECOMPOSITION OF WATER / PRODUCTION OF HYDROGEN

Water consists of the elements hydrogen and oxygen combined in a 2 to 1 ratio. Water can be broken down into hydrogen and oxygen gases by the addition of energy. One way to do this is with a battery or power supply, as shown in (Figure 9).

 A rectangular battery is immersed in a beaker filled with liquid. Each of the battery terminals are covered by an overturned test tube. The test tubes each contain a bubbling liquid. Zoom in areas indicate that the liquid in the beaker is water, 2 H subscript 2 O liquid. The bubbles in the test tube over the negative terminal are hydrogen gas, 2 H subscript 2 gas. The bubbles in the test tube over the positive terminal are oxygen gas, O subscript 2 gas.

The decomposition of water is shown at the macroscopic, microscopic, and symbolic levels. The battery provides an electric current (microscopic) that decomposes water. At the macroscopic level, the liquid separates into the gases hydrogen (on the left) and oxygen (on the right). Symbolically, this change is presented by showing how liquid H₂O separates into H₂ and O₂ gases.

The breakdown of water involves a rearrangement of the atoms in water molecules into different molecules, each composed of two hydrogen atoms and two oxygen atoms, respectively. Two water molecules form one oxygen molecule and two hydrogen

molecules. The representation for what occurs, $\text{2H}_2\text{O(l)} \rightarrow \text{2H}_2\text{(g)} + \text{O}_2\text{(g)}$, will be explored in more depth in later chapters.

The two gases produced have distinctly different properties. Oxygen is not flammable but is required for combustion of a fuel, and hydrogen is highly flammable and a potent energy source. How might this knowledge be applied in our world? One application involves research into more fuel-efficient transportation. Fuel-cell vehicles (FCV) run on hydrogen instead of gasoline (Figure 10). They are more efficient than vehicles with internal combustion engines, are nonpolluting, and reduce greenhouse gas emissions, making us less dependent on fossil fuels. FCVs are not yet economically viable, however, and current hydrogen production depends on natural gas. If we can develop a process to economically decompose water, or produce hydrogen in another environmentally sound way, FCVs may be the way of the future.

 The fuel cell consists of a proton exchange membrane sandwiched between an anode and a cathode. Hydrogen gas enters the battery near the anode. Oxygen gas enters the battery near the cathode. The entering hydrogen gas is broken up into single white spheres that each have a positive charge. These are protons. The protons repel negatively-charged electrons within the anode. These electrons travel through a circuit, providing electricity to anything attached to the battery. The protons continue through the proton exchange membrane and through the cathode to reach the oxygen gas molecules at the opposite end of the battery. There, the oxygen atoms split up into single red spheres. Each oxygen atom takes on two of the incoming protons to form a water molecule.

A fuel cell generates electrical energy from hydrogen and oxygen via an electrochemical process and produces only water as the waste product.

CHEMISTRY OF CELL PHONES

Imagine how different your life would be without cell phones (Figure 11) and other smart devices. Cell phones are made from numerous chemical substances, which are extracted, refined, purified, and assembled using an extensive and in-depth understanding of chemical principles. About 30% of the elements that are found in nature are found within a typical smart phone. The case/body/frame consists of a combination of sturdy, durable polymers comprised primarily of carbon, hydrogen, oxygen, and nitrogen [acrylonitrile butadiene styrene (ABS) and polycarbonate thermoplastics], and light, strong, structural metals, such as aluminum, magnesium, and iron. The display screen is made from a specially toughened glass (silica glass strengthened by the addition of aluminum, sodium, and potassium) and coated with a material to make it conductive (such as indium tin oxide). The circuit board uses a semiconductor material (usually silicon); commonly used metals like copper, tin, silver, and gold; and more unfamiliar elements such as yttrium, praseodymium, and gadolinium. The battery relies upon lithium ions and a variety of other materials, including iron, cobalt, copper, polyethylene oxide, and polyacrylonitrile.

 A cell phone is labeled to show what its components are made of. The case components are made of polymers such as ABS and or metals such as aluminum, iron, and magnesium. The processor components are made of silicon, common metals such as copper, tin and gold, and uncommon elements such as yttrium and gadolinium. The screen components are made of silicon oxide, also known as glass. The glass is strengthened by the addition of aluminum, sodium, and potassium. The battery components contain lithium combined with other metals such as cobalt, iron, and copper.

Almost one-third of naturally occurring elements are used to make a cell phone. (credit: modification of work by John Taylor)

Key Concepts and Summary

Matter is anything that occupies space and has mass. The basic building block of matter is the atom, the smallest unit of an element that can enter into combinations with atoms of the same or other elements. In many substances, atoms are combined into molecules. On earth, matter commonly exists in three states: solids, of fixed shape and volume; liquids, of variable shape but fixed volume; and gases, of variable shape and volume. Under high-temperature conditions, matter also can exist as a plasma. Most matter is a mixture: It is composed of two or more types of matter that can be present in varying amounts and can be separated by physical means. Heterogeneous mixtures vary in composition from point to point; homogeneous mixtures have the same composition from point to point. Pure substances consist of only one type of matter. A pure substance can be an element, which consists of only one type of atom and cannot be broken down by a chemical change, or a compound, which consists of two or more types of atoms.

Chemistry End of Chapter Exercises

1. Why do we use an object's mass, rather than its weight, to indicate the amount of matter it contains?
2. What properties distinguish solids from liquids? Liquids from gases? Solids from gases?
3. How does a heterogeneous mixture differ from a homogeneous mixture? How are they similar?
4. How does a homogeneous mixture differ from a pure substance? How are they similar?
5. How does an element differ from a compound? How are they similar?
6. How do molecules of elements and molecules of compounds differ? In what ways are they similar?
7. How does an atom differ from a molecule? In what ways are they similar?

8. Many of the items you purchase are mixtures of pure compounds. Select three of these commercial products and prepare a list of the ingredients that are pure compounds.
9. Classify each of the following as an element, a compound, or a mixture:
- (a) copper
 - (b) water
 - (c) nitrogen
 - (d) sulfur
 - (e) air
 - (f) sucrose
 - (g) a substance composed of molecules each of which contains two iodine atoms
 - (h) gasoline
10. Classify each of the following as an element, a compound, or a mixture:
- (a) iron
 - (b) oxygen
 - (c) mercury(II) oxide
 - (d) pancake syrup
 - (e) carbon dioxide
 - (f) a substance composed of molecules each of which contains one hydrogen atom and one chlorine atom
 - (g) baking soda
 - (h) baking powder
11. A sulfur atom and a sulfur molecule are not identical. What is the difference?
12. How are the molecules in oxygen gas, the molecules in hydrogen gas and water molecules similar? How do they differ?
13. We refer to astronauts in space as weightless, but not without mass. Why?
14. As we drive an automobile, we don't think about the chemicals consumed and produced. Prepare a list of the principal chemicals consumed and produced during the operation of an automobile.
15. Matter is everywhere around us. Make a list by name of fifteen different kinds of matter that you encounter every day. Your list should include (and label at least one example of each) the following: a solid, a liquid, a gas, an element, a compound, a homogeneous mixture, a heterogeneous mixture and a pure substance.
16. When elemental iron corrodes it combines with oxygen in the air to ultimately form red brown iron(III) oxide which we call rust. (a) If a shiny iron nail with an initial mass of 23.2 g is measured after being coated in a layer of rust, would you expect the mass to have increased, decreased, or remained the same? Explain. (b) If the mass of the iron nail increases to 24.1 g, what mass of oxygen combined with the iron?
17. As stated in the text, convincing examples that demonstrate the law of conservation of matter outside of the laboratory are few and far between. Indicate whether the mass would increase, decrease or stay the same for the following scenarios, where chemical reactions take place:
- (a) Exactly one pound of bread dough is placed in a baking tin. The dough is cooked in an oven at 350 °F releasing a wonderful aroma of freshly baked bread during the cooking process. Is the mass of the baked loaf less than, greater than or the same as the one pound of original dough? Explain.
 - (b) When magnesium burns in air a white flaky ash of magnesium oxide is produced. Is the mass of magnesium oxide less than, greater than, or the same as the original piece of magnesium? Explain.

(c) Antoine Lavoisier, the French scientist credited with first stating the law of conservation of matter, heated a mixture of tin and air in a sealed flask to produce tin oxide. Did the mass of the sealed flask and contents decrease, increase, or remain the same after the heating?

18. Yeast converts glucose to ethanol and carbon dioxide during anaerobic fermentation as depicted in the simple chemical equation here:



- (a) If 200.0 g of glucose is fully converted, what will be the total mass of ethanol and carbon dioxide produced?
- (b) If the fermentation is carried out in an open container, would you expect the mass of the container and contents after fermentation to be less than, greater than or the same as the mass of the container and contents before fermentation? Explain.
- (c) If 97.7 g of carbon dioxide is produced, what mass of ethanol is produced, assuming the experiment started with 200.0 g of glucose?

Glossary

atom

smallest particle of an element that can enter into a chemical combination

compound

pure substance that can be decomposed into two or more elements

diatomic elements

Elements that exist in pairs (two of the same atom) naturally: hydrogen, nitrogen, oxygen, fluorine, chlorine, iodine and bromine

element

substance that is composed of a single type of atom; a substance that cannot be decomposed by a chemical change

gas

state in which matter has neither definite volume nor shape (volume and shape are dependant of the container)

heterogeneous mixture

combination of substances with a composition that varies from point to point

homogeneous mixture

(also, solution) combination of substances with a composition that is uniform throughout

liquid

state of matter that has a definite volume but indefinite shape (dependant of container)

law of conservation of matter

when matter converts from one type to another or changes form, there is no detectable change in the total amount of matter present

mass

fundamental property indicating amount of matter; not to be confused with weight

matter

anything that occupies space and has mass

mixture

matter that can be separated into its components by physical means (not a compound, but may contain one or more compounds)

molecule

bonded collection of two or more atoms of the same or different elements

plasma

gaseous state of matter containing a large number of electrically charged atoms and/or molecules

pure substance

homogeneous substance that has a constant composition

solid

state of matter that is rigid, has a definite shape and has a fairly constant volume

weight

force that gravity exerts on an object; directly proportional to mass, but not the same thing

1.4: Matter is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

2: Atoms and Elements

2.1: Electrons and electronic structure

2: Atoms and Elements is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

2.1: Electrons and electronic structure

5

Harper College Chemistry Department

Electron Shells and the Bohr Model

Note that there is a connection between the number of protons in an element, the atomic number that distinguishes one element from another and the number of electrons it has. In all electrically neutral atoms, the number of electrons is the same as the number of protons. Thus, each element, at least when electrically neutral, has a characteristic number of electrons equal to its atomic number.

In 1913, Danish scientist Niels Bohr (1885–1962) developed an early model of the atom. The Bohr model shows the atom has a central nucleus containing protons and neutrons, with the electrons in circular orbits at specific distances from the nucleus, [Figure 1](#). These orbits form electron shells or energy levels, which are a way of visualizing the number of electrons in the outermost shell. These energy levels are designated by a number and the symbol “n.” For example, 1n represents the first energy level located closest to the nucleus.

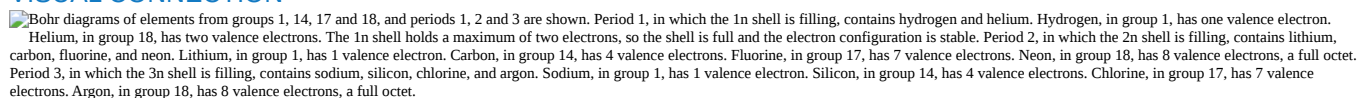
 Three concentric circles around the nucleus of a hydrogen atom represent principal shells. These are named 1 n, 2 n, and 3 n in order of increasing distance from the nucleus. An electron orbits in the shell closest to the nucleus, 1 n.

In the Bohr model, electrons exist within principal shells (n). An electron normally exists in the lowest energy shell available, which is the one closest to the nucleus.

Electrons fill shells in a consistent order: they first fill the shells closest to the nucleus, then they continue to fill shells of increasing energy further from the nucleus. The electrons of the outermost energy level determine the atom’s energetic stability and its tendency to form chemical bonds with other atoms to form molecules.

Under standard conditions, atoms fill the inner shells first, often resulting in a variable number of electrons in the outermost shell. The innermost shell has a maximum of two electrons but the next two electron shells can each have a maximum of eight electrons. This is known as the octet rule, which states, with the exception of the innermost shell, that atoms are more stable energetically when they have eight electrons in their valence shell, the outermost electron shell. [Figure 2](#) shows examples of some neutral atoms and their electron configurations. Notice that helium has a complete outer electron shell, with two electrons filling its first and only shell. Similarly, neon has a complete outer 2n shell containing eight electrons. In contrast, chlorine and sodium have seven and one, respectively, in their outer shells, however, by following the octet rule they would be more energetically stable having eight.

VISUAL CONNECTION

 Bohr diagrams of elements from groups 1, 14, 17 and 18, and periods 1, 2 and 3 are shown. Period 1, in which the 1n shell is filling, contains hydrogen and helium. Hydrogen, in group 1, has one valence electron. Helium, in group 18, has two valence electrons. The 1n shell holds a maximum of two electrons, so the shell is full and the electron configuration is stable. Period 2, in which the 2n shell is filling, contains lithium, carbon, fluorine, and neon. Lithium, in group 1, has 1 valence electron. Carbon, in group 14, has 4 valence electrons. Fluorine, in group 17, has 7 valence electrons. Neon, in group 18, has 8 valence electrons, a full octet. Period 3, in which the 3n shell is filling, contains sodium, silicon, chlorine, and argon. Sodium, in group 1, has 1 valence electron. Silicon, in group 14, has 4 valence electrons. Chlorine, in group 17, has 7 valence electrons. Argon, in group 18, has 8 valence electrons, a full octet.

Bohr diagrams indicate how many electrons fill each principal shell. Group 18 elements (helium, neon, and argon) have a full outer, or valence, shell. A full valence shell is the most stable electron configuration. Elements in other groups have partially filled valence shells and gain or lose electrons to achieve a stable electron configuration.

An atom may give, take or share electrons with another atom to achieve a full valence shell, the most stable electron configuration. Looking at this figure, how many electrons do elements in group 1 need to lose in order to achieve a stable electron configuration? 1 electron. How many electrons do elements in groups 14 and 17 need to gain to achieve a stable configuration? 4 and 1 electron respectively.

Understanding that the periodic table’s organization is based on the total number of protons (and electrons) helps us know how electrons distribute themselves among the energy levels. The periodic table is arranged in columns and rows based on the number of electrons and their location. Examine more closely some of the elements in the table’s far right column in [Figure 3](#). The group 18 atoms helium (He), neon (Ne) and argon (Ar) all have filled outer electron shells, making it unnecessary for them to share electrons with other atoms to attain stability. They are highly stable as single atoms because they are non reactive; scientists coin them inert (or noble gases). Compare this to the group 1 elements in the left-hand column. These elements, including hydrogen (H), lithium (Li) and sodium (Na), all have one electron in their outermost shells. That means that they can achieve a stable configuration and a filled outer shell by donating or sharing one electron with another atom. Hydrogen will donate or share its electron to achieve this

configuration, while lithium and sodium will donate their electron to become stable. As a result of losing a negatively charged electron, they become positively charged ions. Group 17 elements, including fluorine and chlorine, have seven electrons in their outermost shells, so they tend to fill this shell with an electron from other atoms or molecules, making them negatively charged ions. Group 14 elements, of which carbon is the most important to living systems, have four electrons in their outer shell allowing them to make several covalent bonds (discussed below) with other atoms. Thus, the periodic table's columns represent the potential shared state of these elements' outer electron shells that is responsible for their similar chemical characteristics.

 The periodic table consists of eighteen groups and seven periods. Each element has its own square. Within each square is the following information; the atomic number, the symbol, the relative atomic mass, and the name. For example, hydrogen's atomic number is 1, symbol is the letter H; relative atomic mass is 1.01, and name is hydrogen. Two additional rows of elements, known as the lanthanides and actinides, are placed beneath the main table. The lanthanides include elements 57 through 71 and belong in period seven between groups three and four. The actinides include elements 89 through 98 and belong in period eight between the same groups. These elements are placed separately to make the table more compact. For each element, the name, atomic symbol, atomic number, and atomic mass are provided. The atomic number is a whole number that represents the number of protons. The atomic mass, which is the average mass of different isotopes, is estimated to two decimal places. The elements are divided into three categories: metals, nonmetals and metalloids. These form a diagonal line from period two, group thirteen to period seven, group sixteen. All elements to the left of the metalloids are metals, and all elements to the right are nonmetals.

The periodic table shows each element's atomic mass and atomic number. The atomic number appears above the symbol for the element and the approximate atomic mass appears below it.

octet rule

atoms are more stable energetically when they have eight electrons in their valence shell, the outermost electron shell, with the exception of the innermost shell

valence electrons

Electrons found in the outermost energy level and responsible for chemical bonding

2.1: Electrons and electronic structure is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

3: Ionic and Covalent Compounds

[3.1: Ionic and Covalent Compounds](#)

[3.2: Nomenclature](#)

3: Ionic and Covalent Compounds is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

3.1: Ionic and Covalent Compounds

6

Harper College Chemistry Department

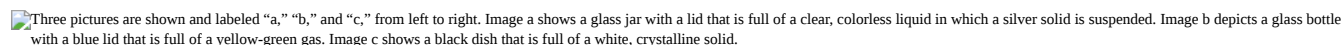
[latexpage]

Ionic Compounds

As you have learned, ions are atoms or molecules bearing an electrical charge. A cation (a positive ion) forms when a neutral atom loses one or more electrons from its valence shell, and an anion (a negative ion) forms when a neutral atom gains one or more electrons in its valence shell.

Compounds composed of ions are called ionic compounds (or salts). The ions are held together by ionic bonds: electrostatic forces of attraction between oppositely charged cations and anions. The properties of ionic compounds shed some light on the nature of ionic bonds. Ionic solids exhibit a crystalline structure and tend to be rigid and brittle. They also tend to have high melting and boiling points, which suggests that ionic bonds are very strong. Ionic solids are also poor conductors of electricity for the same reason—the strength of ionic bonds prevents ions from moving freely in the solid state. Most ionic solids, however, dissolve readily in water. Once dissolved or melted (two different processes), ionic compounds are excellent conductors of electricity and heat because the ions can move about freely.

Neutral atoms and their associated ions have very different physical and chemical properties. Sodium *atoms* form sodium metal, a soft, silvery-white metal that burns vigorously in air and reacts explosively with water. Chlorine *atoms* form chlorine gas, Cl_2 , a yellow-green gas that is extremely corrosive to most metals and very poisonous to animals and plants. The vigorous reaction between the elements sodium and chlorine forms the white, crystalline compound sodium chloride, common table salt, which contains sodium *cations* and chloride *anions*, [Figure 1](#). The compound composed of these ions exhibits properties entirely different from the properties of the elements sodium and chlorine. Chlorine is poisonous, but sodium chloride is essential to life; sodium atoms react vigorously with water, but sodium chloride simply dissolves in water.

 Three pictures are shown and labeled “a,” “b,” and “c,” from left to right. Image a shows a glass jar with a lid that is full of a clear, colorless liquid in which a silver solid is suspended. Image b depicts a glass bottle with a blue lid that is full of a yellow-green gas. Image c shows a black dish that is full of a white, crystalline solid.

(a) Sodium is a soft metal that must be stored in mineral oil to prevent reacting with air or water. (b) Chlorine is a pale yellow-green poisonous gas. (c) When combined, they form white crystals of sodium chloride (table salt). (credit a: modification of work by “Jurii”/Wikimedia Commons)

The Formation of Ionic Compounds

Binary ionic compounds are composed of just two elements: a metal (which forms the cations) and a nonmetal (which forms the anions). For example, NaCl is a binary ionic compound. We can think about the formation of such compounds in terms of the periodic properties of the elements. Many metallic elements have relatively low ionization potentials and lose electrons easily. These elements lie to the left in a period or near the bottom of a group on the periodic table. Nonmetal atoms have relatively high electron affinities and thus readily gain electrons lost by metal atoms, thereby filling their valence shells. Nonmetallic elements are found in the upper-right corner of the periodic table.

As all substances must be electrically neutral, the total number of positive charges on the cations of an ionic compound must equal the total number of negative charges on its anions. The formula of an ionic compound represents the simplest ratio of the numbers of ions necessary to give identical numbers of positive and negative charges. For example, the formula for aluminum oxide, Al_2O_3 , indicates that this ionic compound contains two aluminum cations, Al^{3+} , for every three oxide anions, O^{2-} [thus, $(2 \times +3) + (3 \times -2) = 0$].

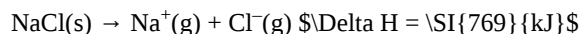
It is important to note, however, that the formula for an ionic compound does *not* represent the physical arrangement of its ions. It is incorrect to refer to a sodium chloride (NaCl) as a “molecule” because there is not a single ionic bond, per se, between any specific pair of sodium and chloride ions. The attractive forces between ions are isotropic—the same in all directions—meaning that any particular ion is equally attracted to all of the nearby ions of opposite charge. This results in the ions arranging themselves into a tightly bound, three-dimensional lattice structure. Sodium chloride, for example, consists of a regular arrangement of equal numbers of Na^+ cations and Cl^- anions ([Figure 2](#)).

Two diagrams are shown and labeled “a” and “b.” Diagram a shows a cube made up of twenty-seven alternating purple and green spheres. The purple spheres are smaller than the green spheres. Diagram b shows the same spheres, but this time, they are spread out and connected in three dimensions by white rods. The purple spheres are labeled “N superscript positive sign” while the green are labeled “Cl superscript negative sign.”

The atoms in sodium chloride (common table salt) are arranged to (a) maximize opposite charges interacting. The smaller spheres represent sodium ions, the larger ones represent chloride ions. In the expanded view (b), the geometry can be seen more clearly.

Note that each ion is “bonded” to all of the surrounding ions—six in this case.

The strong electrostatic attraction between Na^+ and Cl^- ions holds them tightly together in solid NaCl. It requires 769 kJ of energy to dissociate one mole of solid NaCl into separate gaseous Na^+ and Cl^- ions:



Covalent Compounds

Ionic bonding results from the electrostatic attraction of oppositely charged ions that are typically produced by the transfer of electrons between metallic and nonmetallic atoms. A different type of bonding results from the mutual attraction of atoms for a “shared” pair of electrons. Such bonds are called covalent bonds. Covalent bonds are formed between two atoms when both have similar tendencies to attract electrons to themselves (i.e., when both atoms have identical or fairly similar ionization energies and electron affinities). For example, two hydrogen atoms bond covalently to form an H_2 molecule; each hydrogen atom in the H_2 molecule has two electrons stabilizing it, giving each atom the same number of valence electrons as the noble gas He.

Compounds that contain covalent bonds exhibit different physical properties than ionic compounds. Because the attraction between molecules, which are electrically neutral, is weaker than that between electrically charged ions, covalent compounds generally have much lower melting and boiling points than ionic compounds. In fact, many covalent compounds are liquids or gases at room temperature and, in their solid states, they are typically much softer than ionic solids. Furthermore, whereas ionic compounds are good conductors of electricity when dissolved in water, most covalent compounds are insoluble in water; since they are electrically neutral, they are poor conductors of electricity in any state.

Key Concepts and Summary

Atoms gain or lose electrons to form ions with particularly stable electron configurations. The charges of cations formed by the representative metals may be determined readily because, with few exceptions, the electronic structures of these ions have either a noble gas configuration or a completely filled electron shell. The charges of anions formed by the nonmetals may also be readily determined because these ions form when nonmetal atoms gain enough electrons to fill their valence shells.

Chemistry End of Chapter Exercises

- Does a cation gain protons to form a positive charge or does it lose electrons?
- Iron(III) sulfate $[\text{Fe}_2(\text{SO}_4)_3]$ is composed of Fe^{3+} and SO_4^{2-} ions. Explain why a sample of iron(III) sulfate is uncharged.
- Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: P, I, Mg, Cl, In, Cs, O, Pb, Co?
- Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: Br, Ca, Na, N, F, Al, Sn, S, Cd?
- Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:
 - P
 - Mg
 - Al
 - O
 - Cl
 - Cs
- Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:
 - I
 - Sr

- (c) K
- (d) N
- (e) S
- (f) In

7. Using the periodic table, predict whether the following chlorides are ionic or covalent: KCl, NCl₃, ICl, MgCl₂, PCl₅, and CCl₄.
8. Using the periodic table, predict whether the following chlorides are ionic or covalent: SiCl₄, PCl₃, CaCl₂, CsCl, CuCl₂, and CrCl₃.
9. For each of the following compounds, state whether it is ionic or covalent. If it is ionic, write the symbols for the ions involved:

- (a) NF₃
- (b) BaO
- (c) (NH₄)₂CO₃
- (d) Sr(H₂PO₄)₂
- (e) IBr
- (f) Na₂O

10. For each of the following compounds, state whether it is ionic or covalent, and if it is ionic, write the symbols for the ions involved:

- (a) KClO₄
- (b) Mg(C₂H₃O₂)₂
- (c) H₂S
- (d) Ag₂S
- (e) N₂Cl₄
- (f) Co(NO₃)₂

11. For each of the following pairs of ions, write the formula of the compound they will form:

- (a) Ca²⁺, S²⁻
- (b) ⁺NH₄, ⁺NH₄, ⁻SO₄, ⁻SO₄
- (c) Al³⁺, Br⁻
- (d) Na⁺, ⁻HPO₄, ⁻HPO₄
- (e) Mg²⁺, ⁻PO₄, ⁻PO₄

12. For each of the following pairs of ions, write the formula of the compound they will form:

- (a) K⁺, O²⁻
- (b) ⁺NH₄, ⁺NH₄, ⁻PO₄, ⁻PO₄
- (c) Al³⁺, O²⁻
- (d) Na⁺, ⁻CO₃, ⁻CO₃
- (e) Ba²⁺, ⁻PO₄

13. From the labels of several commercial products, prepare a list of six ionic compounds in the products. For each compound, write the formula. (You may need to look up some formulas in a suitable reference.)

Glossary

covalent bond

attractive force between the nuclei of a molecule's atoms and pairs of electrons between the atoms

covalent compound

(also, molecular compound) composed of molecules formed by atoms of two or more different elements

inert pair effect

tendency of heavy atoms to form ions in which their valence s electrons are not lost

ionic bond

electrostatic forces of attraction between the oppositely charged ions of an ionic compound

ionic compound

compound composed of cations and anions combined in ratios, yielding an electrically neutral substance

molecular compound

(also, covalent compound) composed of molecules formed by atoms of two or more different elements

monatomic ion

ion composed of a single atom

oxyanion

polyatomic anion composed of a central atom bonded to oxygen atoms

polyatomic ion

ion composed of more than one atom

3.1: Ionic and Covalent Compounds is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

3.2: Nomenclature

7

Harper College Chemistry Department

Nomenclature, a collection of rules for naming things, is important in science and in many other situations. This module describes an approach that is used to name simple ionic and molecular compounds, such as NaCl , CaCO_3 , and N_2O_4 . The simplest of these are binary compounds, those containing only two elements, but we will also consider how to name ionic compounds containing polyatomic ions and acids (subsequent chapters in this text will focus on these compounds in great detail). We will limit our attention here to inorganic compounds, compounds that are composed principally of elements other than carbon, and will follow the nomenclature guidelines proposed by the International Union of Pure and Applied Chemistry (IUPAC). The rules for organic compounds, in which carbon is the principle element, will be treated in a later chapter on organic chemistry.

Naming Ionic Compounds

To name an inorganic compound, we need to consider the answers to several questions. First, is the compound ionic or molecular? If the compound is ionic, does the metal form ions of only one type (fixed charge) or more than one type (variable charge)? Are the ions monatomic or polyatomic? If the compound is molecular, does it contain hydrogen? If so, does it also contain oxygen? From the answers we derive, we place the compound in an appropriate category and then name it accordingly.

Compounds Containing Only Monatomic Ions

The name of a binary compound containing monatomic ions consists of the name of the cation (the name of the metal) followed by the name of the anion (the name of the nonmetallic element with its ending replaced by the suffix *-ide*). Some examples are given in [Table 1](#).

Names of Some Ionic Compounds	
NaCl , sodium chloride	Na_2O , sodium oxide
KBr , potassium bromide	CdS , cadmium sulfide
CaI_2 , calcium iodide	Mg_3N_2 , magnesium nitride
CsF , cesium fluoride	Ca_3P_2 , calcium phosphide
LiCl , lithium chloride	Al_4C_3 , aluminum carbide

Compounds Containing Polyatomic Ions

Compounds containing polyatomic ions (Table 2a) are named similarly to those containing only monatomic ions, i.e., by naming first the cation and then the anion. Examples are shown in [Table 2](#).

Table 2a: Some Common Polyatomic Ions		
Ammonium NH_4^+	Hypochlorite ClO^-	Nitrite NO_2^-
Acetate $\text{C}_2\text{H}_3\text{O}_2^-$	Hydroxide OH^-	Phosphate PO_4^{3-}
Bicarbonate HCO_3^- Hydrogen carbonate	Hydrogen phosphate HPO_4^{2-}	Phosphite PO_3^{3-}
Carbonate CO_3^{2-}	Hydrogen sulfate HSO_4^-	Sulfate SO_4^{2-}
Chlorate ClO_3^-	Nitrate NO_3^-	Sulfite SO_3^{2-}

Names of Some Polyatomic Ionic Compounds

$\text{KC}_2\text{H}_3\text{O}_2$, potassium acetate	NH_4Cl , ammonium chloride
NaHCO_3 , sodium bicarbonate	CaSO_4 , calcium sulfate
$\text{Al}_2(\text{CO}_3)_3$, aluminum carbonate	$\text{Mg}_3(\text{PO}_4)_2$, magnesium phosphate

IONIC COMPOUNDS IN YOUR CABINETS

Every day you encounter and use a large number of ionic compounds. Some of these compounds, where they are found, and what they are used for are listed in [Table 3](#). Look at the label or ingredients list on the various products that you use during the next few days, and see if you run into any of those in this table, or find other ionic compounds that you could now name or write as a formula.

Everyday Ionic Compounds

Ionic Compound	Use
NaCl , sodium chloride	ordinary table salt
KI , potassium iodide	added to “iodized” salt for thyroid health
NaF , sodium fluoride	ingredient in toothpaste
NaHCO_3 , sodium bicarbonate	baking soda; used in cooking (and as antacid)
Na_2CO_3 , sodium carbonate	washing soda; used in cleaning agents
NaOCl , sodium hypochlorite	active ingredient in household bleach
CaCO_3 calcium carbonate	ingredient in antacids
$\text{Mg}(\text{OH})_2$, magnesium hydroxide	ingredient in antacids
$\text{Al}(\text{OH})_3$, aluminum hydroxide	ingredient in antacids
NaOH , sodium hydroxide	lye; used as drain cleaner
K_3PO_4 , potassium phosphate	food additive (many purposes)
MgSO_4 , magnesium sulfate	added to purified water
Na_2HPO_4 , sodium hydrogen phosphate	anti-caking agent; used in powdered products
Na_2SO_3 , sodium sulfite	preservative

Compounds Containing a Metal Ion with a Variable Charge

Most of the transition metals can form two or more cations with different charges. Compounds of these metals with nonmetals are named with the same method as compounds in the first category, except the charge of the metal ion is specified by a Roman numeral in parentheses after the name of the metal. The charge of the metal ion is determined from the formula of the compound and the charge of the anion. For example, consider binary ionic compounds of iron and chlorine. Iron typically exhibits a charge of either 2+ or 3+ (see [link](#)), and the two corresponding compound formulas are FeCl_2 and FeCl_3 . The simplest name, “iron chloride,” will, in this case, be ambiguous, as it does not distinguish between these two compounds. In cases like this, the charge of the metal ion is included as a Roman numeral in parentheses immediately following the metal name. These two compounds are then unambiguously named iron(II) chloride and iron(III) chloride, respectively. Other examples are provided in [Table 4](#).

Names of Some Transition Metal Ionic Compounds

Transition Metal Ionic Compound	Name
FeCl_3	iron(III) chloride
Hg_2O	mercury(I) oxide
HgO	mercury(II) oxide
$\text{Cu}_3(\text{PO}_4)_2$	copper(II) phosphate

Out-of-date nomenclature used the suffixes *-ic* and *-ous* to designate metals with higher and lower charges, respectively: Iron(III) chloride, FeCl_3 , was previously called ferric chloride, and iron(II) chloride, FeCl_2 , was known as ferrous chloride. Though this naming convention has been largely abandoned by the scientific community, it remains in use by some segments of industry. For example, you may see the words *stannous fluoride* on a tube of toothpaste. This represents the formula SnF_2 , which is more properly named tin(II) fluoride. The other fluoride of tin is SnF_4 , which was previously called stannic fluoride but is now named tin(IV) fluoride.

Naming Ionic Compounds: Name the following ionic compounds, which contain a metal that can have more than one positive charge:

- (a) Fe_2S_3
- (b) CuSe
- (c) GaN
- (d) CrCl_3
- (e) $\text{Ti}_2(\text{SO}_4)_3$

Solution: The anions in these compounds have a fixed negative charge (S^{2-} , Se^{2-} , N^{3-} , Cl^- , and SO_4^{2-}), and the compounds must be neutral. Because the total number of positive charges in each compound must equal the total number of negative charges, the positive ions must be Fe^{3+} , Cu^{2+} , Ga^{3+} , Cr^{3+} , and Ti^{3+} . These charges are used in the names of the metal ions:

- (a) iron(III) sulfide
- (b) copper(II) selenide
- (c) gallium(III) nitride
- (d) chromium(III) chloride
- (e) titanium(III) sulfate

Check Your Learning

Write the formulas of the following ionic compounds:

- (a) chromium(III) phosphide
- (b) mercury(II) sulfide
- (c) manganese(II) phosphate
- (d) copper(I) oxide
- (e) chromium(VI) fluoride

Answer:

(a) CrP ; (b) HgS ; (c) $\text{Mn}_3(\text{PO}_4)_2$; (d) Cu_2O ; (e) CrF_6

Molecular (Covalent) Compounds

The bonding characteristics of molecular compounds are different from ionic compounds and they are named using a different system as well. The charges of cations and anions dictate their ratios in ionic compounds, so specifying the names of the ions provides sufficient information to determine chemical formulas. However, because covalent bonding allows for significant variation in the combination ratios of the atoms in a molecule, the names for molecular compounds must explicitly identify these ratios. Remember ions are only present in ionic compounds.

Compounds Composed of Two Nonmetal Elements

When two nonmetallic elements form a molecular compound, several ratio combinations are often possible. For example, carbon and oxygen can form the compounds CO and CO₂. Since these are different substances with different properties, they cannot both have the same name (they cannot both be called carbon oxide). To deal with this situation, we use a naming method that is somewhat similar to that used for ionic compounds, but with added prefixes to specify the numbers of atoms of each element. The name of the more metallic element (the one farther to the left and/or bottom of the periodic table) is first, followed by the name of the more nonmetallic element (the one farther to the right and/or top) with its ending changed to the suffix *-ide*. The numbers of atoms of each element are designated by the Greek prefixes shown in Table 5.

Number	Prefix	Number	Prefix
1 (sometimes omitted)	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

When only one atom of the first element is present, the prefix *mono-* is usually deleted from that part. Thus, CO is named carbon monoxide and CO₂ is called carbon dioxide. When two vowels are adjacent, the “a” in the Greek prefix is usually dropped. Some other examples are shown in Table 6.

Compound	Name	Compound	Name
SO ₂	sulfur dioxide	BCl ₃	boron trichloride
SO ₃	sulfur trioxide	SF ₆	sulfur hexafluoride
NO ₂	nitrogen dioxide	PF ₅	phosphorus pentafluoride
N ₂ O ₄	dinitrogen tetroxide	P ₄ O ₁₀	tetraphosphorus decaoxide
N ₂ O ₅	dinitrogen pentoxide	IF ₇	iodine heptafluoride

There are a few common names that you will encounter as you continue your study of chemistry. For example, although NO is often called nitric oxide, its proper name is nitrogen monoxide. Similarly, N₂O is known as nitrous oxide even though our rules would specify the name dinitrogen monoxide. (And H₂O is usually called water, not dihydrogen monoxide.)

Naming Covalent Compounds: Name the following covalent compounds:

- (a) SF₆
- (b) N₂O₃
- (c) Cl₂O₇

(d) P_4O_6

Solution: Because these compounds consist solely of nonmetals, we use prefixes to designate the number of atoms of each element:

- (a) sulfur hexafluoride
- (b) dinitrogen trioxide
- (c) dichlorine heptoxide
- (d) tetraphosphorus hexoxide

Check Your Learning Write the formulas for the following compounds:

- (a) phosphorus pentachloride
- (b) dinitrogen monoxide
- (c) iodine heptafluoride
- (d) carbon tetrachloride

Answer:

(a) PCl_5 ; (b) N_2O ; (c) IF_7 ; (d) CCl_4

The following [website](#) provides practice with naming chemical compounds and writing chemical formulas. You can choose binary, polyatomic, and variable charge ionic compounds, as well as molecular compounds.

Binary Acids

Some compounds containing hydrogen are members of an important class of substances known as acids. The chemistry of these compounds is explored in more detail in later chapters of this text, but for now, it will suffice to note that many acids release hydrogen ions, H^+ , when dissolved in water. To denote this distinct chemical property, a mixture of water with an acid is given a name derived from the compound's name. If the compound is a binary acid (comprised of hydrogen and one other nonmetallic element):

1. The word "hydrogen" is changed to the prefix *hydro-*
2. The other nonmetallic element name is modified by adding the suffix *-ic*
3. The word "acid" is added as a second word

For example, when the gas HCl (hydrogen chloride) is dissolved in water, the solution is called *hydrochloric acid*. Several other examples of this nomenclature are shown in [Table 7](#).

Table 7: Names of Some Simple Acids

Name of Gas	Name of Acid
$\text{HF}(g)$, hydrogen fluoride	$\text{HF}(aq)$, hydrofluoric acid
$\text{HCl}(g)$, hydrogen chloride	$\text{HCl}(aq)$, hydrochloric acid
$\text{HBr}(g)$, hydrogen bromide	$\text{HBr}(aq)$, hydrobromic acid
$\text{HI}(g)$, hydrogen iodide	$\text{HI}(aq)$, hydroiodic acid
$\text{H}_2\text{S}(g)$, hydrogen sulfide	$\text{H}_2\text{S}(aq)$, hydrosulfuric acid

Oxyacids

Many compounds containing three or more elements (such as organic compounds or coordination compounds) are subject to specialized nomenclature rules that you will learn later. However, we will briefly discuss the important compounds known as oxyacids, compounds that contain hydrogen, oxygen, at least one other element and are bonded in such a way as to impart acidic

properties to the compound (you will learn the details of this in a later chapter). Typical oxyacids consist of hydrogen combined with a polyatomic, oxygen-containing ion. To name oxyacids:

1. Omit “hydrogen”
2. Start with the root name of the anion
3. Replace *-ate* with *-ic*, or *-ite* with *-ous*
4. Add “acid”

For example, consider H_2CO_3 (which you might be tempted to call “hydrogen carbonate”). To name this correctly, “hydrogen” is omitted; the *-ate* of carbonate is replaced with *-ic*; and acid is added—so its name is carbonic acid. Other examples are given in [Table 8](#). There are some exceptions to the general naming method (e.g., H_2SO_4 is called sulfuric acid, not sulfic acid, and H_2SO_3 is sulfurous, not sulfous, acid).

Table 8: Names of Common Oxyacids

Formula	Anion Name	Acid Name
$\text{HC}_2\text{H}_3\text{O}_2$	acetate	acetic acid
HNO_3	nitrate	nitric acid
HNO_2	nitrite	nitrous acid
HClO_4	perchlorate	perchloric acid
H_2CO_3	carbonate	carbonic acid
H_2SO_4	sulfate	sulfuric acid
H_2SO_3	sulfite	sulfurous acid
H_3PO_4	phosphate	phosphoric acid

[In case you are wondering](#)

ERIN BROKOVICH AND CHROMIUM CONTAMINATION

In the early 1990s, legal file clerk Erin Brockovich ([Figure 1](#)) discovered a high rate of serious illnesses in the small town of Hinckley, California. Her investigation eventually linked the illnesses to groundwater contaminated by Cr(VI) used by Pacific Gas & Electric (PG&E) to fight corrosion in a nearby natural gas pipeline. As dramatized in the film *Erin Brockovich* (for which Julia Roberts won an Oscar), Erin and lawyer Edward Masry sued PG&E for contaminating the water near Hinckley in 1993. The settlement they won in 1996—\$333 million—was the largest amount ever awarded for a direct-action lawsuit in the US at that time.

Figure A shows a photo of Erin Brockovich. Figure B shows a 3-D ball-and-stick model of chromate. Chromate has a chromium atom at its center that forms bonds with four oxygen atoms each. Two of the oxygen atoms form single bonds with the chromium atom while the other two form double bonds each. The structure of dichromate consists of two chromate ions that are bonded and share one of their oxygen atoms to which each chromate atom has a single bond.

(a) Erin Brockovich found that Cr(VI), used by PG&E, had contaminated the Hinckley, California, water supply. (b) The Cr(VI) ion is often present in water as the polyatomic ions chromate, CrO_4^{2-} (left), and dichromate, $\text{Cr}_2\text{O}_7^{2-}$ (right).

Chromium compounds are widely used in industry, such as for chrome plating, in dye-making, as preservatives, and to prevent corrosion in cooling tower water, as occurred near Hinckley. In the environment, chromium exists primarily in either the Cr(III) or Cr(VI) forms. Cr(III), an ingredient of many vitamin and nutritional supplements, forms compounds that are not very soluble in water, and it has low toxicity. But Cr(VI) is much more toxic and forms compounds that are reasonably soluble in water. Exposure to small amounts of Cr(VI) can lead to damage of the respiratory, gastrointestinal, and immune systems, as well as the kidneys, liver, blood, and skin.

Despite cleanup efforts, Cr(VI) groundwater contamination remains a problem in Hinckley and other locations across the globe. A 2010 study by the Environmental Working Group found that of 35 US cities tested, 31 had higher levels of Cr(VI) in their tap water than the public health goal of 0.02 parts per billion set by the California Environmental Protection Agency.

Key Concepts and Summary

Chemists use nomenclature rules to clearly name compounds. Ionic and molecular compounds are named using somewhat-different methods. Binary ionic compounds typically consist of a metal and a nonmetal. The name of the metal is written first, followed by the

name of the nonmetal with its ending changed to *-ide*. For example, K_2O is called potassium oxide. If the metal can form ions with different charges, a Roman numeral in parentheses follows the name of the metal to specify its charge. Thus, $FeCl_2$ is iron(II) chloride and $FeCl_3$ is iron(III) chloride. Some compounds contain polyatomic ions; the names of common polyatomic ions should be memorized. Molecular compounds can form compounds with different ratios of their elements, so prefixes are used to specify the numbers of atoms of each element in a molecule of the compound. Examples include SF_6 , sulfur hexafluoride, and N_2O_4 , dinitrogen tetroxide. Acids are an important class of compounds containing hydrogen and having special nomenclature rules. Binary acids are named using the prefix *hydro-*, changing the *-ide* suffix to *-ic*, and adding “*acid*,” HCl is hydrochloric acid. Oxyacids are named by changing the ending of the anion (*-ate* to *-ic* and *-ite* to *-ous*) and adding “*acid*,” H_2CO_3 is carbonic acid.

Chemistry End of Chapter Exercises

1. Name the following compounds:

- (a) $CsCl$
- (b) BaO
- (c) K_2S
- (d) $BeCl_2$
- (e) HBr
- (f) AlF_3

2. Name the following compounds:

- (a) NaF
- (b) Rb_2O
- (c) BCl_3
- (d) H_2Se
- (e) P_4O_6
- (f) ICl_3

3. Write the formulas of the following compounds:

- (a) rubidium bromide
- (b) magnesium selenide
- (c) sodium oxide
- (d) calcium chloride
- (e) hydrogen fluoride
- (f) gallium phosphide
- (g) aluminum bromide
- (h) ammonium sulfate

4. Write the formulas of the following compounds:

- (a) lithium carbonate
- (b) sodium perchlorate
- (c) barium hydroxide
- (d) ammonium carbonate
- (e) sulfuric acid

- (f) calcium acetate
- (g) magnesium phosphate
- (h) sodium sulfite

5. Write the formulas of the following compounds:

- (a) chlorine dioxide
- (b) dinitrogen tetraoxide
- (c) potassium phosphide
- (d) silver(I) sulfide
- (e) aluminum nitride
- (f) silicon dioxide

6. Write the formulas of the following compounds:

- (a) barium chloride
- (b) magnesium nitride
- (c) sulfur dioxide
- (d) nitrogen trichloride
- (e) dinitrogen trioxide
- (f) tin(IV) chloride

7. Each of the following compounds contains a metal that can exhibit more than one ionic charge. Name these compounds:

- (a) Cr_2O_3
- (b) FeCl_2
- (c) CrO_3
- (d) TiCl_4
- (e) CoO
- (f) MoS_2

8. Each of the following compounds contains a metal that can exhibit more than one ionic charge. Name these compounds:

- (a) NiCO_3
- (b) MoO_3
- (c) $\text{Co}(\text{NO}_3)_2$
- (d) V_2O_5
- (e) MnO_2
- (f) Fe_2O_3

9. The following ionic compounds are found in common household products. Write the formulas for each compound:

- (a) potassium phosphate
- (b) copper(II) sulfate
- (c) calcium chloride
- (d) titanium dioxide
- (e) ammonium nitrate

(f) sodium bisulfate (the common name for sodium hydrogen sulfate)

10. The following ionic compounds are found in common household products. Name each of the compounds:

(a) $\text{Ca}(\text{H}_2\text{PO}_4)_2$

(b) FeSO_4

(c) CaCO_3

(d) MgO

(e) NaNO_2

(f) KI

11. What are the IUPAC names of the following compounds?

(a) manganese dioxide

(b) mercurous chloride (Hg_2Cl_2)

(c) ferric nitrate [$\text{Fe}(\text{NO}_3)_3$]

(d) titanium tetrachloride

(e) cupric bromide (CuBr_2)

Glossary

binary acid

compound that contains hydrogen and one other element, bonded in a way that imparts acidic properties to the compound (ability to release H^+ ions) when dissolved in water

binary compound

compound containing two different elements

nomenclature

system of rules for naming objects of interest

oxyacid

compound that contains hydrogen, oxygen, and one other element, bonded in a way that imparts acidic properties to the compound (ability to release H^+ ions when dissolved in water)

3.2: Nomenclature is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

4: Chemical Bonding

4.1: Lewis structures

4.2: Covalent bonding and electronegativity

4.3: Valence Bond Theory

4.4: Molecular Structure and Polarity

4.5: Atomic Structure

4: Chemical Bonding is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

4.1: Lewis structures

8

Harper College Chemistry Department

Page by: OpenStax Chemistry: Atom First

▼ Summary

By the end of this section, you will be able to:

- Write Lewis symbols for neutral atoms and ions
- Draw Lewis structures depicting the bonding in simple molecules

We have discussed the various types of bonds that form between atoms. These bonds involve the sharing or transfer of valence shell electrons between atoms. In this section, we will explore the typical method for depicting valence shell electrons and chemical bonds, namely Lewis symbols and Lewis structures.

Lewis Symbols

We use Lewis symbols to describe valence electron configurations of atoms and monatomic ions. A Lewis symbol consists of an elemental symbol surrounded by one dot for each of its valence electrons:

 A Lewis structure of calcium is shown. A lone pair of electrons are shown to the right of the symbol.

Figure 1 shows the Lewis symbols for the elements of the third period of the periodic table.

Atoms	Lewis Symbol
sodium	Na •
magnesium	•Mg•
aluminum	•Al•
silicon	•Si•
phosphorus	•P•
sulfur	:S•
chlorine	:Cl•
argon	:Ar:

Lewis symbols illustrating the number of valence electrons for each element in the third period of the periodic table.

Lewis symbols can also be used to illustrate the formation of cations from atoms, as shown here for sodium and calcium:

 Two diagrams are shown. The left diagram shows a Lewis dot structure of sodium with one dot, then a right-facing arrow leading to a sodium symbol with a superscripted plus sign, a plus sign, and the letter "e" with a superscripted negative sign. The terms below this diagram read "Sodium atom" and "Sodium cation." The right diagram shows a Lewis dot structure of calcium with two dots, then a right-facing arrow leading to a calcium symbol with a superscripted two and a plus sign, a plus sign, and the value "2e" with a superscripted negative sign. The terms below this diagram read "Calcium atom" and "Calcium cation."

Likewise, they can be used to show the formation of anions from atoms, as shown here for chlorine and sulfur:

 Two diagrams are shown. The left diagram shows a Lewis dot structure of chlorine with seven dots and the letter "e" with a superscripted negative sign, then a right-facing arrow leading to a chlorine symbol with eight dots and a superscripted negative sign. The terms below this diagram read "Chlorine atom," and "Chlorine anion." The right diagram shows a Lewis dot structure of sulfur with six dots and the symbol "2e" with a superscripted negative sign, then a right-facing arrow leading to a sulfur symbol with eight dots and a superscripted two and negative sign. The terms below this diagram read "Sulfur atom," and "Sulfur anion."

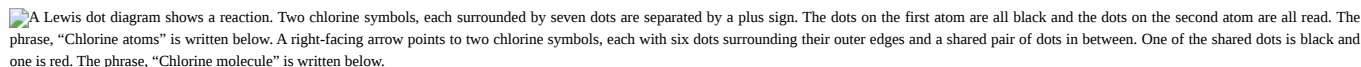
Figure 2 demonstrates the use of Lewis symbols to show the transfer of electrons during the formation of ionic compounds.

 A table is shown with four rows. The header row reads "Metal," "Nonmetal," and "Ionic Compound." The second row shows the Lewis structures of a reaction. A sodium symbol with one dot, a plus sign, and a chlorine symbol with seven dots lie to the left of a right-facing arrow. To the right of the arrow a sodium symbol with a superscripted plus sign is drawn next to a chlorine symbol with eight dots surrounded by brackets with a superscripted negative sign. One of the dots on the Cl atom is red. The terms "sodium atom," "chlorine atom," and "sodium chloride (sodium ion and chloride ion)" are written under the reaction. The third row shows the Lewis structures of a reaction. A magnesium symbol with two red dots, a plus sign, and an oxygen symbol with six dots lie to the left of a right-facing arrow. To the right of the arrow a magnesium symbol with a superscripted two and a plus sign is drawn next to an oxygen symbol with eight dots, two of which are red, surrounded by brackets with a superscripted two and a negative sign. The terms "magnesium atom," "oxygen atom," and "magnesium oxide (magnesium ion and oxide ion)" are written under the reaction. The fourth row shows the Lewis structures of a reaction. A calcium symbol with two red dots, a plus sign, and a fluorine symbol with seven dots lie to the left of a right-facing arrow. To the right of the arrow a calcium symbol with a superscripted two and a plus sign is drawn next to a fluorine symbol with eight dots, one of which is red, surrounded by brackets with a superscripted negative sign and a subscripted two. The terms "calcium atom," "fluorine atoms," and "calcium fluoride (calcium ion and two fluoride ions)" are written under the reaction.

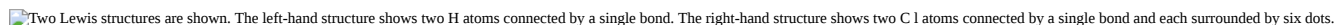
Cations are formed when atoms lose electrons, represented by fewer Lewis dots, whereas anions are formed by atoms gaining electrons. The total number of electrons does not change.

Lewis Structures

We also use Lewis symbols to indicate the formation of covalent bonds, which are shown in Lewis structures, drawings that describe the bonding in molecules and polyatomic ions. For example, when two chlorine atoms form a chlorine molecule, they share one pair of electrons:

A Lewis dot diagram shows a reaction. Two chlorine symbols, each surrounded by seven dots are separated by a plus sign. The dots on the first atom are all black and the dots on the second atom are all red. The phrase, "Chlorine atoms" is written below. A right-facing arrow points to two chlorine symbols, each with six dots surrounding their outer edges and a shared pair of dots in between. One of the shared dots is black and one is red. The phrase, "Chlorine molecule" is written below.

The Lewis structure indicates that each Cl atom has three pairs of electrons that are not used in bonding (called lone pairs) and one shared pair of electrons (written between the atoms). A dash (or line) is sometimes used to indicate a shared pair of electrons:

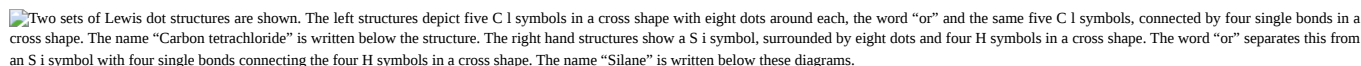
Two Lewis structures are shown. The left-hand structure shows two H atoms connected by a single bond. The right-hand structure shows two Cl atoms connected by a single bond and each surrounded by six dots.

A single shared pair of electrons is called a single bond. Each Cl atom interacts with eight valence electrons: the six in the lone pairs and the two in the single bond.

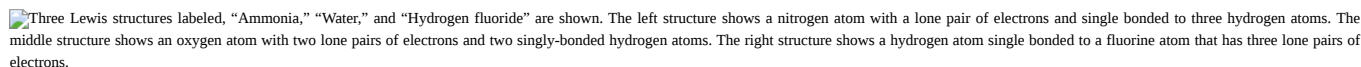
The Octet Rule

The other halogen molecules (F_2 , Br_2 , I_2 and At_2) form bonds like those in the chlorine molecule: one single bond between atoms and three lone pairs of electrons per atom. This allows each halogen atom to have a noble gas electron configuration. The tendency of main group atoms to form enough bonds to obtain eight valence electrons is known as the octet rule.

The number of bonds that an atom can form can often be predicted from the number of electrons needed to reach an octet (eight valence electrons); this is especially true of the nonmetals of the second period of the periodic table (C, N, O, and F). For example, each atom of a group 14 element has four electrons in its outermost shell and therefore requires four more electrons to reach an octet. These four electrons can be gained by forming four covalent bonds, as illustrated here for carbon in CCl_4 (carbon tetrachloride) and silicon in SiH_4 (silane). Because hydrogen only needs two electrons to fill its valence shell, it is an exception to the octet rule. The transition elements and inner transition elements also do not follow the octet rule:

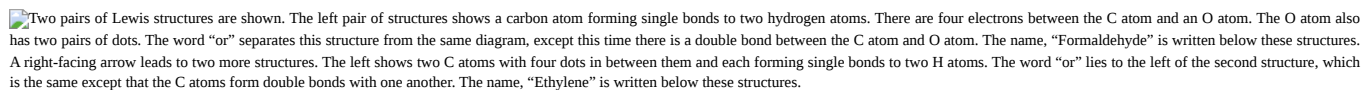
Two sets of Lewis dot structures are shown. The left structures depict five C l symbols in a cross shape with eight dots around each, the word "or" and the same five C l symbols, connected by four single bonds in a cross shape. The name "Carbon tetrachloride" is written below the structure. The right hand structures show a S i symbol, surrounded by eight dots and four H symbols in a cross shape. The word "or" separates this from an S i symbol with four single bonds connecting the four H symbols in a cross shape. The name "Silane" is written below these diagrams.

Group 15 elements such as nitrogen have five valence electrons in the atomic Lewis symbol: one lone pair and three unpaired electrons. To obtain an octet, these atoms form three covalent bonds, as in NH_3 (ammonia). Oxygen and other atoms in group 16 obtain an octet by forming two covalent bonds:

Three Lewis structures labeled, "Ammonia," "Water," and "Hydrogen fluoride" are shown. The left structure shows a nitrogen atom with a lone pair of electrons and single bonded to three hydrogen atoms. The middle structure shows an oxygen atom with two lone pairs of electrons and two singly-bonded hydrogen atoms. The right structure shows a hydrogen atom single bonded to a fluorine atom that has three lone pairs of electrons.

Double and Triple Bonds

As previously mentioned, when a pair of atoms shares one pair of electrons, we call this a single bond. However, a pair of atoms may need to share more than one pair of electrons in order to achieve the requisite octet. A double bond forms when two pairs of electrons are shared between a pair of atoms, as between the carbon and oxygen atoms in CH_2O (formaldehyde) and between the two carbon atoms in C_2H_4 (ethylene):

Two pairs of Lewis structures are shown. The left pair of structures shows a carbon atom forming single bonds to two hydrogen atoms. There are four electrons between the C atom and an O atom. The O atom also has two pairs of dots. The word "or" separates this structure from the same diagram, except this time there is a double bond between the C atom and O atom. The name, "Formaldehyde" is written below these structures. A right-facing arrow leads to two more structures. The left shows two C atoms with four dots in between them and each forming single bonds to two H atoms. The word "or" lies to the left of the second structure, which is the same except that the C atoms form double bonds with one another. The name, "Ethylene" is written below these structures.

A triple bond forms when three electron pairs are shared by a pair of atoms, as in carbon monoxide (CO) and the cyanide ion (CN^-):

Two pairs of Lewis structures are shown and connected by a right-facing arrow. The left pair of structures show a C atom and an O atom with six dots in between them and a lone pair on each. The word "or" and the same structure with a triple bond in between the C atom and O atom also are shown. The name "Carbon monoxide" is written below this structure. The right pair of structures show a C atom and an N atom with six dots in between them and a lone pair on each. The word "or" and the same structure with a triple bond in between the C atom and N atom also are shown. The name "Cyanide ion" is written below this structure.

Writing Lewis Structures with the Octet Rule

For very simple molecules and molecular ions, we can write the Lewis structures by merely pairing up the unpaired electrons on the constituent atoms. See these examples:

Three reactions are shown with Lewis dot diagrams. The first shows a hydrogen with one red dot, a plus sign and a bromine with seven dots, one of which is red, connected by a right-facing arrow to a hydrogen and bromine with a pair of red dots in between them. There are also three lone pairs on the bromine. The second reaction shows a hydrogen with a coefficient of two and one red dot, a plus sign, and a sulfur atom with six dots, two of which are red, connected by a right facing arrow to two hydrogen atoms and one sulfur atom. There are two red dots in between the two hydrogen atoms and the sulfur atom. Both pairs of these dots are red. The sulfur atom also has two lone pairs of dots. The third reaction shows two nitrogen atoms each with five dots, three of which are red, separated by a plus sign, and connected by a right-facing arrow to two nitrogen atoms with six red electron dots in between one another. Each nitrogen atom also has one lone pair of electrons.

For more complicated molecules and molecular ions, it is helpful to follow the step-by-step procedure outlined here:

1. Determine the total number of valence (outer shell) electrons. For cations, subtract one electron for each positive charge. For anions, add one electron for each negative charge.
2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom. (Generally, the least electronegative element should be placed in the center.) Connect each atom to the central atom with a single bond (one electron pair).
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen), completing an octet around each atom; following the octet rule.
4. Place all remaining electrons on the central atom.
5. Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

Let us determine the Lewis structures of SiH_4 , CHO_2^- , NO^+ , and OF_2 as examples in following this procedure:

1. Determine the total number of valence (outer shell) electrons in the molecule or ion. For a molecule, we add the number of valence electrons (VE) on each atom in the molecule:

$$\text{SiH}_4 (\text{Si}: 4 \text{ VE/atom} \times 1 \text{ atom} = 4 \text{ VE}) + (\text{H}: 1 \text{ VE/atom} \times 4 \text{ atoms} = 4 \text{ VE}) = 8 \text{ VE}$$

For a *negative ion*, such as CHO_2^- , we add the number of valence electrons on the atoms to the number of negative charges on the ion (one electron is gained for each single negative charge):

$$\text{CHO}_2^- (\text{C}: 4 \text{ VE/atom} \times 1 \text{ atom} = 4 \text{ VE}) + (\text{H}: 1 \text{ VE/atom} \times 1 \text{ atom} = 1 \text{ VE}) + (\text{O}: 6 \text{ VE/atom} \times 2 \text{ atoms} = 12 \text{ VE}) + (1 \text{ additional electron} = 1 \text{ VE}) = 18 \text{ VE}$$

For a *positive ion*, such as NO^+ , we add the number of valence electrons on the atoms in the ion and then subtract the number of positive charges on the ion (one electron is lost for each single positive charge) from the total number of valence electrons:

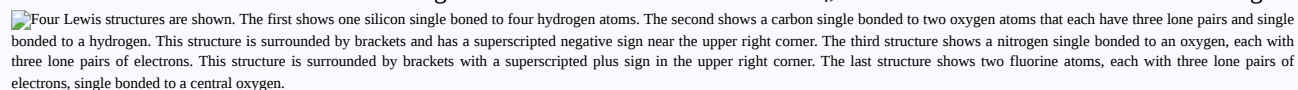
$$\text{NO}^+ (\text{N}: 5 \text{ VE/atom} \times 1 \text{ atom} = 5 \text{ VE}) + (\text{O}: 6 \text{ VE/atom} \times 1 \text{ atom} = 6 \text{ VE}) - 1 \text{ electron (positive charge)} = 10 \text{ VE}$$

Since OF_2 is a neutral molecule, we simply add the number of valence electrons:

$$\text{OF}_2 (\text{O}: 6 \text{ VE/atom} \times 1 \text{ atom} = 6 \text{ VE}) + (\text{F}: 7 \text{ VE/atom} \times 2 \text{ atoms} = 14 \text{ VE}) = 20 \text{ VE}$$

2. When several arrangements of atoms are possible, as for CHO_2^- , the less electronegative carbon atom occupies the central position with the oxygen and hydrogen atoms surrounding it. Other examples include P in POCl_3 , S in SO_2 , and Cl in ClO_4^- .
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen) to complete their valence shells with an octet of electrons.

There are no remaining electrons on SiH_4 , so it is unchanged:

Four Lewis structures are shown. The first shows one silicon single bonded to four hydrogen atoms. The second shows a carbon single bonded to two oxygen atoms that each have three lone pairs and single bonded to a hydrogen. This structure is surrounded by brackets and has a superscripted negative sign near the upper right corner. The third structure shows a nitrogen single bonded to an oxygen, each with three lone pairs of electrons. This structure is surrounded by brackets with a superscripted plus sign in the upper right corner. The last structure shows two fluorine atoms, each with three lone pairs of electrons, single bonded to a central oxygen.

- Place all remaining electrons on the central atom.

For SiH_4 , CHO_2^- , and NO^+ , there are no remaining electrons; we already placed all of the electrons determined in Step 1.

For OF_2 , we had 16 electrons remaining in Step 3, and we placed 12, leaving 4 to be placed on the central atom:

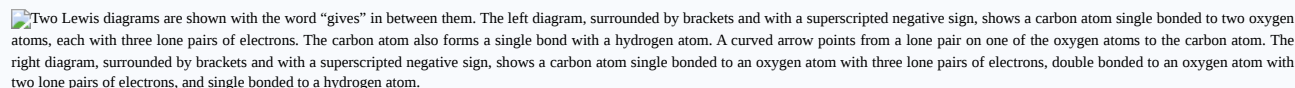
A Lewis structure shows two fluorine atoms, each with three lone pairs of electrons, single bonded to a central oxygen which has two lone pairs of electrons.

- Rearrange the electrons of the outer atoms to make multiple bonds with the central atom in order to obtain octets wherever possible.

SiH_4 : Si already has an octet, so nothing needs to be done.

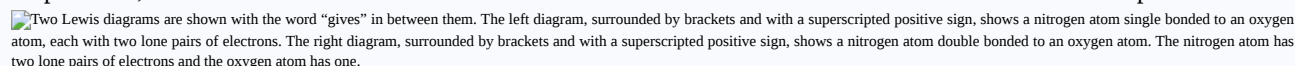
CHO_2^- : We have distributed the valence electrons as lone pairs on the oxygen atoms,

but the carbon atom lacks an octet:

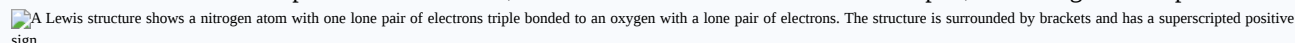
Two Lewis diagrams are shown with the word "gives" in between them. The left diagram, surrounded by brackets and with a superscripted negative sign, shows a carbon atom single bonded to two oxygen atoms, each with three lone pairs of electrons. The carbon atom also forms a single bond with a hydrogen atom. A curved arrow points from a lone pair on one of the oxygen atoms to the carbon atom. The right diagram, surrounded by brackets and with a superscripted negative sign, shows a carbon atom single bonded to an oxygen atom with three lone pairs of electrons, double bonded to an oxygen atom with two lone pairs of electrons, and single bonded to a hydrogen atom.

NO^+ : For this ion, we added eight valence electrons, but neither atom has an octet. We cannot add any more electrons since we have already used the total that we found in

Step 1, so we must move electrons to form a multiple bond:

Two Lewis diagrams are shown with the word "gives" in between them. The left diagram, surrounded by brackets and with a superscripted positive sign, shows a nitrogen atom single bonded to an oxygen atom, each with two lone pairs of electrons. The right diagram, surrounded by brackets and with a superscripted positive sign, shows a nitrogen atom double bonded to an oxygen atom. The nitrogen atom has two lone pairs of electrons and the oxygen atom has one.

This still does not produce an octet, so we must move another pair, forming a triple bond:

A Lewis structure shows a nitrogen atom with one lone pair of electrons triple bonded to an oxygen with a lone pair of electrons. The structure is surrounded by brackets and has a superscripted positive sign.

In OF_2 , each atom has an octet as drawn, so nothing changes.

Writing Lewis Structures: NASA's Cassini-Huygens mission detected a large cloud of toxic hydrogen cyanide (HCN) on Titan, one of Saturn's moons. Titan also contains ethane (H_3CCH_3), acetylene (HCCH), and ammonia (NH_3). What are the Lewis structures of these molecules?

Solution

- Calculate the number of valence electrons. HCN : $(1 \times 1) + (4 \times 1) + (5 \times 1) = 10$

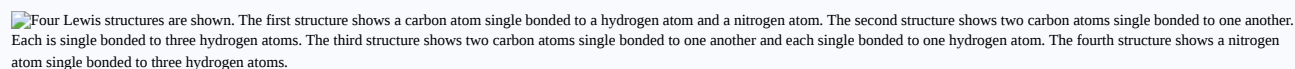
H_3CCH_3 : $(1 \times 3) + (2 \times 4) + (1 \times 3) = 14$

HCCH : $(1 \times 1) + (2 \times 4) + (1 \times 1) = 10$

NH_3 : $(5 \times 1) + (3 \times 1) = 8$

- Draw a skeleton and connect the atoms with single bonds.

Remember that H is never a central atom:

Four Lewis structures are shown. The first structure shows a carbon atom single bonded to a hydrogen atom and a nitrogen atom. The second structure shows two carbon atoms single bonded to one another. Each is single bonded to three hydrogen atoms. The third structure shows two carbon atoms single bonded to one another and each single bonded to one hydrogen atom. The fourth structure shows a nitrogen atom single bonded to three hydrogen atoms.

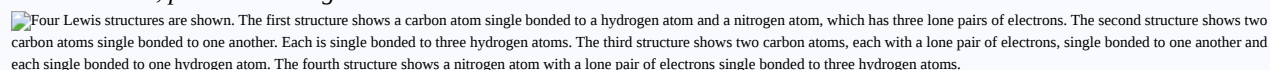
- HCN : six electrons placed on N

H_3CCH_3 : no electrons remain

HCCH : no terminal atoms capable of accepting electrons

NH_3 : no terminal atoms capable of accepting electrons

- Where needed, place remaining electrons on the central atom:

Four Lewis structures are shown. The first structure shows a carbon atom single bonded to a hydrogen atom and a nitrogen atom, which has three lone pairs of electrons. The second structure shows two carbon atoms single bonded to one another. Each is single bonded to three hydrogen atoms. The third structure shows two carbon atoms, each with a lone pair of electrons, single bonded to one another and each single bonded to one hydrogen atom. The fourth structure shows a nitrogen atom with a lone pair of electrons single bonded to three hydrogen atoms.

HCN: no electrons remain

H₃CCH₃: no electrons remain

HCCH: four electrons placed on carbon

NH₃: two electrons placed on nitrogen

5. Where needed, rearrange electrons to form multiple bonds in order to obtain an octet on each atom:

HCN: form two more C–N bonds H₃CCH₃: all atoms have the correct number of electrons

HCCH: form a triple bond between the two carbon atoms

NH₃: all atoms have the correct number of electrons

 Four Lewis structures are shown. The first structure shows a carbon atom single bonded to a hydrogen atom and a nitrogen atom, which has three lone pairs of electrons. Two curved arrows point from the nitrogen to the carbon. Below this structure is the word "gives" and below that is the same structure, but this time there is a triple bond between the carbon and nitrogen. The second structure shows two carbons single bonded to one another and each single bonded to three hydrogen atoms. The third structure shows two carbon atoms, each with a lone pair of electrons, single bonded to one another and each single bonded to one hydrogen atom. Two curved arrows point from the carbon atoms to the space in between the two. Below this structure is the word "gives" and the same structure, but this time with a triple bond between the two carbons. The fourth structure shows a nitrogen atom with a lone pair of electrons single bonded to three hydrogen atoms.

Check Your Learning: Both carbon monoxide, CO, and carbon dioxide, CO₂, are products of the combustion of fossil fuels. Both of these gases also cause problems: CO is toxic and CO₂ has been implicated in global climate change. What are the Lewis structures of these two molecules?

Answer:

 Two Lewis structures are shown. The left shows a carbon triple bonded to an oxygen, each with a lone electron pair. The right structure shows a carbon double bonded to an oxygen on each side. Each oxygen has two lone pairs of electrons.

In case you are wondering

FULLERENE CHEMISTRY

Carbon soot has been known to man since prehistoric times, but it was not until fairly recently that the molecular structure of the main component of soot was discovered. In 1996, the Nobel Prize in Chemistry was awarded to Richard Smalley ([Figure 3](#)), Robert Curl, and Harold Kroto for their work in discovering a new form of carbon, the C₆₀ buckminsterfullerene molecule ([link](#)). An entire class of compounds, including spheres and tubes of various shapes, were discovered based on C₆₀. This type of molecule, called a fullerene, shows promise in a variety of applications. Because of their size and shape, fullerenes can encapsulate other molecules, so they have shown potential in various applications from hydrogen storage to targeted drug delivery systems. They also possess unique electronic and optical properties that have been put to good use in solar powered devices and chemical sensors.

 A photo of Richard Smalley is shown.

Richard Smalley (1943–2005), a professor of physics, chemistry, and astronomy at Rice University, was one of the leading advocates for fullerene chemistry. Upon his death in 2005, the US Senate honored him as the “Father of Nanotechnology.” (credit: United States Department of Energy)

Key Concepts and Summary

Valence electronic structures can be visualized by drawing Lewis symbols (for atoms and monatomic ions) and Lewis structures (for molecules and polyatomic ions). Lone pairs, unpaired electrons and single, double or triple bonds are used to indicate where the valence electrons are located around each atom in a Lewis structure. Most structures—especially those containing second row elements—obey the octet rule, in which every atom (except H) is surrounded by eight electrons. Exceptions to the octet rule occur for odd-electron molecules (free radicals), electron-deficient molecules and hypervalent molecules.

Chemistry End of Chapter Exercises

1. Write the Lewis symbols for each of the following ions:

- (a) As^{3−}
- (b) I[−]
- (c) Be²⁺
- (d) O^{2−}
- (e) Ga³⁺

(f) Li^+

(g) N^{3-}

2. Many monatomic ions are found in seawater, including the ions formed from the following list of elements. Write the Lewis symbols for the monatomic ions formed from the following elements:

(a) Cl

(b) Na

(c) Mg

(d) Ca

(e) K

(f) Br

(g) Sr

(h) F

3. Write the Lewis symbols of the ions in each of the following ionic compounds and the Lewis symbols of the atom from which they are formed:

(a) MgS

(b) Al_2O_3

(c) GaCl_3

(d) K_2O

(e) Li_3N

(f) KF

5. Write the Lewis structure for the diatomic molecule P_2 , an unstable form of phosphorus found in high-temperature phosphorus vapor.

6. Write Lewis structures for the following:

(a) H_2

(b) HBr

(c) PCl_3

(d) SF_2

(e) H_2CCH_2

(f) HNNH

(g) H_2CNH

(h) NO^-

(i) N_2

(j) CO

(k) CN^-

7. Write Lewis structures for the following:

(a) O_2

(b) H_2CO

(c) AsF_3

(d) ClNO

(e) SiCl_4

(f) H_3O^+

(g) NH_4^+

(h) BF_4^-

(i) HCCH

(j) ClCN

(k) C_2^{2-}

8. Write Lewis structures for the following:

(a) ClF_3

(b) PCl_5

(c) BF_3

(d) PF_6^-

13. Methanol, H_3COH , is used as the fuel in some race cars. Ethanol, $\text{C}_2\text{H}_5\text{OH}$, is used extensively as motor fuel in Brazil. Both methanol and ethanol produce CO_2 and H_2O when they burn. Write the chemical equations for these combustion reactions using Lewis structures instead of chemical formulas.

14. Many planets in our solar system contain organic chemicals including methane (CH_4) and traces of ethylene (C_2H_4), ethane (C_2H_6), propyne (H_3CCCH), and diacetylene (HCCCH). Write the Lewis structures for each of these molecules.

15. Carbon tetrachloride was formerly used in fire extinguishers for electrical fires. It is no longer used for this purpose because of the formation of the toxic gas phosgene, Cl_2CO . Write the Lewis structures for carbon tetrachloride and phosgene.

17. The arrangement of atoms in several biologically important molecules is given here. Complete the Lewis structures of these molecules by adding multiple bonds and lone pairs. Do not add any more atoms.

(a) the amino acid serine:

 A Lewis structure is shown. A nitrogen atom is single bonded to two hydrogen atoms and a carbon atom. The carbon atom is single bonded to a hydrogen atom and two other carbon atoms. One of these carbon atoms is single bonded to two hydrogen atoms and an oxygen atom. The oxygen atom is bonded to a hydrogen atom. The other carbon atom is single bonded to two oxygen atoms, one of which is bonded to a hydrogen atom.

(b) urea:

 A Lewis structure is shown. A nitrogen atom is single bonded to two hydrogen atoms and a carbon atom. The carbon atom is single bonded to an oxygen atom and another nitrogen atom. That nitrogen atom is then single bonded to two hydrogen atoms.

(c) pyruvic acid:

 A Lewis structure is shown. A carbon atom is single bonded to three hydrogen atoms and another carbon atom. The second carbon atom is single bonded to an oxygen atom and a third carbon atom. This carbon is then single bonded to two oxygen atoms, one of which is single bonded to a hydrogen atom.

(d) uracil:

 A Lewis hexagonal ring structure is shown. From the top of the ring (moving clockwise), three carbon atoms, one nitrogen atom, a carbon atom, and a nitrogen atom are single bonded to each another. The top carbon atom is single bonded to an oxygen atom. The second and third carbons and the nitrogen atom are each single bonded to a hydrogen atom. The next carbon atom is single bonded to an oxygen atom, and the last nitrogen atom is single bonded to a hydrogen atom.

(e) carbonic acid:

 A Lewis structure is shown. A carbon atom is single bonded to three oxygen atoms. Two of those oxygen atoms are each single bonded to a hydrogen atom.

21. How are single, double, and triple bonds similar? How do they differ?

Glossary

double bond

covalent bond in which two pairs of electrons are shared between two atoms

free radical

molecule that contains an odd number of electrons

hypervalent molecule

molecule containing at least one main group element that has more than eight electrons in its valence shell

Lewis structure

diagram showing lone pairs and bonding pairs of electrons in a molecule or an ion

Lewis symbol

symbol for an element or monatomic ion that uses a dot to represent each valence electron in the element or ion

lone pair

two (a pair of) valence electrons that are not used to form a covalent bond

octet rule

guideline that states main group atoms will form structures in which eight valence electrons interact with each nucleus, counting bonding electrons as interacting with both atoms connected by the bond

single bond

bond in which a single pair of electrons is shared between two atoms

triple bond

bond in which three pairs of electrons are shared between two atoms

4.1: Lewis structures is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

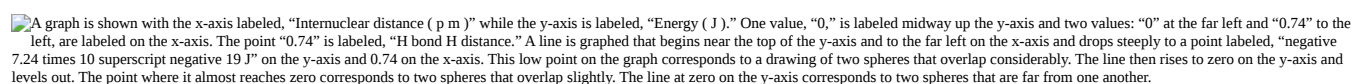
4.2: Covalent bonding and electronegativity

9

Harper College Chemistry Department

Formation of Covalent Bonds

Nonmetal atoms frequently form covalent bonds with other nonmetal atoms. For example, the hydrogen molecule, H_2 , contains a covalent bond between its two hydrogen atoms. [Figure 1](#) illustrates why this bond is formed. Starting on the far right, we have two separate hydrogen atoms with a particular potential energy, indicated by the red line. Along the x -axis is the distance between the two atoms. As the two atoms approach each other (moving left along the x -axis), their valence orbitals ($1s$) begin to overlap. The single electrons on each hydrogen atom then interact with both atomic nuclei, occupying the space around both atoms. The strong attraction of each shared electron to both nuclei stabilizes the system and the potential energy decreases as the bond distance decreases. If the atoms continue to approach each other, the positive charges in the two nuclei begin to repel each other and the potential energy increases. The bond length is determined by the distance at which the lowest potential energy is achieved.

A graph is shown with the x -axis labeled, "Internuclear distance (p m)" while the y -axis is labeled, "Energy (J)." One value, "0," is labeled midway up the y -axis and two values: "0" at the far left and "0.74" to the left, are labeled on the x -axis. The point "0.74" is labeled, "H bond H distance." A line is graphed that begins near the top of the y -axis and to the far left on the x -axis and drops steeply to a point labeled, "negative 7.24 times 10 superscript negative 19 J" on the y -axis and 0.74 on the x -axis. This low point on the graph corresponds to a drawing of two spheres that overlap considerably. The line then rises to zero on the y -axis and levels out. The point where it almost reaches zero corresponds to two spheres that overlap slightly. The line at zero on the y -axis corresponds to two spheres that are far from one another.

The potential energy of two separate hydrogen atoms (right) decreases as they approach each other and the single electrons on each atom are shared to form a covalent bond. The bond length is the internuclear distance at which the lowest potential energy is achieved.

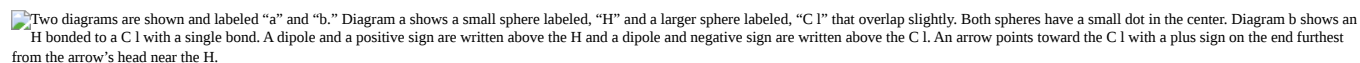
Polar vs. Nonpolar Covalent Bonds

If the atoms that form a covalent bond are identical, as in H_2 , Cl_2 and other diatomic molecules, then the electrons in the bond must be shared equally. We refer to this as a nonpolar covalent bond. Electrons shared in nonpolar covalent bonds have an equal probability of being near each nucleus.

In the case of Cl_2 , each atom starts off with seven valence electrons and each Cl shares one electron with the other, forming one covalent bond:

The total number of electrons around each individual atom consists of six nonbonding electrons and two shared (i.e., bonding) electrons for eight total electrons, matching the number of valence electrons in the noble gas argon. Since the bonding atoms are identical, Cl_2 also features a nonpolar covalent bond.

When the atoms linked by a covalent bond are different, the bonding electrons are shared, but no longer equally. Instead, the bonding electrons are more attracted to one atom than the other, giving rise to a shift of electron density toward that atom. This unequal distribution of electrons is known as a polar covalent bond, characterized by a partial positive charge on one atom and a partial negative charge on the other. The atom that attracts the electrons more strongly acquires the partial negative charge and vice versa. For example, the electrons in the $H-Cl$ bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom. Thus, in an HCl molecule, the chlorine atom carries a partial negative charge and the hydrogen atom has a partial positive charge. [Figure 2](#) shows the distribution of electrons in the $H-Cl$ bond. Note that the shaded area around Cl is much larger than it is around H. Compare this to [Figure 1](#), which shows the even distribution of electrons in the H_2 nonpolar bond.

Two diagrams are shown and labeled "a" and "b." Diagram a shows a small sphere labeled, "H" and a larger sphere labeled, "Cl" that overlap slightly. Both spheres have a small dot in the center. Diagram b shows an H bonded to a Cl with a single bond. A dipole and a positive sign are written above the H and a dipole and negative sign are written above the Cl. An arrow points toward the Cl with a plus sign on the end furthest from the arrow's head near the H.

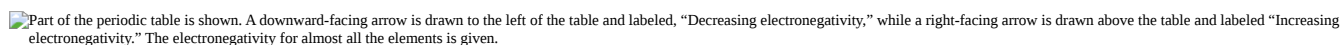
(a) The distribution of electron density in the HCl molecule is uneven. The electron density is greater around the chlorine nucleus. The small, black dots indicate the location of the hydrogen and chlorine nuclei in the molecule. (b) Symbols δ^+ and δ^- indicate the polarity of the $H-Cl$ bond.

We sometimes designate the positive and negative atoms in a polar covalent bond using a lowercase Greek letter "delta," δ , with a plus sign or minus sign to indicate whether the atom has a partial positive charge (δ^+) or a partial negative charge (δ^-). This symbolism is shown for the $H-Cl$ molecule in [Figure 2](#).

Electronegativity

Whether a bond is nonpolar or polar covalent is determined by a property of the bonding atoms called electronegativity. Electronegativity is a measure of the tendency of an atom to attract electrons (or electron density) towards itself. It determines how the shared electrons are distributed between the two atoms in a bond. The more strongly an atom attracts the electrons in its bonds, the larger its electronegativity. Electrons in a polar covalent bond are shifted toward the more electronegative atom; thus, the more electronegative atom is the one with the partial negative charge. The greater the difference in electronegativity, the more polarized the electron distribution and the larger the partial charges of the atoms.

Figure 3 shows the electronegativity values of the elements as proposed by one of the most famous chemists of the twentieth century: Linus Pauling (Figure 4). In general, electronegativity increases from left to right across a period in the periodic table and decreases down a group. Thus, the nonmetals, which lie in the upper right, tend to have the highest electronegativities, with fluorine the most electronegative element of all ($EN = 4.0$). Metals tend to be less electronegative elements, and the group 1 metals have the lowest electronegativities. Note that noble gases are excluded from this figure because these atoms usually do not share electrons with others atoms since they have a full valence shell.

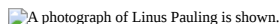
Part of the periodic table is shown. A downward-facing arrow is drawn to the left of the table and labeled, "Decreasing electronegativity," while a right-facing arrow is drawn above the table and labeled "Increasing electronegativity." The electronegativity for almost all the elements is given.

The electronegativity values derived by Pauling follow predictable periodic trends with the higher electronegativities toward the upper right of the periodic table.

In case you are wondering

LINUS PAULING

Linus Pauling, shown in Figure 4, is the only person to have received two unshared (individual) Nobel Prizes: one for chemistry in 1954 for his work on the nature of chemical bonds and one for peace in 1962 for his opposition to weapons of mass destruction. He developed many of the theories and concepts that are foundational to our current understanding of chemistry, including electronegativity and resonance structures.

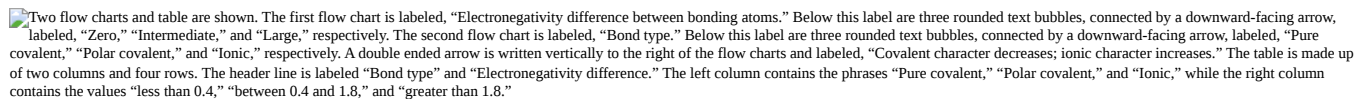
A photograph of Linus Pauling is shown.

Linus Pauling (1901–1994) made many important contributions to the field of chemistry. He was also a prominent activist, publicizing issues related to health and nuclear weapons.

Pauling also contributed to many other fields besides chemistry. His research on sickle cell anemia revealed the cause of the disease—the presence of a genetically inherited abnormal protein in the blood—and paved the way for the field of molecular genetics. His work was also pivotal in curbing the testing of nuclear weapons; he proved that radioactive fallout from nuclear testing posed a public health risk.

Electronegativity and Bond Type

The absolute value of the difference in electronegativity (ΔEN) of two bonded atoms provides a rough measure of the polarity to be expected in the bond and, thus, the bond type. When the difference is very small or zero, the bond is covalent and nonpolar. When it is large, the bond is polar covalent or ionic. The absolute values of the electronegativity differences between the atoms in the bonds H–H, H–Cl, and Na–Cl are 0 (nonpolar), 0.9 (polar covalent), and 2.1 (ionic), respectively. The degree to which electrons are shared between atoms varies from completely equal (nonpolar covalent bonding) to not at all (ionic bonding). Figure 5 shows the relationship between electronegativity difference and bond type.

Two flow charts and table are shown. The first flow chart is labeled, "Electronegativity difference between bonding atoms." Below this label are three rounded text bubbles, connected by a downward-facing arrow, labeled, "Zero," "Intermediate," and "Large," respectively. The second flow chart is labeled, "Bond type." Below this label are three rounded text bubbles, connected by a downward-facing arrow, labeled, "Pure covalent," "Polar covalent," and "Ionic," respectively. A double ended arrow is written vertically to the right of the flow charts and labeled, "Covalent character decreases; ionic character increases." The table is made up of two columns and four rows. The header line is labeled "Bond type" and "Electronegativity difference." The left column contains the phrases "Pure covalent," "Polar covalent," and "Ionic," while the right column contains the values "less than 0.4," "between 0.4 and 1.8," and "greater than 1.8."

As the electronegativity difference increases between two atoms, the bond becomes more ionic.

A rough approximation of the electronegativity differences associated with covalent, polar covalent, and ionic bonds is shown in Figure 5. This table is just a general guide, however, with many exceptions. For example, the H and F atoms in HF have an electronegativity difference of 1.9, and the N and H atoms in NH_3 a difference of 0.9, yet both of these compounds form bonds that are considered polar covalent. Likewise, the Na and Cl atoms in NaCl have an electronegativity difference of 2.1, and the Mn and I atoms in MnI_2 have a difference of 1.0, yet both of these substances form ionic compounds.

The best guide to the covalent or ionic character of a bond is to consider the types of atoms involved and their relative positions in the periodic table. Bonds between two nonmetals are generally covalent; bonding between a metal and a nonmetal is often ionic.

Some compounds contain both covalent and ionic bonds. The atoms in polyatomic ions, such as OH^- , and are held together by polar covalent bonds. However, these polyatomic ions form ionic compounds by combining with ions of opposite charge. For example, potassium nitrate, KNO_3 , contains the K^+ cation and the polyatomic nitrate anion. Thus, bonding in potassium nitrate is ionic, resulting from the electrostatic attraction between the ions K^+ and NO_3^- as well as covalent between the nitrogen and oxygen atoms in

Electronegativity and Bond Polarity: Bond polarities play an important role in determining the structure of proteins. Using the electronegativity values in [Figure 3](#), arrange the following covalent bonds—all commonly found in amino acids—in order of increasing polarity. Then designate the positive and negative atoms using the symbols δ^+ and δ^- :

C–H, C–N, C–O, N–H, O–H, S–H

Solution: The polarity of these bonds increases as the absolute value of the electronegativity difference increases. The atom with the δ^- designation is the more electronegative of the two. [Table 1](#) shows these bonds in order of increasing polarity.

Bond	ΔEN	Polarity
C–H	0.4	
S–H	0.4	
C–N	0.5	
N–H	0.9	
C–O	1.0	
O–H	1.4	

Check Your Learning: Silicones are polymeric compounds containing, among others, the following types of covalent bonds: Si–O, Si–C, C–H, and C–C. Using the electronegativity values in [Figure 3](#), arrange the bonds in order of increasing polarity and designate the positive and negative atoms using the symbols δ^+ and δ^- .

Answer:

Bond	Electronegativity Difference	Polarity
C–C	0.0	nonpolar
C–H	0.4	
Si–C	0.7	
Si–O	1.7	

Key Concepts and Summary

Covalent bonds form when electrons are shared between atoms and are attracted by the nuclei of both atoms. In nonpolar covalent bonds, the electrons are shared equally. In polar covalent bonds, the electrons are shared unequally, as one atom exerts a stronger force of attraction on the electrons than the other. The ability of an atom to attract a pair of electrons in a chemical bond is called its electronegativity. The difference in electronegativity between two atoms determines how polar a bond will be. In a diatomic molecule with two identical atoms, there is no difference in electronegativity, so the bond is nonpolar or pure covalent. When the electronegativity difference is very large, as is the case between metals and nonmetals, the bonding is characterized as ionic.

Chemistry End of Chapter Exercises

1. Why is it incorrect to speak of a molecule of solid NaCl?
2. What information can you use to predict whether a bond between two atoms is covalent or ionic?
3. Predict which of the following compounds are ionic and which are covalent, based on the location of their constituent atoms in the periodic table:
 - (a) Cl_2CO
 - (b) MnO
 - (c) NCl_3
 - (d) CoBr_2
 - (e) K_2S
 - (f) CO
 - (g) CaF_2
 - (h) HI
 - (i) CaO
 - (j) IBr
 - (k) CO_2
4. Explain the difference between a nonpolar covalent bond, a polar covalent bond and an ionic bond.
5. From its position in the periodic table, determine which atom in each pair is more electronegative:
 - (a) Br or Cl
 - (b) N or O
 - (c) S or O
 - (d) P or S
 - (e) Si or N
 - (f) Ba or P
 - (g) N or K
6. From its position in the periodic table, determine which atom in each pair is more electronegative:
 - (a) N or P
 - (b) N or Ge
 - (c) S or F
 - (d) Cl or S
 - (e) H or C
 - (f) Se or P
 - (g) C or Si
7. From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:
 - (a) C, F, H, N, O
 - (b) Br, Cl, F, H, I

- (c) F, H, O, P, S
- (d) Al, H, Na, O, P
- (e) Ba, H, N, O, As

8. From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:

- (a) As, H, N, P, Sb
- (b) Cl, H, P, S, Si
- (c) Br, Cl, Ge, H, Sr
- (d) Ca, H, K, N, Si
- (e) Cl, Cs, Ge, H, Sr

10. Which is the most polar bond?

- (a) C–C
- (b) C–H
- (c) N–H
- (d) O–H
- (e) Se–H

11. Identify the more polar bond in each of the following pairs of bonds:

- (a) HF or HCl
- (b) NO or CO
- (c) SH or OH
- (d) PCl or SCl
- (e) CH or NH
- (f) SO or PO
- (g) CN or NN

12. Which of the following molecules or ions contain polar bonds?

- (a) O₃
- (b) S₈
- (c) PO₄³⁻
- (d) NH₄⁺
- (e) CO₂
- (f) H₂S
- (g) Hg₂²⁺

Glossary

bond length

distance between the nuclei of two bonded atoms at which the lowest potential energy is achieved

covalent bond

bond formed when electrons are shared between atoms

electronegativity

tendency of an atom to attract electrons in a bond to itself

polar covalent bond

covalent bond between atoms of different electronegativities; a covalent bond with a positive end and a negative end

pure covalent bond

(also, nonpolar covalent bond) covalent bond between atoms of identical electronegativities

4.2: Covalent bonding and electronegativity is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

4.3: Valence Bond Theory

10

Harper College Chemistry Department

Derived from [Valence Bond Theory](#) by OpenStax

Atomic Orbital Overlap

There are successful theories that describe the electronic structure of atoms. We can use quantum mechanics to predict the specific regions around an atom where electrons are likely to be located: A spherical shape for an s orbital, a dumbbell shape for a p orbital, and so forth. However, these predictions only describe the orbitals around free atoms. When atoms bond to form molecules, atomic orbitals are not sufficient to describe the regions where electrons will be located in the molecule. A more complete understanding of electron distributions requires a model that can account for the electronic structure of molecules. One popular theory holds that a covalent bond forms when a pair of electrons is shared by two atoms and is simultaneously attracted by the nuclei of both atoms. In the following sections, we will discuss how such bonds are described by valence bond theory and hybridization.

Valence bond theory describes a covalent bond as the overlap of half-filled atomic orbitals (each containing a single electron) that yield a pair of electrons shared between the two bonded atoms. We say that orbitals on two different atoms overlap when a portion of one orbital and a portion of a second orbital occupy the same region of space. According to valence bond theory, a covalent bond results when two conditions are met: (1) an orbital on one atom overlaps an orbital on a second atom and (2) the single electrons in each orbital combine to form an electron pair. The mutual attraction between this negatively charged electron pair and the two atoms' positively charged nuclei serves to physically link the two atoms through a force we define as a covalent bond. The strength of a covalent bond depends on the extent of overlap of the orbitals involved. Orbitals that overlap extensively form bonds that are stronger than those that have less overlap.

In addition to the distance between two orbitals, the orientation of orbitals also affects their overlap (other than for two s orbitals, which are spherically symmetric). Greater overlap is possible when orbitals are oriented such that they overlap on a direct line between the two nuclei. [Figure 1](#) illustrates this for two p orbitals from different atoms; the overlap is greater when the orbitals overlap end to end rather than at an angle.

Two diagrams are shown. Diagram a contains two molecules whose p orbitals, which are depicted as two balloon-shaped structures that meet together to form a peanut shape, are laid end over end, creating an area of overlap. In diagram b, the same two molecules are shown, but this time, they are laid out in a way so as to form a near-ninety degree angle. In this diagram, the ends of two of these peanut-shaped orbitals do not overlap nearly as much.

(a) The overlap of two p orbitals is greatest when the orbitals are directed end to end. (b) Any other arrangement results in less overlap. The dots indicate the locations of the nuclei.

The overlap of two s orbitals (as in H_2), the overlap of an s orbital and a p orbital (as in HCl), and the end-to-end overlap of two p orbitals (as in Cl_2) all produce sigma bonds (σ bonds), as illustrated in [Figure](#). A σ bond is a covalent bond in which the electron density is concentrated in the region along the internuclear axis; that is, a line between the nuclei would pass through the center of the overlap region. Single bonds in Lewis structures are described as σ bonds in valence bond theory.

Three diagrams are shown and labeled "a," "b," and "c." Diagram a shows two spherical orbitals lying side by side and overlapping. Diagram b shows one spherical and one peanut-shaped orbital lying near one another so that the spherical orbital overlaps with one end of the peanut-shaped orbital. Diagram c shows two peanut-shaped orbitals lying end to end so that one end of each orbital overlaps the other.

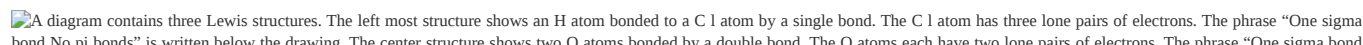
Sigma (σ) bonds form from the overlap of the following: (a) two s orbitals, (b) an s orbital and a p orbital, and (c) two p orbitals. The dots indicate the locations of the nuclei.

A pi bond (π bond) is a type of covalent bond that results from the side-by-side overlap of two p orbitals, as illustrated in [Figure](#). In a π bond, the regions of orbital overlap lie on opposite sides of the internuclear axis. Along the axis itself, there is a node, that is, a plane with no probability of finding an electron.

 Two peanut-shaped orbitals are shown, lying vertically and parallel with one another. They overlap one another along the top and bottom of the orbital.

Pi (π) bonds form from the side-by-side overlap of two p orbitals. The dots indicate the location of the nuclei.

While all single bonds are σ bonds, multiple bonds consist of both σ and π bonds. As the Lewis structures below suggest, O_2 contains a double bond, and N_2 contains a triple bond. The double bond consists of one σ bond and one π bond, and the triple bond consists of one σ bond and two π bonds. Between any two atoms, the first bond formed will always be a σ bond, but there can only be one σ bond in any one location. In any multiple bond, there will be one σ bond and the remaining one or two bonds will be π bonds. These bonds are described in more detail later in this chapter.

 A diagram contains three Lewis structures. The left most structure shows an H atom bonded to a C l atom by a single bond. The C l atom has three lone pairs of electrons. The phrase "One sigma bond No pi bonds" is written below the drawing. The center structure shows two O atoms bonded by a double bond. The O atoms each have two lone pairs of electrons. The phrase "One sigma bond One pi bond" is written below the drawing. The right most structure shows two N atoms bonded by a triple bond. Each N atom has a lone pairs of electrons. The phrase "One sigma bond Two pi bonds" is written below the drawing.

Counting σ and π Bonds

 This figure shows a molecule composed of four carbon atoms. There is a double bond between carbons one and two and three and four, while a single bond holds carbon two and three together. Carbons one and four are also bonded to two hydrogens with a single bond while carbons two and three are each bonded to one hydrogen each by a single bond.

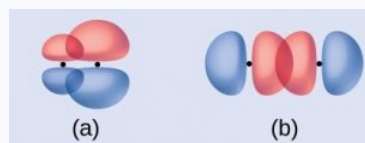
Butadiene, C_4H_6 , is used to make synthetic rubber. Identify the number of σ and π bonds contained in this molecule.

Solution There are six σ C–H bonds and one σ C–C bond, for a total of seven from the single bonds. There are two double bonds that each have a π bond in addition to the σ bond. This gives a total nine σ and two π bonds overall.

Check Your Learning

Identify each illustration as depicting a σ or π bond:

(a) side-by-side overlap of two p orbitals (b) end-to-end overlap of two p orbitals



Answer:

(a) is a π bond with a node along the axis connecting the nuclei while (b) is σ bonds that overlap along the axis.

Key Concepts and Summary

Valence bond theory describes bonding as a consequence of the overlap of two separate atomic orbitals on different atoms that creates a region with one pair of electrons shared between the two atoms. When the orbitals overlap along an axis containing the nuclei, they form a σ bond. When they overlap in a fashion that creates a node along this axis, they form a π bond. Dipole moments can be used to determine partial separations of charges between atoms.

Chemistry End of Chapter Exercises

1. Explain how σ and π bonds are similar and how they are different.
2. Use valence bond theory to explain the bonding in F_2 , HF , and $ClBr$. Sketch the overlap of the atomic orbitals involved in the bonds.
3. Use valence bond theory to explain the bonding in O_2 . Sketch the overlap of the atomic orbitals involved in the bonds in O_2 .
4. How many σ and π bonds are present in the molecule HCN ?
5. A friend tells you N_2 has three π bonds due to overlap of the three p -orbitals on each N atom. Do you agree?
6. Draw the Lewis structures for CO_2 and CO , and predict the number of σ and π bonds for each molecule.

(a) CO_2

(b) CO

Glossary

overlap

coexistence of orbitals from two different atoms sharing the same region of space, leading to the formation of a covalent bond

node

plane separating different lobes of orbitals, where the probability of finding an electron is zero

pi bond (π bond)

covalent bond formed by side-by-side overlap of atomic orbitals; the electron density is found on opposite sides of the internuclear axis

sigma bond (σ bond)

covalent bond formed by overlap of atomic orbitals along the internuclear axis

valence bond theory

description of bonding that involves atomic orbitals overlapping to form σ or π bonds, within which pairs of electrons are shared

4.3: Valence Bond Theory is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

4.4: Molecular Structure and Polarity

11

Harper College Chemistry Department

Derived from [Molecular Structure and Polarity](#) by OpenStax

Learning Objectives

By the end of this section, you will be able to:

- Predict the structures of small molecules using valence shell electron pair repulsion (VSEPR) theory
- Explain the concepts of polar covalent bonds and molecular polarity
- Assess the polarity of a molecule based on its bonding and structure

Thus far, we have used two-dimensional Lewis structures to represent molecules. However, molecular structure is actually three-dimensional, and it is important to be able to describe molecular bonds in terms of their distances, angles, and relative arrangements in space ([Figure 1](#)). A bond angle is the angle between any two bonds that include a common atom, usually measured in degrees. A bond distance (or bond length) is the distance between the nuclei of two bonded atoms along the straight line joining the nuclei. Bond distances are measured in Ångstroms ($1 \text{ Å} = 10^{-10} \text{ m}$) or picometers ($1 \text{ pm} = 10^{-12} \text{ m}$, $100 \text{ pm} = 1 \text{ Å}$).

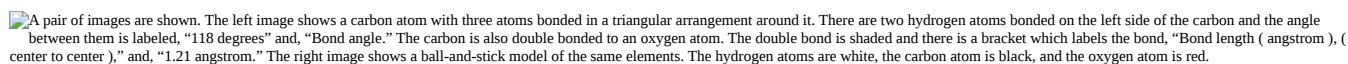


Figure 1: Bond distances (lengths) and angles are shown for the formaldehyde molecule, H_2CO .

VSEPR Theory

Valence shell electron-pair repulsion theory (VSEPR theory) enables us to predict the molecular structure, including approximate bond angles around a central atom, of a molecule from an examination of the number of bonds and lone electron pairs in its Lewis structure. The VSEPR model assumes that electron pairs in the valence shell of a central atom will adopt an arrangement that minimizes repulsions between these electron pairs by maximizing the distance between them. The electrons in the valence shell of a central atom form either bonding pairs of electrons, located primarily between bonded atoms, or lone pairs. The electrostatic repulsion of these electrons is reduced when the various regions of high electron density assume positions as far from each other as possible.

VSEPR theory predicts the arrangement of electron pairs around each central atom and, usually, the correct arrangement of atoms in a molecule. We should understand, however, that the theory only considers electron-pair repulsions. Other interactions, such as nuclear-nuclear repulsions and nuclear-electron attractions, are also involved in the final arrangement that atoms adopt in a particular molecular structure.

As a simple example of VSEPR theory, let us predict the structure of a gaseous BeF_2 molecule. The Lewis structure of BeF_2 ([Figure 2](#)) shows only two electron pairs around the central beryllium atom. With two bonds and no lone pairs of electrons on the central atom, the bonds are as far apart as possible, and the electrostatic repulsion between these regions of high electron density is reduced to a minimum when they are on opposite sides of the central atom. The bond angle is 180° ([Figure 2](#)).

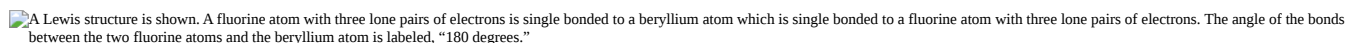


Figure 2 The BeF_2 molecule adopts a linear structure in which the two bonds are as far apart as possible, on opposite sides of the Be atom.

[Figure 3](#) illustrates this and other electron-pair geometries that minimize the repulsions among regions of high electron density (bonds and/or lone pairs). Two regions of electron density around a central atom in a molecule form a linear geometry; three regions form a trigonal planar geometry; four regions form a tetrahedral geometry; five regions form a trigonal bipyramidal geometry; and six regions form an octahedral geometry.

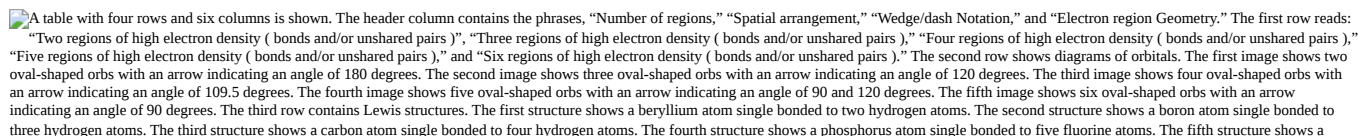


Figure 3 The basic electron-pair geometries predicted by VSEPR theory maximize the space around any region of electron density (bonds or lone pairs). For the purpose of CHM 100 focus on **linear, trigonal planar, and tetrahedral**.

Electron-pair Geometry versus Molecular Structure

It is important to note that electron-pair geometry around a central atom is *not* the same thing as its molecular structure. The electron-pair geometries shown in Figure 3 describe all regions where electrons are located, bonds as well as lone pairs. Molecular structure describes the location of the *atoms*, not the electrons.

We differentiate between these two situations by naming the geometry that includes *all* electron pairs the electron-pair geometry. The structure that includes only the placement of the atoms in the molecule is called the molecular structure. The electron-pair geometries will be the same as the molecular structures when there are no lone electron pairs around the central atom, but they will be different when there are lone pairs present on the central atom.

For example, the methane molecule, CH_4 , which is the major component of natural gas, has four bonding pairs of electrons around the central carbon atom; the electron-pair geometry is tetrahedral, as is the molecular structure (Figure 4). On the other hand, the ammonia molecule, NH_3 , also has four electron pairs associated with the nitrogen atom, and thus has a tetrahedral electron-pair geometry. One of these regions, however, is a lone pair, which is not included in the molecular structure, and this lone pair influences the shape of the molecule (Figure 5).

 A Lewis structure shows a carbon atom single bonded to four hydrogen atoms. This structure uses wedges and dashes to give it a three dimensional appearance.

Figure 4 The molecular structure of the methane molecule, CH_4 , is shown with a tetrahedral arrangement of the hydrogen atoms. VSEPR structures like this one are often drawn using the wedge and dash notation, in which solid lines represent bonds in the plane of the page, solid wedges represent bonds coming up out of the plane, and dashed lines represent bonds going down into the plane.

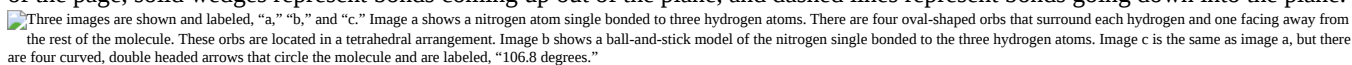
 Three images are shown and labeled, "a," "b," and "c." Image a shows a nitrogen atom single bonded to three hydrogen atoms. There are four oval-shaped orbs that surround each hydrogen and one facing away from the rest of the molecule. These orbs are located in a tetrahedral arrangement. Image b shows a ball-and-stick model of the nitrogen single bonded to the three hydrogen atoms. Image c is the same as image a, but there are four curved, double headed arrows that circle the molecule and are labeled, "106.8 degrees."

Figure 5 (a) The electron-pair geometry for the ammonia molecule is tetrahedral with one lone pair and three single bonds. (b) The trigonal pyramidal molecular structure is determined from the electron-pair geometry. (c) The actual bond angles deviate slightly from the idealized angles because the lone pair takes up a larger region of space than do the single bonds, causing the HNH angle to be slightly smaller than 109.5° .

As seen in Figure 5, small distortions from the ideal angles in Figure 3 can result from differences in repulsion between various regions of electron density. VSEPR theory predicts these distortions by establishing an order of repulsions and an order of the amount of space occupied by different kinds of electron pairs. The order of electron-pair repulsions from greatest to least repulsion is:

">lone pair-lone pair>lone pair-bonding pair>bonding pair-bonding pair lone pair-lone pair>lone pair-bonding pair>bonding pair-bonding pair"

This order of repulsions determines the amount of space occupied by different regions of electrons. A lone pair of electrons occupies a larger region of space than the electrons in a triple bond; in turn, electrons in a triple bond occupy more space than those in a double bond, and so on. The order of sizes from largest to smallest is:

">lone pair>triple bond>double bond>single bond lone pair>triple bond>double bond>single bond"

Consider formaldehyde, H_2CO , which is used as a preservative for biological and anatomical specimens (Figure 1). This molecule has regions of high electron density that consist of two single bonds and one double bond. The basic geometry is trigonal planar with 120° bond angles, but we see that the double bond causes slightly larger angles (121°), and the angle between the single bonds is slightly smaller (118°).

In the ammonia molecule, the three hydrogen atoms attached to the central nitrogen are not arranged in a flat, trigonal planar molecular structure, but rather in a three-dimensional trigonal pyramid (Figure 5) with the nitrogen atom at the apex and the three hydrogen atoms forming the base. The ideal bond angles in a trigonal pyramid are based on the tetrahedral electron pair geometry. Again, there are slight deviations from the ideal because lone pairs occupy larger regions of space than do bonding electrons. The H-N-H bond angles in NH_3 are slightly smaller than the 109.5° angle in a regular tetrahedron (Figure 4) because the lone pair-bonding pair repulsion is greater than the bonding pair-bonding pair repulsion (Figure 5). Figure 6 illustrates the ideal molecular structures, which are predicted based on the electron-pair geometries for various combinations of lone pairs and bonding pairs.

A table is shown that is comprised of six rows and six columns. The header row reads: "Number of Electron Regions," "Electron region geometries; 0 lone pair," "1 lone pair," "2 lone pairs," "3 lone pairs," and "4 lone pairs." The first column contains the numbers 2, 3, 4, 5, and 6. The first space in the second column contains a structure in which the letter E is single bonded to the letter X on each side. The angle of the bonds is labeled with a curved, double headed arrow and the value, "180 degrees." The structure is labeled, "Linear." The second space in the second column contains a structure in which the letter E is single bonded to the letter X on three sides. The angle between the bonds is labeled with a curved, double headed arrow and the value, "120 degrees." The structure is labeled, "Trigonal planar." The third space in the second column contains a structure in which the letter E is single bonded to the letter X four times. The angle between the bonds is labeled with a curved, double headed arrow and the value, "109 degrees." The structure is labeled, "Tetrahedral." The fourth space in the second column contains a structure in which the letter E is single bonded to the letter X on five sides. The angle between the bonds is labeled with a curved, double headed arrow and the values "90 and 120 degrees." The structure is labeled, "Trigonal bipyramid." The fifth space in the second column contains a structure in which the letter E is single bonded to the letter X on six sides. The angle between the bonds is labeled with a curved, double headed arrow and the value, "90 degrees." The structure is labeled, "Octahedral." The first space in the third column is empty while the second contains a structure in which the letter E is single bonded to the letter X on each side and has a lone pair of electrons. The angle between the bonds is labeled with a curved, double headed arrow and the value, "less than 120 degrees." The structure is labeled, "Bent or angular." The third space in the third column contains a structure in which the letter E is single bonded to the letter X three times and to a lone pair of electrons. It is labeled with a curved, double headed arrow and the value, "less than 109 degrees." The structure is labeled, "Trigonal pyramidal." The fourth space in the third column contains a structure in which the letter E is single bonded to the letter X on four sides and has a lone pair of electrons. The bond angle is labeled with a curved, double headed arrow and the values, "less than 90 and less than 120 degrees." The structure is labeled, "Sawhorse or seesaw." The fifth space in the third column contains a structure in which the letter E is single bonded to the letter X on five sides and has a lone pair of electrons. The bond angle is labeled with a curved, double headed arrow and the value, "less than 90 degrees." The structure is labeled, "Square pyramidal." The first and second spaces in the fourth column are empty while the third contains a structure in which the letter E is single bonded to the letter X on each side and has two lone pairs of electrons. The bond angle is labeled with a curved, double headed arrow and the value, "less than less than 109 degrees." The structure is labeled, "Bent or angular." The fourth space in the fourth column contains a structure in which the letter E is single bonded to the letter X three times and to two lone pairs of electrons. The bond angle is labeled with a curved, double headed arrow and the value, "less than 90 degrees." The structure is labeled, "T - shape." The fifth space in the fourth column contains a structure in which the letter E is single bonded to the letter X on four sides and has two lone pairs of electrons. The bond angle is labeled with a curved, double headed arrow and the value "90 degrees." The structure is labeled, "Square planar." The first, second and third spaces in the fifth column are empty while the fourth contains a structure in which the letter E is single bonded to the letter X on each side and has three lone pairs of electrons. The bond angle is labeled with a curved, double headed arrow and the value, "180 degrees." The structure is labeled, "Linear." The fifth space in the fifth column contains a structure in which the letter E is single bonded to the letter X three times and to three lone pairs of electrons. The bond angle is labeled with a curved, double headed arrow and the value, "less than 90 degrees." The structure is labeled, "T - shape." The first, second, third, and fourth spaces in the sixth column are empty while the fifth contains a structure in which the letter E is single bonded to the letter X on each side and has four lone pairs of electrons. The bond angle is labeled with a curved, double headed arrow and the value "180 degrees." The structure is labeled, "Linear." All the structures use wedges and dashes to give them three dimensional appearances.

Figure 6 The molecular structures are identical to the electron-pair geometries when there are no lone pairs present (first column). For a particular number of electron pairs (row), the molecular structures for one or more lone pairs are determined based on modifications of the corresponding electron-pair geometry. For the purpose of CHM 100 only worry about molecules with **2, 3, and 4 electron regions**.

According to VSEPR theory, the terminal atom locations (Xs in Figure 6) are equivalent within the linear, trigonal planar, and tetrahedral electron-pair geometries (the first three rows of the table). It does not matter which X is replaced with a lone pair because the molecules can be rotated to convert positions.

Predicting Electron Pair Geometry and Molecular Structure

The following procedure uses VSEPR theory to determine the electron pair geometries and the molecular structures:

1. Write the Lewis structure of the molecule or polyatomic ion.
2. Count the number of regions of electron density (lone pairs and bonds) around the central atom. A single, double, or triple bond counts as one region of electron density.
3. Identify the electron-pair geometry based on the number of regions of electron density: linear, trigonal planar, or tetrahedral (Figure 6, first column).
4. Use the number of lone pairs to determine the molecular structure (Figure 6). If more than one arrangement of lone pairs and chemical bonds is possible, choose the one that will minimize repulsions, remembering that lone pairs occupy more space than multiple bonds, which occupy more space than single bonds. In trigonal bipyramidal arrangements, repulsion is minimized when every lone pair is in an equatorial position. In an octahedral arrangement with two lone pairs, repulsion is minimized when the lone pairs are on opposite sides of the central atom.

The following examples illustrate the use of VSEPR theory to predict the molecular structure of molecules or ions that have no lone pairs of electrons. In this case, the molecular structure is identical to the electron pair geometry.

EXAMPLE 1

Predicting Electron-pair Geometry and Molecular Structure: CO_2 and BCl_3

Predict the electron-pair geometry and molecular structure for each of the following:

- (a) carbon dioxide, CO_2 , a molecule produced by the combustion of fossil fuels
- (b) boron trichloride, BCl_3 , an important industrial chemical

Solution

(a) We write the Lewis structure of CO_2 as:

A Lewis structure shows a carbon atom double bonded on both the left and right sides to oxygen atoms that each have two lone pairs of electrons.

This shows us two regions of high electron density around the carbon atom—each double bond counts as one region, and there are no lone pairs on the carbon atom. Using VSEPR theory, we predict that the two regions of electron density arrange themselves on opposite sides of the central atom with a bond angle of 180° . The electron-pair geometry and molecular structure are identical, and CO_2 molecules are linear.

(b) We write the Lewis structure of BCl_3 as:

 A Lewis structure depicts a boron atom that is single bonded to three chlorine atoms, each of which has three lone pairs of electrons.

Thus we see that BCl_3 contains three bonds, and there are no lone pairs of electrons on boron. The arrangement of three regions of high electron density gives a trigonal planar electron-pair geometry. The B–Cl bonds lie in a plane with 120° angles between them. BCl_3 also has a trigonal planar molecular structure (Figure 7.21).

 A Lewis structure depicts a boron atom that is single bonded to three chlorine atoms, each of which is oriented in the same flat plane. This figure uses dashes and wedges to give it a three-dimensional appearance.

Figure 7.21

The electron-pair geometry and molecular structure of BCl_3 are both trigonal planar. Note that the VSEPR geometry indicates the correct bond angles (120°), unlike the Lewis structure shown above.

Check Your Learning

Carbonate, CO_3^{2-} , is a common polyatomic ion found in various materials from eggshells to antacids. What are the electron-pair geometry and molecular structure of this polyatomic ion?

Answer:

The electron-pair geometry is trigonal planar and the molecular structure is trigonal planar. Due to resonance, all three C–O bonds are identical. Whether they are single, double, or an average of the two, each bond counts as one region of electron density.

EXAMPLE 2

Predicting Electron-pair Geometry and Molecular Structure: Ammonium

Two of the top 50 chemicals produced in the United States, ammonium nitrate and ammonium sulfate, both used as fertilizers, contain the ammonium ion. Predict the electron-pair geometry and molecular structure of the NH_4^+ cation.

Solution

We write the Lewis structure of NH_4^+ as:

 A Lewis structure depicts a nitrogen atom that is single bonded to four hydrogen atoms. The structure is surrounded by brackets and has a superscripted positive sign.

We can see that NH_4^+ contains four bonds from the nitrogen atom to hydrogen atoms and no lone pairs. We expect the four regions of high electron density to arrange themselves so that they point to the corners of a tetrahedron with the central nitrogen atom in the middle (Figure 7.19). Therefore, the electron pair geometry of NH_4^+ is tetrahedral, and the molecular structure is also tetrahedral (Figure 7.22).

 A Lewis structure depicts a nitrogen atom that is single bonded to four hydrogen atoms. The structure is surrounded by brackets and has a superscripted positive sign. This figure uses dashes and wedges to display its three planes in a tetrahedral shape.

Figure 7.22 The ammonium ion displays a tetrahedral electron-pair geometry as well as a tetrahedral molecular structure.

Check Your Learning

Identify a molecule with trigonal bipyramidal molecular structure.

Answer:

Any molecule with five electron pairs around the central atoms including no lone pairs will be trigonal bipyramidal. PF_5 is a common example.

The next several examples illustrate the effect of lone pairs of electrons on molecular structure.

EXAMPLE 3

Predicting Electron-pair Geometry and Molecular Structure: Lone Pairs on the Central Atom

Predict the electron-pair geometry and molecular structure of a water molecule.

Solution

The Lewis structure of H_2O indicates that there are four regions of high electron density around the oxygen atom: two lone pairs and two chemical bonds:
 A Lewis structure depicts an oxygen atom with two lone pairs of electrons single bonded to two hydrogen atoms.

We predict that these four regions are arranged in a tetrahedral fashion (Figure 7.23), as indicated in Figure 6. Thus, the electron-pair geometry is tetrahedral and the molecular structure is bent with an angle slightly less than 109.5° . In fact, the bond angle is 104.5° .

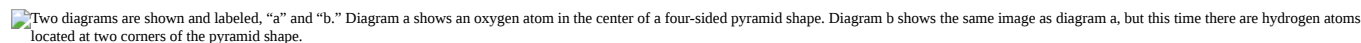
Two diagrams are shown and labeled, "a" and "b." Diagram a shows an oxygen atom in the center of a four-sided pyramid shape. Diagram b shows the same image as diagram a, but this time there are hydrogen atoms located at two corners of the pyramid shape.

Figure 7.23 (a) H_2O has four regions of electron density around the central atom, so it has a tetrahedral electron-pair geometry. (b) Two of the electron regions are lone pairs, so the molecular structure is bent.

Check Your Learning

The hydronium ion, H_3O^+ , forms when acids are dissolved in water. Predict the electron-pair geometry and molecular structure of this cation.

Answer:

electron pair geometry: tetrahedral; molecular structure: trigonal pyramidal

Check Your Learning

In a certain molecule, the central atom has three lone pairs and two bonds. What will the electron pair geometry and molecular structure be?

Answer:

electron pair geometry: trigonal bipyramidal; molecular structure: linear

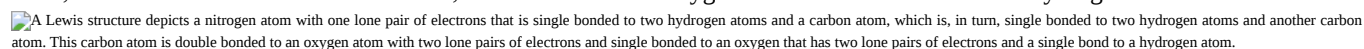
Molecular Structure for Multicenter Molecules

When a molecule or polyatomic ion has only one central atom, the molecular structure completely describes the shape of the molecule. Larger molecules do not have a single central atom, but are connected by a chain of interior atoms that each possess a "local" geometry. The way these local structures are oriented with respect to each other also influences the molecular shape, but such considerations are largely beyond the scope of this introductory discussion. For our purposes, we will only focus on determining the local structures.

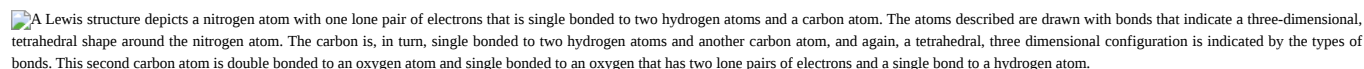
EXAMPLE 7.16

Predicting Structure in Multicenter Molecules

The Lewis structure for the simplest amino acid, glycine, $\text{H}_2\text{NCH}_2\text{CO}_2\text{H}$, is shown here. Predict the local geometry for the nitrogen atom, the two carbon atoms, and the oxygen atom with a hydrogen atom attached:

A Lewis structure depicts a nitrogen atom with one lone pair of electrons that is single bonded to two hydrogen atoms and a carbon atom, which is, in turn, single bonded to two hydrogen atoms and another carbon atom. This carbon atom is double bonded to an oxygen atom with two lone pairs of electrons and single bonded to an oxygen that has two lone pairs of electrons and a single bond to a hydrogen atom.

Solution

A Lewis structure depicts a nitrogen atom with one lone pair of electrons that is single bonded to two hydrogen atoms and a carbon atom. The atoms described are drawn with bonds that indicate a three-dimensional, tetrahedral shape around the nitrogen atom. The carbon is, in turn, single bonded to two hydrogen atoms and another carbon atom, and again, a tetrahedral, three dimensional configuration is indicated by the types of bonds. This second carbon atom is double bonded to an oxygen atom and single bonded to an oxygen that has two lone pairs of electrons and a single bond to a hydrogen atom.

Consider each central atom independently. The electron-pair geometries:

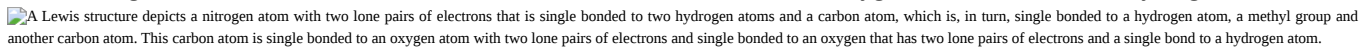
- nitrogen—four regions of electron density; tetrahedral
- carbon (CH_2)—four regions of electron density; tetrahedral
- carbon (CO_2)—three regions of electron density; trigonal planar
- oxygen (OH)—four regions of electron density; tetrahedral

The local structures:

- nitrogen—three bonds, one lone pair; trigonal pyramidal
- carbon (CH_2)—four bonds, no lone pairs; tetrahedral
- carbon (CO_2)—three bonds (double bond counts as one bond), no lone pairs; trigonal planar
- oxygen (OH)—two bonds, two lone pairs; bent (109°)

Check Your Learning

Another amino acid is alanine, which has the Lewis structure shown here. Predict the electron-pair geometry and local structure of the nitrogen atom, the three carbon atoms, and the oxygen atom with hydrogen attached:

A Lewis structure depicts a nitrogen atom with two lone pairs of electrons that is single bonded to two hydrogen atoms and a carbon atom, which is, in turn, single bonded to a hydrogen atom, a methyl group and another carbon atom. This carbon atom is single bonded to an oxygen atom with two lone pairs of electrons and single bonded to an oxygen that has two lone pairs of electrons and a single bond to a hydrogen atom.

Answer:

electron-pair geometries: nitrogen—tetrahedral; carbon (CH)—tetrahedral; carbon (CH_3)—tetrahedral; carbon (CO_2)—trigonal planar; oxygen (OH)—tetrahedral; local structures: nitrogen—trigonal pyramidal; carbon (CH)—tetrahedral; carbon (CH_3)—tetrahedral; carbon (CO_2)—trigonal planar; oxygen (OH)—bent (109°)

LINK TO LEARNING

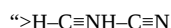
The [molecular shape simulator](#) lets you build various molecules and practice naming their electron-pair geometries and molecular structures.

EXAMPLE 7.17

Molecular Simulation

Using [molecular shape simulator](#) allows us to control whether bond angles and/or lone pairs are displayed by checking or unchecking the boxes under “Options” on the right. We can also use the “Name” checkboxes at bottom-left to display or hide the electron pair geometry (called “electron geometry” in the simulator) and/or molecular structure (called “molecular shape” in the simulator).

Build the molecule HCN in the simulator based on the following Lewis structure:



Click on each bond type or lone pair at right to add that group to the central atom. Once you have the complete molecule, rotate it to examine the predicted molecular structure. What molecular structure is this?

Solution

The molecular structure is linear.

Check Your Learning

Build a more complex molecule in the simulator. Identify the electron-group geometry, molecular structure, and bond angles. Then try to find a chemical formula that would match the structure you have drawn.

Answer:

Answers will vary. For example, an atom with four single bonds, a double bond, and a lone pair has an octahedral electron-group geometry and a square pyramidal molecular structure. XeOF_4 is a molecule that adopts this structure.

Molecular Polarity and Dipole Moment

As discussed previously, polar covalent bonds connect two atoms with differing electronegativities, leaving one atom with a partial positive charge (δ^+) and the other atom with a partial negative charge (δ^-), as the electrons are pulled toward the more electronegative atom. This separation of charge gives rise to a bond dipole moment. The magnitude of a bond dipole moment is represented by the Greek letter mu (μ) and is given by the formula shown here, where Q is the magnitude of the partial charges (determined by the electronegativity difference) and r is the distance between the charges:

$$\mu = Qr$$

This bond moment can be represented as a vector, a quantity having both direction and magnitude ([Figure 7.26](#)). Dipole vectors are shown as arrows pointing along the bond from the less electronegative atom toward the more electronegative atom. A small plus sign is drawn on the less electronegative end to indicate the partially positive end of the bond. The length of the arrow is proportional to the magnitude of the electronegativity difference between the two atoms.

Two images are shown and labeled, "a" and "b." Image a shows a large sphere labeled, "C," a left-facing arrow with a crossed end, and a smaller sphere labeled "H." Image b shows a large sphere labeled, "B," a right-facing arrow with a crossed end, and a smaller sphere labeled "F."

Figure 7.26 (a) There is a small difference in electronegativity between C and H, represented as a short vector. (b) The electronegativity difference between B and F is much larger, so the vector representing the bond moment is much longer.

A whole molecule may also have a separation of charge, depending on its molecular structure and the polarity of each of its bonds. If such a charge separation exists, the molecule is said to be a polar molecule (or dipole); otherwise the molecule is said to be nonpolar. The dipole moment measures the extent of net charge separation in the molecule as a whole. We determine the dipole moment by adding the bond moments in three-dimensional space, taking into account the molecular structure.

For diatomic molecules, there is only one bond, so its bond dipole moment determines the molecular polarity. Homonuclear diatomic molecules such as Br_2 and N_2 have no difference in electronegativity, so their dipole moment is zero. For heteronuclear molecules such as CO , there is a small dipole moment. For HF , there is a larger dipole moment because there is a larger difference in electronegativity.

When a molecule contains more than one bond, the geometry must be taken into account. If the bonds in a molecule are arranged such that their bond moments cancel (vector sum equals zero), then the molecule is nonpolar. This is the situation in CO_2 (Figure 7.27). Each of the bonds is polar, but the molecule as a whole is nonpolar. From the Lewis structure, and using VSEPR theory, we determine that the CO_2 molecule is linear with polar $\text{C}=\text{O}$ bonds on opposite sides of the carbon atom. The bond moments cancel because they are pointed in opposite directions. In the case of the water molecule (Figure 7.27), the Lewis structure again shows that there are two bonds to a central atom, and the electronegativity difference again shows that each of these bonds has a nonzero bond moment. In this case, however, the molecular structure is bent because of the lone pairs on O, and the two bond moments do not cancel. Therefore, water does have a net dipole moment and is a polar molecule (dipole).

Two images are shown and labeled, "a" and "b." Image a shows a carbon atom bonded to two oxygen atoms in a ball-and-stick representation. Two arrows face away from the center of the molecule in opposite directions and are drawn horizontally like the molecule. These arrows are labeled, "Bond moments," and the image is labeled, "Overall dipole moment equals 0." Image b shows an oxygen atom bonded to two hydrogen atoms in a downward-facing v-shaped arrangement. An upward-facing, vertical arrow is drawn below the molecule while two upward and inward facing arrows are drawn above the molecule. The upper arrows are labeled, "Bond moments," while the image is labeled, "Overall dipole moment."

Figure 7.27 The overall dipole moment of a molecule depends on the individual bond dipole moments and how they are arranged. (a) Each CO bond has a bond dipole moment, but they point in opposite directions so that the net CO_2 molecule is nonpolar. (b) In contrast, water is polar because the OH bond moments do not cancel out.

The OCS molecule has a structure similar to CO_2 , but a sulfur atom has replaced one of the oxygen atoms. To determine if this molecule is polar, we draw the molecular structure. VSEPR theory predicts a linear molecule:

An image shows a carbon atom double bonded to a sulfur atom and an oxygen atom which are arranged in a horizontal plane. Two arrows face away from the center of the molecule in opposite directions and are drawn horizontally like the molecule. The left-facing arrow is larger than the right-facing arrow. These arrows are labeled, "Bond moments," and a left-facing arrow below the molecule is labeled, "Overall dipole moment."

The $\text{C}-\text{O}$ bond is considerably polar. Although C and S have very similar electronegativity values, S is slightly more electronegative than C, and so the $\text{C}-\text{S}$ bond is just slightly polar. Because oxygen is more electronegative than sulfur, the oxygen end of the molecule is the negative end.

Chloromethane, CH_3Cl , is a tetrahedral molecule with three slightly polar $\text{C}-\text{H}$ bonds and a more polar $\text{C}-\text{Cl}$ bond. The relative electronegativities of the bonded atoms is $\text{H} < \text{C} < \text{Cl}$, and so the bond moments all point toward the Cl end of the molecule and sum to yield a considerable dipole moment (the molecules are relatively polar).

An image shows a carbon atom single bonded to three hydrogen atoms and a chlorine atom. There are arrows with crossed ends pointing from the hydrogen to the carbon near each bond, and one pointing from the carbon to the chlorine along that bond. The carbon and chlorine arrow is longer. This image uses dashes and wedges to give it a three-dimensional appearance.

For molecules of high symmetry such as BF_3 (trigonal planar), CH_4 (tetrahedral), PF_5 (trigonal bipyramidal), and SF_6 (octahedral), all the bonds are of identical polarity (same bond moment) and they are oriented in geometries that yield nonpolar molecules (dipole moment is zero). Molecules of less geometric symmetry, however, may be polar even when all bond moments are identical. For these molecules, the directions of the equal bond moments are such that they sum to give a nonzero dipole moment and a polar molecule. Examples of such molecules include hydrogen sulfide, H_2S (nonlinear), and ammonia, NH_3 (trigonal pyramidal).

Two Lewis structures are shown. The left structure shows a sulfur atom with two lone pairs of electrons single bonded to two hydrogen atoms. Near the sulfur is a dipole symbol with a superscripted negative sign. Near each hydrogen is a dipole symbol with a superscripted positive sign. The right structure shows a nitrogen atom with one lone pair of electrons single bonded to three hydrogen atoms. Near the nitrogen is a dipole symbol with a superscripted negative sign. Near each hydrogen is a dipole symbol with a superscripted positive sign.

To summarize, to be polar, a molecule must:

1. Contain at least one polar covalent bond.
2. Have a molecular structure such that the sum of the vectors of each bond dipole moment does not cancel.

LINK TO LEARNING

The [molecule polarity simulation](#) provides many ways to explore dipole moments of bonds and molecules.

EXAMPLE 7.18

Polarity Simulations

Open the [molecule polarity simulation](#) and select the “Three Atoms” tab at the top. This should display a molecule ABC with three electronegativity adjusters. You can display or hide the bond moments, molecular dipoles, and partial charges at the right. Turning on the Electric Field will show whether the molecule moves when exposed to a field, similar to [Figure 7.28](#).

Use the electronegativity controls to determine how the molecular dipole will look for the starting bent molecule if:

- (a) A and C are very electronegative and B is in the middle of the range.
- (b) A is very electronegative, and B and C are not.

Solution

- (a) Molecular dipole moment points immediately between A and C.
- (b) Molecular dipole moment points along the A–B bond, toward A.

Check Your Learning

Determine the partial charges that will give the largest possible bond dipoles.

Answer:

The largest bond moments will occur with the largest partial charges. The two solutions above represent how unevenly the electrons are shared in the bond. The bond moments will be maximized when the electronegativity difference is greatest. The controls for A and C should be set to one extreme, and B should be set to the opposite extreme. Although the magnitude of the bond moment will not change based on whether B is the most electronegative or the least, the direction of the bond moment will.

4.4: Molecular Structure and Polarity is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

4.5: Atomic Structure

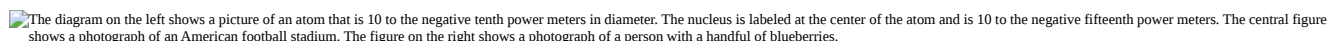
12

Harper College Chemistry Department

[latexpage]

Composition of an Atom

The development of the modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (Figure 1).



If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by “babyknight”/Wikimedia Commons; credit right: modification of work by Paxson Woelber)

Atoms—and the protons, neutrons and electrons that compose them—are extremely small. For example, a carbon atom weighs less than 2×10^{-23} g and an electron has a charge of less than 2×10^{-19} C (coulomb). When describing the properties of tiny objects such as atoms, we use appropriately small units of measure, such as the atomic mass unit (amu) and the fundamental unit of charge (e). The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant isotope of carbon, atoms of which are assigned masses of exactly 12 amu. This isotope is known as “carbon-12” as will be discussed later in this module. Thus, one amu is exactly $\frac{1}{12}$ of the mass of one carbon-12 atom: $1 \text{ amu} = 1.6605 \times 10^{-24}$ g. The fundamental unit of charge (also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19}$ C.

A proton has a mass of 1.0073 amu and a charge of $1+$. A neutron is a slightly heavier particle with a mass 1.0087 amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of $1-$ and is a much lighter particle with a mass of about 0.00055 amu. It would take about 1800 electrons to equal the mass of one proton. The properties of these fundamental particles are summarized in Table 1. An observant student might notice that the sum of an atom's subatomic particles does not equal the atom's actual mass: The total mass of six protons, six neutrons and six electrons is 12.0993 amu, slightly larger than 12.00 amu. This “missing” mass is known as the mass defect, and you will learn about it in the chapter on nuclear chemistry.

Properties of Subatomic Particles					
Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	$1-$	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	$1+$	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

The number of protons in the nucleus of an atom is its atomic number (Z). This is the defining trait of an element: Its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. The total number of protons and neutrons in an atom is called its mass number (A). The number of neutrons is therefore the difference between the mass number and the atomic number: $A - Z = \text{number of neutrons}$.

Erroneous nesting of equation structures

Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an ion. The charge of an atom is defined as follows:

Atomic charge = number of protons – number of electrons

As will be discussed in more detail later in this chapter, atoms (and molecules) typically acquire a charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an anion. Positively charged atoms called cations are formed when an atom loses one or more electrons. For example, a neutral sodium atom ($Z = 11$) has 11 protons and 11 electrons. If this atom loses one electron, it will become a cation with a $1+$ charge ($11 - 10 = 1+$). A neutral oxygen atom ($Z = 8$) has eight protons and electrons, if it gains two electrons it will become an anion with a $2-$ charge ($8 - 10 = 2-$).

Iodine is an essential trace element in our diet; it is needed to produce thyroid hormone. Insufficient iodine in the diet can lead to the development of a goiter, an enlargement of the thyroid gland ([Figure 2](#)).

 Figure A shows a photo of a person who has a very swollen thyroid in his or her neck. Figure B shows a photo of a canister of iodized salt.

(a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification of work by "Almazi"/Wikimedia Commons; credit b: modification of work by Mike Mozart)

The addition of small amounts of iodine to table salt (iodized salt) has essentially eliminated this health concern in the United States, but as much as 40% of the world's population is still at risk of iodine deficiency. The iodine atoms are added as anions and each has a $1-$ charge and a mass number of 127.

Check Your Learning

Determine the numbers of protons, neutrons, and electrons in one of these iodine anions.

Answer

The atomic number of iodine (53) tells us that a neutral iodine atom contains 53 protons in its nucleus and 53 electrons outside its nucleus. Because the sum of the number of protons and neutrons equals the mass number (127), the number of neutrons is 74 ($127 - 53 = 74$). Since the iodine is added as a $1-$ anion, the number of electrons is 54 [$53 - (1-) = 54$].

Check Your Learning

An ion of platinum has a mass number of 195 and contains 74 electrons. How many protons and neutrons does it contain, and what is its charge?

Answer

78 protons; 117 neutrons; charge is $4+$

Chemical Symbols

A chemical symbol is an abbreviation that we use to indicate an element or an atom of an element. For example, the symbol for mercury is Hg ([Figure 3](#)). We use the same symbol to indicate one atom of mercury (microscopic domain) or to label a container of many atoms of the element mercury (macroscopic domain).

 A jar labeled "Hg" is shown with a small amount of liquid mercury in it.

The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

The symbols for several common elements and their atoms are listed in [Table 2](#). Some symbols are derived from the common name of the element; others come from the name in another language (usually Latin). Symbols have one or two letters. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the formula for the compound carbon monoxide, which contains atoms of the elements carbon (C) and oxygen (O). All known elements and their symbols are in the [periodic table](#).

Some Common Elements and Their Symbols			
Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from <i>ferrum</i>)
bromine	Br	lead	Pb (from <i>plumbum</i>)
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from <i>hydrargyrum</i>)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from <i>kalium</i>)
copper	Cu (from <i>cuprum</i>)	silicon	Si
fluorine	F	silver	Ag (from <i>argentum</i>)
gold	Au (from <i>aurum</i>)	sodium	Na (from <i>natrium</i>)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from <i>stannum</i>)
iodine	I	zinc	Zn

Traditionally, the discoverer(s) of a new element names the element. However, until the name is recognized by the International Union of Pure and Applied Chemistry (IUPAC), the recommended name of the new element is based on the Latin word(s) for its atomic number. For example, element 106 was called unnilhexium (Unh), element 107 was called unnilseptium (Uns), and element 108 was called unniloctium (Uno) for several years. These elements are now named after scientists (or occasionally locations); for example, element 106 is now known as *seaborgium* (Sg) in honor of Glenn Seaborg, a Nobel Prize winner who was active in the discovery of several heavy elements.

Visit the IUPAC [site](#) to learn more about the International Union of Pure and Applied Chemistry and explore its periodic table at the bottom of the web page.

Isotopes

The symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol (Figure 4). The atomic number is sometimes written as a subscript also on the left of the symbol. However, since this number defines the element's identity, as does its symbol, it is often omitted. For example, magnesium exists as a mixture of three isotopes, each with an atomic number of 12 and with mass numbers of 24, 25 and 26, respectively. These isotopes can be identified as ^{24}Mg , ^{25}Mg and ^{26}Mg . These isotope symbols are read as "element, mass number" and can be symbolized consistent with this reading. For instance, ^{24}Mg is read as "magnesium 24," and can be written as "magnesium-24" or "Mg-24." ^{25}Mg is read as "magnesium 25," and can be written as "magnesium-25" or "Mg-25." All magnesium atoms have 12 protons in their nucleus. They differ only because a ^{24}Mg atom has 12 neutrons in its nucleus, a ^{25}Mg atom has 13 neutrons and a ^{26}Mg atom has 14 neutrons.

This diagram shows the symbol for helium, "He." The number to the upper left of the symbol is the mass number, which is 4. The number to the upper right of the symbol is the charge which is positive 2. The number to the lower left of the symbol is the atomic number, which is 2. This number is often omitted. Also shown is "Mg" which stands for magnesium. It has a mass number of 24, a charge of positive 2, and an atomic number of 12.

The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted) and the charge as a right superscript.

Information about the naturally occurring isotopes of elements with atomic numbers 1 through 10 is given in Table 3. Note that in addition to standard names and symbols, the isotopes of hydrogen are often referred to using common names and accompanying symbols. Hydrogen-2, symbolized ^2H , is also called deuterium and sometimes symbolized D. Hydrogen-3, symbolized ^3H , is also called tritium and sometimes symbolized T.

Nuclear Compositions of Atoms of the Very Light Elements						
Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
hydrogen	${}^1_1\text{H}$ (protium)	1	1	0	1.0078	99.989
	${}^2_1\text{H}$ (deuterium)	1	1	1	2.0141	0.0115
	${}^3_1\text{H}$ (tritium)	1	1	2	3.01605	(trace amounts)
helium	${}^3_2\text{He}$	2	2	1	3.01603	0.00013
	${}^4_2\text{He}$	2	2	2	4.0026	100
lithium	${}^6_3\text{Li}$	3	3	3	6.0151	7.59
	${}^7_3\text{Li}$	3	3	4	7.0160	92.41
beryllium	${}^9_4\text{Be}$	4	4	5	9.0122	100
boron	${}^{10}_5\text{B}$	5	5	5	10.0129	19.9
	${}^{11}_5\text{B}$	5	5	6	11.0093	80.1
carbon	${}^{12}_6\text{C}$	6	6	6	12.0000	98.89
	${}^{13}_6\text{C}$	6	6	7	13.0034	1.11
	${}^{14}_6\text{C}$	6	6	8	14.0032	(trace amounts)
nitrogen	${}^{14}_7\text{N}$	7	7	7	14.0031	99.63
	${}^{15}_7\text{N}$	7	7	8	15.0001	0.37
oxygen	${}^{16}_8\text{O}$	8	8	8	15.9949	99.757
	${}^{17}_8\text{O}$	8	8	9	16.9991	0.038
	${}^{18}_8\text{O}$	8	8	10	17.9992	0.205
fluorine	${}^{19}_9\text{F}$	9	9	10	18.9984	100
neon	${}^{20}_{10}\text{Ne}$	10	10	10	19.9924	90.48

Nuclear Compositions of Atoms of the Very Light Elements

Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
	${}_{10}^{21}\text{Ne}$	10	10	11	20.9938	0.27
	${}_{10}^{22}\text{Ne}$	10	10	12	21.9914	9.25

Use this [Build an Atom simulator](#) to build atoms of the first 10 elements, see which isotopes exist, check nuclear stability and gain experience with isotope symbols.

Atomic Mass

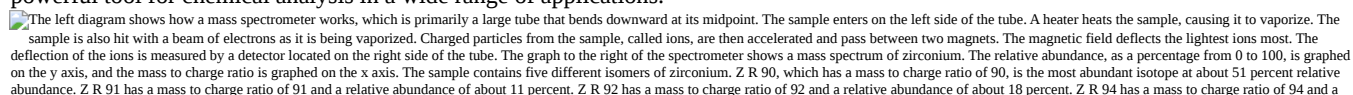
The atomic mass of a single atom is approximately equal to its mass number (a whole number) due to the mass of each proton and each neutron contributing approximately one amu each to an atom, while each electron contributes far less. However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more isotopes.

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted-average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance.

In case you are wondering

MASS SPECTROMETRY

The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer. Mass spectrometry (MS) is widely used in chemistry, forensics, medicine, environmental science, and many other fields to analyze and help identify the substances in a sample of material. In a typical mass spectrometer ([Figure 5](#)), the sample is vaporized and exposed to a high-energy electron beam that causes the sample's atoms (or molecules) to become electrically charged, typically by losing one or more electrons. These cations then pass through a (variable) electric or magnetic field that deflects each cation's path to an extent that depends on both its mass and charge (similar to how the path of a large steel ball bearing rolling past a magnet is deflected to a lesser extent than that of a small steel BB). The ions are detected, and a plot of the relative number of ions generated versus their mass-to-charge ratios (a *mass spectrum*) is made. The height of each vertical feature or peak in a mass spectrum is proportional to the fraction of cations with the specified mass-to-charge ratio. Since its initial use during the development of modern atomic theory, MS has evolved to become a powerful tool for chemical analysis in a wide range of applications.

The left diagram shows how a mass spectrometer works, which is primarily a large tube that bends downward at its midpoint. The sample enters on the left side of the tube. A heater heats the sample, causing it to vaporize. The sample is also hit with a beam of electrons as it is being vaporized. Charged particles from the sample, called ions, are then accelerated and pass between two magnets. The magnetic field deflects the lightest ions most. The deflection of the ions is measured by a detector located on the right side of the tube. The graph to the right of the spectrometer shows a mass spectrum of zirconium. The relative abundance, as a percentage from 0 to 100, is graphed on the y axis, and the mass to charge ratio is graphed on the x axis. The sample contains five different isotopes of zirconium. Z R 90, which has a mass to charge ratio of 90, is the most abundant isotope at about 51 percent relative abundance. Z R 91 has a mass to charge ratio of 91 and a relative abundance of about 11 percent. Z R 92 has a mass to charge ratio of 92 and a relative abundance of about 18 percent. Z R 94 has a mass to charge ratio of 94 and a relative abundance of about 18 percent. Z R 96, which has a mass to charge ratio of 96, is the least abundant zirconium isotope with a relative abundance of about 2 percent.

Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.

See an [animation](#) that explains mass spectrometry. Watch this [video](#) from the Royal Society for Chemistry for a brief description of the rudiments of mass spectrometry.

Chemistry End of Chapter Exercises

1. In what way are isotopes of a given element always different? In what way(s) are they always the same?
2. Write the symbol for each of the following ions:
 - (a) the ion with a 1+ charge, atomic number 55, and mass number 133
 - (b) the ion with 54 electrons, 53 protons, and 74 neutrons
 - (c) the ion with atomic number 15, mass number 31, and a 3- charge
 - (d) the ion with 24 electrons, 30 neutrons, and a 3+ charge
3. Write the symbol for each of the following ions:
 - (a) the ion with a 3+ charge, 28 electrons, and a mass number of 71
 - (b) the ion with 36 electrons, 35 protons, and 45 neutrons
 - (c) the ion with 86 electrons, 142 neutrons, and a 4+ charge
 - (d) the ion with a 2+ charge, atomic number 38, and mass number 87

4. Open the [Build an Atom simulation](#) and click on the Atom icon.
 - (a) Pick any one of the first 10 elements that you would like to build and state its symbol.
 - (b) Drag protons, neutrons, and electrons onto the atom template to make an atom of your element.
 - (c) State the numbers of protons, neutrons, and electrons in your atom, as well as the net charge and mass number.
 - (d) Click on "Net Charge" and "Mass Number," check your answers to (c), and correct, if needed.
 - (e) Predict whether your atom will be stable or unstable. State your reasoning.
 - (f) Check the "Stable/Unstable" box. Was your answer to (e) correct? If not, first predict what you can do to make a stable atom of your element, and then do it and see if it works. Explain your reasoning.
5. Open the [Build an Atom simulation](#)
 - (a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Oxygen-16 and give the isotope symbol for this atom.
 - (b) Now add two more electrons to make an ion and give the symbol for the ion you have created.
6. Open the [Build an Atom simulation](#)
 - (a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Lithium-6 and give the isotope symbol for this atom.
 - (b) Now remove one electron to make an ion and give the symbol for the ion you have created.
7. Determine the number of protons, neutrons, and electrons in the following isotopes that are used in medical diagnoses:
 - (a) atomic number 9, mass number 18, charge of 1⁻
 - (b) atomic number 43, mass number 99, charge of 7⁺
 - (c) atomic number 53, atomic mass number 131, charge of 1⁻
 - (d) atomic number 81, atomic mass number 201, charge of 1⁺
 - (e) Name the elements in parts (a), (b), (c), and (d).
8. The following are properties of isotopes of two elements that are essential in our diet. Determine the number of protons, neutrons and electrons in each and name them.
 - (a) atomic number 26, mass number 58, charge of 2⁺
 - (b) atomic number 53, mass number 127, charge of 1⁻
9. Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:
 - (a) $^{10}_{5}\text{B}$
 - (b) $^{199}_{80}\text{Hg}$
 - (c) $^{63}_{29}\text{Cu}$
 - (d) $^{13}_{6}\text{C}$
 - (e) $^{77}_{34}\text{Se}$
10. Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:
 - (a) ^7_3Li
 - (b) $^{125}_{52}\text{Te}$
 - (c) $^{109}_{47}\text{Ag}$
 - (d) $^{15}_7\text{N}$
 - (e) $^{31}_{15}\text{P}$

Glossary

anion

negatively charged atom or molecule (contains more electrons than protons)

atomic mass

average mass of atoms of an element, expressed in amu

atomic mass unit (amu)

(also, unified atomic mass unit, u, or Dalton, Da) unit of mass equal to 1/12 of the mass of a ^{12}C atom

atomic number (Z)

number of protons in the nucleus of an atom

cation

positively charged atom or molecule (contains fewer electrons than protons)

chemical symbol

one-, two-, or three-letter abbreviation used to represent an element or its atoms

Dalton (Da)

alternative unit equivalent to the atomic mass unit

fundamental unit of charge

(also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19} \text{ C}$

ion

electrically charged atom or molecule (contains unequal numbers of protons and electrons)

isotope

same number of protons and different number of neutrons

mass number (A)

sum of the numbers of neutrons and protons in the nucleus of an atom

unified atomic mass unit (u)

alternative unit equivalent to the atomic mass unit

4.5: Atomic Structure is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

5: Unit Conversions

5.1: Unit conversions

5.2: Density

5: Unit Conversions is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

5.1: Unit conversions

13

Harper College Chemistry Department

Page by: OpenStax

▼ Summary

By the end of this section, you will be able to:

- Explain the dimensional analysis (factor label) approach to mathematical calculations involving quantities
- Use dimensional analysis to carry out unit conversions for a given property and computations involving two or more properties

It is often the case that a quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships. For example, consider measuring the average speed of an athlete running sprints. This is typically accomplished by measuring the *time* required for the athlete to run from the starting line to the finish line, and the *distance* between these two lines, and then computing *speed* from the equation that relates these three properties:

$$\text{speed} = \frac{\text{distance}}{\text{time}}$$

An Olympic-quality sprinter can run 100 m in approximately 10 s, corresponding to an average speed of

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100 \text{ m}}{10 \text{ s}} = \frac{10 \text{ m}}{1 \text{ s}} = 10 \text{ m/s}$$

Note that this simple arithmetic involves dividing the numbers of each measured quantity to yield the number of the computed quantity ($100/10 = 10$) *and likewise* dividing the units of each measured quantity to yield the unit of the computed quantity ($\text{m/s} = \text{m/s}$). Now, consider using this same relation to predict the time required for a person running at this speed to travel a distance of 25 m. The same relation between the three properties is used, but in this case, the two quantities provided are a speed (10 m/s) and a distance (25 m). To yield the sought property, time, the equation must be rearranged appropriately:

$$\text{time} = \frac{\text{distance}}{\text{speed}} = \frac{25 \text{ m}}{10 \text{ m/s}} = 2.5 \text{ s}$$

or

$$\text{time} = (\text{distance m}) \times (\text{speed}^{-1} \text{ s/m}) = \text{s}$$

The time can then be computed as:

$$\text{time s} = \text{distance m} \times \frac{\text{time s}}{\text{distance m}}$$

$$\text{time s} = 25 \text{ m} \times \frac{1 \text{ s}}{10 \text{ m}} = 2.5 \text{ s}$$

Again, arithmetic on the numbers ($25/10 = 2.5$) was accompanied by the same arithmetic on the units ($\text{m/m} = 1$) to yield the number and unit of the result, 2.5 s. Note that, just as for numbers, when a unit is divided by an identical unit (in this case, m/m), the result is “1”—or, as commonly phrased, the units “cancel.”

These calculations are examples of a versatile mathematical approach known as dimensional analysis (or the factor-label method). Dimensional analysis is based on this premise: *the units of quantities must be subjected to the same mathematical operations as their associated numbers*. This method can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities.

Conversion Factors and Dimensional Analysis

A ratio of two equivalent quantities expressed with different measurement units can be used as a unit conversion factor. For example, the lengths of 2.54 cm and 1 in. are equivalent (by definition), and so a unit conversion factor may be derived from the ratio. So from

$$2.54 \text{ cm} = 1 \text{ in}$$

we get

$\frac{\text{2.54 cm}}{\text{1 in}}$ and $\frac{\text{1 in}}{\text{2.54 cm}}$ which can be used as conversion factors.

Several other commonly used conversion factors are given in [Table 1](#).

Table 1: Common Conversion Factors		
Length	Volume	Mass
1 m = 1.0936 yd	1 L = 1.0567 qt	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft ³ = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1 tbsp = 14.787 mL	1 (troy) oz = 31.103 g

When we multiply a quantity (such as distance given in inches) by an appropriate unit conversion factor, we convert the quantity to an equivalent value with different units (such as distance in centimeters). For example, a basketball player's vertical jump of 34 inches can be converted to centimeters by:

$$34 \cancel{\text{in.}} \times \frac{2.54 \text{ cm}}{\cancel{\text{in.}}} = 86 \text{ cm}$$

Since this simple arithmetic involves *quantities*, the premise of dimensional analysis requires that we multiply both *numbers and units*. The numbers of these two quantities are multiplied to yield the number of the product quantity, 86, whereas the units are multiplied to yield $\frac{\text{in.} \times \text{cm}}{\text{in.}}$. Just as for numbers, a ratio of identical units is also numerically equal to one, $\frac{\text{in.}}{\text{in.}} = 1$, and the unit product thus simplifies to *cm*. (When identical units divide to yield a factor of 1, they are said to “cancel.”) Using dimensional analysis, we can determine that a unit conversion factor has been set up correctly by checking to confirm that the original unit will cancel, and the result will contain the sought (converted) unit.

Using a Unit Conversion Factor: The mass of a competition frisbee is 125 g. Convert its mass to ounces using the unit conversion factor derived from the relationship 1 oz = 28.349 g ([Table 1](#)).

Solution: If we have the conversion factor, we can determine the mass in ounces using an equation similar the one used for converting length from inches to centimeters.

$$125 \text{ g} \times \frac{1 \text{ oz}}{28.349 \text{ g}} = 4.41 \text{ oz}$$

We write the unit conversion factor in its two forms:

$$\frac{1 \text{ oz}}{28.349 \text{ g}} \text{ and } \frac{28.349 \text{ g}}{1 \text{ oz}}$$

The correct unit conversion factor is the ratio that cancels the units of grams and leaves ounces.

$$\begin{aligned} x \text{ oz} &= 125 \cancel{\text{g}} \times \frac{1 \text{ oz}}{28.349 \cancel{\text{g}}} \\ &= \left(\frac{125}{28.349} \right) \text{ oz} \\ &= 4.41 \text{ oz (three significant figures)} \end{aligned}$$

Check Your Learning: Convert a volume of 9.345 qt to liters.

Answer:

8.844 L

Beyond simple unit conversions, the factor-label method can be used to solve more complex problems involving computations. Regardless of the details, the basic approach is the same—all the *factors* involved in the calculation must be appropriately oriented to insure that their *labels* (units) will appropriately cancel and/or combine to yield the desired unit in the result. This is why it is referred to as the factor-label method. As your study of chemistry continues, you will encounter many opportunities to apply this approach.

Computing Quantities from Measurement Results and Known Mathematical Relations:

What is the density of common antifreeze in units of g/mL? A 4.00-qt sample of the antifreeze weighs 9.26 lb.

Solution: Since $\text{density} = \frac{\text{mass}}{\text{volume}}$, we need to divide the mass in grams by the volume in milliliters. In general: the number of units of B = the number of units of A $\times$ unit conversion factor. The necessary conversion factors are given in Table 1: 1 lb = 453.59 g; 1 L = 1.0567 qt; 1 L = 1,000 mL. We can convert mass from pounds to grams in one step:

$$9.26 \text{ lb} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 4.20 \times 10^3 \text{ g} \text{ (need 3 significant figures in the answer)}$$

We need to use two steps to convert volume from quarts to milliliters.

$$1. \text{ Convert quarts to liters. } 4.00 \text{ qt} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} = 3.78 \text{ L}$$

2. Convert liters to milliliters.

$$3.78 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 3.78 \times 10^3 \text{ mL} = 3780 \text{ mL}$$

Then,

$$\text{density} = \frac{4.20 \times 10^3 \text{ g}}{3.78 \times 10^3 \text{ mL}} = 1.11 \text{ g/mL}$$

Alternatively, the calculation could be set up in a way that uses three unit conversion factors sequentially as follows:

$$\frac{9.26 \text{ lb}}{4.00 \text{ qt}} \times \frac{453.59 \text{ g}}{1 \text{ lb}} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.11 \text{ g/mL}$$

Check Your Learning: What is the volume in liters of 1.000 oz, given that 1 L = 1.0567 qt and 1 qt = 32 oz (exactly)?

Answer:

$$2.956 \times 10^{-2} \text{ L}$$

Computing Quantities from Measurement Results and Known Mathematical Relations:

While being driven from Philadelphia to Atlanta, a distance of about 1250 km, a 2014 Lamborghini Aventador Roadster uses 213 L gasoline.

(a) What (average) fuel economy, in miles per gallon, did the Roadster get during this trip?

(b) If gasoline costs \$3.80 per gallon, what was the fuel cost for this trip?

Solution

(a) We first convert distance from kilometers to miles:

$$1250 \text{ km} \times \frac{0.62137 \text{ mi}}{1 \text{ km}} = 777 \text{ mi}$$

and then convert volume from liters to gallons:

$$213 \text{ L} \times \frac{1.0567 \text{ qt}}{1 \text{ L}} \times \frac{1 \text{ gal}}{4 \text{ qt}} = 56.3 \text{ gal}$$

$$213 \text{ L} \times 1.0567 \text{ qt/L} \times 1 \text{ gal/4qt} = 56.3 \text{ gal}$$

Then,

$$\text{(average) mileage} = \frac{777 \text{ mi}}{56.3 \text{ gal}} = 13.8 \text{ miles/gallon} = 13.8 \text{ mpg (average)}$$

$$\text{(average) mileage} = 777 \text{ mi} / 56.3 \text{ gal} = 13.8 \text{ miles/gallon} = 13.8 \text{ mpg (average)}$$

Alternatively, the calculation could be set up in a way that uses all the conversion factors sequentially, as follows:

$$\frac{1250 \text{ km}}{213 \text{ L}} \times \frac{0.62137 \text{ mi}}{1 \text{ km}} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} \times \frac{4 \text{ qt}}{1 \text{ gal}} = 13.8 \text{ mpg}$$

$$1250 \text{ km} / 213 \text{ L} \times 0.62137 \text{ mi/km} \times 1 \text{ L} / 1.0567 \text{ qt} \times 4 \text{ qt} / 1 \text{ gal} = 13.8 \text{ mpg}$$

(b) Using the previously calculated volume in gallons, we find:

$$\cancel{\text{SI}}56.3\text{gal} \times \frac{\$3.80}{\cancel{\text{SI}}1\text{gal}} = \$214 \quad (5.1.1)$$

Check Your Learning: A Toyota Prius Hybrid uses 59.7 L gasoline to drive from San Francisco to Seattle, a distance of 1300 km (two significant digits).

(a) What (average) fuel economy, in miles per gallon, did the Prius get during this trip?

(b) If gasoline costs \$ 3.90 per gallon, what was the fuel cost for this trip?

Answer:

(a) 51 mpg; (b) \$ 62

Key Concepts and Summary

Measurements are made using a variety of units. It is often useful or necessary to convert a measured quantity from one unit into another. These conversions are accomplished using unit conversion factors, which are derived by simple applications of a mathematical approach called the factor-label method or dimensional analysis. This strategy is also employed to calculate sought quantities using measured quantities and appropriate mathematical relations.

Key Equations

- $T^{\circ}\text{C} = (5/9 \times T^{\circ}\text{F}) - 32$ $T^{\circ}\text{C} = (5/9 \times T^{\circ}\text{F}) - 32$
- $T^{\circ}\text{F} = (9/5 \times T^{\circ}\text{C}) + 32$ $T^{\circ}\text{F} = (9/5 \times T^{\circ}\text{C}) + 32$
- $\text{TK} = ^{\circ}\text{C} + 273.15$ $\text{TK} = ^{\circ}\text{C} + 273.15$
- $T^{\circ}\text{C} = \text{K} - 273.15$ $T^{\circ}\text{C} = \text{K} - 273.15$

Chemistry End of Chapter Exercises

1. Write conversion factors (as ratios) for the number of:

- (a) yards in 1 meter
- (b) liters in 1 liquid quart
- (c) pounds in 1 kilogram

2. Write conversion factors (as ratios) for the number of:

- (a) kilometers in 1 mile
- (b) liters in 1 cubic foot
- (c) grams in 1 ounce

3. The label on a soft drink bottle gives the volume in two units: 2.0 L and 67.6 fl oz. Use this information to derive a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

4. The label on a box of cereal gives the mass of cereal in two units: 978 grams and 34.5 oz. Use this information to find a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

5. Soccer is played with a round ball having a circumference between 27 and 28 in. and a weight between 14 and 16 oz. What are these specifications in units of centimeters and grams?

6. A woman's basketball has a circumference between 28.5 and 29.0 inches and a maximum weight of 20 ounces (two significant figures). What are these specifications in units of centimeters and grams?

7. How many milliliters of a soft drink are contained in a 12.0-oz can?

8. A barrel of oil is exactly 42 gal. How many liters of oil are in a barrel?

9. The diameter of a red blood cell is about 3×10^{-4} in. What is its diameter in centimeters?

10. The distance between the centers of the two oxygen atoms in an oxygen molecule is 1.21×10^{-8} cm. What is this distance in inches?

11. Is a 197-lb weight lifter light enough to compete in a class limited to those weighing 90 kg or less?
12. A very good 197-lb weight lifter lifted 192 kg in a move called the clean and jerk. What was the mass of the weight lifted in pounds?
13. Many medical laboratory tests are run using 5.0 μL blood serum. What is this volume in milliliters?
14. If an aspirin tablet contains 325 mg aspirin, how many grams of aspirin does it contain?
15. Use scientific (exponential) notation to express the following quantities in terms of the SI base units in [\[link\]](#):
 - (a) 0.13 g
 - (b) 232 Gg
 - (c) 5.23 pm
 - (d) 86.3 mg
 - (e) 37.6 cm
 - (f) 54 μm
 - (g) 1 Ts
 - (h) 27 ps
 - (i) 0.15 mK
16. Complete the following conversions between SI units.
 - (a) 612 g = _____ mg
 - (b) 8.160 m = _____ cm
 - (c) 3779 μg = _____ g
 - (d) 781 mL = _____ L
 - (e) 4.18 kg = _____ g
 - (f) 27.8 m = _____ km
 - (g) 0.13 mL = _____ L
 - (h) 1738 km = _____ m
 - (i) 1.9 Gg = _____ g
17. Gasoline is sold by the liter in many countries. How many liters are required to fill a 12.0-gal gas tank?
18. Milk is sold by the liter in many countries. What is the volume of exactly $\frac{1}{2}$ gal of milk in liters?
19. A long ton is defined as exactly 2240 lb. What is this mass in kilograms?
20. Make the conversion indicated in each of the following:
 - (a) the men's world record long jump, 29 ft $4\frac{1}{4}$ in., to meters
 - (b) the greatest depth of the ocean, about 6.5 mi, to kilometers
 - (c) the area of the state of Oregon, 96,981 mi^2 , to square kilometers
 - (d) the volume of 1 gill (exactly 4 oz) to milliliters
 - (e) the estimated volume of the oceans, 330,000,000 mi^3 , to cubic kilometers.
 - (f) the mass of a 3525-lb car to kilograms
 - (g) the mass of a 2.3-oz egg to grams
21. Make the conversion indicated in each of the following:

- (a) the length of a soccer field, 120 m (three significant figures), to feet
- (b) the height of Mt. Kilimanjaro, at 19,565 ft the highest mountain in Africa, to kilometers
- (c) the area of an 8.5×11 -inch sheet of paper in cm^2
- (d) the displacement volume of an automobile engine, 161 in.^3 , to liters
- (e) the estimated mass of the atmosphere, 5.6×10^{15} tons, to kilograms
- (f) the mass of a bushel of rye, 32.0 lb, to kilograms
- (g) the mass of a 5.00-grain aspirin tablet to milligrams (1 grain = 0.00229 oz)

22. Many chemistry conferences have held a 50-Trillion Angstrom Run (two significant figures). How long is this run in kilometers and in miles? ($1 \text{ \AA} = 1 \times 10^{-10} \text{ m}$)

23. A chemist's 50-Trillion Angstrom Run (see [Exercise](#)) would be an archeologist's 10,900 cubit run. How long is one cubit in meters and in feet? ($1 \text{ \AA} = 1 \times 10^{-8} \text{ cm}$)

24. The gas tank of a certain luxury automobile holds 22.3 gallons according to the owner's manual. If the density of gasoline is 0.8206 g/mL, determine the mass in kilograms and pounds of the fuel in a full tank.

25. As an instructor is preparing for an experiment, he requires 225 g phosphoric acid. The only container readily available is a 150-mL Erlenmeyer flask. Is it large enough to contain the acid, whose density is 1.83 g/mL?

26. To prepare for a laboratory period, a student lab assistant needs 125 g of a compound. A bottle containing 1/4 lb is available. Did the student have enough of the compound?

27. A chemistry student is 159 cm tall and weighs 45.8 kg. What is her height in inches and weight in pounds?

28. In a recent Grand Prix, the winner completed the race with an average speed of 229.8 km/h. What was his speed in miles per hour, meters per second, and feet per second?

29. Solve these problems about lumber dimensions.

(a) To describe to a European how houses are constructed in the US, the dimensions of "two-by-four" lumber must be converted into metric units. The thickness $\times$ width $\times$ length dimensions are 1.50 in. $\times$ 3.50 in. $\times$ 8.00 ft in the US. What are the dimensions in cm $\times$ cm $\times$ m?

(b) This lumber can be used as vertical studs, which are typically placed 16.0 in. apart. What is that distance in centimeters?

30. The mercury content of a stream was believed to be above the minimum considered safe—1 part per billion (ppb) by weight. An analysis indicated that the concentration was 0.68 parts per billion. What quantity of mercury in grams was present in 15.0 L of the water, the density of which is 0.998 g/mL? (1 ppb Hg = 1 ng Hg/1 g water)

31. Calculate the density of aluminum if 27.6 cm^3 has a mass of 74.6 g.

32. Osmium is one of the densest elements known. What is its density if 2.72 g has a volume of 0.121 cm^3 ?

33. Calculate these masses.

(a) What is the mass of 6.00 cm^3 of mercury, density = 13.5939 g/ cm^3 ?

(b) What is the mass of 25.0 mL octane, density = 0.702 g/ cm^3 ?

34. Calculate these masses.

(a) What is the mass of 4.00 cm^3 of sodium, density = 0.97 g/ cm^3 ?

(b) What is the mass of 125 mL gaseous chlorine, density = 3.16 g/L?

Calculate these volumes.

(a) What is the volume of 25 g iodine, density = 4.93 g/ cm^3 ?

(b) What is the volume of 3.28 g gaseous hydrogen, density = 0.089 g/L?

35. Calculate these volumes.

(a) What is the volume of 11.3 g graphite, density = 2.25 g/cm^3 ?

(b) What is the volume of 39.657 g bromine, density = 2.928 g/cm^3 ?

36. Convert the boiling temperature of gold, 2966°C , into degrees Fahrenheit and kelvin.

37. Convert the temperature of scalding water, 54°C , into degrees Fahrenheit and kelvin.

38. Convert the temperature of the coldest area in a freezer, -10°F , to degrees Celsius and kelvin.

39. Convert the temperature of dry ice, -77°C , into degrees Fahrenheit and kelvin.

40. Convert the boiling temperature of liquid ammonia, -28.1°F , into degrees Celsius and kelvin.

41. The label on a pressurized can of spray disinfectant warns against heating the can above 130°F . What are the corresponding temperatures on the Celsius and kelvin temperature scales?

42. The weather in Europe was unusually warm during the summer of 1995. The TV news reported temperatures as high as 45°C . What was the temperature on the Fahrenheit scale?

Glossary

dimensional analysis

(also, factor-label method) versatile mathematical approach that can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities

Fahrenheit

unit of temperature; water freezes at 32°F and boils at 212°F on this scale

unit conversion factor

ratio of equivalent quantities expressed with different units; used to convert from one unit to a different unit

5.1: Unit conversions is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

5.2: Density

14

Harper College Chemistry Department

We use the mass and volume of a substance to determine its density. Thus, the units of density are defined by the base units of mass and length.

The density of a substance is the ratio of the mass of a sample of the substance to its volume. The SI unit for density is the kilogram per cubic meter (kg/m^3). For many situations, however, this is an inconvenient unit, and we often use grams per cubic centimeter (g/cm^3) for the densities of solids and liquids, and grams per liter (g/L) for gases. Although there are exceptions, most liquids and solids have densities that range from about 0.7 g/cm^3 (the density of gasoline) to 19 g/cm^3 (the density of gold). The density of air is about 1.2 g/L . Table 1 shows the densities of some common substances.

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

While there are many ways to determine the density of an object, perhaps the most straightforward method involves separately finding the mass and volume of the object, and then dividing the mass of the sample by its volume.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (5.2.1)$$

\]

In the following example, the mass is found directly by weighing, but the volume is found indirectly through length measurements.

Calculation of Density Gold—in bricks, bars, and coins—has been a form of currency for centuries. In order to swindle people into paying for a brick of gold without actually investing in a brick of gold, people have considered filling the centers of hollow gold bricks with lead to fool buyers into thinking that the entire brick is gold. It does not work: Lead is a dense substance, but its density is not as great as that of gold, 19.3 g/cm^3 .

What is the density of lead if a cube of lead has an edge length of 2.00 cm and a mass of 90.7 g ?

Solution The density of a substance can be calculated by dividing its mass by its volume. The volume of a cube is calculated by cubing the edge length.

$$\text{volume of lead cube} = 2.00 \text{ cm} \times 2.00 \text{ cm} \times 2.00 \text{ cm} = 8.00 \text{ cm}^3 \quad (5.2.2)$$

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{90.7 \text{ g}}{8.00 \text{ cm}^3} = \frac{11.3 \text{ g}}{1.00 \text{ cm}^3} = 11.3 \text{ g/cm}^3 \quad (5.2.3)$$

\]

(We will discuss the reason for rounding to the first decimal place in the next section.)

Check Your Learning

- (a) To three decimal places, what is the volume of a cube (cm^3) with an edge length of 0.843 cm?
- (b) If the cube in part (a) is copper and has a mass of 5.34 g, what is the density of copper to two decimal places?

Answer: (a) 0.599 cm^3 ; (b) 8.91 g/cm^3

To learn more about the relationship between mass, volume, and density, use this [interactive simulator](#) to explore the density of different materials, like wood, ice, brick, and aluminum.

Using Displacement of Water to Determine Density

This [PhET simulation](#) illustrates another way to determine density, using displacement of water. Determine the density of the red and yellow blocks.

Solution

When you open the density simulation and select Same Mass, you can choose from several 5.00-kg colored blocks that you can drop into a tank containing 100.00 L water. The yellow block floats (it is less dense than water), and the water level rises to 105.00 L. While floating, the yellow block displaces 5.00 L water, an amount equal to the weight of the block. The red block sinks (it is more dense than water, which has density = 1.00 kg/L), and the water level rises to 101.25 L.

The red block therefore displaces 1.25 L water, an amount equal to the volume of the block. The density of the red block is:

$$\rho_{\text{density}} = \frac{\text{mass}}{\text{volume}} = \frac{5.00 \text{ kg}}{1.25 \text{ L}} = 4.00 \text{ kg/L}$$

Note that since the yellow block is not completely submerged, you cannot determine its density from this information. But if you hold the yellow block on the bottom of the tank, the water level rises to 110.00 L, which means that it now displaces 10.00 L water, and its density can be found:

$$\rho_{\text{density}} = \frac{\text{mass}}{\text{volume}} = \frac{5.00 \text{ kg}}{10.00 \text{ L}} = 0.500 \text{ kg/L}$$

Check Your Learning

Remove all of the blocks from the water and add the green block to the tank of water, placing it approximately in the middle of the tank. Determine the density of the green block.

Answer: 2.00 kg/L

Experimental Determination of Density Using Water Displacement

A piece of rebar is weighed and then submerged in a graduated cylinder partially filled with water, with results as shown.

 This diagram shows the initial volume of water in a graduated cylinder as 13.5 milliliters. A 69.658 gram piece of metal rebar is added to the graduated cylinder, causing the water to reach a final volume of 22.4 milliliters.

- (a) Use these values to determine the density of this piece of rebar.
- (b) Rebar is mostly iron. Does your result in (a) support this statement? How?

Solution

The volume of the piece of rebar is equal to the volume of the water displaced:

$$\text{volume} = 22.4 \text{ mL} - 13.5 \text{ mL} = 8.9 \text{ mL} = 8.9 \text{ cm}^3$$

(rounded to the nearest 0.1 mL, per the rule for addition and subtraction)

The density is the mass-to-volume ratio:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{69.658 \text{ g}}{8.9 \text{ cm}^3} = 7.8 \text{ g/cm}^3$$

(rounded to two significant figures, per the rule for multiplication and division)

From [link](#), the density of iron is 7.9 g/cm^3 , very close to that of rebar, which lends some support to the fact that rebar is mostly iron.

Check Your Learning

An irregularly shaped piece of a shiny yellowish material is weighed and then submerged in a graduated cylinder, with results as shown.

 This diagram shows the initial volume of water in a graduated cylinder as 17.1 milliliters. A 51.842 gram gold colored rock is added to the graduated cylinder, causing the water to reach a final volume of 19.8 milliliters

(a) Use these values to determine the density of this material.

(b) Do you have any reasonable guesses as to the identity of this material? Explain your reasoning.

Answer: (a) 19 g/cm^3 ; (b) It is likely gold; the right appearance for gold and very close to the density given for gold in [\[link\]](#).

Key Concepts and Summary

Measurements provide quantitative information that is critical in studying and practicing chemistry. Each measurement has an amount, a unit for comparison, and an uncertainty. Measurements can be represented in either decimal or scientific notation. Scientists primarily use the SI (International System) or metric systems. We use base SI units such as meters, seconds, and kilograms, as well as derived units, such as liters (for volume) and g/cm^3 (for density). In many cases, we find it convenient to use unit prefixes that yield fractional and multiple units, such as microseconds (10^{-6} seconds) and megahertz (10^6 hertz), respectively.

[\[1\]](#)

Key Equations

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

-
1. OpenStax Chemistry: Atom First, Chemistry: Atoms First. OpenStax CNX. Nov 13, 2018. Download for free at <http://cnx.org/contents/4539ae23-1ccc-421e-9b25-843acbb6c4b0@9.17>. [↵](#)
-

5.2: Density is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

6: Moles

[6.1: Moles](#)

[6.2: Formula mass](#)

[6.3: Converting between grams and moles](#)

6: Moles is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

6.1: Moles

15

Harper College Chemistry Department

Page by: OpenStax Chemistry: Atoms First

The Mole

The identity of a substance is defined not only by the types of atoms or ions it contains, but by the quantity of each type of atom or ion. For example, water, H_2O , and hydrogen peroxide, H_2O_2 , are alike in that their respective molecules are composed of hydrogen and oxygen atoms. However, because a hydrogen peroxide molecule contains two oxygen atoms, as opposed to the water molecule, which has only one, the two substances exhibit very different properties. Today, we possess sophisticated instruments that allow the direct measurement of these defining microscopic traits; however, the same traits were originally derived from the measurement of macroscopic properties (the masses and volumes of bulk quantities of matter) using relatively simple tools (balances and volumetric glassware). This experimental approach required the introduction of a new unit for the amount of substances, the *mole*, which remains indispensable in modern chemical science.

The mole is an amount unit similar to familiar units like pair, dozen, gross, etc. It provides a specific measure of *the number* of atoms or molecules in a bulk sample of matter. A mole is defined as the amount of substance containing the same number of discrete entities (such as atoms, molecules and ions) as the number of atoms in a sample of pure ^{12}C whose mass is exactly 12.000 g. One Latin connotation for the word “mole” is “large mass” or “bulk,” which is consistent with its use as the name for this unit. The mole provides a link between an easily measured macroscopic property, bulk mass and an extremely important fundamental property, number of atoms, molecules and so forth.

The number of entities composing a mole has been experimentally determined to be $6.02214076 \times 10^{23}$, a fundamental constant named Avogadro’s number (N_A) or the Avogadro constant in honor of Italian scientist Amedeo Avogadro. This constant is properly reported with an explicit unit of “per mole,” a conveniently rounded version being $6.022 \times 10^{23}/\text{mol}$.

Consistent with its definition as an amount unit, 1 mole of any element contains the same number of atoms as 1 mole of any other element. The masses of 1 mole of different elements, however, are different; since the masses of the individual atoms are drastically different. The molar mass of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol) (see [Figure 1](#)).

 This figure contains eight different substances displayed on white circles. The amount of each substance is visibly different.

Figure 1 Each sample contains 6.022×10^{23} atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead and 118.7 g tin. (credit: modification of work by Mark Ott)

Because the definitions of both the mole and the atomic mass unit are based on the same reference substance, ^{12}C , the molar mass of any substance is numerically equivalent to its atomic or formula weight in amu. Per the amu definition, a single ^{12}C atom weighs 12 amu (its atomic mass is 12 amu). According to the definition of the mole, 12 g of ^{12}C contains 1 mole of ^{12}C atoms (its molar mass is 12 g/mol). This relationship holds for all elements, since their atomic masses are measured relative to that of the amu-reference substance, ^{12}C . Extending this principle, the molar mass of a compound in grams is likewise numerically equivalent to its formula mass in amu ([Figure 2](#)).

 This photo shows two vials filled with a colorless liquid. It also shows two bowls: one filled with an off-white powder and one filled with a bright red powder.

Figure 2 Each sample contains 6.02×10^{23} molecules or formula units—1.00 mol of the compound or element. Clock-wise from the upper left: 130.2 g of $\text{C}_8\text{H}_{17}\text{OH}$ (1-octanol, formula mass 130.2 amu), 454.4 g of HgI_2 (mercury(II) iodide, formula mass 454.4 amu), 32.0 g of CH_3OH (methanol, formula mass 32.0 amu) and 256.5 g of S_8 (sulfur, formula mass 256.5 amu). (credit: Sahar Atwa)

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022×10^{23}
H	1.008	1.008	6.022×10^{23}
O	16.00	16.00	6.022×10^{23}
Na	22.99	22.99	6.022×10^{23}
Cl	35.45	35.45	6.022×10^{23}

While atomic mass and molar mass are numerically equivalent, keep in mind that they are vastly different in terms of scale, as represented by the vast difference in the magnitudes of their respective units (amu versus g). To appreciate the enormity of the mole, consider a small drop of water with a mass of about 0.03 g (see [Figure 3](#)). Although this represents just a tiny fraction of 1 mole of water (~18.02 g), it contains more water molecules than can be clearly imagined. If the molecules were distributed equally among the roughly seven billion people on earth, each person would receive more than 100 billion molecules.

 A close-up photo of a water droplet on a leaf is shown. The water droplet is not perfectly spherical.

Figure 3 The number of molecules in a single droplet of water is roughly 100 billion times greater than the number of people on earth. (credit: "tanakawho"/Wikimedia commons)

LINK TO LEARNING

The mole is used in chemistry to represent 6.022×10^{23} of something, but it can be difficult to conceptualize such a large number. Watch this [video](#) and then complete the "Think" questions that follow. Explore more about the mole by reviewing the information under "Dig Deeper."

The relationships between formula mass, the mole and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds. For example, if we know the mass and chemical composition of a substance, we can determine the number of moles and calculate number of atoms or molecules in the sample. Likewise, if we know the number of moles of a substance, we can derive the number of atoms or molecules and calculate the substance's mass.

6.1: Moles is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

6.2: Formula mass

16

Harper College Chemistry Department

Page by: OpenStax Chemistry: Atoms First

▼ Summary

By the end of this section, you will be able to:

- Calculate formula masses for covalent and ionic compounds

We can argue that modern chemical science began when scientists started exploring the quantitative as well as the qualitative aspects of chemistry. For example, Dalton's atomic theory was an attempt to explain the results of measurements that allowed him to calculate the relative masses of elements combined in various compounds. Understanding the relationship between the masses of atoms and the chemical formulas of compounds allows us to quantitatively describe the composition of substances.

Formula Mass

In an earlier chapter, we described the development of the atomic mass unit, the concept of average atomic masses and the use of chemical formulas to represent the elemental makeup of substances. These ideas can be extended to calculate the formula mass of a substance by adding the average atomic masses of all the atoms represented in the substance's formula.

Formula Mass for Covalent Substances

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a molecular mass. Consider chloroform (CHCl_3), a covalent compound once used as a surgical anesthetic and now primarily used in the production of the "anti-stick" polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms. [Figure 1](#) outlines the calculations used to derive the molecular mass of chloroform, which is 119.37 amu.

A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "C," "H," "Cl" and a blank, merged cell that runs the width of the first five columns. The second column contains the numbers "1," "1," and "3" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "12.01," "1.008," and "35.45" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "12.01," "1.008," "106.35," and "119.37." There is a thick black line below the number 106.35. The merged cell under the first five columns reads "Molecular mass." To the left of the table is a diagram of a molecule. Three green spheres are attached to a slightly smaller black sphere, which is also attached to a smaller white sphere. The green spheres lie beneath and to the sides of the black sphere while the white sphere is located straight up from the black sphere.

The average mass of a chloroform molecule, CHCl_3 , is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.

Likewise, the molecular mass of an aspirin molecule, $\text{C}_9\text{H}_8\text{O}_4$, is the sum of the atomic masses of nine carbon atoms, eight hydrogen atoms and four oxygen atoms, which amounts to 180.15 amu ([Figure 2](#)).

A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "C," "H," "O," and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers "9," "8," and "4" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "12.01," "1.008," and "16.00" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "108.09," "8.064," "64.00," and "180.15." There is a thick black line below the number 64.00. The merged cell under the first five columns reads "Molecular mass." To the left of the table is a diagram of a molecule. Six black spheres are located in a six-sided ring and connected by alternating double and single black bonds. Attached to each of the four black spheres is one smaller white sphere. Attached to the farthest right black sphere is a red sphere, connected to two more black spheres, all in a row. Attached to the last black sphere of that row are two more white spheres. Attached to the first black sphere of that row is another red sphere. A black sphere, attached to two red spheres and a white sphere is attached to the black sphere on the top right of the six-sided ring.

The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $\text{C}_9\text{H}_8\text{O}_4$.

Computing Molecular Mass for a Covalent Compound:

Ibuprofen, $\text{C}_{13}\text{H}_{18}\text{O}_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Advil and Motrin. What is the molecular mass (amu) for this compound?

Solution: Molecules of this compound are comprised of 13 carbon atoms, 18 hydrogen atoms, and 2 oxygen atoms. Following the approach described above, the average molecular mass for this compound is therefore:

A table is shown that is made up of six columns and five rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "C," "H," "O," and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers "13," "8," and "2" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "12.01," "1.008," and "16.00" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "156.13," "18.114," "32.00," and "206.27." There is a thick black line below the number 32.00. The merged cell under the first five columns reads "Molecular mass." To the right is a ball-and-stick model of the structure. At the center, it shows six black spheres arranged in a six-sided ring with alternating double bonds. The two black spheres at the top and bottom of the six-sided ring are each bonded to one, smaller, white sphere. The black sphere on the left side of the six-sided ring is connected to another black sphere. This sphere is connected to two smaller, white spheres and another black sphere. This black sphere is connected to one, smaller white sphere, and two other black spheres. Each of these last two black spheres is connected to two smaller, white spheres. The black sphere on the right side of the six-sided ring is connected to another black sphere. This black sphere is connected to one smaller, white sphere and two other black spheres. The black sphere that is connected to it and is situated to the top right is connected to two smaller, white spheres. The black sphere connected towards the bottom right is connected to two red spheres. It forms a double bond with one of these red spheres and the other red sphere is connected to a smaller, white sphere.

Check Your Learning

Acetaminophen, $C_8H_9NO_2$, is a covalent compound and the active ingredient in several popular nonprescription pain medications, such as Tylenol. What is the molecular mass (amu) for this compound?

Answer: 151.16 amu

Formula Mass for Ionic Compounds

Ionic compounds are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter. The formula mass for an ionic compound is calculated in the same way as the formula mass for covalent compounds: by adding the average atomic masses of all the atoms in the compound's formula. Keep in mind, however, that the formula for an ionic compound does not represent the composition of a discrete molecule, so it may not correctly be referred to as the "molecular mass."

As an example, consider sodium chloride, NaCl, the chemical name for common table salt. Sodium chloride is an ionic compound composed of sodium cations, Na^+ , and chloride anions, Cl^- , combined in a 1:1 ratio. The formula mass for this compound is computed as 58.44 amu (Figure 3).

A table and diagram are shown. The table is made up of six columns and four rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "Na," "Cl," and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers "1" and "1" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "22.99" and "35.45" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "22.99," "35.45," and "58.44." There is a thick black line below the number "35.45." The merged cell under the first five columns reads "Formula mass." To the left of the table is a diagram of a chemical structure. The diagram shows green and purple spheres placed in an alternating pattern, making up the corners of eight stacked cubes to form one larger cube. The green spheres are slightly smaller than the purple spheres.

Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.

Note that the average masses of neutral sodium and chlorine atoms were used in this computation, rather than the masses for sodium cations and chlorine anions. This approach is perfectly acceptable when computing the formula mass of an ionic compound. Even though, a sodium cation has a slightly smaller mass than a sodium atom (since it is missing an electron), this difference will be offset by the fact that a chloride anion is slightly more massive than a chloride atom (due to the extra electron). Moreover, the mass of an electron is negligibly small with respect to the mass of a typical atom. Even when calculating the mass of an isolated ion, the missing or additional electrons can generally be ignored, since their contribution to the overall mass is negligible, reflected only in the non-significant digits that will be lost when the computed mass is properly rounded. The few exceptions to this guideline are very light ions derived from elements with precisely known atomic masses.

Computing Formula Mass for an Ionic Compound:

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution:

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. Three sulfate ions have 3 sulfurs and 12 oxygens (3 ions $\times$ 4 oxygens per ion). For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

A table is shown that is made up of six columns and five rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "Al," "S," "O," and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers "2," "3," and "12" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "26.98," "32.06," and "16.00" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "53.96," "96.18," "192.00," and "342.14." There is a thick black line under the number 192.00. The merged cell under the first five columns reads "Molecular mass." To the right of this table is a ball-and-stick structure. It shows yellow and grey spheres connected to red spheres in a complex pattern. The yellow and grey spheres are similar in size, but the red spheres appear to be smaller by comparison.

Check Your Learning

Calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$, is an ionic compound and a common anti-caking agent added to food products. What is the formula mass (amu) of calcium phosphate?

Answer: 310.18 amu

Key Concepts and Summary

The formula mass of a substance is the sum of the average atomic masses of each atom represented in the chemical formula and is expressed in atomic mass units. The formula mass of a covalent compound is also called the molecular mass. Due to the use of the same reference substance in defining the atomic mass unit and the mole, the formula mass (amu) and molar mass (g/mol) for any substance are numerically equivalent (for example, one H_2O molecule weighs approximately 18 amu and 1 mole of H_2O molecules weighs approximately 18 g).

Chemistry End of Chapter Exercises

1. What is the total mass (amu) of carbon in each of the following molecules?

- (a) CH_4
- (b) CHCl_3
- (c) $\text{C}_{12}\text{H}_{10}\text{O}_6$
- (d) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

2. What is the total mass of hydrogen in each of the molecules?

- (a) CH_4
- (b) CHCl_3
- (c) $\text{C}_{12}\text{H}_{10}\text{O}_6$
- (d) $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$

3. Calculate the molecular or formula mass of each of the following:

- (a) P_4
- (b) H_2O
- (c) $\text{Ca}(\text{NO}_3)_2$
- (d) $\text{CH}_3\text{CO}_2\text{H}$ (acetic acid)
- (e) $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ (sucrose, cane sugar).

4. Determine the molecular mass of the following compounds:

(a)

 A structure is shown. A C atom is bonded to two H atoms and forms a double bond with one O atom.

(b)

 A structure is shown. Two C atoms form a triple bond with each other. Each C atom also forms a single bond with an H atom.

(c)

 A structure is shown. Two C atoms form double bonds with each other. Each C atom also forms a single bond with an H atom and a Br atom.

(d)

 A structure is shown. An S atom forms double bonds with two O atoms. The S atom also forms a single bond with an O atom which forms a single bond with an H atom. The S atom also forms a single bond with another O atom which forms a single bond with another H atom.

5. Determine the molecular mass of the following compounds:

(a)

 A structure is shown. Two C atoms form double bonds with each other. The C atom on the left forms a single bond with two H atoms each. The C atom on the right forms a single bond with an H atom and with a C H subscript 2 C H subscript 3 group.

(b)

 A structure is shown. There is a C atom which forms single bonds with three H atoms each. This C atom is bonded to another C atom. This second C atom forms a triple bond with another C atom which forms a single bond with a fourth C atom. The fourth C atom forms single bonds with three H atoms each.

(c)

 A structure is shown. An S i atom forms a single bond with a C l atom, a single bond with a C l atom, a single bond with an H atom, and a single bond with another S i atom. The second S i atom forms a single bond with a C l atom, a single bond with a C l atom, and a single bond with an H atom.

(d)

 A structure is shown. A P atom forms a double bond with an O atom. It also forms a single bond with an O atom which forms a single bond with an H atom. It also forms a single bond with another O atom which forms a single bond with an H atom. It also forms a single bond with another O atom which forms a single bond with an H atom.

6. Which molecule has a molecular mass of 28.05 amu?

(a)

 A structure is shown. A C atom forms a triple bond with another C atom. Each C atom also forms a single bond with an H atom.

(b)

 A structure is shown. Two C atoms form a double bond with each other. Each C atom also forms a single bond with two H atoms.

(c)

 A structure is shown. A C atom forms a single bond with three H atoms each and with another C atom. The second C atom also forms a single bond with three H atoms each.

Glossary

Avogadro's number (N_A)

experimentally determined value of the number of entities comprising 1 mole of substance, equal to $6.022 \times 10^{23} \text{ mol}^{-1}$

formula mass

sum of the average masses for all atoms represented in a chemical formula; for covalent compounds, this is also the molecular mass

molar mass

mass in grams of 1 mole of a substance

mole

amount of substance containing the same number of atoms, molecules, ions, or other entities as the number of atoms in exactly 12 grams of ^{12}C

6.2: Formula mass is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

6.3: Converting between grams and moles

17

Harper College Chemistry Department

Moles to Mass

Where can you find the molar mass of the elements? [The periodic table](#). You should always have one on hand—don't expect to get very far without one!

How heavy is 1.5 mol of lead? How many moles in 22.34g of water? Calculating the mass of a sample from the number of moles it contains is quite simple. We use the **molar mass** (mass of one mole) of the substance to convert between mass and moles. When writing calculations, we denote the molar mass of a substance by an upper case “M” (e.g. M(Ne) means “the molar mass of neon”). As always, “n” stands for the number of moles and “m” indicates the mass of a substance. To find the solutions to the two questions we just asked, let's apply some dimensional analysis:

$$1.5 \text{ mol Pb} \times \left(\frac{207.2 \text{ g Pb}}{1 \text{ mol Pb}} \right) = 310.8 \text{ g Pb}$$

Can you see how the units cancel to give you the answer you want? All you needed to know was that you had 1.5 mol Pb (lead), and that 1 mol Pb weighs 207.2 grams. Thus, multiplying 1.5 mol Pb by 207.2 g Pb and dividing by 1 mol Pb gives you 310.8 g Pb, your answer.

Deriving Grams from Moles for an Element

A liter of air contains 9.2×10^{-4} mol argon. What is the mass of Ar in a liter of air?

Solution

The molar amount of Ar is provided and must be used to derive the corresponding mass in grams. Since the amount of Ar is less than 1 mole, the mass will be less than the mass of 1 mole of Ar, approximately 40 g. The molar amount in question is approximately one-one thousandth ($\sim 10^{-3}$) of a mole, and so the corresponding mass should be roughly one-one thousandth of the molar mass (~ 0.04 g):

A diagram of two boxes connected by a right-facing arrow is shown. The box on the left contains the phrase, “Moles of Ar atoms (mol)” while the one on the right contains the phrase, “Mass of Ar atoms (g).” There is a phrase under the arrow that says “Multiply by molar mass (g / mol).”

In this case, logic dictates (and the factor-label method supports) multiplying the provided amount (mol) by the molar mass (g/mol):

$$9.2 \times 10^{-4} \text{ mol Ar} (39.95 \text{ g / mol Ar}) = 0.037 \text{ g Ar}$$

The result is in agreement with our expectations, around 0.04 g Ar.

Check Your Learning

What is the mass of 2.561 mol of gold?

Answer:

504.4 g

Deriving Grams from Moles for a Compound

Vitamin C is a covalent compound with the molecular formula $\text{C}_6\text{H}_8\text{O}_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is 1.42×10^{-4} mol. What is the mass of this allowance in grams?

Solution

As for elements, the mass of a compound can be derived from its molar amount as shown:

A diagram of two boxes connected by a right-facing arrow is shown. The box on the left contains the phrase, “Moles of vitamin C (mol)” while the one the right contains the phrase, “Mass of vitamin C (g).” There is a phrase under the arrow that says “Multiply by molar mass (g / mol).”

The molar mass for this compound is computed to be 176.124 g/mol. The given number of moles is a very small fraction of a mole ($\sim 10^{-4}$ or one-ten thousandth); therefore, we would expect the corresponding mass to be about one-ten thousandth of the molar mass (~ 0.02 g). Performing the calculation, we get:

$$>1.42 \times 10^{-4} \text{ mol vitamin C (176.124 g/mol vitamin C)} = 0.0250 \text{ g vitamin C}$$

This is consistent with the anticipated result.

Check Your Learning

What is the mass of 0.443 mol of hydrazine, N_2H_4 ?

Answer: 14.2 g

Mass to Moles

But we had one more question: “How many moles in 22.34 g of water?” This is just as easy:

$$22.34 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g H}_2\text{O}} = 1.24 \text{ mol H}_2\text{O}$$

Where did the 18 g H_2O come from? We looked at the periodic table and simply added up the atomic masses of two hydrogens and an oxygen to get the molecular weight of water. This turned out to be 18, and since all the masses on the periodic table are given with respect to 1 mole, we knew that 1 mol of water weighed 18 grams. This gave us the relationship above, which is really just (again) watching units cancel out!

Deriving Moles from Grams for an Element

According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles?

Solution

The mass of K is provided, and the corresponding amount of K in moles is requested. Referring to the periodic table, the atomic mass of K is 39.10 amu, and so its molar mass is 39.10 g/mol. The given mass of K (4.7 g) is a bit more than one-tenth the molar mass (39.10 g), so a reasonable “ballpark” estimate of the number of moles would be slightly greater than 0.1 mol.

The molar amount of a substance may be calculated by dividing its mass (g) by its molar mass (g/mol):

A diagram of two boxes connected by a right-facing arrow is shown. The box on the left contains the phrase, “Mass of K atoms (g)” while the one on the right contains the phrase, “Moles of K atoms (mol).” There is a phrase under the arrow that says, “Divide by molar mass (g / mol).”

The factor-label method supports this mathematical approach since the unit “g” cancels and the answer has units of “mol:”

$$>4.7 \text{ g K (mol K / 39.10 g)} = 0.12 \text{ mol K}$$

The calculated magnitude (0.12 mol K) is consistent with our ballpark expectation, since it is a bit greater than 0.1 mol.

Check Your Learning

Beryllium is a light metal used to fabricate transparent X-ray windows for medical imaging instruments. How many moles of Be are in a thin-foil window weighing 3.24 g?

Answer: 0.360 mol

Deriving Moles from Grams for a Compound

Our bodies synthesize protein from amino acids. One of these amino acids is glycine, which has the molecular formula $\text{C}_2\text{H}_5\text{O}_2\text{N}$. How many moles of glycine molecules are contained in 28.35 g of glycine?

Solution

We can derive the number of moles of a compound from its mass following the same procedure we used for an element in [Example 2.7](#):

A diagram of two boxes connected by a right-facing arrow is shown. The box on the left contains the phrase, "Mass of C subscript 2 H subscript 5 O subscript 2 N (g)" while the box on the right contains the phrase, "Moles of C subscript 2 H subscript 5 O subscript 2 N (mol)." There is a phrase under the arrow that says "Divide by molar mass (g / mol)."

The molar mass of glycine is required for this calculation, and it is computed in the same fashion as its molecular mass. One mole of glycine, $\text{C}_2\text{H}_5\text{O}_2\text{N}$, contains 2 moles of carbon, 5 moles of hydrogen, 2 moles of oxygen, and 1 mole of nitrogen:

A table is shown that is made up of six columns and six rows. The header row reads: "Element," "Quantity (mol element / mol compound)," a blank space, "Molar mass (g / mol element)," a blank space, and "Subtotal (a m u)." The first column contains the symbols "C," "H," "O," "N," and a merged cell. The merged cell runs the width of the first five columns. The second column contains the numbers "2," "5," "2," and "1" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "12.01," "1.008," "16.00," and "14.007" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "24.02," "5.040," "32.00," "14.007," and "75.07." There is a thick black line under the number 14.007. The merged cell under the first five columns reads "Molar mass (g / mol compound)." There is a ball-and-stick drawing to the right of this table. It shows a black sphere that forms a double bond with a slightly smaller red sphere, a single bond with another red sphere, and a single bond with another black sphere. The red sphere that forms a single bond with the black sphere also forms a single bond with a smaller, white sphere. The second black sphere forms a single bond with a smaller, white sphere and a smaller blue sphere. The blue sphere forms a single bond with two smaller, white spheres each.

The provided mass of glycine (~28 g) is a bit more than one-third the molar mass (~75 g/mol), so we would expect the computed result to be a bit greater than one-third of a mole (~0.33 mol). Dividing the compound's mass by its molar mass yields:

$$>28.35 \text{ g glycine (mol glycine / 75.07 g)} = 0.378 \text{ mol glycine}$$

This result is consistent with our rough estimate.

Check Your Learning

How many moles of sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$, are in a 25-g sample of sucrose?

Answer: 0.073 mol

Calculating Molar Masses

Before we can do these types of calculations, we first have to know the molar mass. Fortunately, this is not difficult, as the molar mass of an element is exactly the same as the **atomic weight** of that element. A table of atomic weights can be used to find the molar mass of elements (this information is often included in the periodic table). For example, the atomic weight of oxygen is **16.00** amu, so its molar mass is 16.00 g/mol.

For species with more than one element, we simply add up the atomic weights of each element to obtain the molar mass of the compound. For example, sulfur trioxide gas is made up of sulfur and oxygen, whose atomic weights are 32.06 and 16.00 respectively.

$$\{$$

$$\begin{aligned} M(\text{SO}_3) &= 32.06 + (3 \times 16.00) \\ &= 80.06 \text{ g/mol} \end{aligned} \quad (6.3.1)$$

$$\}\}$$

The procedure for more complex compounds is essentially the same. Aluminium carbonate, for example, contains aluminium, carbon, and oxygen. To find the molar mass, we have to be careful to find the *total* number of atoms of each element. Three carbonate ions each containing three oxygen atoms gives a total of nine oxygens. The atomic weights of aluminium and carbon are 26.98 and 12.01 respectively.

$$\begin{aligned} M(\text{Al}_2(\text{CO}_3)_3) &= 2 \times 26.98 + 3 \times 12.01 + 9 \times 16.00 \\ &= 233.99 \text{ g/mol} \end{aligned}$$

As described in the previous section, molar mass is expressed as "grams *per* mole". The word *per* in this context implies a mathematical relationship between grams and mole. Think of this as a ratio. The fact that a *per* relationship, ratio, exists between grams and moles implies that you can use dimensional analysis to interconvert between the two. For example, if we wanted to know the mass of 0.50 mole of molecular hydrogen (H_2) we could set up the following equations:

The **known** molar mass of H_2 is:

$$2.03 \text{ g H}_2 / \text{mol H}_2$$

We are **given** that we have 0.50 moles of H₂ and we want to **find** the number of grams of H₂ that this represents. To perform the dimensional analysis, we arrange the *known* and the *given* so that the units cancel, leaving only the units of the item we want to *find*.

$$0.5 \text{ mol} \times \frac{2.03 \text{ g H}_2}{\text{mol H}_2} = 1.01 \text{ g H}_2$$

EXERCISE

1. Determine the mass of 0.752 mol of H₂ gas.
 2. How many *moles* of molecular hydrogen are present in 6.022 grams of H₂?
 3. If you have 22.414 grams of Cl₂, how many *moles* of molecular chlorine do you have?
-

6.3: Converting between grams and moles is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

7: Chemical Equations and Reactions

[7.1: Balancing Chemical Equations](#)

[7.2: Energy in chemical reactions](#)

[7.3: Equilibrium](#)

[7.4: Catalysis](#)

7: Chemical Equations and Reactions is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

7.1: Balancing Chemical Equations

18

Harper College Chemistry Department

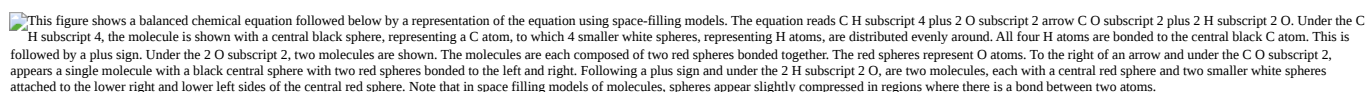
Page by: OpenStax

▼ Summary

By the end of this section, you will be able to:

- Derive chemical equations from narrative descriptions of chemical reactions.
- Write and balance chemical equations in molecular, total ionic, and net ionic formats.

A preceding chapter introduced the use of element symbols to represent individual atoms. When atoms gain or lose electrons to yield ions or combine with other atoms to form molecules, their symbols are modified or combine to generate chemical formulas that appropriately represent these species. Extending this symbolism to represent both the identities and the relative quantities of substances undergoing a chemical (or physical) change involves writing and balancing a chemical equation. Consider as an example the reaction between one methane molecule (CH_4) and two diatomic oxygen molecules (O_2) to produce one carbon dioxide molecule (CO_2) and two water molecules (H_2O). The chemical equation representing this process is provided in the upper half of [Figure 1](#), with space-filling molecular models shown in the lower half of the figure.

 This figure shows a balanced chemical equation followed below by a representation of the equation using space-filling models. The equation reads $\text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}$. Under the CH_4 subscript 4, the molecule is shown with a central black sphere, representing a C atom, to which 4 smaller white spheres, representing H atoms, are distributed evenly around. All four H atoms are bonded to the central black C atom. This is followed by a plus sign. Under the 2O_2 subscript 2, two molecules are shown. The molecules are each composed of two red spheres bonded together. The red spheres represent O atoms. To the right of an arrow and under the CO_2 subscript 2, appears a single molecule with a black central sphere with two red spheres bonded to the left and right. Following a plus sign and under the $2 \text{H}_2\text{O}$ subscript 2, are two molecules, each with a central red sphere and two smaller white spheres attached to the lower right and lower left sides of the central red sphere. Note that in space filling models of molecules, spheres appear slightly compressed in regions where there is a bond between two atoms.

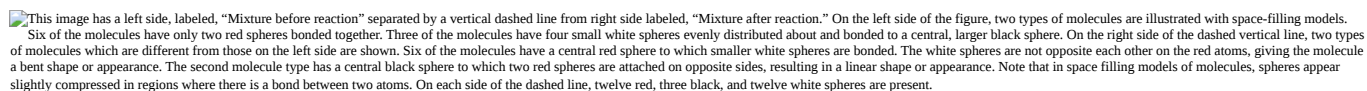
The reaction between methane and oxygen to yield carbon dioxide and water (shown at bottom) may be represented by a chemical equation using formulas (top).

This example illustrates the fundamental aspects of any chemical equation:

1. The substances undergoing reaction are called reactants and their formulas are placed on the left side of the equation.
2. The substances generated by the reaction are called products and their formulas are placed on the right side of the equation.
3. Plus signs (+) separate individual reactant and product formulas and an arrow ($\rightarrow$) separates the reactant and product (left and right) sides of the equation.
4. The relative numbers of reactant and product species are represented by coefficients (numbers placed immediately to the left of each formula). A coefficient of 1 is typically omitted.

It is common practice to use the smallest possible whole-number coefficients in a chemical equation, as is done in this example. Realize, however, that these coefficients represent the *relative* numbers of reactants and products and, therefore, they may be correctly interpreted as ratios. Methane and oxygen react to yield carbon dioxide and water in a 1:2:1:2 ratio. This ratio is satisfied if the numbers of these molecules are, respectively, 1-2-1-2, or 2-4-2-4 or 3-6-3-6 and so on ([Figure 2](#)). Likewise, these coefficients may be interpreted with regard to any amount (number) unit and so this equation may be correctly read in many ways, including:

- One methane molecule and two oxygen molecules react to yield one carbon dioxide molecule and two water molecules.
- One dozen methane molecules and two dozen oxygen molecules react to yield one dozen carbon dioxide molecules and two dozen water molecules.
- One mole of methane molecules and 2 moles of oxygen molecules react to yield 1 mole of carbon dioxide molecules and 2 moles of water molecules. The mole concept will be discussed in a different chapter.

 This image has a left side, labeled, "Mixture before reaction" separated by a vertical dashed line from right side labeled, "Mixture after reaction." On the left side of the figure, two types of molecules are illustrated with space-filling models. Six of the molecules have only two red spheres bonded together. Three of the molecules have four small white spheres evenly distributed about and bonded to a central, larger black sphere. On the right side of the dashed vertical line, two types of molecules which are different from those on the left side are shown. Six of the molecules have a central red sphere to which smaller white spheres are bonded. The white spheres are not opposite each other on the red atoms, giving the molecule a bent shape or appearance. The second molecule type has a central black sphere to which two red spheres are attached on opposite sides, resulting in a linear shape or appearance. Note that in space filling models of molecules, spheres appear slightly compressed in regions where there is a bond between two atoms. On each side of the dashed line, twelve red, three black, and twelve white spheres are present.

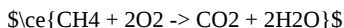
Regardless of the absolute numbers of molecules involved, the ratios between numbers of molecules of each species that react (the reactants) and molecules of each species that form (the products) are the same and are given by the chemical reaction equation.

Balancing Equations

The chemical equation described in the previous section is balanced, meaning that equal numbers of atoms for each element involved in the reaction are represented on the reactant and product sides. This is a requirement the equation must satisfy to be consistent with the law of conservation of matter. It may be confirmed by simply summing the numbers of atoms on either side of the arrow and comparing these sums to ensure they are equal. Note that the number of atoms for a given element is calculated by multiplying the coefficient of any formula containing that element by the element's subscript in the formula. If an element appears in more than one formula on a given side of the equation, the number of atoms represented in each must be computed and then added together. For example, both product species in the example reaction, CO_2 and H_2O , contain the element oxygen, and so the number of oxygen atoms on the product side of the equation is

$$\left(1 \times \text{CO}_2\right) \text{ molecule} \times \frac{2}{1} \text{ O atoms} \left(\text{CO}_2 \right) + \left(2 \times \text{H}_2\text{O}\right) \text{ molecules} \times \frac{1}{1} \text{ O atoms} \left(\text{H}_2\text{O} \right) = 4 \text{ O atoms}$$

The equation for the reaction between methane and oxygen to yield carbon dioxide and water is confirmed to be balanced per this approach, as shown here:



Element	Reactants	Products	Balanced?
C	$1 \times 1 = 1$	$1 \times 1 = 1$	$1 = 1$, yes
H	$4 \times 1 = 4$	$2 \times 2 = 4$	$4 = 4$, yes
O	$2 \times 2 = 4$	$(1 \times 2) + (2 \times 1) = 4$	$4 = 4$, yes

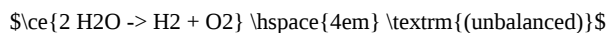
A balanced chemical equation often may be derived from a qualitative description of some chemical reaction by a fairly simple approach known as balancing by inspection. Consider as an example the decomposition of water to yield molecular hydrogen and oxygen. This process is represented qualitatively by an *unbalanced* chemical equation:



Comparing the number of H and O atoms on either side of this equation confirms its imbalance:

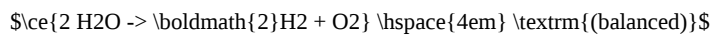
Element	Reactants	Products	Balanced?
H	$1 \times 2 = 2$	$1 \times 2 = 2$	$2 = 2$, yes
O	$1 \times 1 = 1$	$1 \times 2 = 2$	$1 \neq 2$, no

The numbers of H atoms on the reactant and product sides of the equation are equal, but the numbers of O atoms are not. To achieve balance, the *coefficients* of the equation may be changed as needed. Keep in mind, of course, that the *formula subscripts* define, in part, the identity of the substance, and so these cannot be changed without altering the qualitative meaning of the equation. For example, changing the reactant formula from H_2O to H_2O_2 would yield balance in the number of atoms, but doing so also changes the reactant's identity (it's now hydrogen peroxide and not water). The O atom balance may be achieved by changing the coefficient for H_2O to 2.



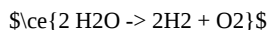
Element	Reactants	Products	Balanced?
H	$2 \times 2 = 4$	$1 \times 2 = 2$	$4 \neq 2$, no
O	$2 \times 1 = 2$	$1 \times 2 = 2$	$2 = 2$, yes

The H atom balance was upset by this change, but it is easily reestablished by changing the coefficient for the H_2 product to 2.



Element	Reactants	Products	Balanced?
H	$2 \times 2 = 4$	$2 \times 2 = 4$	$4 = 4$, yes
O	$2 \times 1 = 2$	$1 \times 2 = 2$	$2 = 2$, yes

These coefficients yield equal numbers of both H and O atoms on the reactant and product sides and the balanced equation is, therefore:



Balancing Chemical Equations

Write a balanced equation for the reaction of molecular nitrogen (N_2) and oxygen (O_2) to form dinitrogen pentoxide.

Solution:

First, write the unbalanced equation.



Next, count the number of each type of atom present in the unbalanced equation.

Element	Reactants	Products	Balanced?
N	$1 \times 2 = 2$	$1 \times 2 = 2$	$2 = 2$, yes
O	$1 \times 2 = 2$	$1 \times 5 = 5$	$2 \neq 5$, no

Though nitrogen is balanced, changes in coefficients are needed to balance the number of oxygen atoms. To balance the number of oxygen atoms, a reasonable first attempt would be to change the coefficients for the O_2 and N_2O_5 to integers that will yield 10 O atoms (the least common multiple for the O atom subscripts in these two formulas).



Element	Reactants	Products	Balanced?
N	$1 \times 2 = 2$	$2 \times 2 = 4$	$2 \neq 4$, no
O	$5 \times 2 = 10$	$2 \times 5 = 10$	$10 = 10$, yes

The N atom balance has been upset by this change; it is restored by changing the coefficient for the reactant N_2 to 2.



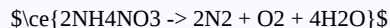
Element	Reactants	Products	Balanced?
N	$2 \times 2 = 4$	$2 \times 2 = 4$	$4 = 4$, yes
O	$5 \times 2 = 10$	$2 \times 5 = 10$	$10 = 10$, yes

The numbers of N and O atoms on either side of the equation are now equal, and so the equation is balanced.

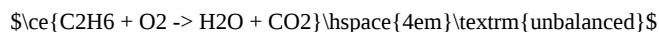
Check Your Learning

Write a balanced equation for the decomposition of ammonium nitrate to form molecular nitrogen, molecular oxygen and water. (Hint: Balance oxygen last, since it is present in more than one molecule on the right side of the equation.)

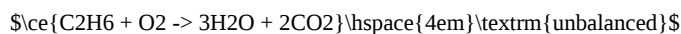
Answer:



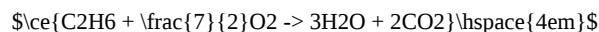
It is sometimes convenient to use fractions instead of integers as intermediate coefficients in the process of balancing a chemical equation. When balance is achieved, all the equation's coefficients may then be multiplied by a whole number to convert the fractional coefficients to integers without upsetting the atom balance. For example, consider the reaction of ethane (C_2H_6) with oxygen to yield H_2O and CO_2 , represented by the unbalanced equation:



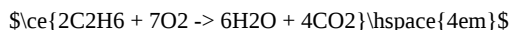
Following the usual inspection approach, one might first balance C and H atoms by changing the coefficients for the two product species, as shown:



This results in seven O atoms on the product side of the equation, an odd number—no integer coefficient can be used with the O_2 reactant to yield an odd number, so a fractional coefficient, $\frac{7}{2}$, is used instead to yield a provisional balanced equation:



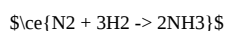
A conventional balanced equation with integer-only coefficients is derived by multiplying each coefficient by 2:



Finally with regard to balanced equations, recall that convention dictates use of the *smallest whole-number coefficients*. Although the equation for the reaction between molecular nitrogen and molecular hydrogen to produce ammonia is, indeed, balanced,



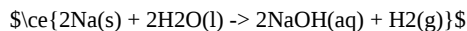
the coefficients are not the smallest possible integers representing the relative numbers of reactant and product molecules. Dividing each coefficient by the greatest common factor, 3, gives the preferred equation:



Use this interactive [tutorial](#) for additional practice balancing equations.

Additional Information in Chemical Equations

The physical states of reactants and products in chemical equations very often are indicated with a parenthetical abbreviation following the formulas. Common abbreviations include *s* for solids, *l* for liquids, *g* for gases and *aq* for substances dissolved in water (*aqueous solutions*). These notations are illustrated in the example equation here:



This equation represents the reaction that takes place when sodium metal is placed in water. The solid sodium reacts with liquid water to produce molecular hydrogen gas and the ionic compound sodium hydroxide (a solid in pure form, but readily dissolved in water).

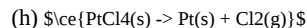
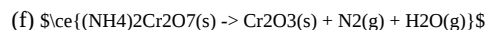
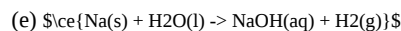
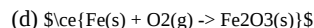
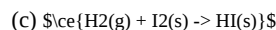
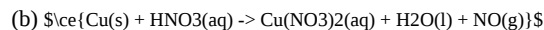
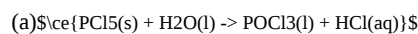
Key Concepts and Summary

Chemical equations are symbolic representations of chemical and physical changes. Formulas for the substances undergoing the change (reactants) and substances generated by the change (products) are separated by an arrow and preceded by integer coefficients indicating their relative numbers. Balanced equations are those whose coefficients result in equal numbers of atoms for each element in the reactants and products.

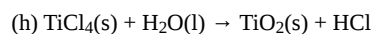
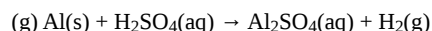
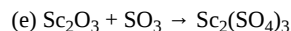
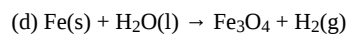
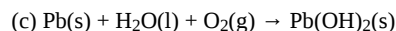
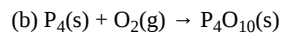
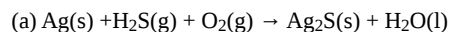
End of Chapter Exercises

1. What does it mean to say an equation is balanced? Why is it important for an equation to be balanced?

2. Balance the following equations:



3. Balance the following equations:



4. Write a balanced molecular equation describing each of the following chemical reactions.

(a) Solid calcium carbonate is heated and decomposes to solid calcium oxide and carbon dioxide gas.

(b) Gaseous butane, C_4H_{10} , reacts with diatomic oxygen gas to yield gaseous carbon dioxide and water vapor.

(c) Aqueous solutions of magnesium chloride and sodium hydroxide react to produce solid magnesium hydroxide and aqueous sodium chloride.

(d) Water vapor reacts with sodium metal to produce solid sodium hydroxide and hydrogen gas.

5. Write a balanced equation describing each of the following chemical reactions.

(a) Solid potassium chlorate, KClO_3 , decomposes to form solid potassium chloride and diatomic oxygen gas.

(b) Solid aluminum metal reacts with solid diatomic iodine to form solid Al_2I_6 .

(c) When solid sodium chloride is added to aqueous sulfuric acid, hydrogen chloride gas and aqueous sodium sulfate are produced.

(d) Aqueous solutions of phosphoric acid and potassium hydroxide react to produce aqueous potassium dihydrogen phosphate and liquid water.

6. Colorful fireworks often involve the decomposition of barium nitrate and potassium chlorate and the reaction of the metals magnesium, aluminum, and iron with oxygen.
- (a) Write the formulas of barium nitrate and potassium chlorate.
 - (b) The decomposition of solid potassium chlorate leads to the formation of solid potassium chloride and diatomic oxygen gas. Write an equation for the reaction.
 - (c) The decomposition of solid barium nitrate leads to the formation of solid barium oxide, diatomic nitrogen gas, and diatomic oxygen gas. Write an equation for the reaction.
7. A novel process for obtaining magnesium from sea water involves several reactions. Write a balanced chemical equation for each step of the process.
- (a) The first step is the decomposition of solid calcium carbonate from seashells to form solid calcium oxide and gaseous carbon dioxide.
 - (b) The second step is the formation of solid calcium hydroxide as the only product from the reaction of the solid calcium oxide with liquid water.
 - (c) Solid calcium hydroxide is then added to the seawater, reacting with dissolved magnesium chloride to yield solid magnesium hydroxide and aqueous calcium chloride.
 - (d) The solid magnesium hydroxide is added to a hydrochloric acid solution, producing dissolved magnesium chloride and liquid water.
 - (e) Finally, the magnesium chloride is melted and electrolyzed to yield liquid magnesium metal and diatomic chlorine gas.

Glossary

balanced equation

chemical equation with equal numbers of atoms for each element in the reactant and product

chemical equation

symbolic representation of a chemical reaction

coefficient

number placed in front of symbols or formulas in a chemical equation to indicate their relative amount

molecular equation

chemical equation in which all reactants and products are represented as neutral substances

product

substance formed by a chemical or physical change; shown on the right side of the arrow in a chemical equation

reactant

substance undergoing a chemical or physical change; shown on the left side of the arrow in a chemical equation

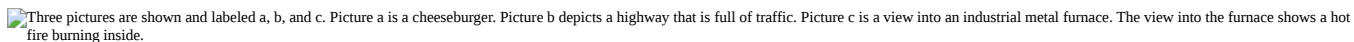
7.1: Balancing Chemical Equations is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

7.2: Energy in chemical reactions

19

Harper College Chemistry Department

Chemical changes and their accompanying changes in energy are important parts of our everyday world (Figure 1). The macronutrients in food (proteins, fats, and carbohydrates) undergo metabolic reactions that provide the energy to keep our bodies functioning. We burn a variety of fuels (gasoline, natural gas, coal) to produce energy for transportation, heating and the generation of electricity. Industrial chemical reactions use enormous amounts of energy to produce raw materials (such as iron and aluminum). Energy is then used to manufacture those raw materials into useful products, such as cars, skyscrapers, and bridges.

Three pictures are shown and labeled a, b, and c. Picture a is a cheeseburger. Picture b depicts a highway that is full of traffic. Picture c is a view into an industrial metal furnace. The view into the furnace shows a hot fire burning inside.

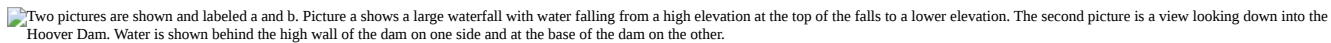
The energy involved in chemical changes is important to our daily lives: (a) A cheeseburger for lunch provides the energy you need to get through the rest of the day; (b) the combustion of gasoline provides the energy that moves your car (and you) between home, work, and school; and (c) coke, a processed form of coal, provides the energy needed to convert iron ore into iron, which is essential for making many of the products we use daily. (credit a: modification of work by “Pink Sherbet Photography”/Flickr; credit b: modification of work by Jeffery Turner)

This chapter will introduce the basic ideas of an important area of science concerned with the amount of heat absorbed or released during chemical and physical changes—an area called thermochemistry. The concepts introduced in this chapter are widely used in almost all scientific and technical fields. Food scientists use them to determine the energy content of foods. Biologists study the energetics of living organisms, such as the metabolic combustion of sugar into carbon dioxide and water. The oil, gas, and transportation industries, renewable energy providers, and many others endeavor to find better methods to produce energy for our commercial and personal needs. Engineers strive to improve energy efficiency, find better ways to heat and cool our homes, refrigerate our food and drinks, and meet the energy and cooling needs of computers and electronics, among other applications. Understanding thermochemical principles is essential for chemists, physicists, biologists, geologists, every type of engineer, and just about anyone who studies or does any kind of science.

Energy

Energy can be defined as the capacity to supply heat or do work. One type of work (w) is the process of causing matter to move against an opposing force. For example, we do work when we inflate a bicycle tire—we move matter (the air in the pump) against the opposing force of the air already in the tire.

Like matter, energy comes in different types. One scheme classifies energy into two types: potential energy, the energy an object has because of its relative position, composition or condition and kinetic energy, the energy that an object possesses because of its motion. Water at the top of a waterfall or dam has potential energy because of its position; when it flows downward through generators, it has kinetic energy that can be used to do work and produce electricity in a hydroelectric plant (Figure 2). A battery has potential energy because the chemicals within it can produce electricity that can do work.

Two pictures are shown and labeled a and b. Picture a shows a large waterfall with water falling from a high elevation at the top of the falls to a lower elevation. The second picture is a view looking down into the Hoover Dam. Water is shown behind the high wall of the dam on one side and at the base of the dam on the other.

(a) Water that is higher in elevation, for example, at the top of Victoria Falls, has a higher potential energy than water at a lower elevation. As the water falls, some of its potential energy is converted into kinetic energy. (b) If the water flows through generators at the bottom of a dam, such as the Hoover Dam shown here, its kinetic energy is converted into electrical energy. (credit a: modification of work by Steve Jurvetson; credit b: modification of work by “curimedia”/Wikimedia commons)

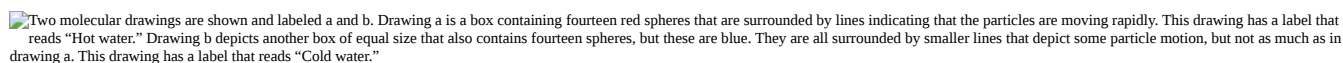
Energy can be converted from one form into another, but all of the energy present before a change occurs always exists in some form after the change is completed. This observation is expressed in the law of conservation of energy: during a chemical or physical change, energy can be neither created nor destroyed, although it can be changed in form. (This is also one version of the first law of thermodynamics, as you will learn later.)

When one substance is converted into another, there is always an associated conversion of one form of energy into another. Heat is usually released or absorbed, but sometimes the conversion involves light, electrical energy or some other form of energy. For example, chemical energy (a type of potential energy) is stored in the molecules that compose gasoline. When gasoline is combusted within the cylinders of a car’s engine, the rapidly expanding gaseous products of this chemical reaction generate mechanical energy (a type of kinetic energy) when they move the cylinders’ pistons.

According to the law of conservation of matter (seen in an earlier chapter), there is no detectable change in the total amount of matter during a chemical change. When chemical reactions occur, the energy changes are relatively modest and the mass changes are too small to measure, so the laws of conservation of matter and energy hold well. However, in nuclear reactions, the energy changes are much larger (by factors of a million or so), the mass changes are measurable, and matter-energy conversions are significant. This will be examined in more detail in a later chapter on nuclear chemistry. To encompass both chemical and nuclear changes, we combine these laws into one statement: The total quantity of matter and energy in the universe is fixed.

Thermal Energy, Temperature and Heat

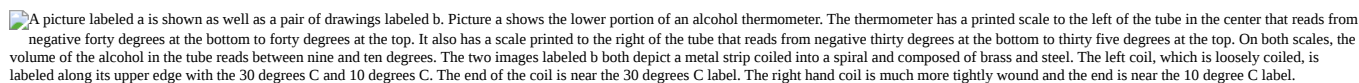
Thermal energy is kinetic energy associated with the random motion of atoms and molecules. Temperature is a quantitative measure of “hot” or “cold.” When the atoms and molecules in an object are moving or vibrating quickly, they have a higher average kinetic energy (KE) and we say that the object is “hot.” When the atoms and molecules are moving slowly, they have lower average KE and we say that the object is “cold” (Figure 3). Assuming that no chemical reaction or phase change (such as melting or vaporizing) occurs, increasing the amount of thermal energy in a sample of matter will cause its temperature to increase. In addition, assuming that no chemical reaction or phase change (such as condensation or freezing) occurs, decreasing the amount of thermal energy in a sample of matter will cause its temperature to decrease.

Two molecular drawings are shown and labeled a and b. Drawing a is a box containing fourteen red spheres that are surrounded by lines indicating that the particles are moving rapidly. This drawing has a label that reads “Hot water.” Drawing b depicts another box of equal size that also contains fourteen spheres, but these are blue. They are all surrounded by smaller lines that depict some particle motion, but not as much as in drawing a. This drawing has a label that reads “Cold water.”

(a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.

Click on this [interactive simulation](#) to view the effects of temperature on molecular motion.

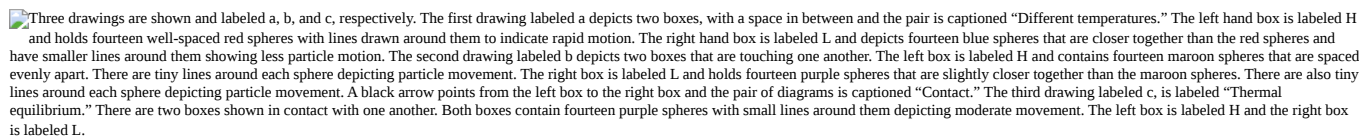
Most substances expand as their temperature increases and contract as their temperature decreases. This property can be used to measure temperature changes, as shown in Figure 4. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

A picture labeled a is shown as well as a pair of drawings labeled b. Picture a shows the lower portion of an alcohol thermometer. The thermometer has a printed scale to the left of the tube in the center that reads from negative forty degrees at the bottom to forty degrees at the top. It also has a scale printed to the right of the tube that reads from negative thirty degrees at the bottom to thirty five degrees at the top. On both scales, the volume of the alcohol in the tube reads between nine and ten degrees. The two images labeled b both depict a metal strip coiled into a spiral and composed of brass and steel. The left coil, which is loosely coiled, is labeled along its upper edge with the 30 degrees C and 10 degrees C. The end of the coil is near the 30 degrees C label. The right hand coil is much more tightly wound and the end is near the 10 degree C label.

(a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. When heated or cooled, one of the metals (brass) expands or contracts more than the other metal (steel), causing the strip to coil or uncoil. Both types of thermometers have a calibrated scale that indicates the temperature. (credit a: modification of work by “dwstucke”/Flickr)

The following [demonstration](#) allows one to view the effects of heating and cooling a coiled bimetallic strip.

Heat (q) is the transfer of thermal energy between two bodies at different temperatures. Heat flow (a redundant term, but one commonly used) increases the thermal energy of one body and decreases the thermal energy of the other. Suppose we initially have a high temperature (and high thermal energy) substance (H) and a low temperature (and low thermal energy) substance (L). The atoms and molecules in H have a higher average KE than those in L. If we place substance H in contact with substance L, the thermal energy will flow spontaneously from substance H to substance L. The temperature of substance H will decrease, as will the average KE of its molecules; the temperature of substance L will increase, along with the average KE of its molecules. Heat flow will continue until the two substances are at the same temperature (Figure 5).

Three drawings are shown and labeled a, b, and c, respectively. The first drawing labeled a depicts two boxes, with a space in between and the pair is captioned “Different temperatures.” The left hand box is labeled H and holds fourteen well-spaced red spheres with lines drawn around them to indicate rapid motion. The right hand box is labeled L and depicts fourteen blue spheres that are closer together than the red spheres and have smaller lines around them showing less particle motion. The second drawing labeled b depicts two boxes that are touching one another. The left box is labeled H and contains fourteen maroon spheres that are spaced evenly apart. There are tiny lines around each sphere depicting particle movement. A black arrow points from the left box to the right box and the pair of diagrams is captioned “Contact.” The third drawing labeled c, is labeled “Thermal equilibrium.” There are two boxes shown in contact with one another. Both boxes contain fourteen purple spheres with small lines around them depicting moderate movement. The left box is labeled H and the right box is labeled L.

(a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they are put into contact with each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two objects reach “thermal equilibrium” when both substances are at the same temperature, and their molecules have the same average kinetic energy.

Matter undergoing chemical reactions and physical changes can release or absorb heat. A change that releases heat is called an exothermic process. For example, the combustion reaction that occurs when using an oxyacetylene torch is an exothermic process—this process also releases energy in the form of light as evidenced by the torch’s flame (Figure 6). A reaction or change that absorbs heat is an endothermic process. A cold pack used to treat muscle strains provides an example of an endothermic process.

When the substances in the cold pack (water and a salt like ammonium nitrate) are brought together, the resulting process absorbs heat, leading to the sensation of cold.

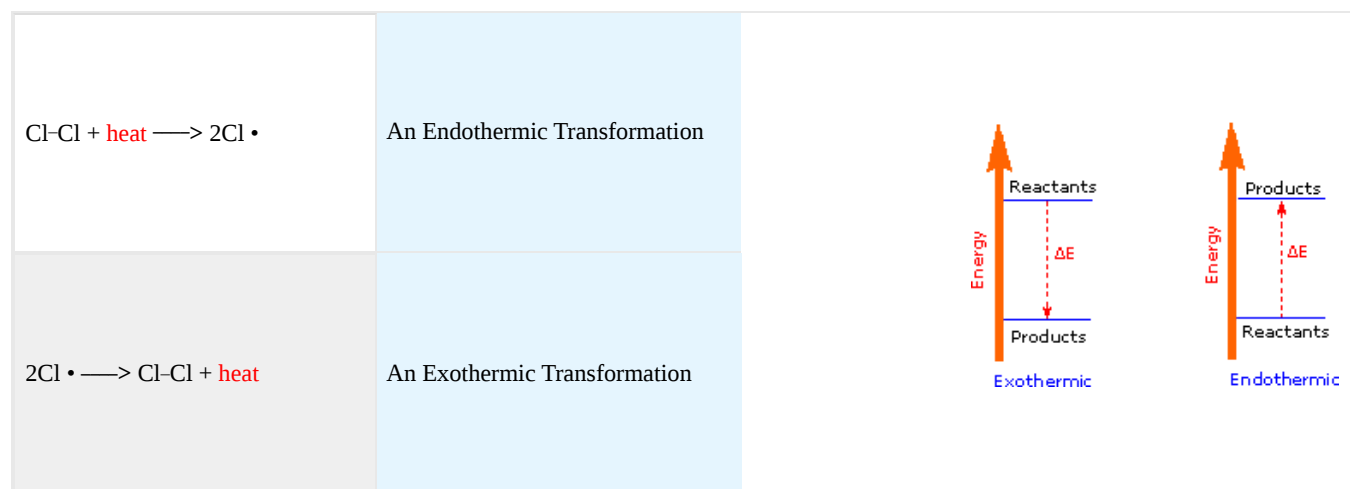
 Two pictures are shown and labeled a and b. Picture a shows a metal railroad tie being cut with the flame of an acetylene torch. Picture b shows a chemical cold pack containing ammonium nitrate.

(a) An oxyacetylene torch produces heat by the combustion of acetylene in oxygen. The energy released by this exothermic reaction heats and then melts the metal being cut. The sparks are tiny bits of the molten metal flying away. (b) A cold pack uses an endothermic process to create the sensation of cold. (credit a: modification of work by "Skatebiker"/Wikimedia commons)

Historically, energy was measured in units of calories (cal). A calorie is the amount of energy required to raise one gram of water by 1 degree C (1 kelvin). However, this quantity depends on the atmospheric pressure and the starting temperature of the water. The ease of measurement of energy changes in calories has meant that the calorie is still frequently used. The Calorie (with a capital C), or large calorie, commonly used in quantifying food energy content, is a kilocalorie. The SI unit of heat, work and energy is the joule. A joule (J) is defined as the amount of energy used when a force of 1 newton moves an object 1 meter. It is named in honor of the English physicist James Prescott Joule. One joule is equivalent to $1 \text{ kg m}^2/\text{s}^2$, which is also called 1 newton-meter. A kilojoule (kJ) is 1000 joules. To standardize its definition, 1 calorie has been set to equal 4.184 joules.

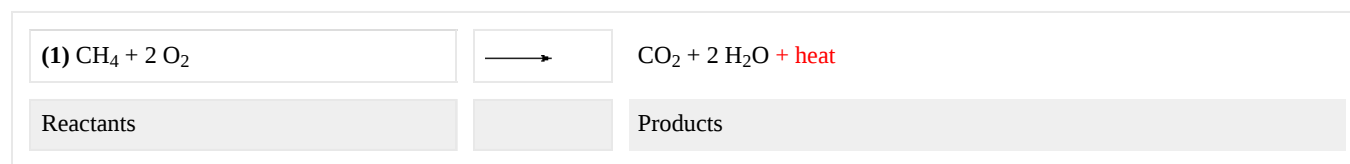
Chemists are concerned with both physical and chemical change, and the energy associated with each. Chemical reactions involve the making and breaking of covalent bonds and these transformations have energy consequences. Heat is the most common form of energy introduced into or released from chemical reactions, but light, electrical current and sound may also be involved. Also, if gases are produced in a reaction, the work done on the surroundings by volume expansion becomes part of the energy balance, and in explosive reactions may be the major consequence. Of course, reactions that consume a gas undergo an opposite volume change and work is done on the system by the surroundings.

The covalent bond of a chlorine molecule provides a simple example of the energy changes associated with bond breaking and bond making. This bond may be broken by the introduction of heat or light energy and it has been determined that 57.9 kcal/mol (242.3 kJ/mol) is required to do so. This is shown in the first of the following reactions; the second reaction describes the reverse bond-forming process. The energy absorbed or released in these reactions is referred to as the bond dissociation energy. If the bond dissociation energy is introduced from the surroundings in the form of heat, the transformation is said to be endothermic. If heat passes from the system to the surroundings, the transformation is termed exothermic. Using our initial terminology, we may say that the covalently bonded system has a lower potential energy than the unbonded diatomic system. Indeed, it is helpful to think of exothermic reactions as proceeding from a higher energy (less stable) reactant state to a lower energy (more stable) product state, as shown in the diagram on the right. Some basic principles of reaction energetics were [discussed earlier](#).



In more complex chemical reactions some (or even all) of the bonds that hold together the atoms of reactant and product molecules may be broken while other bonds are formed. Energy is required to break bonds and since different bonds have different bond dissociation energies, there is often a significant overall energy change in the course of a reaction. In the combustion of methane,

for example, all six bonds in the reactant molecules are broken and six new bonds are formed in the product molecules (equation 1). To analyze such reactions we need to keep track of and evaluate heat changes in a precise and systematic manner.



Glossary

bond dissociation energy

The energy absorbed or released in these reactions

exothermic process

A change that releases heat

endothermic process

A change that absorbs heat

kinetic energy

the energy that an object possesses because of its motion

potential energy

the energy an object has because of its relative position, composition or condition

7.2: Energy in chemical reactions is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

7.3: Equilibrium

20

Harper College Chemistry Department

Le Chatelier's Principle

Derived from [Shifting Equilibria: Le Chatelier's Principle](#) by OpenStax

Page by: OpenStax Chemistry: Atom First

▼ Summary

Reactions proceed in both directions (reactants go to products and products go to reactants). We can tell a reaction is at equilibrium if the reaction quotient (Q) is equal to the equilibrium constant (K). We next address what happens when a system at equilibrium is disturbed so that Q is no longer equal to K . If a system at equilibrium is subjected to a perturbation or stress (such as a change in concentration) the position of equilibrium changes. Since this stress affects the concentrations of the reactants and the products, the value of Q will no longer equal the value of K . To re-establish equilibrium, the system will either shift toward the products (if $Q < K$) or the reactants (if $Q > K$) until Q returns to the same value as K .

This process is described by Le Chatelier's principle: When a chemical system at equilibrium is disturbed, it returns to equilibrium by counteracting the disturbance. As described in the previous paragraph, the disturbance causes a change in Q ; the reaction will shift to re-establish $Q = K$.

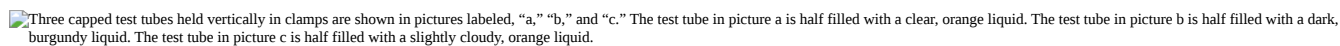
Predicting the Direction of a Reversible Reaction

Le Chatelier's principle can be used to predict changes in equilibrium concentrations when a system that is at equilibrium is subjected to a stress. However, if we have a mixture of reactants and products that have not yet reached equilibrium, the changes necessary to reach equilibrium may not be so obvious. In such a case, we can compare the values of Q and K for the system to predict the changes.

Effect of Change in Concentration on Equilibrium

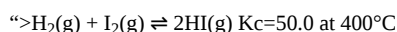
A chemical system at equilibrium can be temporarily shifted out of equilibrium by adding or removing one or more of the reactants or products. The concentrations of both reactants and products then undergo additional changes to return the system to equilibrium.

The stress on the system in [Figure 1](#) is the reduction of the equilibrium concentration of SCN^- (lowering the concentration of one of the reactants would cause Q to be larger than K). As a consequence, Le Chatelier's principle leads us to predict that the concentration of $\text{Fe}(\text{SCN})^{2+}$ should decrease, increasing the concentration of SCN^- part way back to its original concentration and increasing the concentration of Fe^{3+} above its initial equilibrium concentration.



(a) The test tube contains 0.1 M Fe^{3+} . (b) Thiocyanate ion has been added to solution in (a), forming the red $\text{Fe}(\text{SCN})^{2+}$ ion, " $\text{Fe}^{3+}(\text{aq}) + \text{SCN}^-(\text{aq}) \rightleftharpoons \text{Fe}(\text{SCN})^{2+}(\text{aq})$ ". (c) Silver nitrate has been added to the solution in (b), precipitating some of the SCN^- as the white solid AgSCN , " $\text{Ag}^+(\text{aq}) + \text{SCN}^-(\text{aq}) \rightleftharpoons \text{AgSCN}(\text{s})$ ". The decrease in the SCN^- concentration shifts the first equilibrium in the solution to the left, decreasing the concentration (and lightening color) of the $\text{Fe}(\text{SCN})^{2+}$ (credit: modification of work by Mark Ott).

The effect of a change in concentration on a system at equilibrium is illustrated further by the equilibrium of this chemical reaction:



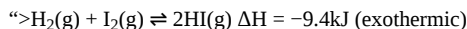
The numeric values for this example have been determined experimentally. A mixture of gases at 400°C with $[\text{H}_2] = [\text{I}_2] = 0.221 \text{ M}$ and $[\text{HI}] = 1.563 \text{ M}$ is at equilibrium; for this mixture, $Q_c = K_c = 50.0$. If H_2 is introduced into the system so quickly that its concentration doubles before it begins to react (new $[\text{H}_2] = 0.442 \text{ M}$), the reaction will shift so that a new equilibrium is reached, at which $[\text{H}_2] = 0.374 \text{ M}$, $[\text{I}_2] = 0.153 \text{ M}$ and $[\text{HI}] = 1.692 \text{ M}$. ">

We have stressed this system by introducing additional H_2 . The stress is relieved when the reaction shifts to the right, using up some (but not all) of the excess H_2 , reducing the amount of uncombined I_2 and forming additional HI .

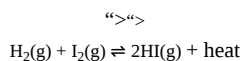
Effect of Change in Temperature on Equilibrium

Changing concentration or pressure disrupts an equilibrium because the reaction quotient is shifted away from the equilibrium value. Changing the temperature of a system at equilibrium has a different effect. A change in temperature actually changes the value of the equilibrium constant. However, we can qualitatively predict the effect of the temperature change by treating it as a stress on the system and applying Le Chatelier's principle.

When hydrogen reacts with gaseous iodine, heat is evolved.



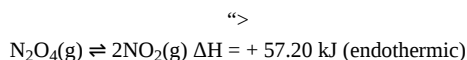
Because this reaction is exothermic, we can write it with heat as a product.



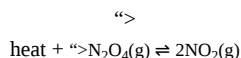
Increasing the temperature of the reaction increases the internal energy of the system. Thus, increasing the temperature has the effect of increasing the amount of one of the products of this reaction. The reaction shifts to the left to relieve the stress and there is an increase in the concentration of H_2 and I_2 and a reduction in the concentration of HI . Lowering the temperature of this system reduces the amount of energy present, favors the production of heat and favors the formation of hydrogen iodide.

When we change the temperature of a system at equilibrium, the equilibrium constant for the reaction changes. Lowering the temperature in the HI system increases the equilibrium constant. At the new equilibrium the concentration of HI has increased and the concentrations of H_2 and I_2 decreased. Raising the temperature decreases the value of the equilibrium constant, from 67.5 at 357°C to 50.0 at 400°C .

Temperature affects the equilibrium between NO_2 and N_2O_4 in this reaction



The positive ΔH value tells us that the reaction is endothermic and could be written



At higher temperatures, the gas mixture has a deep brown color, indicative of a significant amount of brown NO_2 molecules. However, if we put a stress on the system by cooling the mixture (withdrawing energy), the equilibrium shifts to the left to supply some of the energy lost by cooling. The concentration of colorless N_2O_4 increases and the concentration of brown NO_2 decreases, causing the brown color to fade.

This interactive [animation](#) allows you to apply Le Chatelier's principle to predict the effects of changes in concentration, pressure, and temperature on reactant and product concentrations.

Key Concepts and Summary

Systems at equilibrium can be disturbed by changes to temperature and concentration. The system's response to these disturbances is described by Le Chatelier's principle: The system will respond in a way that counteracts the disturbance.

Effects of Disturbances of Equilibrium and K			
Disturbance	Observed Change as Equilibrium is Restored	Direction of Shift	Effect on K
reactant added	added reactant is partially consumed	toward products	none
product added	added product is partially consumed	toward reactants	none
temperature increase	heat is absorbed	toward products for endothermic, toward reactants for exothermic	changes
temperature decrease	heat is given off	toward reactants for endothermic, toward products for exothermic	changes

Chemistry End of Chapter Exercises

- What property of a reaction can we use to predict the effect of a change in temperature on the value of an equilibrium constant?
- Suggest four ways in which the concentration of hydrazine, N_2H_4 , could be increased in an equilibrium described by the following equation:

$$\text{N}_2(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{N}_2\text{H}_4(\text{g}) \quad \Delta H = 95 \text{ kJ}$$
- Suggest four ways in which the concentration of PH_3 could be increased in an equilibrium described by the following equation:

$$\text{P}_4(\text{g}) + 6\text{H}_2(\text{g}) \rightleftharpoons 4\text{PH}_3(\text{g}) \quad \Delta H = 110.5 \text{ kJ}$$
- How will an increase in temperature affect each of the following equilibria? How will a decrease in the volume of the reaction vessel affect each?
 - $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \quad \Delta H = -92 \text{ kJ}$
 - $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) \quad \Delta H = +181 \text{ kJ}$
 - $2\text{O}_3(\text{g}) \rightleftharpoons 3\text{O}_2(\text{g}) \quad \Delta H = -285 \text{ kJ}$
 - $\text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{CaCO}_3(\text{s}) \quad \Delta H = -176 \text{ kJ}$
- How will an increase in temperature affect each of the following equilibria? How will a decrease in the volume of the reaction vessel affect each?
 - $2\text{H}_2\text{O}(\text{g}) \rightleftharpoons 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = +484 \text{ kJ}$
 - $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \quad \Delta H = -92.2 \text{ kJ}$
 - $2\text{Br}(\text{g}) \rightleftharpoons \text{Br}_2(\text{g}) \quad \Delta H = -224 \text{ kJ}$
 - $\text{H}_2(\text{g}) + \text{I}_2(\text{s}) \rightleftharpoons 2\text{HI}(\text{g}) \quad \Delta H = +53 \text{ kJ}$
- Ammonia is a weak base that reacts with water according to this equation:

$$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$$
 Will any of the following increase the percent of ammonia that is converted to the ammonium ion in water?
 - Addition of NaOH
 - Addition of HCl
 - Addition of NH_4Cl
- Acetic acid is a weak acid that reacts with water according to this equation:

$$\text{CH}_3\text{CO}_2\text{H}(\text{aq}) + \text{H}_2\text{O}(\text{aq}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CH}_3\text{CO}_2^-(\text{aq})$$
 Will any of the following increase the percent of acetic acid that reacts and produces CH_3CO_2^- ion?
 - Addition of HCl
 - Addition of NaOH
 - Addition of NaCH_3CO_2

Glossary

Le Chatelier's principle

when a chemical system at equilibrium is disturbed, it returns to equilibrium by counteracting the disturbance

position of equilibrium

concentrations or partial pressures of components of a reaction at equilibrium (commonly used to describe conditions before a disturbance)

stress

change to a reaction's conditions that may cause a shift in the equilibrium

7.3: [Equilibrium](#) is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

7.4: Catalysis

21

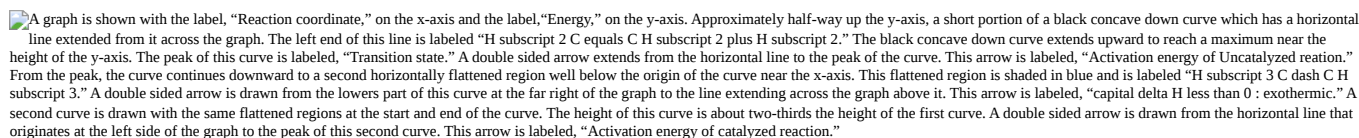
Harper College Chemistry Department

Derived from [Catalysis](#) by OpenStax

Page by: OpenStax Chemistry: Atom First

▼ Summary

A catalyst speeds up the rate of a reaction by lowering the activation energy; in addition, the catalyst is regenerated in the process. Several reactions that are energetically favorable in the absence of a catalyst only occur at a reasonable rate when a catalyst is present. One example is hydrogenation, a process used in food industries to convert unsaturated fats to saturated fats. A comparison of the reaction coordinate diagrams (also known as energy diagrams) for catalyzed and uncatalyzed hydrogenation of a simple hydrocarbon molecule is shown in [Figure](#).

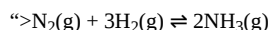
A graph is shown with the label, "Reaction coordinate," on the x-axis and the label, "Energy," on the y-axis. Approximately half-way up the y-axis, a short portion of a black concave down curve which has a horizontal line extended from it across the graph. The left end of this line is labeled "H subscript 2 C equals C H subscript 2 plus H subscript 2." The black concave down curve extends upward to reach a maximum near the height of the y-axis. The peak of this curve is labeled, "Transition state." A double sided arrow extends from the horizontal line to the peak of the curve. This arrow is labeled, "Activation energy of Uncatalyzed reaction." From the peak, the curve continues downward to a second horizontally flattened region well below the origin of the curve near the x-axis. This flattened region is shaded in blue and is labeled "H subscript 3 C dash C H subscript 3." A double sided arrow is drawn from the lower part of this curve at the far right of the graph to the line extending across the graph above it. This arrow is labeled, "capital delta H less than 0 : exothermic." A second curve is drawn with the same flattened regions at the start and end of the curve. The height of this curve is about two-thirds the height of the first curve. A double sided arrow is drawn from the horizontal line that originates at the left side of the graph to the peak of this second curve. This arrow is labeled, "Activation energy of catalyzed reaction."

This graph compares the reaction coordinates for catalyzed and uncatalyzed alkene hydrogenation.

Catalysts Do Not Affect Equilibrium

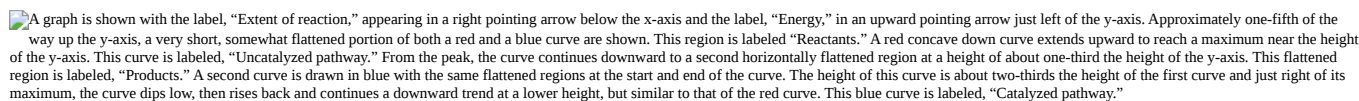
A catalyst can speed up the rate of a reaction. Though this increase in reaction rate may cause a system to reach equilibrium more quickly (by speeding up the forward and reverse reactions), a catalyst has no effect on the value of an equilibrium constant nor on equilibrium concentrations.

The interplay of changes in concentration or pressure, temperature, and the lack of an influence of a catalyst on a chemical equilibrium is illustrated in the industrial synthesis of ammonia from nitrogen and hydrogen according to the equation:



">A large quantity of ammonia is manufactured by this reaction. Each year, ammonia is among the top 10 chemicals, by mass, manufactured in the world. About 2 billion pounds are manufactured in the United States each year. Ammonia plays a vital role in our global economy. It is used in the production of fertilizers and is, itself, an important fertilizer for the growth of corn, cotton and other crops. Large quantities of ammonia are converted to nitric acid, which plays an important role in the production of fertilizers, explosives, plastics, dyes, and fibers, and is also used in the steel industry.

Catalysts function by providing an alternate reaction mechanism that has a lower activation energy than would be found in the absence of the catalyst. In some cases, the catalyzed mechanism may include additional steps, as depicted in the reaction diagrams shown in [Figure 2](#). This lower activation energy results in an increase in rate as described by the Arrhenius equation. Note that a catalyst decreases the activation energy for both the forward and the reverse reactions and hence *accelerates both the forward and the reverse reactions*. Consequently, the presence of a catalyst will permit a system to reach equilibrium more quickly, but it has no effect on the position of the equilibrium as reflected in the value of its equilibrium constant (see the chapter on chemical equilibrium).

A graph is shown with the label, "Extent of reaction," appearing in a right pointing arrow below the x-axis and the label, "Energy," in an upward pointing arrow just left of the y-axis. Approximately one-fifth of the way up the y-axis, a very short, somewhat flattened portion of both a red and a blue curve are shown. This region is labeled "Reactants." A red concave down curve extends upward to reach a maximum near the height of the y-axis. This curve is labeled, "Uncatalyzed pathway." From the peak, the curve continues downward to a second horizontally flattened region at a height of about one-third the height of the y-axis. This flattened region is labeled, "Products." A second curve is drawn in blue with the same flattened regions at the start and end of the curve. The height of this curve is about two-thirds the height of the first curve and just right of its maximum, the curve dips low, then rises back and continues a downward trend at a lower height, but similar to that of the red curve. This blue curve is labeled, "Catalyzed pathway."

This potential energy diagram shows the effect of a catalyst on the activation energy. The catalyst provides a different reaction path with a lower activation energy. As shown, the catalyzed pathway involves a two-step mechanism (note the presence of two transition states) and an intermediate species (represented by the valley between the two transitions states).

Using Reaction Diagrams to Compare Catalyzed Reactions The two reaction diagrams here represent the same reaction: one without a catalyst and one with a catalyst. Identify which diagram suggests the presence of a catalyst, and determine the activation energy for the catalyzed reaction:

 In this figure, two graphs are shown. The x-axes are labeled, "Extent of reaction," and the y-axes are labeled, "Energy (k J)." The y-axes are marked off from 0 to 50 in intervals of five. In a, a blue curve is shown. It begins with a horizontal segment at about 6. The curve then rises sharply near the middle to reach a maximum of about 32 and similarly falls to another horizontal segment at about 10. In b, the curve begins and ends similarly, but the maximum reached near the center of the graph is only 20.

Solution:

A catalyst does not affect the energy of reactant or product, so those aspects of the diagrams can be ignored; they are, as we would expect, identical in that respect. There is, however, a noticeable difference in the transition state, which is distinctly lower in diagram (b) than it is in (a). This indicates the use of a catalyst in diagram (b). The activation energy is the difference between the energy of the starting reagents and the transition state—a maximum on the reaction coordinate diagram. The reagents are at 6 kJ and the transition state is at 20 kJ, so the activation energy can be calculated as follows, $E_a = 20\text{kJ} - 6\text{kJ} = 14\text{kJ}$.

Check Your Learning

Determine which of the two diagrams here (both for the same reaction) involves a catalyst, and identify the activation energy for the catalyzed reaction:

 In this figure, two graphs are shown. The x-axes are labeled, "Extent of reaction," and the y-axes are labeled, "Energy (k J)." The y-axes are marked off from 0 to 100 at intervals of 10. In a, a blue curve is shown. It begins with a horizontal segment at about 10. The curve then rises sharply near the middle to reach a maximum of about 91, then sharply falls to about 52, again rises sharply to about 73 and falls to another horizontal segment at about 5. In b, the curve begins and ends similarly, but the first peak reaches about 81, drops to about 55, then rises to about 77 before falling to the horizontal region at about 5.

Answer:

Diagram (b) is a catalyzed reaction with an activation energy of about 70 kJ.

Key Concepts and Summary

Catalysts affect the rate of a chemical reaction by altering its mechanism to provide a lower activation energy, but they do not affect equilibrium.

Chemistry End of Chapter Exercises

- Water gas is a 1:1 mixture of carbon monoxide and hydrogen gas and is called water gas because it is formed from steam and hot carbon in the following reaction: $\text{H}_2\text{O}(\text{g}) + \text{C}(\text{s}) \rightleftharpoons \text{H}_2(\text{g}) + \text{CO}(\text{g})$. Methanol, a liquid fuel that could possibly replace gasoline, can be prepared from water gas and hydrogen at high temperature and pressure in the presence of a suitable catalyst. What will happen to the concentrations of H_2 , CO , and CH_3OH at equilibrium if more catalyst is added?
- Nitrogen and oxygen react at high temperatures. What will happen to the concentrations of N_2 , O_2 and NO at equilibrium if a catalyst is added?
- For each of the following pairs of reaction diagrams, identify which of the pair is catalyzed:

(a)

 In this figure, two graphs are shown. The x-axes are labeled, "Extent of reaction," and the y-axes are labeled, "Energy (k J)." The y-axis of the first graph is marked off from 0 to 30 in intervals of 5. The y-axis of the second graph is marked off from 0 to 25 by intervals of 5. In a, a blue curve is shown. It begins with a horizontal region at about 12. The curve then rises sharply near the middle to reach a maximum of about 24 and similarly falls to another horizontal segment at about 5. In b, the curve begins and ends similarly, but the maximum reached near the center of the graph is only 20.

(b)

 In this figure, two graphs are shown. The x-axes are labeled, "Extent of reaction," and the y-axes are labeled, "Energy." The y-axes are marked off from 0 to 50 in intervals of 5. In a, a blue curve is shown. It begins with a horizontal region at about 2. The curve then rises sharply near the middle to reach a maximum of about 43 and similarly falls to another horizontal segment at about 15. In b, the curve begins and ends similarly, but the maximum reached near the center of the graph is only about 32.

(c)

 In this figure, two graphs are shown. The x-axes are labeled, "Extent of reaction" and the y-axes are labeled, "Energy (k J)." The y-axes are marked off from 0 to 50 at intervals of 5. In a, a blue curve is shown. It begins with a horizontal segment at about 2J. The curve then rises sharply near the middle to reach a maximum of about 46, then sharply falls to about 35, again rises to about 38 and falls to another horizontal segment at about 15. In b, the curve begins and ends similarly, but the first peak reaches about 46, drops to about 35, then rises to about 43 before falling to the horizontal region at about 15.

(d)

 In this figure, two graphs are shown. The x-axes are labeled, "Extent of reaction," and the y-axes are labeled, "Energy (k J)." The y-axes are marked off from 0 to 50 at intervals of 5. In a, a blue curve is shown. It begins with a horizontal segment at about 34. The curve then rises sharply near the middle to reach a maximum of about 45, then sharply falls to about 25, again rises sharply to about 35 and falls to another horizontal segment at about 15. In b, the curve begins and ends similarly, but the first peak reaches about 40, drops to 25, then rises to 35 before falling to the horizontal region at about 15.

Footnotes

- 1 Herrlich, P. "The Responsibility of the Scientist: What Can History Teach Us About How Scientists Should Handle Research That Has the Potential to Create Harm?" *EMBO Reports* 14 (2013): 759–764.

- 2 “The Nobel Prize in Chemistry 1995,” Nobel Prize.org, accessed February 18, 2015, http://www.nobelprize.org/nobel_priz...aureates/1995/.

Glossary

catalyst

something that speeds up the rate of a reaction by lowering the activation energy

activation energy

the energy required to change reactants into products

7.4: Catalysis is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

8: Gases, Liquids, and Solids

8.1: Kinetic Molecular Theory

8.2: Pressure

8.3: The Gas Laws

8.4: Intermolecular forces

8.5: Types of solutions

8.6: Electrolytes

8.7: Concentration

8.8: Osmosis, diffusion, and dialysis

8: Gases, Liquids, and Solids is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.1: Kinetic Molecular Theory

22

Harper College Chemistry Department

Derived from [Kinetic Molecular Theory](#) by OpenStax

Page by: OpenStax Chemistry 2e

The kinetic molecular theory (KMT) is a simple microscopic model that effectively explains the gas laws described in previous modules of this chapter. This theory is based on the following five postulates described here. (Note: The term “molecule” will be used to refer to the individual chemical species that compose the gas, although some gases are composed of atomic species, for example, the noble gases.)

1. Gases are composed of molecules that are in continuous motion, travelling in straight lines and changing direction only when they collide with other molecules or with the walls of a container.
2. The molecules composing the gas are negligibly small compared to the distances between them.
3. The pressure exerted by a gas in a container results from collisions between the gas molecules and the container walls.
4. Gas molecules exert no attractive or repulsive forces on each other or the container walls; therefore, their collisions are *elastic* (do not involve a loss of energy).
5. The average kinetic energy of the gas molecules is proportional to the kelvin temperature of the gas.

The various gas laws can be derived from the assumptions of the KMT, which have led chemists to believe that the assumptions of the theory accurately represent the properties of gas molecules. We will first look at the individual gas laws (Boyle's, Charles', Amontons', Avogadro's, and Dalton's laws) conceptually to see how the KMT explains them.

8.1: Kinetic Molecular Theory is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.2: Pressure

23

Harper College Chemistry Department

Derived from [Gas Pressure](#) by OpenStax

Page by: OpenStax Chemistry: Atom First

The earth's atmosphere exerts a pressure, as does any other gas. Although we do not normally notice atmospheric pressure, we are sensitive to pressure changes—for example, when your ears “pop” during take-off and landing while flying, or when you dive underwater. Gas pressure is caused by the force exerted by gas molecules colliding with the surfaces of objects ([Figure 1](#)). Although the force of each collision is very small, any surface of appreciable area experiences a large number of collisions in a short time, which can result in a high pressure. In fact, normal air pressure is strong enough to crush a metal container when not balanced by equal pressure from inside the container.

The left side of this figure includes a graphic of the earth with an inverted rectangular prism extending from a point on it. Near the top of the image, the label, “square inch column of air molecules” is connected to the prism with a line segment. This label is also connected with a line segment to a downward pointing arrow at the right side of the figure. Beneath the arrow is a red circle labeled, “atmospheric pressure.” A narrow rectangle with a dashed line border extends from the bottom of the arrow vertically through the circle. Directly beneath this rectangle at the lower edge of the circle is a hand with a thumb appearing to be resting on a tabletop. The thumb is connected with a line segment to the label, “14.7 lbs of pressure on 1 square inch.” The red circle is sitting on top of the thumb.

The atmosphere above us exerts a large pressure on objects at the surface of the earth, roughly equal to the weight of a bowling ball pressing on an area the size of a human thumbnail.

A dramatic [illustration](#) of atmospheric pressure is provided in this brief video, which shows a railway tanker car imploding when its internal pressure is decreased. A smaller scale [demonstration](#) of this phenomenon is briefly explained.

Atmospheric pressure is caused by the weight of the column of air molecules in the atmosphere above an object, such as the tanker car. At sea level, this pressure is roughly the same as that exerted by a full-grown African elephant standing on a doormat, or a typical bowling ball resting on your thumbnail. These may seem like huge amounts, and they are, but life on earth has evolved under such atmospheric pressure. If you actually perch a bowling ball on your thumbnail, the pressure experienced is *twice* the usual pressure, and the sensation is unpleasant.

In general, pressure is defined as the force exerted on a given area: $\text{pressure} = \text{force}/\text{area}$. Note that pressure is directly proportional to force and inversely proportional to area. Thus, pressure can be increased either by increasing the amount of force or by decreasing the area over which it is applied; pressure can be decreased by decreasing the force or increasing the area.

Let's apply this concept to determine which would be more likely to fall through thin ice in [Figure 2](#)—the elephant or the figure skater?

This figure includes two photographs. Figure a is a photo of a large gray elephant on grassy, beige terrain. Figure b is a photo of a figure skater with her right skate on the ice, upper torso lowered, arms extended upward behind her chest, and left leg extended upward behind her.

Although (a) an elephant's weight is large, creating a very large force on the ground, (b) the figure skater exerts a much higher pressure on the ice due to the small surface area of her skates. (credit a: modification of work by Guido da Rozze; credit b: modification of work by Ryosuke Yagi)

A large African elephant can weigh 7 tons, supported on four feet, each with a diameter of about 1.5 ft (footprint area of 250 in²), so the pressure exerted by each foot is about 14 lb/in²:

$$\text{pressure}_{\text{elephant, foot}} = 14000 \frac{\text{lb}}{\text{elephant}} \times \frac{\text{elephant}}{4 \text{ feet}} \times \frac{1 \text{ foot}}{250 \text{ in}^2} = \frac{14 \text{ lb}}{\text{in}^2}$$

The figure skater weighs about 120 lbs, supported on two skate blades, each with an area of about 2 in², so the pressure exerted by each blade is about 30 lb/in²:

$$\text{pressure}_{\text{skate, blade}} = 120 \frac{\text{lb}}{\text{skater}} \times \frac{\text{skater}}{2 \text{ blades}} \times \frac{1 \text{ blade}}{2 \text{ in}^2} = \frac{30 \text{ lb}}{\text{in}^2}$$

Even though the elephant is more than one hundred-times heavier than the skater, it exerts less than one-half of the pressure and would therefore be less likely to fall through thin ice. On the other hand, if the skater removes her skates and stands with bare feet (or regular footwear) on the ice, the larger area over which her weight is applied greatly reduces the pressure exerted:

$$\text{pressure}_{\text{human, foot}} = 120 \frac{\text{lb}}{\text{skater}} \times \frac{\text{skater}}{2 \text{ feet}} \times \frac{1 \text{ foot}}{30 \text{ in}^2} = \frac{2 \text{ lb}}{\text{in}^2}$$

The SI unit of pressure is the pascal (Pa), with 1 Pa = 1 N/m², where N is the newton, a unit of force defined as 1 kg m/s². One pascal is a small pressure; in many cases, it is more convenient to use units of kilopascal (1 kPa = 1000 Pa) or bar (1 bar = 100,000 Pa). In the United States, pressure is often measured in pounds of force on an area of one square inch—pounds per square inch (psi)—for example, in car tires. Pressure can also be measured using the unit atmosphere (atm), which originally represented the average sea level air pressure at the approximate latitude of Paris (45°). [Table 1](#) provides some information on these and a few other common units for pressure measurements

Table 1: Pressure Units

Unit Name and Abbreviation	Definition or Relation to Other Unit
pascal (Pa)	1 Pa = 1 N/m ² recommended IUPAC unit
kilopascal (kPa)	1 kPa = 1000 Pa
pounds per square inch (psi)	air pressure at sea level is ~14.7 psi
atmosphere (atm)	1 atm = 101,325 Pa = 760 mmHg air pressure at sea level is ~1 atm
bar (bar, or b)	1 bar = 100,000 Pa (exactly) commonly used in meteorology
millibar (mbar, or mb)	1000 mbar = 1 bar
inches of mercury (in. Hg)	1 in. Hg = 3386 Pa = 25.4 mm Hg used by aviation industry, also some weather reports
torr	1 torr = 1/760 atm named after Evangelista Torricelli, inventor of the barometer
millimeters of mercury (mm Hg)	1 mm Hg ~1 torr

Conversion of Pressure Units

The United States National Weather Service reports pressure in both inches of Hg and millibars. Convert a pressure of 29.2 in. Hg into:

- (a) torr
- (b) atm
- (c) kPa
- (d) mbar

Solution

This is a unit conversion problem. The relationships between the various pressure units are given in [Table 1](#).

- (a) $29.2 \text{ in. Hg} \times \frac{25.4 \text{ mm Hg}}{1 \text{ in. Hg}} \times \frac{1 \text{ torr}}{2.54 \text{ mm Hg}} = 742 \text{ torr}$
- (b) $29.2 \text{ in. Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 0.976 \text{ atm}$
- (c) $0.976 \text{ atm} \times \frac{101,325 \text{ Pa}}{1 \text{ atm}} \times \frac{1 \text{ kPa}}{1000 \text{ Pa}} = 98.9 \text{ kPa}$
- (d) $98.9 \text{ kPa} \times \frac{1000 \text{ Pa}}{1 \text{ kPa}} \times \frac{1 \text{ bar}}{100,000 \text{ Pa}} \times \frac{1,000 \text{ mbar}}{1 \text{ bar}} = 989 \text{ mbar}$

Check Your Learning

A typical barometric pressure in Kansas City is 740 torr. What is this pressure in a) atmospheres, b) millimeters of mercury, c) kilopascals, and d) bar?

Answer:

- a) 0.974 atm; b) 740 mm Hg; c) 98.7 kPa; d) 0.987 bar

BAROMETER

We can measure atmospheric pressure, the force exerted by the atmosphere on the earth's surface, with a barometer ([Figure 3](#)). A barometer is a glass tube that is closed at one end, filled with a nonvolatile liquid such as mercury and then inverted and immersed in a container of that liquid. The atmosphere exerts pressure on the liquid outside the tube, the column of liquid exerts pressure inside the tube, and the pressure at the liquid surface is the same inside and outside the tube. The height of the liquid in the tube is, therefore, proportional to the pressure exerted by the atmosphere.

This figure shows two barometers. The barometer to the left contains a shallow reservoir, or open container, of mercury. A narrow tube extends upward from the reservoir above the reservoir. This tube is sealed at the top. To the right, a second similar setup is shown with a reservoir filled with water. Line segments connect the label "vacuum" to the tops of the two narrow tubes. The tube on the left shows the mercury in the reservoir extending in a column upward in the narrow tube. Similarly, the tube on the right shows the water in the reservoir extending upward into the narrow tube. Double-headed arrows extend from the surface of each liquid in the reservoir to the top of the liquid in each tube. A narrow column or bar extends from the surface of the reservoir to the same height. This bar is labeled "atmospheric pressure." The level of the water in its tube is significantly higher than the level of mercury in its tube.

In a barometer, the height, h , of the column of liquid is used as a measurement of the air pressure. Using very dense liquid mercury (left) permits the construction of reasonably sized barometers, whereas using water (right) would require a barometer more than 30 feet tall.

If the liquid is water, normal atmospheric pressure will support a column of water over 10 meters high, which is rather inconvenient for making (and reading) a barometer. Because mercury (Hg) is about 13.6-times denser than water, a mercury barometer only needs to be 1/13.6 as tall as a water barometer—a more suitable size. Standard atmospheric pressure of 1 atm at sea level (101,325 Pa) corresponds to a column of mercury that is about 760 mm (29.92 in.) high. The torr was originally intended to be a unit equal to one millimeter of mercury. The pressure exerted by a fluid due to gravity is known as hydrostatic pressure, p :

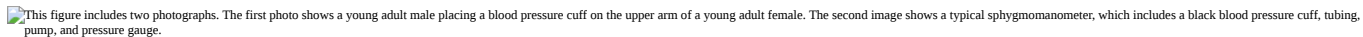
$$p = h \times \rho \times g$$

where h is the height of the fluid, ρ is the density of the fluid, and g is acceleration due to gravity.

[Incase you are wondering](#)

MEASURING BLOOD PRESSURE

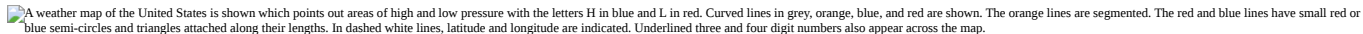
Blood pressure is measured using a device called a sphygmomanometer (Greek *sphygmos* = “pulse”). It consists of an inflatable cuff to restrict blood flow, a manometer to measure the pressure, and a method of determining when blood flow begins and when it becomes impeded (Figure 4). Since its invention in 1881, it has been an essential medical device. There are many types of sphygmomanometers: manual ones that require a stethoscope and are used by medical professionals; mercury ones, used when the most accuracy is required; less accurate mechanical ones; and digital ones that can be used with little training but that have limitations. When using a sphygmomanometer, the cuff is placed around the upper arm and inflated until blood flow is completely blocked, then slowly released. As the heart beats, blood forced through the arteries causes a rise in pressure. This rise in pressure at which blood flow begins is the *systolic pressure*—the peak pressure in the cardiac cycle. When the cuff’s pressure equals the arterial systolic pressure, blood flows past the cuff, creating audible sounds that can be heard using a stethoscope. This is followed by a decrease in pressure as the heart’s ventricles prepare for another beat. As cuff pressure continues to decrease, eventually sound is no longer heard; this is the *diastolic pressure*—the lowest pressure (resting phase) in the cardiac cycle. Blood pressure units from a sphygmomanometer are in terms of millimeters of mercury (mm Hg).

 This figure includes two photographs. The first photo shows a young adult male placing a blood pressure cuff on the upper arm of a young adult female. The second image shows a typical sphygmomanometer, which includes a black blood pressure cuff, tubing, pump, and pressure gauge.

(a) A medical technician prepares to measure a patient’s blood pressure with a sphygmomanometer. (b) A typical sphygmomanometer uses a valved rubber bulb to inflate the cuff and a diaphragm gauge to measure pressure. (credit a: modification of work by Master Sgt. Jeffrey Allen)

METEOROLOGY, CLIMATOLOGY, AND ATMOSPHERIC SCIENCE

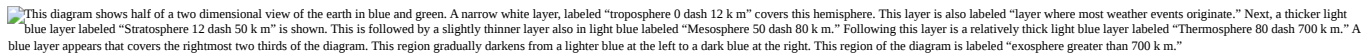
Throughout the ages, people have observed clouds, winds, and precipitation, trying to discern patterns and make predictions: when it is best to plant and harvest; whether it is safe to set out on a sea voyage; and much more. We now face complex weather and atmosphere-related challenges that will have a major impact on our civilization and the ecosystem. Several different scientific disciplines use chemical principles to help us better understand weather, the atmosphere, and climate. These are meteorology, climatology, and atmospheric science. Meteorology is the study of the atmosphere, atmospheric phenomena, and atmospheric effects on earth’s weather. Meteorologists seek to understand and predict the weather in the short term, which can save lives and benefit the economy. Weather forecasts (Figure 5) are the result of thousands of measurements of air pressure, temperature, and the like, which are compiled, modeled, and analyzed in weather centers worldwide.

 A weather map of the United States is shown which points out areas of high and low pressure with the letters H in blue and L in red. Curved lines in grey, orange, blue, and red are shown. The orange lines are segmented. The red and blue lines have small red or blue semi-circles and triangles attached along their lengths. In dashed white lines, latitude and longitude are indicated. Underlined three and four digit numbers also appear across the map.

Meteorologists use weather maps to describe and predict weather. Regions of high (H) and low (L) pressure have large effects on weather conditions. The gray lines represent locations of constant pressure known as isobars. (credit: modification of work by National Oceanic and Atmospheric Administration)

In terms of weather, low-pressure systems occur when the earth’s surface atmospheric pressure is lower than the surrounding environment: Moist air rises and condenses, producing clouds. Movement of moisture and air within various weather fronts instigates most weather events.

The atmosphere is the gaseous layer that surrounds a planet. Earth’s atmosphere, which is roughly 100–125 km thick, consists of roughly 78.1% nitrogen and 21.0% oxygen, and can be subdivided further into the regions shown in Figure 6: the exosphere (furthest from earth, > 700 km above sea level), the thermosphere (80–700 km), the mesosphere (50–80 km), the stratosphere (second lowest level of our atmosphere, 12–50 km above sea level), and the troposphere (up to 12 km above sea level, roughly 80% of the earth’s atmosphere by mass and the layer where most weather events originate). As you go higher in the troposphere, air density and temperature both decrease.

 This diagram shows half of a two dimensional view of the earth in blue and green. A narrow white layer, labeled “troposphere 0 dash 12 k m” covers this hemisphere. This layer is also labeled “layer where most weather events originate.” Next, a thicker light blue layer labeled “Stratosphere 12 dash 50 k m” is shown. This is followed by a slightly thinner layer also in light blue labeled “Mesosphere 50 dash 80 k m.” Following this layer is a relatively thick light blue layer labeled “Thermosphere 80 dash 700 k m.” A blue layer appears that covers the rightmost two thirds of the diagram. This region gradually darkens from a lighter blue at the left to a dark blue at the right. This region of the diagram is labeled “exosphere greater than 700 k m.”

Earth’s atmosphere has five layers: the troposphere, the stratosphere, the mesosphere, the thermosphere, and the exosphere.

Climatology is the study of the climate, averaged weather conditions over long time periods, using atmospheric data. However, climatologists study patterns and effects that occur over decades, centuries, and millennia, rather than shorter time frames of hours, days, and weeks like meteorologists. Atmospheric science is an even broader field, combining meteorology, climatology, and other scientific disciplines that study the atmosphere.

8.2: Pressure is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.3: The Gas Laws

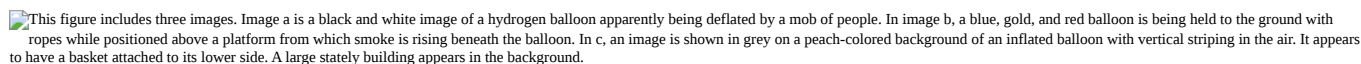
24

Harper College Chemistry Department

Derived from [Relating pressure volume amount and temperature the ideal gas law](#) by OpenStax

Page by: OpenStax Chemistry: Atom First

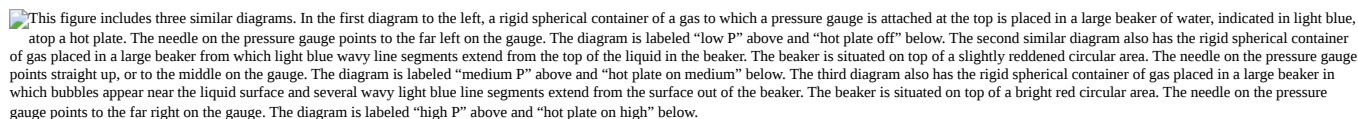
During the seventeenth and especially eighteenth centuries, driven both by a desire to understand nature and a quest to make balloons in which they could fly ([Figure 1](#)), a number of scientists established the relationships between the macroscopic physical properties of gases, that is, pressure, volume, temperature and amount of gas. Although their measurements were not precise by today's standards, they were able to determine the mathematical relationships between pairs of these variables (e.g., pressure and temperature, pressure and volume) that hold for an *ideal* gas—a hypothetical construct that real gases approximate under certain conditions. Eventually, these individual laws were combined into a single equation—the *ideal gas law*—that relates gas quantities for gases and is quite accurate for low pressures and moderate temperatures. We will consider the key developments in individual relationships (for pedagogical reasons not quite in historical order), then put them together in the ideal gas law.

 This figure includes three images. Image a is a black and white image of a hydrogen balloon apparently being deflated by a mob of people. In image b, a blue, gold, and red balloon is being held to the ground with ropes while positioned above a platform from which smoke is rising beneath the balloon. In c, an image is shown in grey on a peach-colored background of an inflated balloon with vertical striping in the air. It appears to have a basket attached to its lower side. A large stately building appears in the background.

In 1783, the first (a) hydrogen-filled balloon flight, (b) manned hot air balloon flight, and (c) manned hydrogen-filled balloon flight occurred. When the hydrogen-filled balloon depicted in (a) landed, the frightened villagers of Gonesse reportedly destroyed it with pitchforks and knives. The launch of the latter was reportedly viewed by 400,000 people in Paris.

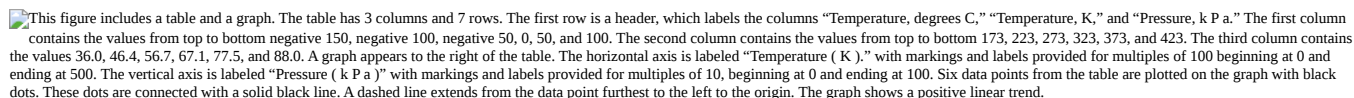
Pressure and Temperature: Gay-Lussac's Law (or Amontons's Law)

Imagine filling a rigid container (this indicates volume is constant) attached to a pressure gauge with gas and then sealing the container so that no gas may escape (this indicates constant moles). If the container is cooled, the gas inside likewise gets colder and its pressure is observed to decrease. Since the container is rigid and tightly sealed, both the volume and number of moles of gas remain constant. If we heat the sphere, the gas inside gets hotter ([Figure 2](#)) and the pressure increases.

 This figure includes three similar diagrams. In the first diagram to the left, a rigid spherical container of a gas to which a pressure gauge is attached at the top is placed in a large beaker of water, indicated in light blue, atop a hot plate. The needle on the pressure gauge points to the far left on the gauge. The diagram is labeled "low P" above and "hot plate off" below. The second similar diagram also has the rigid spherical container of gas placed in a large beaker from which light blue wavy line segments extend from the top of the liquid in the beaker. The beaker is situated on top of a slightly reddened circular area. The needle on the pressure gauge points straight up, or to the middle on the gauge. The diagram is labeled "medium P" above and "hot plate on medium" below. The third diagram also has the rigid spherical container of gas placed in a large beaker in which bubbles appear near the liquid surface and several wavy light blue line segments extend from the surface out of the beaker. The beaker is situated on top of a bright red circular area. The needle on the pressure gauge points to the far right on the gauge. The diagram is labeled "high P" above and "hot plate on high" below.

The effect of temperature on gas pressure: When the hot plate is off, the pressure of the gas in the sphere is relatively low. As the gas is heated, the pressure of the gas in the sphere increases.

This relationship between temperature and pressure is observed for any sample of gas confined to a constant volume. An example of experimental pressure-temperature data is shown for a sample of air under these conditions in [Figure 3](#). We find that temperature and pressure are linearly related, and if the temperature is on the kelvin scale, then P and T are directly proportional (again, when *volume and moles of gas are held constant*); if the temperature on the kelvin scale increases by a certain factor, the gas pressure increases by the same factor.

 This figure includes a table and a graph. The table has 3 columns and 7 rows. The first row is a header, which labels the columns "Temperature, degrees C," "Temperature, K," and "Pressure, kPa." The first column contains the values from top to bottom negative 150, negative 100, negative 50, 0, 50, and 100. The second column contains the values from top to bottom 173, 223, 273, 323, 373, and 423. The third column contains the values 36.0, 46.4, 56.7, 67.1, 77.5, and 88.0. A graph appears to the right of the table. The horizontal axis is labeled "Temperature (K)" with markings and labels provided for multiples of 100 beginning at 0 and ending at 500. The vertical axis is labeled "Pressure (k P a)" with markings and labels provided for multiples of 10, beginning at 0 and ending at 100. Six data points from the table are plotted on the graph with black dots. These dots are connected with a solid black line. A dashed line extends from the data point furthest to the left to the origin. The graph shows a positive linear trend.

For a constant volume and amount of air, the pressure and temperature are directly proportional, provided the temperature is in kelvin. (Measurements cannot be made at lower temperatures because of the condensation of the gas.) When this line is extrapolated to lower pressures, it reaches a pressure of 0 at -273°C , which is 0 on the kelvin scale and the lowest possible temperature, called absolute zero.

Guillaume Amontons was the first to empirically establish the relationship between the pressure and the temperature of a gas (~1700) and Joseph Louis Gay-Lussac determined the relationship more precisely (~1800). Because of this, the pressure-temperature (P - T) relationship for gases is known as either Amontons's law or Gay-Lussac's law. Under either name, it states that *the pressure of a given amount of gas is directly proportional to its temperature on the kelvin scale when the volume is held constant*. Mathematically, this can be written:

$$P \propto T, \text{ or } P = \text{constant} \times T, \text{ or } P = k \times T$$

where $\propto$ means “is proportional to,” and k is a proportionality constant that depends on the identity, amount, and volume of the gas. For a confined, constant volume of gas, the ratio $\frac{P}{T}$ is therefore constant (i.e., $\frac{P}{T} = k$). If the gas is initially in “Condition 1” (with $P = P_1$ and $T = T_1$), and then changes to “Condition 2” (with $P = P_2$ and $T = T_2$), we have that $\frac{P_1}{T_1} = k$ and $\frac{P_2}{T_2} = k$, which reduces to $\frac{P_1}{T_1} = \frac{P_2}{T_2}$. This equation is useful for pressure-temperature calculations for a confined gas at constant volume. Note that temperatures must be on the kelvin scale for any gas law calculations (0 on the kelvin scale and the lowest possible temperature is called absolute zero).

Predicting Change in Pressure with Temperature

A can of hair spray is used until it is empty except for the propellant, isobutane gas.

- On the can is the warning “Store only at temperatures below 120 °F (48.8 °C). Do not incinerate.” Why?
- The gas in the can is initially at 24 °C and 360 kPa, and the can has a volume of 350 mL. If the can is left in a car that reaches 50 °C on a hot day, what is the new pressure in the can?

Solution

- The can contains an amount of isobutane gas at a constant volume, so if the temperature is increased by heating, the pressure will increase proportionately. High temperature could lead to high pressure, causing the can to burst. (Also, isobutane is combustible, so incineration could cause the can to explode.)
- We are looking for a pressure change due to a temperature change at constant volume, so we will use Amontons’/Gay-Lussac’s law. Taking P_1 and T_1 as the initial values, T_2 as the temperature where the pressure is unknown and P_2 as the unknown pressure, and converting °C to K, we have:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ which means that } \frac{360 \text{ kPa}}{297 \text{ K}} = \frac{P_2}{323 \text{ K}}$$

Rearranging and solving gives:

$$P_2 = \frac{360 \text{ kPa} \times 323 \text{ K}}{297 \text{ K}} = 392 \text{ kPa}$$

Check Your Learning

A sample of nitrogen, N_2 , occupies 45.0 mL at 27 °C and 600 torr. What pressure will it have if cooled to –73 °C while the volume remains constant?

Answer:

400 torr (temperature should be in Kelvin)

Volume and Temperature: Charles’ Law

If we fill a balloon with air and seal it, the balloon contains a specific amount of air at atmospheric pressure, let’s say 1 atm. If we put the balloon in a refrigerator, the gas inside gets cold and the balloon shrinks (although both the amount of gas and its pressure remain constant). If we make the balloon very cold, it will shrink a great deal and it expands again when it warms up.

This [video](#) shows how cooling and heating a gas causes its volume to decrease or increase, respectively.

These examples of the effect of temperature on the volume of a given amount of a confined gas at constant pressure are true in general: the volume increases as the temperature increases and decreases as the temperature decreases. Volume-temperature data for a 1-mole sample of methane gas at 1 atm are listed and graphed in [Figure 4](#).

This figure includes a table and a graph. The table has 3 columns and 6 rows. The first row is a header, which labels the columns “Temperature, degrees C,” “Temperature, K,” and “Pressure, kPa.” The first column contains the values from top to bottom negative 100, negative 50, 0, 100, and 200. The second column contains the values from top to bottom 173, 223, 273, 373, and 473. The third column contains the values 14.10, 18.26, 22.40, 30.65, and 38.88. A graph appears to the right of the table. The horizontal axis is labeled “Temperature (K)” with markings and labels provided for multiples of 100 beginning at 0 and ending at 300. The vertical axis is labeled “Volume (L)” with markings and labels provided for multiples of 10, beginning at 0 and ending at 30. Five data points from the table are plotted on the graph with black dots. These dots are connected with a solid black line. The graph shows a positive linear trend.

The volume and temperature are linearly related for 1 mole of methane gas at a constant pressure of 1 atm. If the temperature is in kelvin, volume and temperature are directly proportional. The line stops at 111 K because methane liquefies at this temperature; when extrapolated, it intersects the graph’s origin, representing a temperature of absolute zero.

The relationship between the volume and temperature of a given amount of gas at constant pressure is known as Charles’ law in recognition of the French scientist and balloon flight pioneer Jacques Alexandre César Charles. Charles’ law states that *the volume*

of a given amount of gas is directly proportional to its temperature on the kelvin scale when the pressure is held constant.

Mathematically, this can be written as:

$$V \propto T, \text{ or } V = \text{constant} \times T, \text{ or } V = k \times T, \text{ or } \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

with k being a proportionality constant that depends on the amount and pressure of the gas.

For a confined, constant pressure gas sample, $\frac{V}{T}$ is constant (i.e., the ratio = k), and as seen with the P - T relationship, this leads to another form of Charles's law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Predicting Change in Volume with Temperature

A sample of carbon dioxide, CO_2 , occupies 0.300 L at 10 °C and 750 torr. What volume will the gas have at 30 °C and 750 torr?

Solution

Because we are looking for the volume change caused by a temperature change at constant pressure, this is a job for Charles' law. Taking V_1 and T_1 as the initial values, T_2 as the temperature at which the volume is unknown and V_2 as the unknown volume and converting °C into K we have:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ which means that } \frac{0.300 \text{ L}}{283 \text{ K}} = \frac{V_2}{303 \text{ K}}$$

Rearranging and solving gives:

$$V_2 = \frac{0.300 \text{ L} \times 303 \text{ K}}{283 \text{ K}} = 0.321 \text{ L}$$

This answer supports our expectation from Charles's law, namely, that raising the gas temperature (from 283 K to 303 K) at a constant pressure will yield an increase in its volume (from 0.300 L to 0.321 L).

Check Your Learning

A sample of oxygen, O_2 , occupies 32.2 mL at 30 °C and 452 torr. What volume will it occupy at −70 °C and the same pressure?

Answer:
21.6 mL

Measuring Temperature with a Volume Change

Temperature is sometimes measured with a gas thermometer by observing the change in the volume of the gas as the temperature changes at constant pressure. The hydrogen in a particular hydrogen gas thermometer has a volume of 150.0 cm^3 when immersed in a mixture of ice and water (0.00 °C). When immersed in boiling liquid ammonia, the volume of the hydrogen, at the same pressure, is 131.7 cm^3 . Find the temperature of boiling ammonia on the kelvin and Celsius scales.

Solution

A volume change caused by a temperature change at constant pressure means we should use Charles' law. Taking V_1 and T_1 as the initial values, T_2 as the temperature at which the volume is unknown and V_2 as the unknown volume and converting °C into K we have:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ which means that } \frac{150.0 \text{ cm}^3}{273.15 \text{ K}} = \frac{131.7 \text{ cm}^3}{T_2}$$

Rearrangement gives

$$T_2 = 131.7 \text{ cm}^3 \times \frac{273.15 \text{ K}}{150.0 \text{ cm}^3} = 239.8 \text{ K}$$

Subtracting 273.15 from 239.8 K, we find that the temperature of the boiling ammonia on the Celsius scale is −33.4 °C.

Check Your Learning

What is the volume of a sample of ethane at 467 K and 1.1 atm, if it occupies 405 mL at 298 K and 1.1 atm?

Answer:

Volume and Pressure: Boyle's Law

If we partially fill an airtight syringe with air, the syringe contains a specific amount of air at constant temperature, say 25 °C. If we slowly push in the plunger while keeping temperature constant, the gas in the syringe is compressed into a smaller volume and its pressure increases; if we pull out the plunger, the volume increases and the pressure decreases. This example of the effect of volume on the pressure of a given amount of a confined gas is true in general. Decreasing the volume of a contained gas will increase its pressure, and increasing its volume will decrease its pressure. In fact, if the volume increases by a certain factor, the pressure decreases by the same factor and vice versa. Volume-pressure data for an air sample at room temperature are graphed in Figure 5.

This figure contains a diagram and two graphs. The diagram shows a syringe labeled with a scale in mL or cc with multiples of 5 labeled beginning at 5 and ending at 30. The markings halfway between these measurements are also provided. Attached at the top of the syringe is a pressure gauge with a scale marked by fives from 40 on the left to 5 on the right. The gauge needle rests between 10 and 15, slightly closer to 15. The syringe plunger position indicates a volume measurement about halfway between 10 and 15 mL or cc. The first graph is labeled "V (mL)" on the horizontal axis and "P (psi)" on the vertical axis. Points are labeled at 5, 10, 15, 20, and 25 mL with corresponding values of 39.0, 19.5, 13.0, 9.8, and 6.5 psi. The points are connected with a smooth curve that is declining at a decreasing rate of change. The second graph is labeled "V (mL)" on the horizontal axis and "1 divided by P (psi)" on the vertical axis. The horizontal axis is labeled at multiples of 5, beginning at zero and extending up to 35 mL. The vertical axis is labeled by multiples of 0.02, beginning at 0 and extending up to 0.18. Six points indicated by black dots on this graph are connected with a black line segment showing a positive linear trend.

When a gas occupies a smaller volume, it exerts a higher pressure; when it occupies a larger volume, it exerts a lower pressure (assuming the amount of gas and the temperature do not change). Since P and V are inversely proportional, a graph of $1/P$ vs. V is linear.

Unlike the P - T and V - T relationships, pressure and volume are not directly proportional to each other. Instead, P and V exhibit inverse proportionality: Increasing the pressure results in a decrease of the volume of the gas. Mathematically this can be written:

$$P \propto 1/V \text{ or } P = k \times \frac{1}{V} \text{ or } P \times V = k \text{ or } P_1 V_1 = P_2 V_2$$

with k being a constant. Graphically, this relationship is shown by the straight line that results when plotting the inverse of the pressure ($\frac{1}{P}$) versus the volume (V) or the inverse of volume ($\frac{1}{V}$) versus the pressure (P). Graphs with curved lines are difficult to read accurately at low or high values of the variables and they are more difficult to use in fitting theoretical equations and parameters to experimental data. For those reasons, scientists often try to find a way to "linearize" their data. If we plot P versus V , we obtain a hyperbola (see Figure 6).

This diagram shows two graphs. In a, a graph is shown with volume on the horizontal axis and pressure on the vertical axis. A curved line is shown on the graph showing a decreasing trend with a decreasing rate of change. In b, a graph is shown with volume on the horizontal axis and one divided by pressure on the vertical axis. A line segment, beginning at the origin of the graph, shows a positive, linear trend.

The relationship between pressure and volume is inversely proportional. (a) The graph of P vs. V is a hyperbola, whereas (b) the graph of $\frac{1}{P}$ vs. V is linear.

The relationship between the volume and pressure of a given amount of gas at constant temperature was first published by the English natural philosopher Robert Boyle over 300 years ago. It is summarized in the statement now known as Boyle's law: *The volume of a given amount of gas held at constant temperature is inversely proportional to the pressure under which it is measured.*

Volume of a Gas Sample

The sample of gas in Figure 5 has a volume of 15.0 mL at a pressure of 13.0 psi. Determine the pressure of the gas at a volume of 7.5 mL, using:

- the P - V graph in Figure 5
- the $\frac{1}{P}$ vs. V graph in Figure 5
- the Boyle's law equation

Comment on the likely accuracy of each method.

Solution

- Estimating from the P - V graph gives a value for P somewhere around 27 psi.
- Estimating from the $\frac{1}{P}$ vs. V graph give a value of about 26 psi.
- From Boyle's law, we know that the product of pressure and volume (PV) for a given sample of gas at a constant temperature is always equal to the same value. Therefore we have $P_1 V_1 = k$ and $P_2 V_2 = k$ which means that $P_1 V_1 = P_2 V_2$.

Using P_1 and V_1 as the known values 13.0 psi and 15.0 mL, P_2 as the pressure at which the volume is unknown, and V_2 as the unknown volume, we have:

$$P_1V_1 = P_2V_2 \text{ or } 13.0 \text{ psi} \times 15.0 \text{ mL} = P_2 \times 7.5 \text{ mL}$$

Solving:

$$P_2 = \frac{13.0 \text{ psi} \times 15.0 \text{ mL}}{7.5 \text{ mL}} = 26 \text{ psi}$$

It was more difficult to estimate well from the P - V graph, so (a) is likely more inaccurate than (b) or (c). The calculation will be as accurate as the equation and measurements allow.

Check Your Learning

The sample of gas in [Figure 5](#) has a volume of 30.0 mL at a pressure of 6.5 psi. Determine the volume of the gas at a pressure of 11.0 psi, using:

- (a) the P - V graph in [Figure 5](#)
- (b) the $\frac{1}{P}$ vs. V graph in [Figure 5](#)
- (c) the Boyle's law equation

Comment on the likely accuracy of each method.

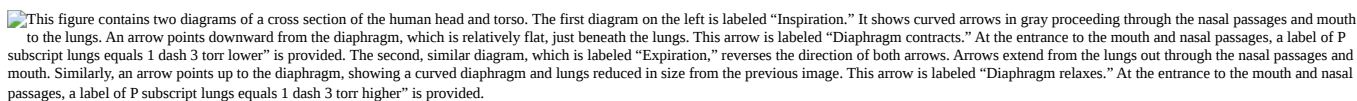
Answer:

- (a) about 17–18 mL;
- (b) ~18 mL;
- (c) 17.7 mL;

it was more difficult to estimate well from the P - V graph, so (a) is likely more inaccurate than (b); the calculation will be as accurate as the equation and measurements allow.

Breathing and Boyle's Law

What do you do about 20 times per minute for your whole life, without a break and often without even being aware of it? The answer, of course, is respiration or breathing. How does it work? It turns out that the gas laws apply here, too. Your lungs take in gas that your body needs (oxygen) and get rid of waste gas (carbon dioxide). Lungs are made of spongy, stretchy tissue that expands and contracts while you breathe. When you inhale, your diaphragm and intercostal muscles (the muscles between your ribs) contract, expanding your chest cavity and making your lung volume larger. The increase in volume leads to a decrease in pressure (Boyle's law). This causes air to flow into the lungs (from high pressure to low pressure). When you exhale, the process reverses: Your diaphragm and rib muscles relax, your chest cavity contracts and your lung volume decreases, causing the pressure to increase (Boyle's law again) and air flows out of the lungs (from high pressure to low pressure). You then breathe in and out again, and again, repeating this Boyle's law cycle for the rest of your life ([Figure 7](#)).

This figure contains two diagrams of a cross section of the human head and torso. The first diagram on the left is labeled "Inspiration." It shows curved arrows in gray proceeding through the nasal passages and mouth to the lungs. An arrow points downward from the diaphragm, which is relatively flat, just beneath the lungs. This arrow is labeled "Diaphragm contracts." At the entrance to the mouth and nasal passages, a label of $P_{\text{lungs}} = 1 \text{ dash } 3 \text{ torr lower}$ is provided. The second, similar diagram, which is labeled "Expiration," reverses the direction of both arrows. Arrows extend from the lungs out through the nasal passages and mouth. Similarly, an arrow points up to the diaphragm, showing a curved diaphragm and lungs reduced in size from the previous image. This arrow is labeled "Diaphragm relaxes." At the entrance to the mouth and nasal passages, a label of $P_{\text{lungs}} = 1 \text{ dash } 3 \text{ torr higher}$ is provided.

Breathing occurs because expanding and contracting lung volume creates small pressure differences between your lungs and your surroundings, causing air to be drawn into and forced out of your lungs.

Moles of Gas and Volume: Avogadro's Law

The Italian scientist Amedeo Avogadro advanced a hypothesis in 1811 to account for the behavior of gases, stating that equal volumes of all gases, measured under the same conditions of temperature and pressure, contain the same number of molecules. Over time, this relationship was supported by many experimental observations as expressed by Avogadro's law: *For a confined gas, the volume (V) and number of moles (n) are directly proportional, if the pressure and temperature both remain constant.*

In equation form, this is written as:

$$V \propto n \text{ or } V = k \times n \text{ or } \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Mathematical relationships can also be determined for the other variable pairs, such as P versus n , and n versus T .

8.3: The Gas Laws is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.4: Intermolecular forces

25

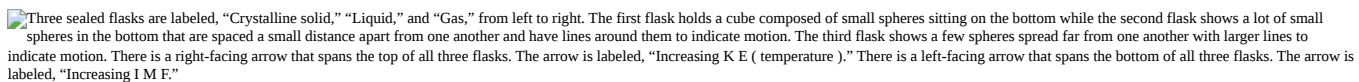
Harper College Chemistry Department

As was the case for gaseous substances, the kinetic molecular theory may be used to explain the behavior of solids and liquids. In the following description, the term *particle* will be used to refer to an atom, molecule, or ion. Note that we will use the popular phrase “intermolecular attraction” to refer to attractive forces between the particles of a substance, regardless of whether these particles are molecules, atoms, or ions. The phrase “intermolecular forces” is also used.

Consider these two aspects of the molecular-level environments in solid, liquid, and gaseous matter:

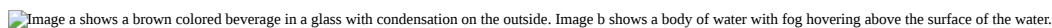
- Particles in a solid are tightly packed together and often arranged in a regular pattern; in a liquid, they are close together with no regular arrangement; in a gas, they are far apart with no regular arrangement.
- Particles in a solid vibrate about fixed positions and do not generally move in relation to one another; in a liquid, they move past each other but remain in essentially constant contact; in a gas, they move independently of one another except when they collide.

The differences in the properties of a solid, liquid, or gas reflect the strengths of the attractive forces between the atoms, molecules, or ions that make up each phase. The phase in which a substance exists depends on the relative extents of its intermolecular forces (IMFs) and the kinetic energies (KE) of its molecules. IMFs are the various forces of attraction that may exist between the atoms and molecules of a substance due to electrostatic phenomena, as will be detailed in this module. These forces serve to hold particles close together, whereas the particles’ KE provides the energy required to overcome the attractive forces and thus increase the distance between particles. [Figure 1](#) illustrates how changes in physical state may be induced by changing the temperature, hence, the average KE, of a given substance.

Three sealed flasks are labeled, “Crystalline solid,” “Liquid,” and “Gas,” from left to right. The first flask holds a cube composed of small spheres sitting on the bottom while the second flask shows a lot of small spheres in the bottom that are spaced a small distance apart from one another and have lines around them to indicate motion. The third flask shows a few spheres spread far from one another with larger lines to indicate motion. There is a right-facing arrow that spans the top of all three flasks. The arrow is labeled, “Increasing K E (temperature).” There is a left-facing arrow that spans the bottom of all three flasks. The arrow is labeled, “Increasing I M F.”

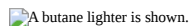
Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

As an example of the processes depicted in this figure, consider a sample of water. When gaseous water is cooled sufficiently, the attractions between H_2O molecules will be capable of holding them together when they come into contact with each other; the gas condenses, forming liquid H_2O . For example, liquid water forms on the outside of a cold glass as the water vapor in the air is cooled by the cold glass, as seen in [Figure 2](#).

Image a shows a brown colored beverage in a glass with condensation on the outside. Image b shows a body of water with fog hovering above the surface of the water.

Condensation forms when water vapor in the air is cooled enough to form liquid water, such as (a) on the outside of a cold beverage glass or (b) in the form of fog. (credit a: modification of work by Jenny Downing; credit b: modification of work by Cory Zanker)

We can also liquefy many gases by compressing them, if the temperature is not too high. The increased pressure brings the molecules of a gas closer together, such that the attractions between the molecules become strong relative to their KE. Consequently, they form liquids. Butane, C_4H_{10} , is the fuel used in disposable lighters and is a gas at standard temperature and pressure. Inside the lighter’s fuel compartment, the butane is compressed to a pressure that results in its condensation to the liquid state, as shown in [Figure 3](#).

A butane lighter is shown.

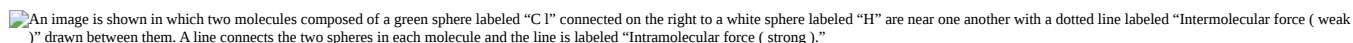
Gaseous butane is compressed within the storage compartment of a disposable lighter, resulting in its condensation to the liquid state. (credit: modification of work by “Sam-Cat”/Flickr)

Finally, if the temperature of a liquid becomes sufficiently low, or the pressure on the liquid becomes sufficiently high, the molecules of the liquid no longer have enough KE to overcome the IMF between them, and a solid forms. A more thorough discussion of these and other changes of state, or phase transitions, is provided in a later module of this chapter.

Access this [interactive simulation](#) on states of matter, phase transitions, and intermolecular forces. This simulation is useful for visualizing concepts introduced throughout this chapter.

Forces between Molecules

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not *intramolecular* forces. *Intramolecular* forces are those *within* the molecule that keep the molecule together, for example, the bonds between the atoms. *Intermolecular* forces are the attractions *between* molecules, which determine many of the physical properties of a substance. [Figure 4](#) illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

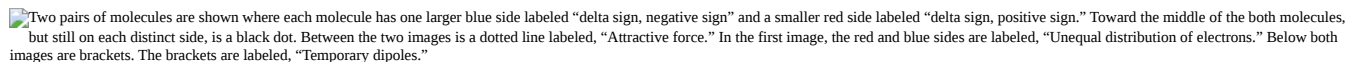


Intramolecular forces keep a molecule intact. *Intermolecular* forces hold multiple molecules together and determine many of a substance's properties.

All of the attractive forces between neutral atoms and molecules are known as van der Waals forces, although they are usually referred to more informally as intermolecular attraction. We will consider the various types of IMFs in the next three sections of this module.

Dispersion Forces

One of the three van der Waals forces is present in all condensed phases, regardless of the nature of the atoms or molecules composing the substance. This attractive force is called the London dispersion force in honor of German-born American physicist Fritz London who, in 1928, first explained it. This force is often referred to as simply the dispersion force. Because the electrons of an atom or molecule are in constant motion (or, alternatively, the electron's location is subject to quantum-mechanical variability), at any moment in time, an atom or molecule can develop a temporary, instantaneous dipole if its electrons are distributed asymmetrically. The presence of this dipole can, in turn, distort the electrons of a neighboring atom or molecule, producing an induced dipole. These two rapidly fluctuating, temporary dipoles thus result in a relatively weak electrostatic attraction between the species—a so-called dispersion force like that illustrated in [Figure 5](#).



Dispersion forces result from the formation of temporary dipoles, as illustrated here for two nonpolar diatomic molecules.

Dispersion forces that develop between atoms in different molecules can attract the two molecules to each other. The forces are relatively weak, however, and become significant only when the molecules are very close. Larger and heavier atoms and molecules exhibit stronger dispersion forces than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in [Table](#).

Melting and Boiling Points of the Halogens				
Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K
chlorine, Cl_2	71 g/mol	99 pm	172 K	238 K
bromine, Br_2	160 g/mol	114 pm	266 K	332 K
iodine, I_2	254 g/mol	133 pm	387 K	457 K
astatine, At_2	420 g/mol	150 pm	575 K	610 K

The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther from the nuclei than in a smaller atom. Thus, they are less tightly held and can more easily form the temporary dipoles that produce the attraction. The measure of how easy or difficult it is for another electrostatic

charge (for example, a nearby ion or polar molecule) to distort a molecule's charge distribution (its electron cloud) is known as polarizability. A molecule that has a charge cloud that is easily distorted is said to be very polarizable and will have large dispersion forces; one with a charge cloud that is difficult to distort is not very polarizable and will have small dispersion forces.

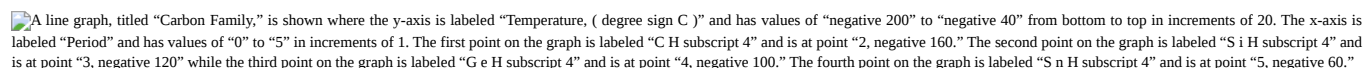
London Forces and Their Effects

Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4 , SiH_4 , GeH_4 , and SnH_4 . Explain your reasoning.

Solution

Applying the skills acquired in the chapter on chemical bonding and molecular geometry, all of these compounds are predicted to be nonpolar, so they may experience only dispersion forces: the smaller the molecule, the less polarizable and the weaker the dispersion forces; the larger the molecule, the larger the dispersion forces. The molar masses of CH_4 , SiH_4 , GeH_4 , and SnH_4 are approximately 16 g/mol, 32 g/mol, 77 g/mol, and 123 g/mol, respectively. Therefore, CH_4 is expected to have the lowest boiling point and SnH_4 the highest boiling point. The ordering from lowest to highest boiling point is expected to be $\text{CH}_4 < \text{SiH}_4 < \text{GeH}_4 < \text{SnH}_4$.

A graph of the actual boiling points of these compounds versus the period of the group 14 element shows this prediction to be correct:

A line graph, titled "Carbon Family," is shown where the y-axis is labeled "Temperature, (degree sign C)" and has values of "negative 200" to "negative 40" from bottom to top in increments of 20. The x-axis is labeled "Period" and has values of "0" to "5" in increments of 1. The first point on the graph is labeled "C H subscript 4" and is at point "2, negative 160." The second point on the graph is labeled "S i H subscript 4" and is at point "3, negative 120" while the third point on the graph is labeled "G e H subscript 4" and is at point "4, negative 100." The fourth point on the graph is labeled "S n H subscript 4" and is at point "5, negative 60."

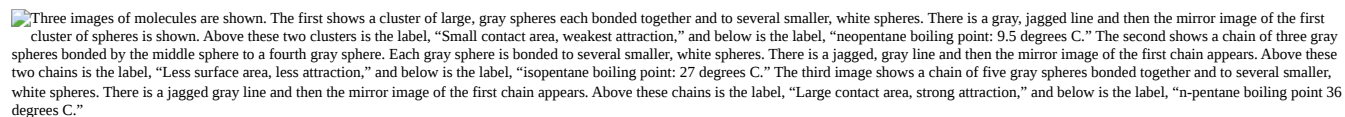
Check Your Learning

Order the following hydrocarbons from lowest to highest boiling point: C_2H_6 , C_3H_8 , and C_4H_{10} .

Answer:

$\text{C}_2\text{H}_6 < \text{C}_3\text{H}_8 < \text{C}_4\text{H}_{10}$. All of these compounds are nonpolar and only have London dispersion forces: the larger the molecule, the larger the dispersion forces and the higher the boiling point. The ordering from lowest to highest boiling point is therefore $\text{C}_2\text{H}_6 < \text{C}_3\text{H}_8 < \text{C}_4\text{H}_{10}$.

The shapes of molecules also affect the magnitudes of the dispersion forces between them. For example, boiling points for the isomers *n*-pentane, isopentane, and neopentane (shown in Figure 6) are 36 °C, 27 °C, and 9.5 °C, respectively. Even though these compounds are composed of molecules with the same chemical formula, C_5H_{12} , the difference in boiling points suggests that dispersion forces in the liquid phase are different, being greatest for *n*-pentane and least for neopentane. The elongated shape of *n*-pentane provides a greater surface area available for contact between molecules, resulting in correspondingly stronger dispersion forces. The more compact shape of isopentane offers a smaller surface area available for intermolecular contact and, therefore, weaker dispersion forces. Neopentane molecules are the most compact of the three, offering the least available surface area for intermolecular contact and, hence, the weakest dispersion forces. This behavior is analogous to the connections that may be formed between strips of VELCRO brand fasteners: the greater the area of the strip's contact, the stronger the connection.

Three images of molecules are shown. The first shows a cluster of large, gray spheres each bonded together and to several smaller, white spheres. There is a gray, jagged line and then the mirror image of the first cluster of spheres is shown. Above these two clusters is the label, "Small contact area, weakest attraction," and below is the label, "neopentane boiling point: 9.5 degrees C." The second shows a chain of three gray spheres bonded by the middle sphere to a fourth gray sphere. Each gray sphere is bonded to several smaller, white spheres. There is a jagged, gray line and then the mirror image of the first chain appears. Above these two chains is the label, "Less surface area, less attraction," and below is the label, "isopentane boiling point: 27 degrees C." The third image shows a chain of five gray spheres bonded together and to several smaller, white spheres. There is a jagged gray line and then the mirror image of the first chain appears. Above these chains is the label, "Large contact area, strong attraction," and below is the label, "n-pentane boiling point 36 degrees C."

The strength of the dispersion forces increases with the contact area between molecules, as demonstrated by the boiling points of these pentane isomers.

GECKOS AND INTERMOLECULAR FORCES

Geckos have an amazing ability to adhere to most surfaces. They can quickly run up smooth walls and across ceilings that have no toe-holds, and they do this without having suction cups or a sticky substance on their toes. And while a gecko can lift its feet easily as it walks along a surface, if you attempt to pick it up, it sticks to the surface. How are geckos (as well as spiders and some other insects) able to do this? Although this phenomenon has been investigated for hundreds of years, scientists only recently uncovered the details of the process that allows geckos' feet to behave this way.

Geckos' toes are covered with hundreds of thousands of tiny hairs known as *setae*, with each seta, in turn, branching into hundreds of tiny, flat, triangular tips called *spatulae*. The huge numbers of spatulae on its setae provide a gecko, shown in Figure 7, with a

large total surface area for sticking to a surface. In 2000, Kellar Autumn, who leads a multi-institutional gecko research team, found that geckos adhered equally well to both polar silicon dioxide and nonpolar gallium arsenide. This proved that geckos stick to surfaces because of dispersion forces—weak intermolecular attractions arising from temporary, synchronized charge distributions between adjacent molecules. Although dispersion forces are very weak, the total attraction over millions of spatulae is large enough to support many times the gecko's weight.

In 2014, two scientists developed a model to explain how geckos can rapidly transition from “sticky” to “non-sticky.” Alex Greaney and Congcong Hu at Oregon State University described how geckos can achieve this by changing the angle between their spatulae and the surface. Geckos' feet, which are normally nonsticky, become sticky when a small shear force is applied. By curling and uncurling their toes, geckos can alternate between sticking and unsticking from a surface, and thus easily move across it. Further investigations may eventually lead to the development of better adhesives and other applications.

 Three figures are shown. The first is a photo of the bottom of a gecko's foot. The second is a bigger version which shows the setae. The third is a bigger version of the setae and shows the spatulae.

Geckos' toes contain large numbers of tiny hairs (setae), which branch into many triangular tips (spatulae). Geckos adhere to surfaces because of van der Waals attractions between the surface and a gecko's millions of spatulae. By changing how the spatulae contact the surface, geckos can turn their stickiness “on” and “off.” (credit photo: modification of work by “JC*+A!”/Flickr)

Watch this [video](#) to learn more about Kellar Autumn's research that determined that van der Waals forces are responsible for a gecko's ability to cling and climb.

Dipole-Dipole Attractions

Recall from the chapter on chemical bonding and molecular geometry that *polar* molecules have a partial positive charge on one side and a partial negative charge on the other side of the molecule—a separation of charge called a *dipole*. Consider a polar molecule such as hydrogen chloride, HCl. In the HCl molecule, the more electronegative Cl atom bears the partial negative charge, whereas the less electronegative H atom bears the partial positive charge. An attractive force between HCl molecules results from the attraction between the positive end of one HCl molecule and the negative end of another. This attractive force is called a dipole-dipole attraction—the electrostatic force between the partially positive end of one polar molecule and the partially negative end of another, as illustrated in [Figure 8](#).

 Two pairs of molecules are shown where each molecule has one larger blue side labeled “delta sign, negative sign” and a smaller red side labeled “delta sign, positive sign. In the first pair, the red sides of the two molecules both face the left and the blue side to the right. A horizontal dotted line lies in between the two. In the second pair, the molecules face up and down, with the red and blue ends aligning. A horizontal dotted line lies between the red and blue ends facing upward and another lies between the red and blue ends facing downward.

This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

The effect of a dipole-dipole attraction is apparent when we compare the properties of HCl molecules to nonpolar F₂ molecules. Both HCl and F₂ consist of the same number of atoms and have approximately the same molecular mass. At a temperature of 150 K, molecules of both substances would have the same average KE. However, the dipole-dipole attractions between HCl molecules are sufficient to cause them to “stick together” to form a liquid, whereas the relatively weaker dispersion forces between nonpolar F₂ molecules are not, and so this substance is gaseous at this temperature. The higher normal boiling point of HCl (188 K) compared to F₂ (85 K) is a reflection of the greater strength of dipole-dipole attractions between HCl molecules, compared to the attractions between nonpolar F₂ molecules. We will often use values such as boiling or freezing points, or enthalpies of vaporization or fusion, as indicators of the relative strengths of IMFs of attraction present within different substances.

Dipole-Dipole Forces and Their Effects

Predict which will have the higher boiling point: N₂ or CO. Explain your reasoning.

Solution

CO and N₂ are both diatomic molecules with masses of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N₂ is nonpolar, its molecules cannot exhibit dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N₂ molecules, so CO is expected to have the higher boiling point.

Check Your Learning

Predict which will have the higher boiling point: ICl or Br₂. Explain your reasoning.

Answer:

ICl. ICl and Br₂ have similar masses (~160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br₂ is nonpolar and does not. The relatively stronger dipole-dipole attractions require more energy to overcome, so ICl will have the higher boiling point.

Hydrogen Bonding

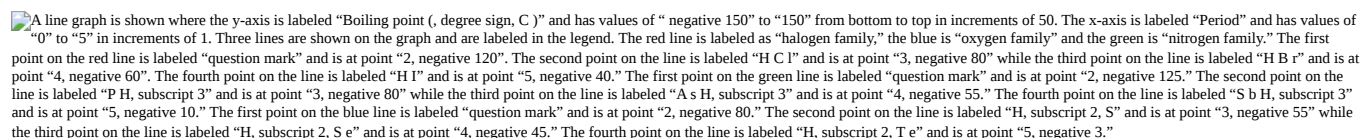
Nitrosyl fluoride (ONF, molecular mass 49 amu) is a gas at room temperature. Water (H₂O, molecular mass 18 amu) is a liquid, even though it has a lower molecular mass. We clearly cannot attribute this difference between the two compounds to dispersion forces. Both molecules have about the same shape and ONF is the heavier and larger molecule. It is, therefore, expected to experience more significant dispersion forces. Additionally, we cannot attribute this difference in boiling points to differences in the dipole moments of the molecules. Both molecules are polar and exhibit comparable dipole moments. The large difference between the boiling points is due to a particularly strong dipole-dipole attraction that may occur when a molecule contains a hydrogen atom bonded to a fluorine, oxygen, or nitrogen atom (the three most electronegative elements). The very large difference in electronegativity (EN) between the H atom (EN=2.1) and the atom to which it is bonded (EN=4.0 for an F atom, EN=3.5 for an O atom, or EN=3.0 for a N atom), combined with the very small size of a H atom and the relatively small sizes of F, O, or N atoms, leads to *highly concentrated partial charges* with these atoms. Molecules with F-H, O-H, or N-H moieties are very strongly attracted to similar moieties in nearby molecules, a particularly strong type of dipole-dipole attraction called hydrogen bonding. Examples of hydrogen bonds include HF···HF, H₂O···HOH, and H₃N···HNH₂, in which the hydrogen bonds are denoted by dots. [Figure 9](#) illustrates hydrogen bonding between water molecules.

 Five water molecules are shown near one another, but not touching. A dotted line lies between many of the hydrogen atoms on one molecule and the oxygen atom on another molecule.

Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

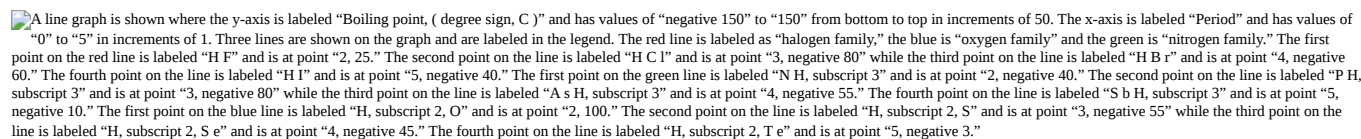
Despite use of the word “bond,” keep in mind that hydrogen bonds are *intermolecular* attractive forces, not *intramolecular* attractive forces (covalent bonds). Hydrogen bonds are much weaker than covalent bonds, only about 5 to 10% as strong, but are generally much stronger than other dipole-dipole attractions and dispersion forces.

Hydrogen bonds have a pronounced effect on the properties of condensed phases (liquids and solids). For example, consider the trends in boiling points for the binary hydrides of group 15 (NH₃, PH₃, AsH₃, and SbH₃), group 16 hydrides (H₂O, H₂S, H₂Se, and H₂Te), and group 17 hydrides (HF, HCl, HBr, and HI). The boiling points of the heaviest three hydrides for each group are plotted in [Figure 10](#). As we progress down any of these groups, the polarities of the molecules decrease slightly, whereas the sizes of the molecules increase substantially. The effect of increasingly stronger dispersion forces dominates that of increasingly weaker dipole-dipole attractions, and the boiling points are observed to increase steadily.

 A line graph is shown where the y-axis is labeled “Boiling point, (degree sign, C)” and has values of “negative 150” to “150” from bottom to top in increments of 50. The x-axis is labeled “Period” and has values of “0” to “5” in increments of 1. Three lines are shown on the graph and are labeled in the legend. The red line is labeled as “halogen family,” the blue is “oxygen family” and the green is “nitrogen family.” The first point on the red line is labeled “question mark” and is at point “2, negative 120.” The second point on the line is labeled “H Cl” and is at point “3, negative 80” while the third point on the line is labeled “H B r” and is at point “4, negative 60.” The fourth point on the line is labeled “H I” and is at point “5, negative 40.” The first point on the green line is labeled “question mark” and is at point “2, negative 125.” The second point on the line is labeled “P H, subscript 3” and is at point “3, negative 80” while the third point on the line is labeled “A s H, subscript 3” and is at point “4, negative 55.” The fourth point on the line is labeled “S b H, subscript 3” and is at point “5, negative 10.” The first point on the blue line is labeled “question mark” and is at point “2, negative 80.” The second point on the line is labeled “H, subscript 2, S” and is at point “3, negative 55” while the third point on the line is labeled “H, subscript 2, S e” and is at point “4, negative 45.” The fourth point on the line is labeled “H, subscript 2, T e” and is at point “5, negative 3.”

For the group 15, 16, and 17 hydrides, the boiling points for each class of compounds increase with increasing molecular mass for elements in periods 3, 4, and 5.

If we use this trend to predict the boiling points for the lightest hydride for each group, we would expect NH₃ to boil at about -120 °C, H₂O to boil at about -80 °C, and HF to boil at about -110 °C. However, when we measure the boiling points for these compounds, we find that they are dramatically higher than the trends would predict, as shown in [Figure 11](#). The stark contrast between our naïve predictions and reality provides compelling evidence for the strength of hydrogen bonding.

 A line graph is shown where the y-axis is labeled “Boiling point, (degree sign, C)” and has values of “negative 150” to “150” from bottom to top in increments of 50. The x-axis is labeled “Period” and has values of “0” to “5” in increments of 1. Three lines are shown on the graph and are labeled in the legend. The red line is labeled as “halogen family,” the blue is “oxygen family” and the green is “nitrogen family.” The first point on the red line is labeled “H F” and is at point “2, 25.” The second point on the line is labeled “H Cl” and is at point “3, negative 80” while the third point on the line is labeled “H B r” and is at point “4, negative 60.” The fourth point on the line is labeled “H I” and is at point “5, negative 40.” The first point on the green line is labeled “N H, subscript 3” and is at point “2, negative 40.” The second point on the line is labeled “P H, subscript 3” and is at point “3, negative 80” while the third point on the line is labeled “A s H, subscript 3” and is at point “4, negative 55.” The fourth point on the line is labeled “S b H, subscript 3” and is at point “5, negative 10.” The first point on the blue line is labeled “H, subscript 2, O” and is at point “2, 100.” The second point on the line is labeled “H, subscript 2, S” and is at point “3, negative 55” while the third point on the line is labeled “H, subscript 2, S e” and is at point “4, negative 45.” The fourth point on the line is labeled “H, subscript 2, T e” and is at point “5, negative 3.”

In comparison to periods 3–5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding.

Effect of Hydrogen Bonding on Boiling Points

Consider the compounds dimethylether (CH₃OCH₃), ethanol (CH₃CH₂OH), and propane (CH₃CH₂CH₃). Their boiling points, not necessarily in order, are -42.1 °C, -24.8 °C, and 78.4 °C. Match each compound with its boiling point. Explain your reasoning.

Solution

The VSEPR-predicted shapes of CH_3OCH_3 , $\text{CH}_3\text{CH}_2\text{OH}$, and $\text{CH}_3\text{CH}_2\text{CH}_3$ are similar, as are their molar masses (46 g/mol, 46 g/mol, and 44 g/mol, respectively), so they will exhibit similar dispersion forces. Since $\text{CH}_3\text{CH}_2\text{CH}_3$ is nonpolar, it may exhibit *only* dispersion forces. Because CH_3OCH_3 is polar, it will also experience dipole-dipole attractions. Finally, $\text{CH}_3\text{CH}_2\text{OH}$ has an $-\text{OH}$ group, and so it will experience the uniquely strong dipole-dipole attraction known as hydrogen bonding. So the ordering in terms of strength of IMFs, and thus boiling points, is $\text{CH}_3\text{CH}_2\text{CH}_3 < \text{CH}_3\text{OCH}_3 < \text{CH}_3\text{CH}_2\text{OH}$. The boiling point of propane is -42.1°C , the boiling point of dimethylether is -24.8°C , and the boiling point of ethanol is 78.5°C .

Check Your Learning

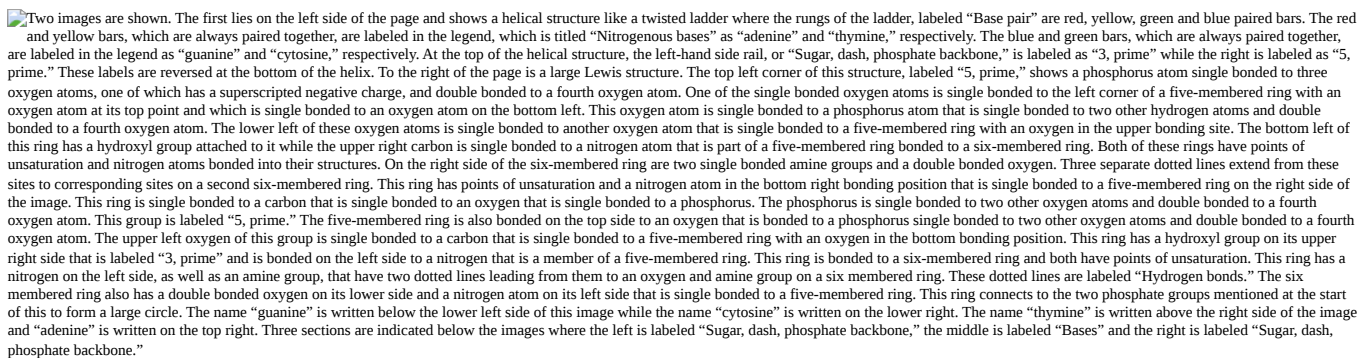
Ethane (CH_3CH_3) has a melting point of -183°C and a boiling point of -89°C . Predict the melting and boiling points for methylamine (CH_3NH_2). Explain your reasoning.

Answer:

The melting point and boiling point for methylamine are predicted to be significantly greater than those of ethane. CH_3CH_3 and CH_3NH_2 are similar in size and mass, but methylamine possesses an $-\text{NH}$ group and therefore may exhibit hydrogen bonding. This greatly increases its IMFs, and therefore its melting and boiling points. It is difficult to predict values, but the known values are a melting point of -93°C and a boiling point of -6°C .

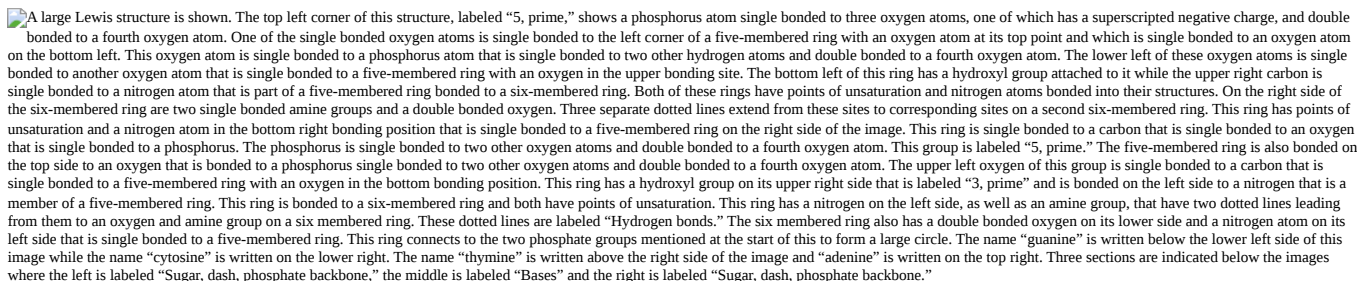
HYDROGEN BONDING AND DNA

Deoxyribonucleic acid (DNA) is found in every living organism and contains the genetic information that determines the organism's characteristics, provides the blueprint for making the proteins necessary for life, and serves as a template to pass this information on to the organism's offspring. A DNA molecule consists of two (anti-)parallel chains of repeating nucleotides, which form its well-known double helical structure, as shown in [Figure 12](#).



Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (credit: modification of work by Jerome Walker, Dennis Myts)

Each nucleotide contains a (deoxyribose) sugar bound to a phosphate group on one side, and one of four nitrogenous bases on the other. Two of the bases, cytosine (C) and thymine (T), are single-ringed structures known as pyrimidines. The other two, adenine (A) and guanine (G), are double-ringed structures called purines. These bases form complementary base pairs consisting of one purine and one pyrimidine, with adenine pairing with thymine, and cytosine with guanine. Each base pair is held together by hydrogen bonding. A and T share two hydrogen bonds, C and G share three, and both pairings have a similar shape and structure [Figure 13](#).



The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called "complementary base pairs."

The cumulative effect of millions of hydrogen bonds effectively holds the two strands of DNA together. Importantly, the two strands of DNA can relatively easily “unzip” down the middle since hydrogen bonds are relatively weak compared to the covalent bonds that hold the atoms of the individual DNA molecules together. This allows both strands to function as a template for replication.

Key Concepts and Summary

The physical properties of condensed matter (liquids and solids) can be explained in terms of the kinetic molecular theory. In a liquid, intermolecular attractive forces hold the molecules in contact, although they still have sufficient KE to move past each other.

Intermolecular attractive forces, collectively referred to as van der Waals forces, are responsible for the behavior of liquids and solids and are electrostatic in nature. Dipole-dipole attractions result from the electrostatic attraction of the partial negative end of one dipolar molecule for the partial positive end of another. The temporary dipole that results from the motion of the electrons in an atom can induce a dipole in an adjacent atom and give rise to the London dispersion force. London forces increase with increasing molecular size. Hydrogen bonds are a special type of dipole-dipole attraction that results when hydrogen is bonded to one of the three most electronegative elements: F, O, or N.

Chemistry End of Chapter Exercises

1. a) In terms of their bulk properties, how do liquids and solids differ? b) How are they similar?
2. a) In terms of the kinetic molecular theory, in what ways are liquids similar to solids? b) In what ways are liquids different from solids?
3. a) In terms of the kinetic molecular theory, in what ways are liquids similar to gases? b) In what ways are liquids different from gases?
4. Explain why liquids assume the shape of any container into which they are poured, whereas solids are rigid and retain their shape.
5. What is the evidence that all neutral atoms and molecules exert attractive forces on each other?
6. Open the [PhET States of Matter Simulation](#) to answer the following questions:
 - (a) Select the Solid, Liquid, Gas tab. Explore by selecting different substances, heating and cooling the systems, and changing the state. What similarities do you notice between the four substances for each phase (solid, liquid, gas)? What differences do you notice?
 - (b) For each substance, select each of the states and record the given temperatures. How do the given temperatures for each state correlate with the strengths of their intermolecular attractions? Explain.
 - (c) Select the Interaction Potential tab, and use the default neon atoms. Move the Ne atom on the right and observe how the potential energy changes. Select the Total Force button, and move the Ne atom as before. When is the total force on each atom attractive and large enough to matter? Then select the Component Forces button, and move the Ne atom. When do the attractive (van der Waals) and repulsive (electron overlap) forces balance? How does this relate to the potential energy versus the distance between atoms graph? Explain.
7. Define the following and give an example of each:
 - (a) dispersion force
 - (b) dipole-dipole attraction
 - (c) hydrogen bond
8. The types of intermolecular forces in a substance are identical whether it is a solid, a liquid, or a gas. Why then does a substance change phase from a gas to a liquid or to a solid?
9. Why do the boiling points of the noble gases increase in the order $\text{He} < \text{Ne} < \text{Ar} < \text{Kr} < \text{Xe}$?
10. Neon and HF have approximately the same molecular masses.

- (a) Explain why the boiling points of Neon and HF differ.
- (b) Compare the change in the boiling points of Ne, Ar, Kr, and Xe with the change of the boiling points of HF, HCl, HBr, and HI, and explain the difference between the changes with increasing atomic or molecular mass.
11. Arrange each of the following sets of compounds in order of increasing boiling point temperature:
- (a) HCl, H₂O, SiH₄
- (b) F₂, Cl₂, Br₂
- (c) CH₄, C₂H₆, C₃H₈
- (d) O₂, NO, N₂
12. The molecular mass of butanol, C₄H₉OH, is 74.14; that of ethylene glycol, CH₂(OH)CH₂OH, is 62.08, yet their boiling points are 117.2 °C and 174 °C, respectively. Explain the reason for the difference.
13. On the basis of intermolecular attractions, explain the differences in the boiling points of *n*-butane (−1 °C) and chloroethane (12 °C), which have similar molar masses.
14. On the basis of dipole moments and/or hydrogen bonding, explain in a qualitative way the differences in the boiling points of acetone (56.2 °C) and 1-propanol (97.4 °C), which have similar molar masses.
15. The melting point of H₂O(s) is 0 °C. Would you expect the melting point of H₂S(s) to be −85 °C, 0 °C, or 185 °C? Explain your answer.
16. Silane (SiH₄), phosphine (PH₃), and hydrogen sulfide (H₂S) melt at −185 °C, −133 °C, and −85 °C, respectively. What does this suggest about the polar character and intermolecular attractions of the three compounds?
17. Explain why a hydrogen bond between two water molecules is weaker than a hydrogen bond between two hydrogen fluoride molecules.
18. Under certain conditions, molecules of acetic acid, CH₃COOH, form “dimers,” pairs of acetic acid molecules held together by strong intermolecular attractions:

 A Lewis structure shows a carbon atom single bonded to three hydrogen atoms and one other carbon atom, that is in turn double bonded to an oxygen atom and single bonded to another oxygen atom that is single bonded to a hydrogen atom. Dotted lines connect the terminal oxygen and hydrogen atoms to a reciprocal Lewis structure to the right, rotated 180 degrees. Each dotted line is labeled “hydrogen bond.”

Draw a dimer of acetic acid, showing how two CH₃COOH molecules are held together, and stating the type of IMF that is responsible.

19. Proteins are chains of amino acids that can form in a variety of arrangements, one of which is a helix. What kind of IMF is responsible for holding the protein strand in this shape? On the protein image, show the locations of the IMFs that hold the protein together:

 Two turns of a helical structure are shown horizontally. Three Lewis structures are superimposed on the helix. The first shows horizontally stacked dashes next to an oxygen atom, with three dots connecting to a hydrogen atom, and a single dash connecting the hydrogen atom to a nitrogen atom. The second shows a carbon atom double bonded to an oxygen atom, then three dots connecting to a hydrogen atom which is bonded to a nitrogen atom. The third shows a carbon atom double bonded to an oxygen atom with three dots extending to the right of the oxygen atom.

20. The density of liquid NH₃ is 0.64 g/mL; the density of gaseous NH₃ at STP is 0.0007 g/mL. Explain the difference between the densities of these two phases.
21. Identify the intermolecular forces present in the following solids:
- (a) CH₃CH₂OH
- (b) CH₃CH₂CH₃
- (c) CH₃CH₂Cl

Glossary

dipole-dipole attraction

intermolecular attraction between two permanent dipoles

dispersion force

(also, London dispersion force) attraction between two rapidly fluctuating, temporary dipoles; significant only when particles are very close together

hydrogen bonding

occurs when exceptionally strong dipoles attract; bonding that exists when hydrogen is bonded to one of the three most electronegative elements: F, O, or N

induced dipole

temporary dipole formed when the electrons of an atom or molecule are distorted by the instantaneous dipole of a neighboring atom or molecule

instantaneous dipole

temporary dipole that occurs for a brief moment in time when the electrons of an atom or molecule are distributed asymmetrically

intermolecular force

noncovalent attractive force between atoms, molecules, and/or ions

polarizability

measure of the ability of a charge to distort a molecule's charge distribution (electron cloud)

van der Waals force

attractive or repulsive force between molecules, including dipole-dipole, dipole-induced dipole, and London dispersion forces; does not include forces due to covalent or ionic bonding, or the attraction between ions and molecules

8.4: Intermolecular forces is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.5: Types of solutions

26

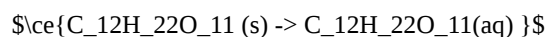
Harper College Chemistry Department

Page by: OpenStax

By the end of this section, you will be able to:

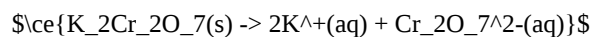
- Describe the basic properties of solutions and how they form
- Predict whether a given mixture will yield a solution based on molecular properties of its components
- Explain why some solutions either produce or absorb heat when they form

An earlier chapter of this text introduced *solutions*, defined as homogeneous mixtures of two or more substances. Often, one component of a solution is present at a significantly greater concentration, in which case it is called the *solvent*. The other components of the solution present in relatively lesser concentrations are called *solutes*. Sugar is a covalent solid composed of sucrose molecules, $C_{12}H_{22}O_{11}$. When this compound dissolves in water, its molecules become uniformly distributed among the molecules of water:



The term “aq” in parentheses signifies that the sucrose molecules are solutes and are therefore *individually dispersed* throughout the *aqueous solution* (water is the solvent). Although sucrose molecules are heavier than water molecules, they remain dispersed throughout the solution; gravity does not cause them to “settle out” over time.

Potassium dichromate, $K_2Cr_2O_7$, is an ionic compound composed of colorless potassium ions, K^+ , and orange dichromate ions, $Cr_2O_7^{2-}$. When a small amount of solid potassium dichromate is added to water, the compound dissolves and dissociates to yield potassium ions and dichromate ions uniformly distributed throughout the mixture (Figure 1), as indicated in this equation:



As with the mixture of sugar and water, this mixture is also an aqueous solution. Its solutes, potassium and dichromate ions, remain individually dispersed among the solvent (water) molecules.

The first photo shows a small mound of an orange crystalline solid. There is a right-facing arrow. The second photo shows a translucent, colorless liquid in a clear container. There is a right-facing arrow. The third photo shows a translucent orange liquid in a clear, covered container.

When potassium dichromate ($K_2Cr_2O_7$) is mixed with water, it forms a homogeneous orange solution. (credit: modification of work by Mark Ott)

Visit this [virtual lab](#) to view simulations of the dissolution of common covalent and ionic substances (sugar and salt) in water.

Water is used so often as a solvent that the word solution has come to imply an aqueous solution to many people. However, almost any gas, liquid, or solid can act as a solvent. Many alloys are solid solutions of one metal dissolved in another; for example, US five-cent coins contain nickel (25%) dissolved in copper (75%). Air is a gaseous solution, a homogeneous mixture of nitrogen, oxygen, and several other gases. Oxygen (a gas), alcohol (a liquid), and sugar (a solid) all dissolve in water (a liquid) to form liquid solutions. Table 1 gives examples of several different solutions and the phases of the solutes and solvents.

Solution	Solute	Solvent
air	$O_2(g)$	$N_2(g)$
soft drinks ¹	$CO_2(g)$	$H_2O(l)$
hydrogen in palladium	$H_2(g)$	$Pd(s)$
rubbing alcohol	$H_2O(l)$	$C_3H_8O(l)$ (2-propanol)
saltwater	$NaCl(s)$	$H_2O(l)$
brass	$Zn(s)$	$Cu(s)$

Solutions exhibit these defining traits:

- They are homogeneous; that is, after a solution is mixed, it has the same composition at all points throughout (its composition is uniform).
- The physical state of a solution—solid, liquid, or gas—is typically the same as that of the solvent, as demonstrated by the examples in [Table 1](#).
- The components of a solution are dispersed on a molecular scale; that is, they consist of a mixture of separated molecules, atoms, and/or ions.
- The dissolved solute in a solution will not settle out or separate from the solvent.
- The composition of a solution, or the concentrations of its components, can be varied continuously, within limits.

The Formation of Solutions

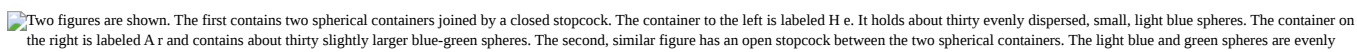
The formation of a solution is an example of a spontaneous process, a process that occurs under specified conditions without the requirement of energy from some external source. Sometimes we stir a mixture to speed up the dissolution process, but this is not necessary; a homogeneous solution would form if we waited long enough. The topic of spontaneity is critically important to the study of chemical thermodynamics, but will not be discussed here. For purposes of this chapter's discussion, it will suffice to consider two criteria that *favor*, but do not guarantee, the spontaneous formation of a solution:

1. a decrease in the internal energy of the system (an exothermic change)
2. an increase in the disorder in the system (which indicates an increase in the *entropy* of the system)

In the process of dissolution, an internal energy change often, but not always, occurs as heat is absorbed or evolved. An increase in disorder always results when a solution forms.

When the strengths of the intermolecular forces of attraction between solute and solvent species in a solution are the same or greater than those present in the separated components, the solution is formed with no accompanying energy change. Such a solution is called an ideal solution. A mixture of ideal gases (or gases such as helium and argon, which closely approach ideal behavior) is an example of an ideal solution, since the entities comprising these gases experience no significant intermolecular attractions.

When containers of helium and argon are connected, the gases spontaneously mix due to diffusion and form a solution ([Figure 2](#)). The formation of this solution clearly involves an increase in disorder, since the helium and argon atoms occupy a volume twice as large as that which each occupied before mixing.

Two figures are shown. The first contains two spherical containers joined by a closed stopcock. The container to the left is labeled H e. It holds about thirty evenly dispersed, small, light blue spheres. The container on the right is labeled A r and contains about thirty slightly larger blue-green spheres. The second, similar figure has an open stopcock between the two spherical containers. The light blue and green spheres are evenly dispersed and present in both containers.

Samples of helium and argon spontaneously mix to give a solution in which the disorder of the atoms of the two gases is increased.

Ideal solutions may also form when structurally similar liquids are mixed. For example, mixtures of the alcohols methanol (CH_3OH) and ethanol ($\text{C}_2\text{H}_5\text{OH}$) form ideal solutions, as do mixtures of the hydrocarbons pentane, C_5H_{12} , and hexane, C_6H_{14} . Placing methanol and ethanol, or pentane and hexane, in the bulbs will result in the same diffusion and subsequent mixing of these liquids as is observed for the He and Ar gases (although at a much slower rate), yielding solutions with no significant change in energy. Unlike a mixture of gases, however, the components of these liquid-liquid solutions do, indeed, experience intermolecular attractive forces. But since the molecules of the two substances being mixed are structurally very similar, the intermolecular attractive forces between like and unlike molecules are essentially the same, and the dissolution process, therefore, does not entail any appreciable increase or decrease in energy. These examples illustrate how diffusion alone can provide the driving force required to cause the spontaneous formation of a solution. In some cases, however, the relative magnitudes of intermolecular forces of attraction between solute and solvent species may prevent dissolution.

Three types of intermolecular attractive forces are relevant to the dissolution process: solute-solute, solvent-solvent, and solute-solvent. As illustrated in [Figure 3](#), the formation of a solution may be viewed as a stepwise process in which energy is consumed to overcome solute-solute and solvent-solvent attractions (endothermic processes) and released when solute-solvent attractions are established (an exothermic process referred to as solvation). The relative magnitudes of the energy changes associated with these stepwise processes determine whether the dissolution process overall will release or absorb energy. In some cases, solutions do not form because the energy required to separate solute and solvent species is so much greater than the energy released by solvation.

The top, central region of the figure shows solute particles as seven blue spheres and solvent particles as 16 red spheres in separate, labeled boxes. The particles in these boxes are touching. An arrow labeled "Step 1" points left of the solute box, and shows the blue spheres no longer touching in another box labeled "expanded solute." An arrow labeled "Step 2" points right from the solvent box and shows the red spheres no longer touching in another box labeled "expanded solvent." Arrows proceed from the bottom of the expanded solute and expanded solvent boxes and join at the bottom of the figure where a step 3 label is shown. The joined arrows point to a box just above in which the red and blue spheres are mixed together and touching. The solute and solvent boxes are joined by another arrow labeled "direct formation of solution" which points downward at the center of the figure. This arrow also points to the box containing mixed red and blue spheres near the bottom of the figure.

This schematic representation of dissolution shows a stepwise process involving the endothermic separation of solute and solvent species (Steps 1 and 2) and exothermic solvation (Step 3).

For example, cooking oils and water will not mix to any appreciable extent to yield solutions (Figure 4). Hydrogen bonding is the dominant intermolecular attractive force present in liquid water; the nonpolar hydrocarbon molecules of cooking oils are not capable of hydrogen bonding, instead being held together by dispersion forces. Forming an oil-water solution would require overcoming the very strong hydrogen bonding in water, as well as the significantly strong dispersion forces between the relatively large oil molecules. And, since the polar water molecules and nonpolar oil molecules would not experience very strong intermolecular attraction, very little energy would be released by solvation.

 This photo shows an array of colors in an oil water mixture.

A mixture of nonpolar cooking oil and polar water does not yield a solution. (credit: Gautam Dogra)

On the other hand, a mixture of ethanol and water will mix in any proportions to yield a solution. In this case, both substances are capable of hydrogen bonding, and so the solvation process is sufficiently exothermic to compensate for the endothermic separations of solute and solvent molecules.

As noted at the beginning of this chapter, spontaneous solution formation is favored, but not guaranteed, by exothermic dissolution processes. While many soluble compounds do, indeed, dissolve with the release of heat, some dissolve endothermically. Ammonium nitrate (NH_4NO_3) is one such example and is used to make instant cold packs for treating injuries like the one pictured in Figure 5. A thin-walled plastic bag of water is sealed inside a larger bag with solid NH_4NO_3 . When the smaller bag is broken, a solution of NH_4NO_3 forms, absorbing heat from the surroundings (the injured area to which the pack is applied) and providing a cold compress that decreases swelling. Endothermic dissolutions such as this one require a greater energy input to separate the solute species than is recovered when the solutes are solvated, but they are spontaneous nonetheless due to the increase in disorder that accompanies formation of the solution.

 This figure shows a single use instant cold pack with labels indicating an inner bag of water surrounded by white particulate ammonium nitrate.

An instant cold pack gets cold when certain salts, such as ammonium nitrate, dissolve in water—an endothermic process.

Watch this brief [video](#) illustrating endothermic and exothermic dissolution processes.

Key Concepts and Summary

A solution forms when two or more substances combine physically to yield a mixture that is homogeneous at the molecular level. The solvent is the most concentrated component and determines the physical state of the solution. The solutes are the other components typically present at concentrations less than that of the solvent. Solutions may form endothermically or exothermically, depending upon the relative magnitudes of solute and solvent intermolecular attractive forces. Ideal solutions form with no appreciable change in energy.

Chemistry End of Chapter Exercises

- How do solutions differ from compounds? From other mixtures?
- Which of the principal characteristics of solutions can we see in the solutions of $\text{K}_2\text{Cr}_2\text{O}_7$ shown in Figure 1?
- When KNO_3 is dissolved in water, the resulting solution is significantly colder than the water was originally.
 - Is the dissolution of KNO_3 an endothermic or an exothermic process?
 - What conclusions can you draw about the intermolecular attractions involved in the process?
 - Is the resulting solution an ideal solution?
- Give an example of each of the following types of solutions:
 - a gas in a liquid
 - a gas in a gas
 - a solid in a solid

5. Indicate the most important types of intermolecular attractions in each of the following solutions:

(a) The solution in [Figure 1](#).

(b) $\text{NO}(l)$ in $\text{CO}(l)$

(c) $\text{Cl}_2(g)$ in $\text{Br}_2(l)$

(d) $\text{HCl}(g)$ in benzene $\text{C}_6\text{H}_6(l)$

(e) Methanol $\text{CH}_3\text{OH}(l)$ in $\text{H}_2\text{O}(l)$

6. Predict whether each of the following substances would be more soluble in water (polar solvent) or in a hydrocarbon such as heptane (C_7H_{16} , nonpolar solvent):

(a) vegetable oil (nonpolar)

(b) isopropyl alcohol (polar)

(c) potassium bromide (ionic)

7. Heat is released when some solutions form; heat is absorbed when other solutions form. Provide a molecular explanation for the difference between these two types of spontaneous processes.

8. Solutions of hydrogen in palladium may be formed by exposing Pd metal to H_2 gas. The concentration of hydrogen in the palladium depends on the pressure of H_2 gas applied, but in a more complex fashion than can be described by Henry's law. Under certain conditions, 0.94 g of hydrogen gas is dissolved in 215 g of palladium metal (solution density = 10.8 g cm^3).

(a) Determine the molarity of this solution.

(b) Determine the molality of this solution.

(c) Determine the percent by mass of hydrogen atoms in this solution.

Footnotes

- 1 If bubbles of gas are observed within the liquid, the mixture is not homogeneous and, thus, not a solution.

Glossary

alloy

solid mixture of a metallic element and one or more additional elements

ideal solution

solution that forms with no accompanying energy change

solvation

exothermic process in which intermolecular attractive forces between the solute and solvent in a solution are established

spontaneous process

physical or chemical change that occurs without the addition of energy from an external source

8.5: Types of solutions is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.6: Electrolytes

27

Harper College Chemistry Department

Page by: OpenStax

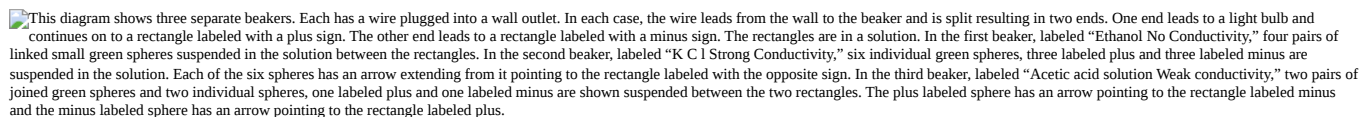
▼ Summary

By the end of this module, you will be able to:

- Define and give examples of electrolytes
- Distinguish between the physical and chemical changes that accompany dissolution of ionic and covalent electrolytes
- Relate electrolyte strength to solute-solvent attractive forces

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**.

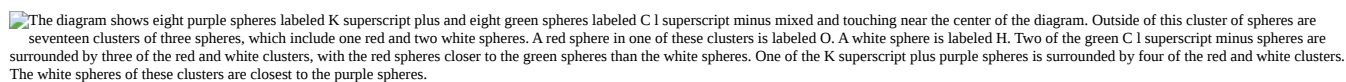
Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, 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 1**).

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 "KCl 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.

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.

Ionic Electrolytes

Water and other polar molecules are attracted to ions, as shown in **Figure 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.

The diagram shows eight purple spheres labeled K superscript plus and eight green spheres labeled Cl 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 Cl 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.

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.

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.

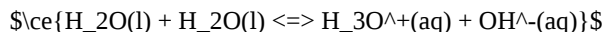
Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the negative (oxygen) ends to the positive potassium ions. 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 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).

Covalent Electrolytes

Pure water is an extremely poor conductor of electricity because it is only very slightly ionized—only about two out of every 1 billion molecules ionize at 25 °C. Water ionizes when one molecule of water gives up a proton (H^+ ion) to another molecule of water, yielding hydronium and hydroxide ions.



In some cases, we find that solutions prepared from covalent compounds conduct electricity because the solute molecules react chemically with the solvent to produce ions. For example, pure hydrochloric acid is a gas consisting of covalent HCl molecules. This gas contains no ions. However, when we dissolve hydrochloric acid in water, we find that the solution is a very good conductor. The water molecules play an essential part in forming ions: Solutions of hydrochloric acid in many other solvents, such as benzene, do not conduct electricity and do not contain ions.

Hydrochloric acid is an *acid*, and so its molecules react with water, transferring H^+ ions to form hydronium ions (H_3O^+) and chloride ions (Cl^-):

A chemical equation is shown. To the left, two hydrogen atoms are linked, each with a single dash to a central oxygen atom to the left and below the oxygen symbol, which has two pairs of dots, above and to the right of the atom. A plus sign is shown to the right, then a hydrogen atom linked to the left side of chlorine atom by a single dash with three pairs of dots, above, to the right, and below the element symbol. An arrow points to the products which are three hydrogen atoms linked by single dashes to a central oxygen atom shown in brackets with superscript plus. The oxygen atom has a single pair of dots above the element symbol. This is followed by a plus and Cl superscript minus. This symbol is surrounded by four pairs of dots, above and below and to the left and right of the element symbol.

This reaction is essentially 100% complete for HCl (i.e., it is a *strong acid* and, consequently, a strong electrolyte). Likewise, weak acids and bases that only react partially generate relatively low concentrations of ions when dissolved in water and are classified as weak electrolytes.

Key Concepts and Summary

Substances that dissolve in water to yield ions are called electrolytes. Electrolytes may be covalent compounds that chemically react with water to produce ions (for example, acids and bases), or they may be ionic compounds that dissociate to yield their constituent cations and anions, when dissolved. Dissolution of an ionic compound is facilitated by ion-dipole attractions between the ions of the compound and the polar water molecules. Soluble ionic substances and strong acids ionize completely and are strong electrolytes, while weak acids and bases ionize to only a small extent and are weak electrolytes. Nonelectrolytes are substances that do not produce ions when dissolved in water.

Chemistry End of Chapter Exercises

1. Explain why the ions Na^+ and Cl^- are strongly solvated in water but not in hexane, a solvent composed of nonpolar molecules.
2. Explain why solutions of HBr in benzene (a nonpolar solvent) are nonconductive, while solutions in water (a polar solvent) are conductive.
3. Consider the solutions presented:

(a) Which of the following sketches best represents the ions in a solution of $\text{Fe}(\text{NO}_3)_3(\text{aq})$?

In this figure, three beakers labeled x, y, and z are shown containing various arrangements of blue and red spheres suspended in solution. In beaker x, three small red spheres surround a single central blue sphere in small clusters which in turn are grouped in threes around a single red sphere, forming four larger clusters. In beaker y, the four large clusters are present without the central red spheres. Four individual red spheres are now present. In beaker z, the large clusters are not present. Twelve of the small clusters of three red and one blue sphere are present along with four single red spheres.

(b) Write a balanced chemical equation showing the products of the dissolution of $\text{Fe}(\text{NO}_3)_3$.

4. Compare the processes that occur when methanol (CH_3OH), hydrogen chloride (HCl), and sodium hydroxide (NaOH) dissolve in water. Write equations and prepare sketches showing the form in which each of these compounds is present in its respective solution.

5. What is the expected electrical conductivity of the following solutions?

- (a) $\text{NaOH}(aq)$
- (b) $\text{HCl}(aq)$
- (c) $\text{C}_6\text{H}_{12}\text{O}_6(aq)$ (glucose)
- (d) $\text{NH}_3(aq)$

6. Why are most *solid* ionic compounds electrically nonconductive, whereas aqueous solutions of ionic compounds are good conductors? Would you expect a *liquid* (molten) ionic compound to be electrically conductive or nonconductive? Explain.

7. Indicate the most important type of intermolecular attraction responsible for solvation in each of the following solutions:

- (a) the solutions in [Figure 2](#)
- (b) methanol, CH_3OH , dissolved in ethanol, $\text{C}_2\text{H}_5\text{OH}$
- (c) methane, CH_4 , dissolved in benzene, C_6H_6
- (d) the polar halocarbon CF_2Cl_2 dissolved in the polar halocarbon $\text{CF}_2\text{ClCFCl}_2$
- (e) $\text{O}_2(l)$ in $\text{N}_2(l)$

Glossary

dissociation

physical process accompanying the dissolution of an ionic compound in which the compound's constituent ions are solvated and dispersed throughout the solution

electrolyte

substance that produces ions when dissolved in water

ion-dipole attraction

electrostatic attraction between an ion and a polar molecule

nonelectrolyte

substance that does not produce ions when dissolved in water

strong electrolyte

substance that dissociates or ionizes completely when dissolved in water

weak electrolyte

substance that ionizes only partially when dissolved in water

8.6: [Electrolytes](#) is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.7: Concentration

28

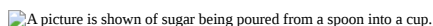
Harper College Chemistry Department

Page by: OpenStax

By the end of this section, you will be able to:

- Describe the fundamental properties of solutions
- Calculate solution concentrations using molarity
- Perform dilution calculations using the dilution equation

In preceding sections, we focused on the composition of substances: samples of matter that contain only one type of element or compound. However, mixtures—samples of matter containing two or more substances physically combined—are more commonly encountered in nature than are pure substances. Similar to a pure substance, the relative composition of a mixture plays an important role in determining its properties. The relative amount of oxygen in a planet’s atmosphere determines its ability to sustain aerobic life. The relative amounts of iron, carbon, nickel, and other elements in steel (a mixture known as an “alloy”) determine its physical strength and resistance to corrosion. The relative amount of the active ingredient in a medicine determines its effectiveness in achieving the desired pharmacological effect. The relative amount of sugar in a beverage determines its sweetness (see Figure 1). In this section, we will describe one of the most common ways in which the relative compositions of mixtures may be quantified.

A picture is shown of sugar being poured from a spoon into a cup.

Sugar is one of many components in the complex mixture known as coffee. The amount of sugar in a given amount of coffee is an important determinant of the beverage’s sweetness. (credit: Jane Whitney)

Solutions

We have previously defined solutions as homogeneous mixtures, meaning that the composition of the mixture (and therefore its properties) is uniform throughout its entire volume. Solutions occur frequently in nature and have also been implemented in many forms of manmade technology. We will explore a more thorough treatment of solution properties in the chapter on solutions and colloids, but here we will introduce some of the basic properties of solutions.

The relative amount of a given solution component is known as its concentration. Often, though not always, a solution contains one component with a concentration that is significantly greater than that of all other components. This component is called the solvent and may be viewed as the medium in which the other components are dispersed, or dissolved. Solutions in which water is the solvent are, of course, very common on our planet. A solution in which water is the solvent is called an aqueous solution.

A solute is a component of a solution that is typically present at a much lower concentration than the solvent. Solute concentrations are often described with qualitative terms such as dilute (of relatively low concentration) and concentrated (of relatively high concentration).

Concentrations may be quantitatively assessed using a wide variety of measurement units, each convenient for particular applications. Molarity (M) is a useful concentration unit for many applications in chemistry. Molarity is defined as the number of moles of solute in exactly 1 liter (1 L) of the solution:

$$M = \frac{\text{mol solute}}{\text{L solution}}$$

Calculating Molar Concentrations

A 355-mL soft drink sample contains 0.133 mol of sucrose (table sugar). What is the molar concentration of sucrose in the beverage?

Solution

Since the molar amount of solute and the volume of solution are both given, the molarity can be calculated using the definition of molarity. Per this definition, the solution volume must be converted from mL to L:

$$M = \frac{\text{mol solute}}{\text{L solution}} = \frac{0.133 \text{ mol}}{355 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.375 \text{ M}$$

Check Your Learning

A teaspoon of table sugar contains about 0.01 mol sucrose. What is the molarity of sucrose if a teaspoon of sugar has been dissolved in a cup of tea with a volume of 200 mL?

Answer:

0.05 M

Deriving Moles and Volumes from Molar Concentrations

How much sugar (mol) is contained in a modest sip (~10 mL) of the soft drink from the [Example](#) above where the concentration is 0.375 M?

Solution

In this case, we can rearrange the definition of molarity to isolate the quantity sought, moles of sugar. We then substitute the value for molarity that we derived in [Example](#), 0.375 M:

$$M = \frac{\text{mol solute}}{L \text{ solution}}$$

$$\text{mol solute} = M \times L \text{ solution}$$

$$\text{mol solute} = \frac{0.375 \text{ mol sugar}}{L} \times 10 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.004 \text{ mol sugar}$$

Check Your Learning

What volume (mL) of the sweetened tea described in the [Example](#) above contains the same amount of sugar (mol) as 10 mL of the soft drink in this example?

Answer

80 mL

Calculating Molar Concentrations from the Mass of Solute

Distilled white vinegar ([Figure 2](#)) is a solution of acetic acid, $\text{CH}_3\text{CO}_2\text{H}$, in water. A 0.500-L vinegar solution contains 25.2 g of acetic acid. What is the concentration of the acetic acid solution in units of molarity?

 A label on a container is shown. The label has a picture of a salad with the words "Distilled White Vinegar," and, "Reduced with water to 5% acidity," written above it.

Distilled white vinegar is a solution of acetic acid in water.

Solution

As in previous examples, the definition of molarity is the primary equation used to calculate the quantity sought. In this case, the mass of solute is provided instead of its molar amount, so we must use the solute's molar mass to obtain the amount of solute in moles:

$$\text{moles } \text{C}_2\text{H}_3\text{CO}_2\text{H} = \frac{25.2 \text{ g } \text{C}_2\text{H}_3\text{CO}_2\text{H}}{60.052 \frac{\text{g}}{\text{mol}}} = 0.420 \text{ mol } \text{C}_2\text{H}_3\text{CO}_2\text{H}$$

$$M = \frac{\text{mol solute}}{L \text{ solution}}$$

$$M = \frac{0.420 \text{ mol } \text{C}_2\text{H}_3\text{CO}_2\text{H}}{0.5 \text{ L solution}} = 0.839 \text{ M } \text{C}_2\text{H}_3\text{CO}_2\text{H}$$

Check Your Learning

Calculate the molarity of 6.52 g of CoCl_2 (128.9 g/mol) dissolved in an aqueous solution with a total volume of 75.0 mL.

Answer:

0.674 M

Determining the Mass of Solute in a Given Volume of Solution

How many grams of NaCl are contained in 0.250 L of a 5.30 M solution?

Solution

The volume and molarity of the solution are specified, so the amount (mol) of solute is easily computed as demonstrated in [Example](#):

$$M = \frac{\text{mol solute}}{L \text{ solution}}$$

$$\text{mol solute} = M \times L \text{ solution}$$

$$\text{mol solute} = 5.30 \frac{\text{mol NaCl}}{L} \times 0.250 L = 1.325 \text{ mol NaCl}$$

Finally, this molar amount is used to derive the mass of NaCl:

$$1.325 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{\text{mol NaCl}} = 77.4 \text{ g NaCl}$$

Check Your Learning

How many grams of CaCl_2 (110.98 g/mol) are contained in 250.0 mL of a 0.200-M solution of calcium chloride?

Answer:

5.55 g CaCl_2

When performing calculations stepwise, it is important to refrain from rounding any intermediate calculation results, which can lead to rounding errors in the final result. In [Example](#), the molar amount of NaCl computed in the first step, 1.325 mol, would be properly rounded to 1.32 mol if it were to be reported; however, although the last digit (5) is not significant, it must be retained as a guard digit in the intermediate calculation. If we had not retained this guard digit, the final calculation for the mass of NaCl would have been 77.1 g, a difference of 0.3 g.

In addition to retaining a guard digit for intermediate calculations, we can also avoid rounding errors by performing computations in a single step (see [Example](#)). This eliminates intermediate steps so that only the final result is rounded.

Determining the Volume of Solution Containing a Given Mass of Solute

In [Example](#), we found the typical concentration of vinegar to be 0.839 M. What volume of vinegar contains 75.6 g of acetic acid?

Solution

First, use the molar mass to calculate moles of acetic acid from the given mass:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g}} = \text{mol solute}$$

Then, use the molarity of the solution to calculate the volume of solution containing this molar amount of solute:

$$\text{mol solute} \times \frac{L \text{ solution}}{\text{mol solute}} = L \text{ solution}$$

Combining these two steps into one yields:

$$\text{g solute} \times \frac{\text{mol solute}}{\text{g solute}} \times \frac{L \text{ solution}}{\text{mol solute}} = L \text{ solution}$$

$$75.6 \text{ g } \text{C}_2\text{H}_3\text{CO}_2\text{H} \times \frac{\text{mol } \text{C}_2\text{H}_3\text{CO}_2\text{H}}{60.05 \text{ g}} \times \frac{L \text{ solution}}{0.839 \text{ mol } \text{C}_2\text{H}_3\text{CO}_2\text{H}} = 1.50 L \text{ solution}$$

Check Your Learning

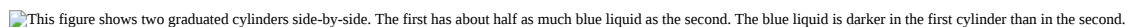
What volume of a 1.50-M KBr solution contains 66.0 g KBr?

Answer:

0.370 L

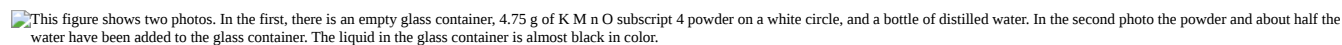
Dilution of Solutions

Dilution is the process whereby the concentration of a solution is lessened by the addition of solvent. For example, we might say that a glass of iced tea becomes increasingly diluted as the ice melts. The water from the melting ice increases the volume of the solvent (water) and the overall volume of the solution (iced tea), thereby reducing the relative concentrations of the solutes that give the beverage its taste ([Figure 3](#)).

 This figure shows two graduated cylinders side-by-side. The first has about half as much blue liquid as the second. The blue liquid is darker in the first cylinder than in the second.

Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott)

Dilution is also a common means of preparing solutions of a desired concentration. By adding solvent to a measured portion of a more concentrated *stock solution*, we can achieve a particular concentration. For example, commercial pesticides are typically sold as solutions in which the active ingredients are far more concentrated than is appropriate for their application. Before they can be used on crops, the pesticides must be diluted. This is also a very common practice for the preparation of a number of common laboratory reagents (Figure 4).

 This figure shows two photos. In the first, there is an empty glass container, 4.75 g of KMnO_4 powder on a white circle, and a bottle of distilled water. In the second photo the powder and about half the water have been added to the glass container. The liquid in the glass container is almost black in color.

A solution of KMnO_4 is prepared by mixing water with 4.74 g of KMnO_4 in a flask. (credit: modification of work by Mark Ott)

A simple mathematical relationship can be used to relate the volumes and concentrations of a solution before and after the dilution process. According to the definition of molarity, the molar amount of solute (n) in a solution is equal to the product of the solution's molarity (M) and its volume (L) in liters:

$$n = M \times L$$

Expressions like these may be written for a solution before and after it is diluted:

$$n_1 = M_1 \times L_1$$

$$n_2 = M_2 \times L_2$$

where the subscripts "1" and "2" refer to the solution before and after the dilution, respectively. Since the dilution process *does not change the amount of solute in the solution*, $n_1 = n_2$. Thus, these two equations may be set equal to one another:

$$M_1 \times L_1 = M_2 \times L_2$$

This relation is commonly referred to as the dilution equation. Although we derived this equation using molarity as the unit of concentration and liters as the unit of volume, other units of concentration and volume may be used, so long as the units properly cancel per the factor-label method. Reflecting this versatility, the dilution equation is often written in the more general form:

$$C_1 V_1 = C_2 V_2$$

where C and V are concentration and volume, respectively.

Use the [simulation](#) to explore the relations between solute amount, solution volume, and concentration and to confirm the dilution equation.

Determining the Concentration of a Diluted Solution

If 0.850 L of a 5.00- M solution of copper nitrate, $\text{Cu}(\text{NO}_3)_2$, is diluted to a volume of 1.80 L by the addition of water, what is the molarity of the diluted solution?

Solution

We are given the volume and concentration of a stock solution, V_1 and C_1 , and the volume of the resultant diluted solution, V_2 . We need to find the concentration of the diluted solution, C_2 . We thus rearrange the dilution equation, by dividing both sides of the equation by V_2 , in order to isolate C_2 :

$$C_1 V_1 = C_2 V_2$$

$$\frac{C_1 V_1}{V_2} = C_2$$

Since the stock solution is being diluted by more than two-fold (volume is increased from 0.85 L to 1.80 L), we would expect the diluted solution's concentration to be less than one-half 5 M . We will compare this ballpark estimate to the calculated result to check for any gross errors in computation (for example, such as an improper substitution of the given quantities). Substituting the given values for the terms on the right side of this equation yields:

$$\frac{5.00 \text{ mol/L} \times 0.850 \text{ L}}{1.80 \text{ L}} = C_2 = 2.36 \text{ M}$$

This result compares well to our ballpark estimate (it's a bit less than one-half the stock concentration, 5 M).

Check Your Learning

What is the concentration of the solution that results from diluting 25.0 mL of a 2.04-M solution of CH_3OH to 500.0 mL?

Answer:

0.102 M CH_3OH

Volume of a Diluted Solution

What volume of 0.12 M HBr can be prepared from 11 mL (0.011 L) of 0.45 M HBr?

Solution

We are given the volume and concentration of a stock solution, V_1 and C_1 , and the concentration of the resultant diluted solution, C_2 . We need to find the volume of the diluted solution, V_2 . We thus rearrange the dilution equation in order to isolate V_2 :

$$C_1V_1 = C_2V_2$$

$$V_2 = \frac{C_1V_1}{C_2}$$

Since the diluted concentration (0.12 M) is slightly more than one-fourth the original concentration (0.45 M), we would expect the volume of the diluted solution to be roughly four times the original volume, or around 44 mL. Substituting the given values and solving for the unknown volume yields:

$$V_2 = \frac{0.011 \text{ L} \times 0.45 \text{ mol/L}}{0.12 \text{ mol/L}} = 0.041 \text{ L}$$

The volume of the 0.12-M solution is 0.041 L (41 mL). The result is reasonable and compares well with our rough estimate.

Check Your Learning

A laboratory experiment calls for 0.125 M HNO_3 . What volume of 0.125 M HNO_3 can be prepared from 0.250 L of 1.88 M HNO_3 ?

Answer:

3.76 L

Volume of a Concentrated Solution Needed for Dilution

What volume of 1.59 M KOH is required to prepare 5.00 L of 0.100 M KOH?

Solution

We are given the concentration of a stock solution, C_1 , and the volume and concentration of the resultant diluted solution, V_2 and C_2 . We need to find the volume of the stock solution, V_1 . We thus rearrange the dilution equation in order to isolate V_1 :

$$C_1V_1 = C_2V_2$$

$$V_1 = \frac{C_2V_2}{C_1}$$

Since the concentration of the diluted solution 0.100 M is roughly one-sixteenth that of the stock solution (1.59 M), we would expect the volume of the stock solution to be about one-sixteenth that of the diluted solution, or around 0.3 liters. Substituting the given values and solving for the unknown volume yields:

$$V_1 = \frac{0.100 \text{ M} \times 5.00 \text{ L}}{1.59 \text{ M}} = 0.314 \text{ L}$$

Thus, we would need 0.314 L of the 1.59-M solution to prepare the desired solution. This result is consistent with our rough estimate.

Check Your Learning

What volume of a 0.575-M solution of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, can be prepared from 50.00 mL of a 3.00-M glucose solution?

Answer:

0.261 L

Key Concepts and Summary

Solutions are homogeneous mixtures. Many solutions contain one component, called the solvent, in which other components, called solutes, are dissolved. An aqueous solution is one for which the solvent is water. The concentration of a solution is a measure of the relative amount of solute in a given amount of solution. Concentrations may be measured using various units, with one very useful unit being molarity, defined as the number of moles of solute per liter of solution. The solute concentration of a solution may be decreased by adding solvent, a process referred to as dilution. The dilution equation is a simple relation between concentrations and volumes of a solution before and after dilution.

Chemistry End of Chapter Exercises

1. Explain what changes and what stays the same when 1.00 L of a solution of NaCl is diluted to 1.80 L.
2. What information do we need to calculate the molarity of a sulfuric acid solution?
3. a) What does it mean when we say that a 200-mL sample and a 400-mL sample of a solution of salt have the same molarity? b) In what ways are the two samples identical? c) In what ways are these two samples different?
4. Determine the molarity for each of the following solutions:
 - (a) 0.444 mol of CoCl_2 in 0.654 L of solution
 - (b) 98.0 g of phosphoric acid, H_3PO_4 , in 1.00 L of solution
 - (c) 0.2074 g of calcium hydroxide, $\text{Ca}(\text{OH})_2$, in 40.00 mL of solution
 - (d) 10.5 kg of $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$ in 18.60 L of solution
 - (e) 7.0×10^{-3} mol of I_2 in 100.0 mL of solution
 - (f) 1.8×10^4 mg of HCl in 0.075 L of solution
5. Determine the molarity of each of the following solutions:
 - (a) 1.457 mol KCl in 1.500 L of solution
 - (b) 0.515 g of H_2SO_4 in 1.00 L of solution
 - (c) 20.54 g of $\text{Al}(\text{NO}_3)_3$ in 1575 mL of solution
 - (d) 2.76 kg of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 1.45 L of solution
 - (e) 0.005653 mol of Br_2 in 10.00 mL of solution
 - (f) 0.000889 g of glycine, $\text{C}_2\text{H}_5\text{NO}_2$, in 1.05 mL of solution
6. Consider this question: What is the mass of the solute in 0.500 L of 0.30 M glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, used for intravenous injection?
 - (a) Outline the steps necessary to answer the question.
 - (b) Answer the question.
7. Consider this question: What is the mass of solute in 200.0 L of a 1.556-M solution of KBr?
 - (a) Outline the steps necessary to answer the question.
 - (b) Answer the question.
8. Calculate the number of moles and the mass of the solute in each of the following solutions:
 - (a) 2.00 L of 18.5 M H_2SO_4 , concentrated sulfuric acid
 - (b) 100.0 mL of 3.8×10^{-5} M NaCN, the minimum lethal concentration of sodium cyanide in blood serum
 - (c) 5.50 L of 13.3 M H_2CO , the formaldehyde used to “fix” tissue samples
 - (d) 325 mL of 1.8×10^{-6} M FeSO_4 , the minimum concentration of iron sulfate detectable by taste in drinking water
9. Calculate the number of moles and the mass of the solute in each of the following solutions:

- (a) 325 mL of $8.23 \times 10^{-5} \text{ M}$ KI, a source of iodine in the diet
 - (b) 75.0 mL of $2.2 \times 10^{-5} \text{ M}$ H_2SO_4 , a sample of acid rain
 - (c) 0.2500 L of 0.1135 M K_2CrO_4 , an analytical reagent used in iron assays
 - (d) 10.5 L of 3.716 M $(\text{NH}_4)_2\text{SO}_4$, a liquid fertilizer
10. Consider this question: What is the molarity of KMnO_4 in a solution of 0.0908 g of KMnO_4 in 0.500 L of solution?
- (a) Outline the steps necessary to answer the question.
 - (b) Answer the question.
11. Consider this question: What is the molarity of HCl if 35.23 mL of a solution of HCl contain 0.3366 g of HCl?
- (a) Outline the steps necessary to answer the question.
 - (b) Answer the question.
12. Calculate the molarity of each of the following solutions:
- (a) 0.195 g of cholesterol, $\text{C}_{27}\text{H}_{46}\text{O}$, in 0.100 L of serum, the average concentration of cholesterol in human serum
 - (b) 4.25 g of NH_3 in 0.500 L of solution, the concentration of NH_3 in household ammonia
 - (c) 1.49 kg of isopropyl alcohol, $\text{C}_3\text{H}_7\text{OH}$, in 2.50 L of solution, the concentration of isopropyl alcohol in rubbing alcohol
 - (d) 0.029 g of I_2 in 0.100 L of solution, the solubility of I_2 in water at 20°C
13. Calculate the molarity of each of the following solutions:
- (a) 293 g HCl in 666 mL of solution, a concentrated HCl solution
 - (b) 2.026 g FeCl_3 in 0.1250 L of a solution used as an unknown in general chemistry laboratories
 - (c) 0.001 mg Cd^{2+} in 0.100 L, the maximum permissible concentration of cadmium in drinking water
 - (d) 0.0079 g $\text{C}_7\text{H}_5\text{SNO}_3$ in one ounce (29.6 mL), the concentration of saccharin in a diet soft drink.
14. There is about 1.0 g of calcium, as Ca^{2+} , in 1.0 L of milk. What is the molarity of Ca^{2+} in milk?
15. What volume of a 1.00-M $\text{Fe}(\text{NO}_3)_3$ solution can be diluted to prepare 1.00 L of a solution with a concentration of 0.250 M?
16. If 0.1718 L of a 0.3556-M $\text{C}_3\text{H}_7\text{OH}$ solution is diluted to a concentration of 0.1222 M, what is the volume of the resulting solution?
17. If 4.12 L of a 0.850 M- H_3PO_4 solution is be diluted to a volume of 10.00 L, what is the concentration of the resulting solution?
18. What volume of a 0.33-M $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ solution can be diluted to prepare 25 mL of a solution with a concentration of 0.025 M?
19. What is the concentration of the NaCl solution that results when 0.150 L of a 0.556-M solution is allowed to evaporate until the volume is reduced to 0.105 L?
20. What is the molarity of the diluted solution when each of the following solutions is diluted to the given final volume?
- (a) 1.00 L of a 0.250-M solution of $\text{Fe}(\text{NO}_3)_3$ is diluted to a final volume of 2.00 L
 - (b) 0.5000 L of a 0.1222-M solution of $\text{C}_3\text{H}_7\text{OH}$ is diluted to a final volume of 1.250 L
 - (c) 2.35 L of a 0.350-M solution of H_3PO_4 is diluted to a final volume of 4.00 L
 - (d) 22.50 mL of a 0.025-M solution of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ is diluted to 100.0 mL
21. What is the final concentration of the solution produced when 225.5 mL of a 0.09988-M solution of Na_2CO_3 is allowed to evaporate until the solution volume is reduced to 45.00 mL?
22. A 2.00-L bottle of a solution of concentrated HCl was purchased for the general chemistry laboratory. The solution contained 868.8 g of HCl. What is the molarity of the solution?

23. An experiment in a general chemistry laboratory calls for a 2.00-M solution of HCl. How many mL of 11.9 M HCl would be required to make 250 mL of 2.00 M HCl?
24. What volume of a 0.20-M K_2SO_4 solution contains 57 g of K_2SO_4 ?
25. The US Environmental Protection Agency (EPA) places limits on the quantities of toxic substances that may be discharged into the sewer system. Limits have been established for a variety of substances, including hexavalent chromium, which is limited to 0.50 mg/L. If an industry is discharging hexavalent chromium as potassium dichromate ($K_2Cr_2O_7$), what is the maximum permissible molarity of that substance?

Key Equations

- $M = \frac{\text{mol \ solute}}{L \ \text{solution}}$
- $C_1V_1 = C_2V_2$

Glossary

aqueous solution

solution for which water is the solvent

concentrated

qualitative term for a solution containing solute at a relatively high concentration

concentration

quantitative measure of the relative amounts of solute and solvent present in a solution

dilute

qualitative term for a solution containing solute at a relatively low concentration

dilution

process of adding solvent to a solution in order to lower the concentration of solutes

dissolved

describes the process by which solute components are dispersed in a solvent

molarity (M)

unit of concentration, defined as the number of moles of solute dissolved in 1 liter of solution

solute

solution component present in a concentration less than that of the solvent

solvent

solution component present in a concentration that is higher relative to other components

8.7: Concentration is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

8.8: Osmosis, diffusion, and dialysis

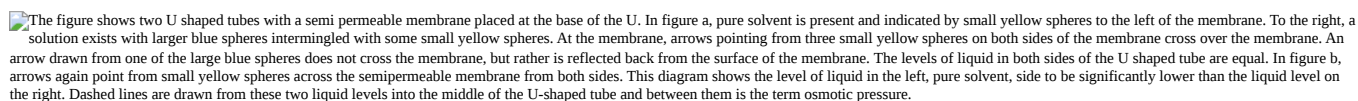
29

Harper College Chemistry Department

Osmosis and Osmotic Pressure of Solutions

A number of natural and synthetic materials exhibit *selective permeation*, meaning that only molecules or ions of a certain size, shape, polarity, charge, and so forth, are capable of passing through (permeating) the material. Biological cell membranes provide elegant examples of selective permeation in nature, while dialysis tubing used to remove metabolic wastes from blood is a more simplistic technological example. Regardless of how they may be fabricated, these materials are generally referred to as semipermeable membranes.

Consider the apparatus illustrated in [Figure 1](#), in which samples of pure solvent and a solution are separated by a membrane that only solvent molecules may permeate. Solvent molecules will diffuse across the membrane in both directions. Since the concentration of *solvent* is greater in the pure solvent than the solution, these molecules will diffuse from the solvent side of the membrane to the solution side at a faster rate than they will in the reverse direction. The result is a net transfer of solvent molecules from the pure solvent to the solution. Transfer of solvent molecules through a semipermeable membrane is a process known as osmosis.

The figure shows two U shaped tubes with a semi permeable membrane placed at the base of the U. In figure a, pure solvent is present and indicated by small yellow spheres to the left of the membrane. To the right, a solution exists with larger blue spheres intermingled with some small yellow spheres. At the membrane, arrows pointing from three small yellow spheres on both sides of the membrane cross over the membrane. An arrow drawn from one of the large blue spheres does not cross the membrane, but rather is reflected back from the surface of the membrane. The levels of liquid in both sides of the U shaped tube are equal. In figure b, arrows again point from small yellow spheres across the semipermeable membrane from both sides. This diagram shows the level of liquid in the left, pure solvent, side to be significantly lower than the liquid level on the right. Dashed lines are drawn from these two liquid levels into the middle of the U-shaped tube and between them is the term osmotic pressure.

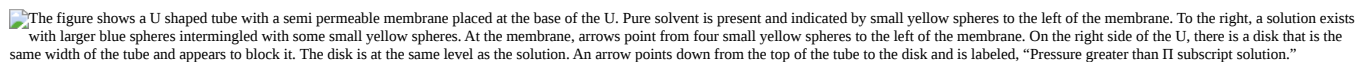
Osmosis results in the transfer of solvent molecules from a sample of low (or zero) solute concentration to a sample of higher solute concentration.

When osmosis is carried out in an apparatus like that shown in [Figure 1](#), the volume of the solution increases as it becomes diluted by accumulation of solvent. This causes the level of the solution to rise, increasing its hydrostatic pressure (due to the weight of the column of solution in the tube) and resulting in a faster transfer of solvent molecules back to the pure solvent side. When the pressure reaches a value that yields a reverse solvent transfer rate equal to the osmosis rate, bulk transfer of solvent ceases. This pressure is called the osmotic pressure (Π) of the solution. The osmotic pressure of a dilute solution is related to its solute molarity, M , and absolute temperature, T , according to the equation

$$\Pi = MRT$$

where R is the universal gas constant.

If a solution is placed in an apparatus like the one shown in [Figure 2](#), applying pressure greater than the osmotic pressure of the solution reverses the osmosis and pushes solvent molecules from the solution into the pure solvent. This technique of reverse osmosis is used for large-scale desalination of seawater and on smaller scales to produce high-purity tap water for drinking.

The figure shows a U shaped tube with a semi permeable membrane placed at the base of the U. Pure solvent is present and indicated by small yellow spheres to the left of the membrane. To the right, a solution exists with larger blue spheres intermingled with some small yellow spheres. At the membrane, arrows point from four small yellow spheres to the left of the membrane. On the right side of the U, there is a disk that is the same width of the tube and appears to block it. The disk is at the same level as the solution. An arrow points down from the top of the tube to the disk and is labeled, "Pressure greater than Π subscript solution."

Applying a pressure greater than the osmotic pressure of a solution will reverse osmosis. Solvent molecules from the solution are pushed into the pure solvent.

REVERSE OSMOSIS WATER PURIFICATION

In the process of osmosis, water moves through a semipermeable membrane from a less concentrated solution to a more concentrated solution. Osmotic pressure is the amount of pressure that must be applied to the more concentrated solution to cause osmosis to stop. If greater pressure is applied, the water will go from the more concentrated solution to a less concentrated (more pure) solution. This is called reverse osmosis. Reverse osmosis (RO) is used to purify water in many applications, from desalination plants in coastal cities, to water-purifying machines in grocery stores ([Figure 3](#)), and smaller reverse-osmosis household units. With a hand-operated pump, small RO units can be used in third-world countries, disaster areas, and in lifeboats. Our military forces have a variety of generator-operated RO units that can be transported in vehicles to remote locations.

 This figure shows two photos of reverse osmosis systems. The first is a small system that appears easily portable. The second is larger and situated outdoors.

Reverse osmosis systems for purifying drinking water are shown here on (a) small and (b) large scales. (credit a: modification of work by Jerry Kirkhart; credit b: modification of work by Willard J. Lathrop)

Examples of osmosis are evident in many biological systems because cells are surrounded by semipermeable membranes. Carrots and celery that have become limp because they have lost water can be made crisp again by placing them in water. Water moves into the carrot or celery cells by osmosis. A cucumber placed in a concentrated salt solution loses water by osmosis and absorbs some salt to become a pickle. Osmosis can also affect animal cells. Solute concentrations are particularly important when solutions are injected into the body. Solutes in body cell fluids and blood serum give these solutions an osmotic pressure of approximately 7.7 atm. Solutions injected into the body must have the same osmotic pressure as blood serum; that is, they should be isotonic with blood serum. If a less concentrated solution, a hypotonic solution, is injected in sufficient quantity to dilute the blood serum, water from the diluted serum passes into the blood cells by osmosis, causing the cells to expand and rupture. This process is called hemolysis. When a more concentrated solution, a hypertonic solution, is injected, the cells lose water to the more concentrated solution, shrivel, and possibly die in a process called crenation. These effects are illustrated in [Figure 4](#).

 This figure shows three scenarios relating to red blood cell membranes. In a, H_2O has two arrows drawn from it pointing into a red disk. Beneath it in a circle are eleven similar disks with a bulging appearance, one of which appears to have burst with blue liquid erupting from it. In b, the image is similar except that rather than having two arrows pointing into the red disk, one points in and a second points out toward the H_2O . In the circle beneath, twelve of the red disks are present. In c, both arrows are drawn from a red shriveled disk toward the H_2O . In the circle below, twelve shriveled disks are shown.

Red blood cell membranes are water permeable and will (a) swell and possibly rupture in a hypotonic solution; (b) maintain normal volume and shape in an isotonic solution; and (c) shrivel and possibly die in a hypertonic solution. (credit a/b/c: modifications of work by "LadyofHats"/Wikimedia commons)

8.8: Osmosis, diffusion, and dialysis is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

9: Acids and Bases

9.1: What are acids and bases?

9.2: pH and pOH

9: Acids and Bases is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

9.1: What are acids and bases?

30

Harper College Chemistry Department

Page by: OpenStax

By the end of this section, you will be able to:

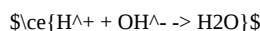
- Identify acids, bases, and conjugate acid-base pairs according to the Brønsted-Lowry definition
- Write equations for acid and base ionization reactions
- Use the ion-product constant for water to calculate hydronium and hydroxide ion concentrations
- Describe the acid-base behavior of amphiprotic substances

Acids and bases have been known for a long time. When Robert Boyle characterized them in 1680, he noted that acids dissolve many substances, change the color of certain natural dyes (for example, they change litmus from blue to red), and lose these characteristic properties after coming into contact with alkalis (bases). In the eighteenth century, it was recognized that acids have a sour taste, react with limestone to liberate a gaseous substance (now known to be CO_2), and interact with alkalis to form neutral substances. In 1815, Humphry Davy contributed greatly to the development of the modern acid-base concept by demonstrating that hydrogen is the essential constituent of acids. Around that same time, Joseph Louis Gay-Lussac concluded that acids are substances that can neutralize bases and that these two classes of substances can be defined only in terms of each other. The significance of hydrogen was reemphasized in 1884 when Svante Arrhenius defined an acid as a compound that dissolves in water to yield hydrogen cations (now recognized to be hydronium ions) and a base as a compound that dissolves in water to yield hydroxide anions.

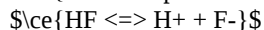
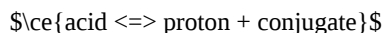
In an earlier chapter on chemical reactions, we defined acids and bases as Arrhenius did: We identified an acid as a compound that dissolves in water to yield hydronium ions (H_3O^+) and a base as a compound that dissolves in water to yield hydroxide ions (OH^-). This definition is not wrong; it is simply limited.

Later, we extended the definition of an acid or a base using the more general definition proposed in 1923 by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry. Their definition centers on the proton, H^+ . A proton is what remains when the most common isotope of hydrogen, ^1_1H , loses an electron. A compound that donates a proton to another compound is called a Brønsted-Lowry acid, and a compound that accepts a proton is called a Brønsted-Lowry base. An acid-base reaction is the transfer of a proton from a proton donor (acid) to a proton acceptor (base). In a subsequent chapter of this text we will introduce the most general model of acid-base behavior introduced by the American chemist G. N. Lewis.

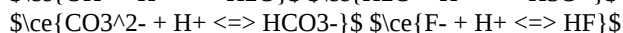
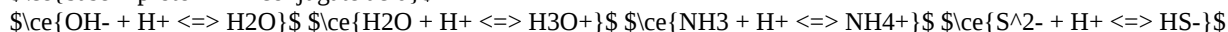
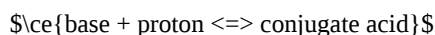
Acids may be compounds such as HCl or H_2SO_4 , organic acids like acetic acid (CH_3COOH) or ascorbic acid (vitamin C), or H_2O . Anions (such as HSO_4^- , H_2PO_4^- , HS^- , and HCO_3^-) and cations (such as H_3O^+ , NH_4^+ , and $[\text{Al}(\text{H}_2\text{O})_6]^{3+}$) may also act as acids. Bases fall into the same three categories. Bases may be neutral molecules (such as H_2O , NH_3 , and CH_3NH_2), anions (such as OH^- , HS^- , HCO_3^- , CO_3^{2-} , F^- , and PO_4^{3-}), or cations (such as $[\text{Al}(\text{H}_2\text{O})_5\text{OH}]^{2+}$). The most familiar bases are ionic compounds such as NaOH and $\text{Ca}(\text{OH})_2$, which contain the hydroxide ion, OH^- . The hydroxide ion in these compounds accepts a proton from acids to form water:



We call the product that remains after an acid donates a proton the conjugate base of the acid. This species is a base because it can accept a proton (to re-form the acid):



We call the product that results when a base accepts a proton the base's conjugate acid. This species is an acid because it can give up a proton (and thus re-form the base):



In these two sets of equations, the behaviors of acids as proton donors and bases as proton acceptors are represented in isolation. 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, OH^- , and the conjugate acid of ammonia, NH_4^+ + 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_2O 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_2O (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, " OH^- (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 " NH_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 " NH_4^+ (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, $\text{H}_2\text{O} + \text{NH}_3 \rightleftharpoons \text{OH}^- + \text{NH}_4^+$ is shown in purple, and is labeled below in purple as "acid," plus NH_3 in green, labeled below in green as "base," followed by a double sided arrow and OH^- in purple, labeled in purple as "conjugate base," plus NH_4^+ 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.

The reaction between a Brønsted-Lowry acid and water is called acid ionization. For example, when hydrogen fluoride dissolves in water and ionizes, protons are transferred from hydrogen fluoride molecules to water molecules, yielding hydronium ions and fluoride ions:

This figure has two rows. In both rows, a chemical reaction is shown. In the first, structural formulas are provided. In this model, in purple, an H atom is connected to an F atom with a single bond. The F atom has pairs of electron dots at the top, right, and bottom. This is followed by a plus sign, which is followed in green by an O atom which has H atoms singly bonded above and to the right. The O atom has pairs of electron dots on its left and lower sides. A double arrow follows. To the right, in brackets is a structure with a central O atom in green, with green H atoms singly bonded above and to the right. A pair of green electron dots is on the lower side of the O atom. To the left of the green O atom, a purple H atom is singly bonded. This is followed by a plus sign and an F atom in purple with pairs of electron dots above, right, below, and to the left. This atom also has a superscript negative sign. The reaction is written in symbolic form below. HF is labeled in purple below as "Acid subscript 1." This is followed by plus H_2O , which is labeled in green below as "Base subscript 2." A double sided arrow follows. To the right is H_3O^+ , which is labeled in green as below in as "Acid subscript 2." This is followed by plus F^- surrounded by 4 pairs of dots and superscript negative. The label below in purple reads, "Base subscript 1." To the right of the reactions is the formula, $\text{K} = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$.

When we add a base to water, a base ionization reaction occurs in which protons are transferred from water molecules to base molecules. For example, adding pyridine to water yields hydroxide ions and pyridinium ions:

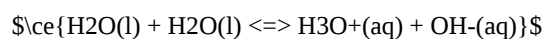
This figure has two rows. In both rows, a chemical reaction is shown. In the first, structural formulas are provided. In this model, in red, is an O atom which has H atoms singly bonded above and to the right. The O atom has lone pairs of electron dots on its left and lower sides. This is followed by a plus sign. The plus sign is followed, in blue, by an N atom with one lone pair of electron dots. The N atom forms a double bond with a C atom, which forms a single bond with a C atom. The second C atom forms a double bond with another C atom, which forms a single bond with another C atom. The fourth C atom forms a double bond with a fifth C atom, which forms a single bond with the N atom. This creates a ring structure. Each C atom is also bonded to an H atom. An equilibrium arrow follows this structure. To the right, in brackets is a structure where an N atom bonded to an H atom, which is red, appears. The N atom forms a double bond with a C atom, which forms a single bond with a C atom. The second C atom forms a double bond with another C atom, which forms a single bond with another C atom. The fourth C atom forms a double bond with a fifth C atom, which forms a single bond with the N atom. This creates a ring structure. Each C atom is also bonded to an H atom. Outside the brackets, to the right, is a superscript positive sign. This structure is followed by a plus sign. Another structure that appears in brackets also appears. An O atom with three lone pairs of electron dots is bonded to an H atom. There is a superscript negative sign outside the brackets. To the right, is the equation: $\text{K} = \frac{[\text{C}_5\text{H}_5\text{NH}^+][\text{OH}^-]}{[\text{C}_5\text{H}_5][\text{H}_2\text{O}]}$. Under the initial equation, is this equation: $\text{H}_2\text{O} + \text{C}_5\text{H}_5\text{N} \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+ + \text{OH}^-$. H_2O is labeled, "acid," in red. $\text{C}_5\text{H}_5\text{N}$ is labeled, "base," in blue. $\text{C}_5\text{H}_5\text{NH}^+$ is labeled, "acid," in blue. OH^- is labeled, "base," in red.

Notice that both these ionization reactions are represented as equilibrium processes. The relative extent to which these acid and base ionization reactions proceed is an important topic treated in a later section of this chapter. In the preceding paragraphs we saw that water can function as either an acid or a base, depending on the nature of the solute dissolved in it. In fact, in pure water or in any aqueous solution, water acts both as an acid and a base. A very small fraction of water molecules donate protons to other water molecules to form hydronium ions and hydroxide ions:

This figure has two rows. In both rows, a chemical reaction is shown. In the first, structural formulas are provided. In this model, in purple, O atom which has H atoms singly bonded above and to the right. The O atom has pairs of electron dots on its left and lower sides. This is followed by a plus sign, which is followed in green by an O atom which has H atoms singly bonded above and to the right. The O atom has pairs of electron dots on its left and lower sides. A double arrow follows. To the right, in brackets is a structure with a central O atom in green, with green H atoms singly bonded above and to the right. A pair of green electron dots is on the lower side of the O atom. To the left of the green O atom, a purple H atom is singly bonded. Outside the brackets to the right is a superscript plus. This is followed by a plus sign and an O atom in purple with pairs of electron dots above, left, and below. An H atom is singly bonded to the right. This atom has a superscript negative sign. The reaction is written in symbolic form below. $\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$. H_2O is labeled in purple below as "Acid subscript 1." This is followed by plus H_2O , which is labeled in green below as "Base subscript 2." A double sided arrow follows. To the right is H_3O^+ , which is labeled in green as below in as "Acid subscript 2." This is followed by plus OH^- with pairs of dots above, below, and to the left with a singly bonded H on the right with a superscript negative. The label below in purple reads, "Base subscript 1."

This type of reaction, in which a substance ionizes when one molecule of the substance reacts with another molecule of the same substance, is referred to as autoionization.

Pure water undergoes autoionization to a very slight extent. Only about two out of every 10^9 molecules in a sample of pure water are ionized at 25 °C. The equilibrium constant for the ionization of water is called the ion-product constant for water (K_w):



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

The slight ionization of pure water is reflected in the small value of the equilibrium constant; at 25 °C, K_w has a value of 1.0×10^{-14} . The process is endothermic, and so the extent of ionization and the resulting concentrations of hydronium ion and hydroxide

ion increase with temperature. For example, at 100 °C, the value for K_w is about 5.6×10^{-13} , roughly 50 times larger than the value at 25 °C.

9.1: What are acids and bases? is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

9.2: pH and pOH

31

Harper College Chemistry Department

As discussed earlier, hydronium and hydroxide ions are present both in pure water and in all aqueous solutions, and their concentrations are inversely proportional as determined by the ion product of water (K_w). The concentrations of these ions in a solution are often critical determinants of the solution's properties and the chemical behaviors of its other solutes, and specific vocabulary has been developed to describe these concentrations in relative terms. A solution is neutral if it contains equal concentrations of hydronium and hydroxide ions; acidic if it contains a greater concentration of hydronium ions than hydroxide ions; and basic if it contains a lesser concentration of hydronium ions than hydroxide ions.

A common means of expressing quantities, the values of which may span many orders of magnitude, is to use a logarithmic scale. One such scale that is very popular for chemical concentrations and equilibrium constants is based on the p-function, defined as shown where "X" is the quantity of interest and "log" is the base-10 logarithm:

$$pX = -\log X$$

The pH of a solution is therefore defined as shown here, where $[H_3O^+]$ is the molar concentration of hydronium ion in the solution:

$$pH = -\log[H_3O^+] = -\log[H^+]$$

Rearranging this equation to isolate the hydronium ion molarity yields the equivalent expression:

$$[H_3O^+] = 10^{-pH}$$

Likewise, the hydroxide ion molarity may be expressed as a p-function, or pOH:

$$pOH = -\log[OH^-]$$

or

$$[OH^-] = 10^{-pOH}$$

Finally, the relation between these two ion concentration expressed as p-functions is easily derived from the K_w expression:

$$K_w = [H_3O^+][OH^-]$$

$$-\log K_w = -\log([H_3O^+][OH^-]) = -\log[H_3O^+] + -\log[OH^-]$$

$$pK_w = pH + pOH$$

At 25 °C, the value of K_w is 1.0×10^{-14} , and so:

$$14.00 = pH + pOH$$

As was shown in [\[link\]](#), 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:

$$pH = -\log[H_3O^+] = -\log(1.0 \times 10^{-7}) = 7$$

$$pOH = -\log[OH^-] = -\log(1.0 \times 10^{-7}) = 7$$

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).

Since the autoionization constant K_w is temperature dependent, these correlations between pH values and the acidic/neutral/basic adjectives will be different at temperatures other than 25 °C. For example, the "Check Your Learning" exercise accompanying [\[link\]](#) showed the hydronium molarity of pure water at 80 °C is $4.9 \times 10^{-7} M$, which corresponds to pH and pOH values of 6.31.

$$pH = -\log[H_3O^+] = -\log(4.9 \times 10^{-7}) = 6.31 \quad pH = -\log[H_3O^+] = -\log(4.9 \times 10^{-7}) = 6.31$$

$$pOH = -\log[OH^-] = -\log(4.9 \times 10^{-7}) = 6.31 \quad pOH = -\log[OH^-] = -\log(4.9 \times 10^{-7}) = 6.31$$

At this temperature, then, neutral solutions exhibit $\text{pH} = \text{pOH} = 6.31$, acidic solutions exhibit pH less than 6.31 and pOH greater than 6.31, whereas basic solutions exhibit pH greater than 6.31 and pOH less than 6.31. This distinction can be important when studying certain processes that occur at nonstandard temperatures, such as enzyme reactions in warm-blooded organisms. Unless otherwise noted, references to pH values are presumed to be those at standard temperature (25°C) (Table).

Table: Summary of Relations for Acidic, Basic and Neutral Solutions		
Classification	Relative Ion Concentrations	pH at 25°C
acidic	$[\text{H}_3\text{O}^+] > [\text{OH}^-]$	$\text{pH} < 7$
neutral	$[\text{H}_3\text{O}^+] = [\text{OH}^-]$	$\text{pH} = 7$
basic	$[\text{H}_3\text{O}^+] < [\text{OH}^-]$	$\text{pH} > 7$

Figure 1 shows the relationships between $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH , and pOH , and gives values for these properties at standard temperatures for some common substances.

A table is provided with 5 columns. The first column is labeled "left bracket H subscript 3 O superscript plus right bracket (M)." Powers of ten are listed in the column beginning at 10^1 , including 10^0 or 1, $10^{\text{superscript negative 1}}$, decreasing by single powers of 10 to $10^{\text{superscript negative 15}}$. The second column is labeled "left bracket O H superscript negative right bracket (M)." Powers of ten are listed in the column beginning at $10^{\text{superscript negative 15}}$, increasing by single powers of 10 to including 10^0 or 1, and 10^1 . The third column is labeled "p H." Values listed in this column are integers beginning at negative 1, increasing by ones up to 14. The fourth column is labeled "p O H." Values in this column are integers beginning at 15, decreasing by ones up to negative 1. The fifth column is labeled "Sample Solution." A vertical line at the left of the column has tick marks corresponding to each p H level in the table. Substances are listed next to this line segment with line segments connecting them to the line to show approximate p H and p O H values. 1 M H C l is listed at a p H of 0. Gastric juices are listed at a p H of about 1.5. Lime juice is listed at a p H of about 2, followed by 1 M C H subscript 3 C O subscript 2 H, followed by stomach acid at a p H value of nearly 3. Wine is listed around 3.5. Coffee is listed just past 5. Pure water is listed at a p H of 7. Pure blood is just beyond 7. Milk of Magnesia is listed just past a p H of 10.5. Household ammonia is listed just before a p H of 12. 1 M N a O H is listed at a p H of 0. To the right of this labeled arrow is an arrow that points up and down through the height of the column. A beige strip passes through the table and to this double headed arrow at p H 7. To the left of the double headed arrow in this beige strip is the label "neutral." A narrow beige strip runs through the arrow. Just above and below this region, the arrow is purple. It gradually turns to a bright red as it extends upward. At the top of the arrow, near the head of the arrow is the label "acidic." Similarly, the lower region changes color from purple to blue moving to the bottom of the column. The head at this end of the arrow is labeled "basic."

The pH and pOH scales represent concentrations of $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$, respectively. The pH and pOH values of some common substances at standard temperature (25°C) are shown in this chart.

Calculation of pH from $[\text{H}_3\text{O}^+]$

What is the pH of stomach acid, a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3} \text{ M}$?

Solution

$$\begin{aligned}\text{pH} &= -\log[\text{H}_3\text{O}^+] \\ &= -\log(1.2 \times 10^{-3}) = -(-2.92) = 2.92\end{aligned}$$

(The use of logarithms is explained in Appendix B. Recall that, as we have done here, when taking the log of a value, keep as many decimal places in the result as there are significant figures in the value.)

Check Your Learning

Water exposed to air contains carbonic acid, H_2CO_3 , due to the reaction between carbon dioxide and water:



Air-saturated water has a hydronium ion concentration caused by the dissolved CO_2 of $2.0 \times 10^{-6} \text{ M}$, about 20-times larger than that of pure water. Calculate the pH of the solution at 25°C .

Answer:

5.70

Calculation of Hydronium Ion Concentration from pH

Calculate the hydronium ion concentration of blood, the pH of which is 7.3 (slightly alkaline).

Solution

$$\begin{aligned}\text{pH} &= -\log[\text{H}_3\text{O}^+] = 7.3 \\ [\text{H}_3\text{O}^+] &= 10^{(-7.3)} = 5 \times 10^{-8} \text{ M}\end{aligned}$$

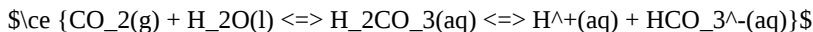
(On a calculator take the antilog, or the "inverse" log, of -7.3 , or calculate $10^{-7.3}$.)

Check Your Learning Calculate the hydronium ion concentration of a solution with a pH of -1.07 .

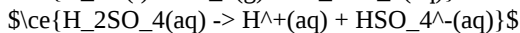
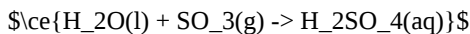
Answer:

ENVIRONMENTAL SCIENCE

Normal rainwater has a pH between 5 and 6 due to the presence of dissolved CO_2 which forms carbonic acid:



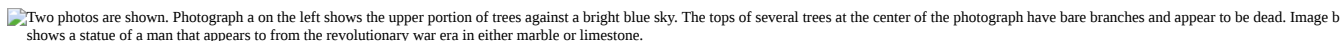
Acid rain is rainwater that has a pH of less than 5, due to a variety of nonmetal oxides, including CO_2 , SO_2 , SO_3 , NO , and NO_2 being dissolved in the water and reacting with it to form not only carbonic acid, but sulfuric acid and nitric acid. The formation and subsequent ionization of sulfuric acid are shown here:



Carbon dioxide is naturally present in the atmosphere because we and most other organisms produce it as a waste product of metabolism. Carbon dioxide is also formed when fires release carbon stored in vegetation or when we burn wood or fossil fuels. Sulfur trioxide in the atmosphere is naturally produced by volcanic activity, but it also stems from burning fossil fuels, which have traces of sulfur, and from the process of “roasting” ores of metal sulfides in metal-refining processes. Oxides of nitrogen are formed in internal combustion engines where the high temperatures make it possible for the nitrogen and oxygen in air to chemically combine.

Acid rain is a particular problem in industrial areas where the products of combustion and smelting are released into the air without being stripped of sulfur and nitrogen oxides. In North America and Europe until the 1980s, it was responsible for the destruction of forests and freshwater lakes, when the acidity of the rain actually killed trees, damaged soil, and made lakes uninhabitable for all but the most acid-tolerant species. Acid rain also corrodes statuary and building facades that are made of marble and limestone (Figure 2). Regulations limiting the amount of sulfur and nitrogen oxides that can be released into the atmosphere by industry and automobiles have reduced the severity of acid damage to both natural and manmade environments in North America and Europe. It is now a growing problem in industrial areas of China and India.

For further information on acid rain, visit this [website](#) hosted by the US Environmental Protection Agency.

Two photos are shown. Photograph a on the left shows the upper portion of trees against a bright blue sky. The tops of several trees at the center of the photograph have bare branches and appear to be dead. Image b shows a statue of a man that appears to be from the revolutionary war era in either marble or limestone.

(a) Acid rain makes trees more susceptible to drought and insect infestation, and depletes nutrients in the soil. (b) It also corrodes statues that are carved from marble or limestone. (credit a: modification of work by Chris M Morris; credit b: modification of work by “Eden, Janine and Jim”/Flickr)

Calculation of pOH

What are the pOH and the pH of a 0.0125 M solution of potassium hydroxide, KOH?

Solution

Potassium hydroxide is a highly soluble ionic compound and completely dissociates when dissolved in dilute solution, yielding $[\text{OH}^-] = 0.0125 \text{ M}$:

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.0125) = -(-1.903) = 1.903$$

The pH can be found from the pOH:

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - \text{pOH} = 14.000 - 1.903 = 12.903$$

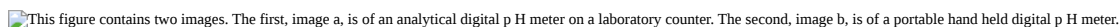
Check Your Learning

The hydronium ion concentration of vinegar is approximately $4 \times 10^{-3} \text{ M}$. What are the corresponding values of pOH and pH?

Answer:

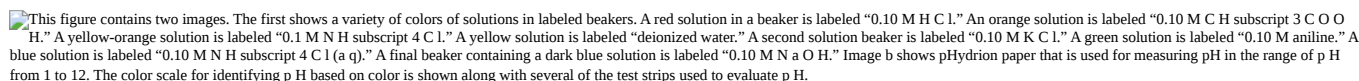
$$\text{pOH} = 11.6, \text{pH} = 2.4$$

The acidity of a solution is typically assessed experimentally by measurement of its pH. The pOH of a solution is not usually measured, as it is easily calculated from an experimentally determined pH value. The pH of a solution can be directly measured using a pH meter (Figure 3).

 This figure contains two images. The first, image a, is of an analytical digital pH meter on a laboratory counter. The second, image b, is of a portable hand-held digital pH meter.

(a) A research-grade pH meter used in a laboratory can have a resolution of 0.001 pH units, an accuracy of ± 0.002 pH units, and may cost in excess of \$1000. (b) A portable pH meter has lower resolution (0.01 pH units), lower accuracy (± 0.2 pH units), and a far lower price tag. (credit b: modification of work by Jacopo Werther)

The pH of a solution may also be visually estimated using colored indicators (Figure 4).

 This figure contains two images. The first shows a variety of colors of solutions in labeled beakers. A red solution in a beaker is labeled "0.10 M HCl." An orange solution is labeled "0.10 M CH₃COOH." A yellow-orange solution is labeled "0.1 M NH₄Cl." A yellow solution is labeled "deionized water." A second solution beaker is labeled "0.10 M KCl." A green solution is labeled "0.10 M aniline." A blue solution is labeled "0.10 M NaOH." A final beaker containing a dark blue solution is labeled "0.10 M NaOH." Image b shows pH paper that is used for measuring pH in the range of pH from 1 to 12. The color scale for identifying pH based on color is shown along with several of the test strips used to evaluate pH.

(a) A universal indicator assumes a different color in solutions of different pH values. Thus, it can be added to a solution to determine the pH of the solution. The eight vials each contain a universal indicator and 0.1 M solutions of progressively weaker acids: HCl (pH = 1), CH₃CO₂H (pH = 3), and NH₄Cl (pH = 5), deionized water, a neutral substance (pH = 7); and 0.1 M solutions of the progressively stronger bases: KCl (pH = 7), aniline, C₆H₅NH₂ (pH = 9), NH₃ (pH = 11), and NaOH (pH = 13). (b) pH paper contains a mixture of indicators that give different colors in solutions of differing pH values. (credit: modification of work by Sahar Atwa)

Key Concepts and Summary

The concentration of hydronium ion in a solution of an acid in water is greater than 1.0×10^{-7} M at 25 °C. The concentration of hydroxide ion in a solution of a base in water is greater than 1.0×10^{-7} M at 25 °C. The concentration of H₃O⁺ in a solution can be expressed as the pH of the solution; $\text{pH} = -\log[\text{H}_3\text{O}^+]$. The concentration of OH[−] can be expressed as the pOH of the solution: $\text{pOH} = -\log[\text{OH}^-]$. In pure water, pH = 7.00 and pOH = 7.00

Key Equations

- $\text{pH} = -\log[\text{H}_3\text{O}^+]$
- $\text{pOH} = -\log[\text{OH}^-]$
- $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$
- $[\text{OH}^-] = 10^{-\text{pOH}}$
- $\text{pH} + \text{pOH} = \text{p}K_w = 14.00$ at 25 °C

Chemistry End of Chapter Exercises

- Explain why a sample of pure water at 40 °C is neutral even though $[\text{H}_3\text{O}^+] = 1.7 \times 10^{-7}$ M. K_w is 2.9×10^{-14} at 40 °C.
- The ionization constant for water (K_w) is 2.9×10^{-14} at 40 °C. Calculate $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, and pOH for pure water at 40 °C.
- The ionization constant for water (K_w) is 9.311×10^{-14} at 60 °C. Calculate $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, and pOH for pure water at 60 °C.
- Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:
 - 0.200 M HCl
 - 0.0143 M NaOH
 - 3.0 M HNO₃
 - 0.0031 M Ca(OH)₂
- Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:
 - 0.000259 M HClO₄
 - 0.21 M NaOH
 - 0.000071 M Ba(OH)₂
 - 2.5 M KOH
- What are the pH and pOH of a solution of 2.0 M HCl, which ionizes completely?
- What are the hydronium and hydroxide ion concentrations in a solution whose pH is 6.52?
- Calculate the hydrogen ion concentration and the hydroxide ion concentration in wine from its pH. See Figure 1 for useful information.

Calculate the hydronium ion concentration and the hydroxide ion concentration in lime juice from its pH. See [Figure 1](#) for useful information.

9. The hydronium ion concentration in a sample of rainwater is found to be $1.7 \times 10^{-6} M$ at 25 °C. What is the concentration of hydroxide ions in the rainwater?

10. The hydroxide ion concentration in household ammonia is $3.2 \times 10^{-3} M$ at 25 °C. What is the concentration of hydronium ions in the solution?

Glossary

acidic

describes a solution in which $[H_3O^+] > [OH^-]$

basic

describes a solution in which $[H_3O^+] < [OH^-]$

neutral

describes a solution in which $[H_3O^+] = [OH^-]$

pH

logarithmic measure of the concentration of hydronium ions in a solution

pOH

logarithmic measure of the concentration of hydroxide ions in a solution

9.2: pH and pOH is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

10: Organic Chemistry

- 10.1: Drawing organic structures
- 10.2: Alkanes and isomers
- 10.3: Functional groups
- 10.4: Organic Reactions
- 10.5: Carbonyl substitution reactions

10: Organic Chemistry is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

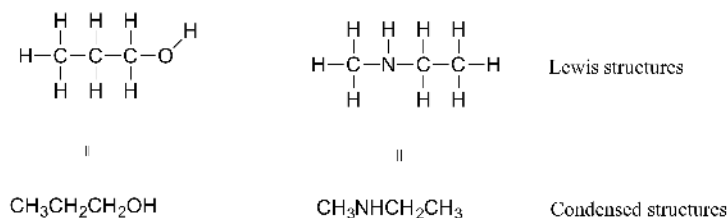
10.1: Drawing organic structures

32

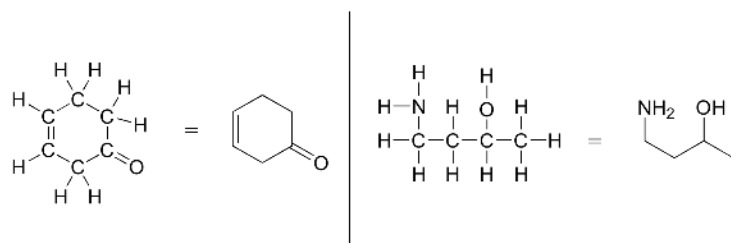
Harper College Chemistry Department

Using the 'line structure' convention

If you look ahead at the way organic compounds are drawn in following chapters, you will see that the figures are somewhat different from the Lewis structures you are used to seeing in previous chapters. In some cases, you will see **condensed structures** for smaller molecules instead of full Lewis structures, such as these:



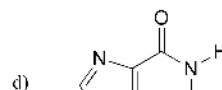
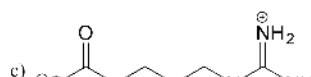
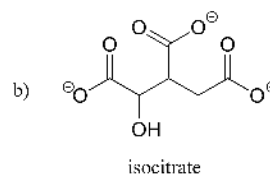
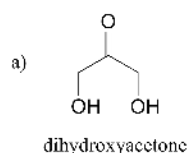
More commonly, organic and biological chemists use an abbreviated drawing convention called **line structures**. The convention is quite simple and makes it easier to draw molecules, but line structures do take a little bit of getting used to. Carbon atoms are depicted not by a capital C, but by a 'corner' between two bonds or a free end of a bond. Open-chain molecules are usually drawn out in a 'zig-zig' shape. Hydrogens attached to carbons are generally not shown: rather, like lone pairs, they are simply implied (unless a positive formal charge is shown, all carbons are assumed to have a full octet of valence electrons). Hydrogens bonded to nitrogen, oxygen, sulfur, or anything other than carbon *are* shown, but are usually drawn without showing the bond. The following examples illustrate the convention.



As you can see, the 'pared down' line structure makes it much easier to see the basic structure of the molecule and the locations where there is something other than C-C and C-H single bonds. For larger, more complex biological molecules, it becomes impractical to use full Lewis structures. Conversely, very small molecules such as ethane should be drawn with their full Lewis or condensed structures.

Sometimes, one or more carbon atoms in a line structure will be depicted with a capital C, if doing so makes an explanation easier to follow. *If you label a carbon with a C, you also must draw in the hydrogens for that carbon.*

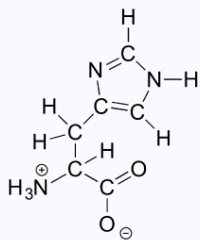
A good way to test your understanding of the line structure convention is to determine the number of hydrogen atoms in a molecule from its line structure. Do this for the structures below.



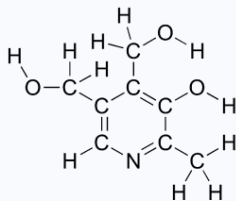
Check Your Learning

Example 1

- a) Draw a line structure for the DNA base 2-deoxycytidine (the full structure was shown earlier)
 b) Draw line structures for histidine (an amino acid) and pyridoxine (Vitamin B₆).



histidine

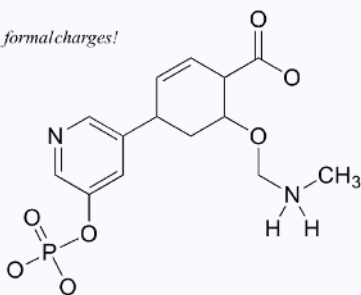


pyridoxine
(Vitamin B₆)

Example 2

Add non-zero formal charges to the structural drawing below:

missing formal charges!



Example 3

Find, anywhere in this textbook, one example of each of the common bonding patterns specified below. Check your answers with your instructor or tutor.

- carbon with one double bond, two single bonds, no lone pairs, and zero formal charge
- oxygen with two single bonds, two lone pairs, and zero formal charge
- oxygen with one double bond, two lone pairs, and zero formal charge
- nitrogen with one double bond, two single bonds, and a +1 formal charge
- oxygen with one single bond, three lone pairs, and a negative formal charge

10.1: Drawing organic structures is shared under a CC BY-SA license and was authored, remixed, and/or curated by LibreTexts.

10.2: Alkanes and isomers

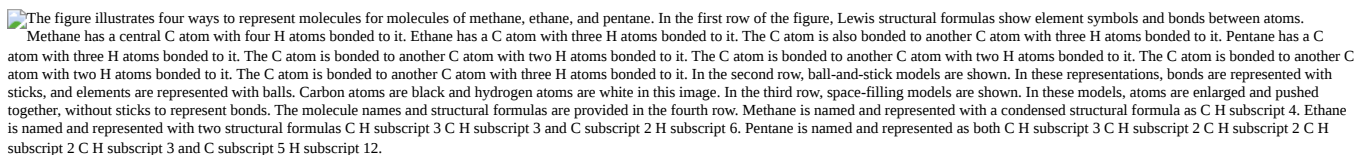
33

Harper College Chemistry Department

Alkanes

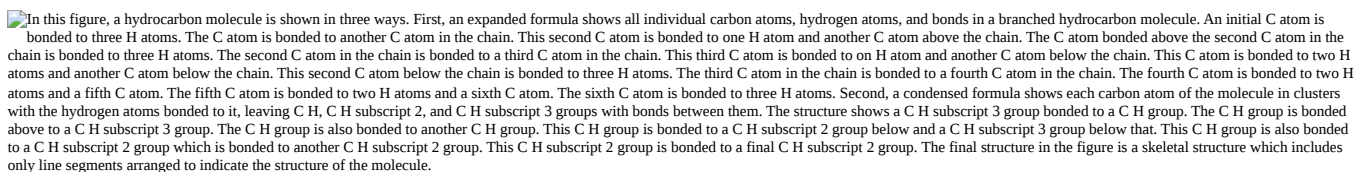
Alkanes, or saturated hydrocarbons, contain only single covalent bonds between carbon atoms. Each of the carbon atoms in an alkane has sp^3 hybrid orbitals and is bonded to four other atoms, each of which is either carbon or hydrogen. The Lewis structures and models of methane, ethane, and pentane are illustrated in Figure 1. Carbon chains are usually drawn as straight lines in Lewis structures, but one has to remember that Lewis structures are not intended to indicate the geometry of molecules. Notice that the carbon atoms in the structural models (the ball-and-stick and space-filling models) of the pentane molecule do not lie in a straight line. Because of the sp^3 hybridization, the bond angles in carbon chains are close to 109.5° , giving such chains in an alkane a zigzag shape.

The structures of alkanes and other organic molecules may also be represented in a less detailed manner by condensed structural formulas (or simply, *condensed formulas*). Instead of the usual format for chemical formulas in which each element symbol appears just once, a condensed formula is written to suggest the bonding in the molecule. These formulas have the appearance of a Lewis structure from which most or all of the bond symbols have been removed. Condensed structural formulas for ethane and pentane are shown at the bottom of Figure 1, and several additional examples are provided in the exercises at the end of this chapter.

The figure illustrates four ways to represent molecules for molecules of methane, ethane, and pentane. In the first row of the figure, Lewis structural formulas show element symbols and bonds between atoms. Methane has a central C atom with four H atoms bonded to it. Ethane has a C atom with three H atoms bonded to it. The C atom is also bonded to another C atom with three H atoms bonded to it. Pentane has a C atom with three H atoms bonded to it. The C atom is bonded to another C atom with two H atoms bonded to it. The C atom is bonded to another C atom with three H atoms bonded to it. The C atom is bonded to another C atom with two H atoms bonded to it. In the second row, ball-and-stick models are shown. In these representations, bonds are represented with sticks, and elements are represented with balls. Carbon atoms are black and hydrogen atoms are white in this image. In the third row, space-filling models are shown. In these models, atoms are enlarged and pushed together, without sticks to represent bonds. The molecule names and structural formulas are provided in the fourth row. Methane is named and represented with a condensed structural formula as CH_4 . Ethane is named and represented with two structural formulas C_2H_6 and $C_2H_5CH_3$. Pentane is named and represented as both C_5H_{12} and $CH_3CH_2CH_2CH_2CH_3$.

Pictured are the Lewis structures, ball-and-stick models, and space-filling models for molecules of methane, ethane, and pentane.

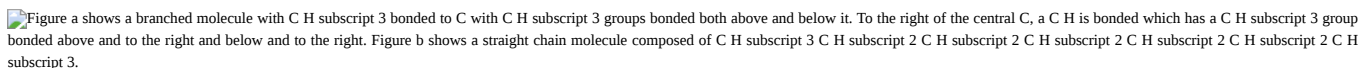
A common method used by organic chemists to simplify the drawings of larger molecules is to use a skeletal structure (also called a line-angle structure). In this type of structure, carbon atoms are not symbolized with a C, but represented by each end of a line or bend in a line. Hydrogen atoms are not drawn if they are attached to a carbon. Other atoms besides carbon and hydrogen are represented by their elemental symbols. Figure 2 shows three different ways to draw the same structure.

In this figure, a hydrocarbon molecule is shown in three ways. First, an expanded formula shows all individual carbon atoms, hydrogen atoms, and bonds in a branched hydrocarbon molecule. An initial C atom is bonded to three H atoms. The C atom is bonded to another C atom in the chain. This second C atom is bonded to one H atom and another C atom above the chain. The C atom bonded above the second C atom in the chain is bonded to three H atoms. The second C atom in the chain is bonded to a third C atom in the chain. This third C atom is bonded to one H atom and another C atom below the chain. This C atom is bonded to two H atoms and another C atom below the chain. This second C atom below the chain is bonded to three H atoms. The third C atom in the chain is bonded to a fourth C atom in the chain. The fourth C atom is bonded to two H atoms and a fifth C atom. The fifth C atom is bonded to two H atoms and a sixth C atom. The sixth C atom is bonded to three H atoms. Second, a condensed formula shows each carbon atom of the molecule in clusters with the hydrogen atoms bonded to it, leaving CH , CH_2 , and CH_3 groups with bonds between them. The structure shows a CH_3 group bonded to a CH group. The CH group is bonded above to a CH_2 group. The CH_2 group is also bonded to another CH group. This CH group is bonded to a CH_2 group below and a CH_3 group below that. This CH_3 group is also bonded to a CH_2 group which is bonded to another CH_2 group. This CH_2 group is bonded to a final CH_3 group. The final structure in the figure is a skeletal structure which includes only line segments arranged to indicate the structure of the molecule.

The same structure can be represented three different ways: an expanded formula, a condensed formula, and a skeletal structure.

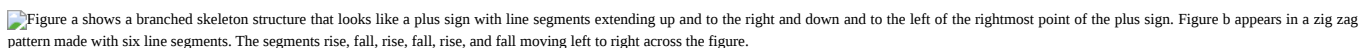
Drawing Skeletal Structures

Draw the skeletal structures for these two molecules:

Figure a shows a branched molecule with CH_3 bonded to C with CH_3 groups bonded both above and below it. To the right of the central C, a CH is bonded which has a CH_3 group bonded above and to the right and below and to the right. Figure b shows a straight chain molecule composed of CH_3 , CH_2 , CH_2 , CH_2 , CH_2 , CH_2 , CH_3 .

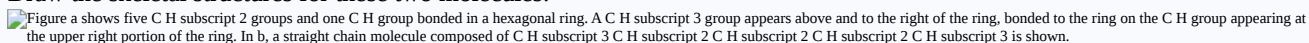
Solution

Each carbon atom is converted into the end of a line or the place where lines intersect. All hydrogen atoms attached to the carbon atoms are left out of the structure (although we still need to recognize they are there):

Figure a shows a branched skeleton structure that looks like a plus sign with line segments extending up and to the right and down and to the left of the rightmost point of the plus sign. Figure b appears in a zigzag pattern made with six line segments. The segments rise, fall, rise, fall, rise, and fall moving left to right across the figure.

Check Your Learning

Draw the skeletal structures for these two molecules:

Figure a shows five CH_2 groups and one CH group bonded in a hexagonal ring. A CH_3 group appears above and to the right of the ring, bonded to the ring on the CH group appearing at the upper right portion of the ring. In b, a straight chain molecule composed of CH_3 , CH_2 , CH_2 , CH_2 , CH_2 , CH_3 is shown.

Answer:

In a, a hexagon with a vertex at the top is shown. The vertex just to the right has a line segment attached that extends up and to the right. In b, a zig zag pattern is shown in which line segments rise, fall, rise, fall, and rise moving left to right.

Interpreting Skeletal Structures

Identify the chemical formula (empirical formula) of the molecule represented here:

This figure shows a pentagon with a vertex pointing right, from which a line segment extends that has two line segments attached at its right end, one extending up and to the right, and the other extending down and to the right.

Solution

There are eight places where lines intersect or end, meaning that there are eight carbon atoms in the molecule. Since we know that carbon atoms tend to make four bonds, each carbon atom will have the number of hydrogen atoms that are required for four bonds to obey the octet rule. This compound contains 16 hydrogen atoms for a molecular formula of C_8H_{16} .

Location of the hydrogen atoms:

In this figure a ring composed of four CH_2 groups and one CH group in a pentagonal shape is shown. From the CH group, which is at the right side of the pentagon, a CH is bonded. From this CH , a CH_3 group is bonded above and to the right and a second is bonded below and to the right.

Check Your Learning

Identify the chemical formula of the molecule represented here:

A skeleton model is shown with a zig zag pattern that rises, falls, rises, and falls again left to right through the center of the molecule. From the two risen points, line segments extend both up and down, creating four branches.

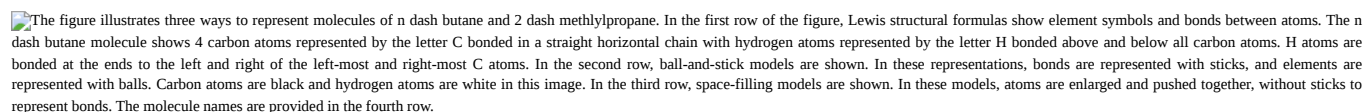
Answer:

C_9H_{20}

All alkanes are composed of carbon and hydrogen atoms and have similar bonds, structures, and formulas; noncyclic alkanes all have a formula of C_nH_{2n+2} . The number of carbon atoms present in an alkane has no limit. Greater numbers of atoms in the molecules will lead to stronger intermolecular attractions (dispersion forces) and correspondingly different physical properties of the molecules. Properties such as melting point and boiling point (Table 1) usually change smoothly and predictably as the number of carbon and hydrogen atoms in the molecules change.

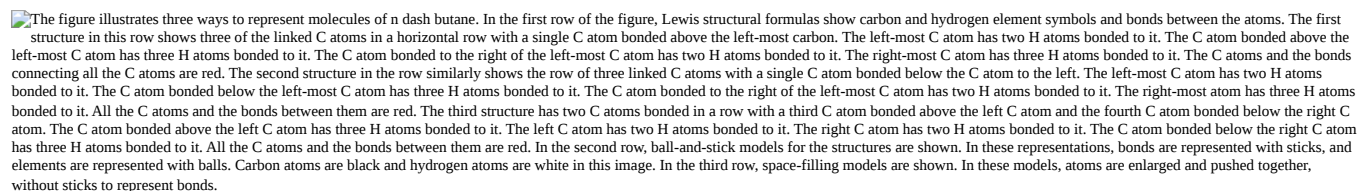
Properties of Some Alkanes ³					
Alkane	Molecular Formula	Melting Point (°C)	Boiling Point (°C)	Phase at STP ⁴	Number of Structural Isomers
methane	CH_4	-182.5	-161.5	gas	1
ethane	C_2H_6	-183.3	-88.6	gas	1
propane	C_3H_8	-187.7	-42.1	gas	1
butane	C_4H_{10}	-138.3	-0.5	gas	2
pentane	C_5H_{12}	-129.7	36.1	liquid	3
hexane	C_6H_{14}	-95.3	68.7	liquid	5
heptane	C_7H_{16}	-90.6	98.4	liquid	9
octane	C_8H_{18}	-56.8	125.7	liquid	18
nonane	C_9H_{20}	-53.6	150.8	liquid	35
decane	$C_{10}H_{22}$	-29.7	174.0	liquid	75
tetradecane	$C_{14}H_{30}$	5.9	253.5	solid	1858
octadecane	$C_{18}H_{38}$	28.2	316.1	solid	60,523

Hydrocarbons with the same formula, including alkanes, can have different structures. For example, two alkanes have the formula C_4H_{10} . They are called *n*-butane and 2-methylpropane (or isobutane), and have the following Lewis structures:

The figure illustrates three ways to represent molecules of *n*-butane and 2-methylpropane. In the first row of the figure, Lewis structural formulas show element symbols and bonds between atoms. The *n*-butane molecule shows 4 carbon atoms represented by the letter C bonded in a straight horizontal chain with hydrogen atoms represented by the letter H bonded above and below all carbon atoms. H atoms are bonded at the ends to the left and right of the left-most and right-most C atoms. In the second row, ball-and-stick models are shown. In these representations, bonds are represented with sticks, and elements are represented with balls. Carbon atoms are black and hydrogen atoms are white in this image. In the third row, space-filling models are shown. In these models, atoms are enlarged and pushed together, without sticks to represent bonds. The molecule names are provided in the fourth row.

The compounds *n*-butane and 2-methylpropane are structural isomers (the term constitutional isomers is also commonly used). Constitutional isomers have the same molecular formula but different bonding arrangements of the atoms in their molecules. The *n*-butane molecule contains an *unbranched chain*, meaning that no carbon atom is bonded to more than two other carbon atoms. We use the term *normal*, or the prefix *n*, to refer to a chain of carbon atoms without branching. The compound 2-methylpropane has a branched chain (the carbon atom in the center of the Lewis structure is bonded to three other carbon atoms)

Identifying isomers from Lewis structures is not as easy as it looks. Lewis structures that look different may actually represent the same isomers. For example, the three structures in [Figure 3](#) all represent the same molecule, *n*-butane, and hence are not different isomers. They are identical because each contains an unbranched chain of four carbon atoms.

The figure illustrates three ways to represent molecules of *n*-butane. In the first row of the figure, Lewis structural formulas show carbon and hydrogen element symbols and bonds between the atoms. The first structure in this row shows three of the linked C atoms in a horizontal row with a single C atom bonded above the left-most carbon. The left-most C atom has two H atoms bonded to it. The C atom bonded above the left-most C atom has three H atoms bonded to it. The C atom bonded to the right of the left-most C atom has two H atoms bonded to it. The right-most C atom has three H atoms bonded to it. The C atoms and the bonds connecting all the C atoms are red. The second structure in the row similarly shows the row of three linked C atoms with a single C atom bonded below the C atom to the left. The left-most C atom has two H atoms bonded to it. The C atom bonded below the left-most C atom has three H atoms bonded to it. The C atom bonded to the right of the left-most C atom has two H atoms bonded to it. The right-most atom has three H atoms bonded to it. All the C atoms and the bonds between them are red. The third structure has two C atoms bonded in a row with a third C atom bonded above the left C atom and the fourth C atom bonded below the right C atom. The C atom bonded above the left C atom has three H atoms bonded to it. The left C atom has two H atoms bonded to it. The right C atom has two H atoms bonded to it. The C atom bonded below the right C atom has three H atoms bonded to it. All the C atoms and the bonds between them are red. In the second row, ball-and-stick models for the structures are shown. In these representations, bonds are represented with sticks, and elements are represented with balls. Carbon atoms are black and hydrogen atoms are white in this image. In the third row, space-filling models are shown. In these models, atoms are enlarged and pushed together, without sticks to represent bonds.

These three representations of the structure of *n*-butane are not isomers because they all contain the same arrangement of atoms and bonds.

10.2: Alkanes and isomers is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

10.3: Functional groups

34

Harper College Chemistry Department

Adapted from LibreTexts: [Organic Chemistry With a Biological Emphasis](#) by [Tim Soderberg](#) (University of Minnesota, Morris)
1.2: Functional Groups and Organic nomenclature

Functional groups in organic compounds

Functional groups are structural units within organic compounds that are defined by specific bonding arrangements between specific atoms. The structure of capsaicin (Figure 1) incorporates several functional groups, labeled in the figure below and explained throughout this section.

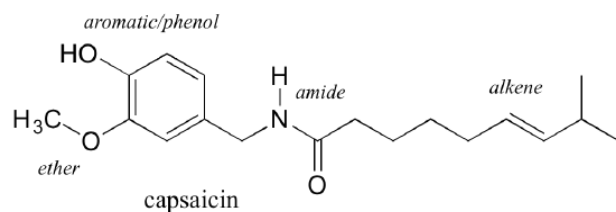
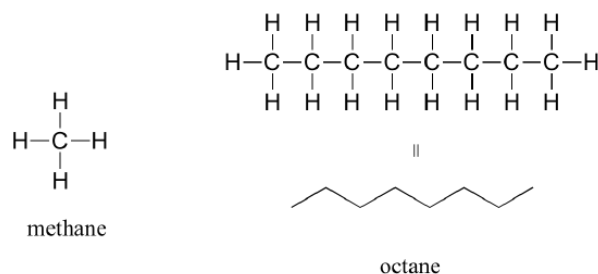


Figure 1. Structure of capsaicin with labeled functional groups.

As we progress in our study of organic chemistry, it will become extremely important to be able to quickly recognize the most common functional groups, because *they are the key structural elements that define how organic molecules react*. For now, we will only worry about drawing and recognizing each functional group, as depicted by Lewis and line structures.

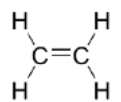
The 'default' in organic chemistry (essentially, the *lack* of any functional groups) is given the term **alkane**, characterized by single bonds between carbon and carbon, or between carbon and hydrogen. Methane, CH_4 , is the natural gas you may burn in your furnace. Octane, C_8H_{18} , is a component of gasoline.

Alkanes



Alkenes (sometimes called **olefins**) have carbon-carbon double bonds, and **alkynes** have carbon-carbon triple bonds. Ethene, the simplest alkene example, is a gas that serves as a cellular signal in fruits to stimulate ripening. If you want bananas to ripen quickly, put them in a paper bag along with an apple – the apple emits ethene gas, setting off the ripening process in the bananas. Ethyne, commonly called acetylene, is used as a fuel in welding blow torches.

Alkenes and alkynes

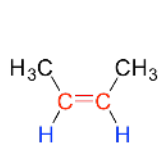


ethene
(an alkene)



ethyne
(an alkyne)

The bonding nature in alkenes is trigonal planar and in alkynes is linear. Furthermore, many alkenes can take two geometric forms: *cis* or *trans*. The *cis* and *trans* forms of a given alkene are different molecules with different physical properties because there is a very high energy barrier to rotation about a double bond. In the example below, the difference between *cis* and *trans* alkenes is readily apparent.



cis-alkene

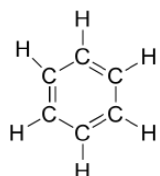


trans-alkene

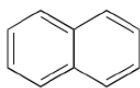
Alkanes, alkenes, and alkynes are all classified as **hydrocarbons**, because they are composed solely of carbon and hydrogen atoms. Alkanes are said to be **saturated hydrocarbons**, because the carbons are bonded to the maximum possible number of hydrogens – in other words, they are *saturated* with hydrogen atoms. The double and triple-bonded carbons in alkenes and alkynes have fewer hydrogen atoms bonded to them – they are thus referred to as **unsaturated hydrocarbons**. Hydrogen can be added to double and triple bonds, in a type of reaction called ‘hydrogenation,’ which will be discussed in a later chapter.

The **aromatic** group is exemplified by benzene (which used to be a commonly used solvent in the organic lab, but was shown to be carcinogenic later), and naphthalene, a compound with a distinctive ‘mothball’ smell. Aromatic groups are planar (flat) ring structures and are widespread in nature.

Aromatics

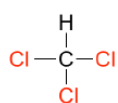


benzene

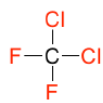


naphthalene

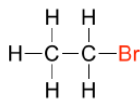
When the carbon of an alkane is bonded to one or more halogens, the group is referred to as a **alkyl halide** or **haloalkane**. Chloroform is a useful solvent in the laboratory and was one of the earlier anesthetics used in surgery. Chlorodifluoromethane was used as a refrigerant and a propellant in aerosol sprays until the late twentieth century; but its use was discontinued after it was found to have harmful effects on the ozone layer. Bromoethane is a simple alkyl halide often used in organic synthesis. Alkyl halides groups are quite rare in biomolecules.



trichloromethane
(chloroform)



dichlorodifluoromethane
(Freon-12)



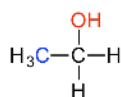
bromoethane

Haloalkanes

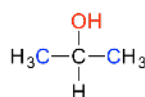
In the **alcohol** functional group, a single-bonded carbon is bonded to an OH group (the OH group, by itself, is referred to as a **hydroxyl**) not to be confused with the hydroxide polyatomic ion, OH^- . Except for methanol, all alcohols can be classified as primary, secondary, or tertiary. In a **primary alcohol**, the carbon bonded to the hydroxyl group is also bonded to only one other carbon. In a **secondary alcohol** and **tertiary alcohol**, the carbon is bonded to two or three other carbons, respectively. When the hydroxyl group is *directly* attached to an aromatic ring, the resulting group is called a **phenol**. The sulfur analog of an alcohol is called a **thiol** (from the Greek *thio*, for sulfur).



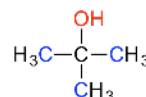
methanol



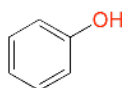
primary alcohol



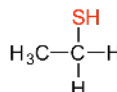
secondary alcohol



tertiary alcohol



phenol

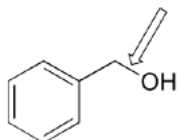


thiol

Alcohols, phenols, and thiols

Note that the definition of a phenol states that the hydroxyl oxygen must be *directly* attached to one of the carbons of the aromatic ring. The compound below, therefore, is *not* a phenol – it is a primary alcohol.

hydroxyl is not attached to carbon
in aromatic ring

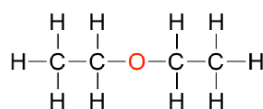


primary alcohol
(not a phenol)

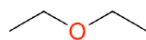
The distinction is important, as there is a significant difference in the reactivity of alcohols and phenols.

In an **ether** functional group, a central oxygen is bonded to two carbons. Below is the structure of diethyl ether, a common laboratory solvent and also one of the first compounds to be used as an anesthetic during operations. The sulfur analog of an ether is called a **thioether** or **sulfide**.

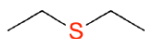
Ethers and sulfides



||



an ether



a sulfide

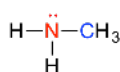
Amines are characterized by nitrogen atoms with single bonds to hydrogen and carbon. Just as there are primary, secondary and tertiary alcohols, there are primary, secondary and tertiary amines. Ammonia is a special case with no carbon atoms, only hydrogen.

One of the most important properties of amines is that they are basic and are readily protonated to form **ammonium** cations. In the case where a nitrogen has four bonds to carbon (which is somewhat unusual in biomolecules), it is called a quaternary ammonium ion.

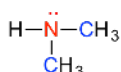
Amines



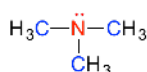
ammonia



a primary amine



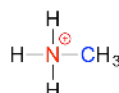
a secondary amine



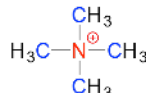
a tertiary amine



ammonium ion



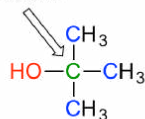
a primary ammonium ion



a quaternary ammonium ion

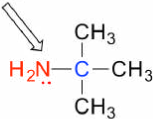
Note: Do not be confused by how the terms, ‘primary’, ‘secondary’, and ‘tertiary’ are applied to alcohols and amines – the definitions are different. In alcohols, what matters is how many other carbons the alcohol *carbon* is bonded to, while in amines, what matters is how many carbons the *nitrogen* is bonded to.

carbon is bonded
to three carbons



a tertiary alcohol

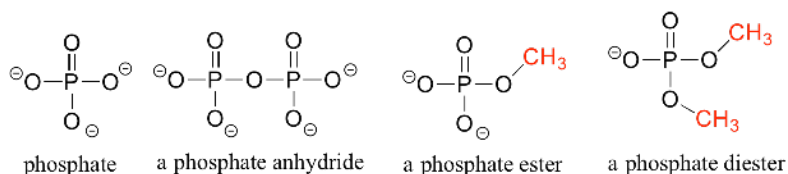
nitrogen is bonded
to one carbon



a primary amine

Phosphate and its derivative functional groups are ubiquitous in biomolecules. A phosphate linked to a single organic group is called a **phosphate ester**; when it has links to two organic groups it is called a **phosphate diester**. A linkage between two phosphates creates a **phosphate anhydride**.

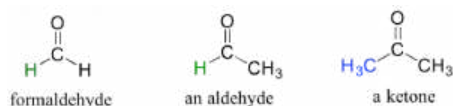
Organic phosphates



There are a number of functional groups that contain a carbon-oxygen double bond, which is commonly referred to as a **carbonyl**. Ketones and aldehydes are two closely related carbonyl-based functional groups that react in very similar ways. In a **ketone**, the carbon atom of a carbonyl is bonded to two other carbons. In an **aldehyde**, the carbonyl carbon is bonded on one side to a hydrogen, and on the other side to a carbon. The exception to this definition is formaldehyde, in which the carbonyl carbon has bonds to two hydrogens.

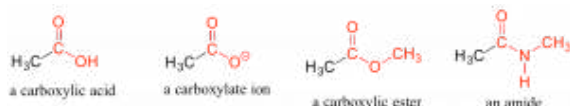
A group with a carbon-nitrogen double bond is called an **imine**, or sometimes a **Schiff base**.

Aldehydes, ketones, and imines

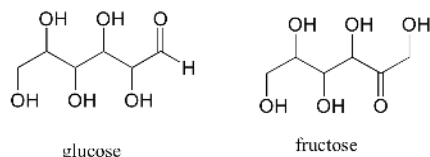


When a carbonyl carbon is bonded on one side to a carbon (or hydrogen) and on the other side to an oxygen, nitrogen, or sulfur, the functional group is considered to be one of the ‘**carboxylic acid derivatives**’, a designation that describes a set of related functional groups. The eponymous member of this family is the **carboxylic acid** functional group, in which the carbonyl is bonded to a hydroxyl group. The conjugate base of a carboxylic acid is a **carboxylate**. Other derivatives are carboxylic esters (usually just called ‘**esters**’) and **amides**, to name a few. These carboxylic acid derivatives are very common in biological molecules and metabolic pathways.

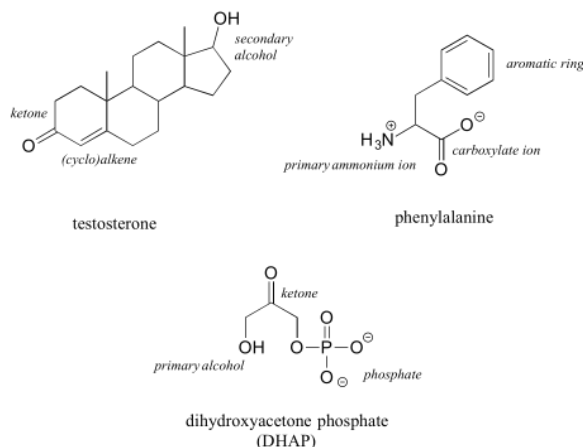
Carboxylic acid derivatives



A single compound often contains several functional groups, particularly in biological organic chemistry. The six-carbon sugar molecules glucose and fructose, for example, contain aldehyde and ketone groups, respectively, and both contain five alcohol groups (a compound with several alcohol groups is often referred to as a ‘**polyol**’).



The hormone testosterone, the amino acid phenylalanine, and the glycolysis metabolite dihydroxyacetone phosphate all contain multiple functional groups, as labeled below.



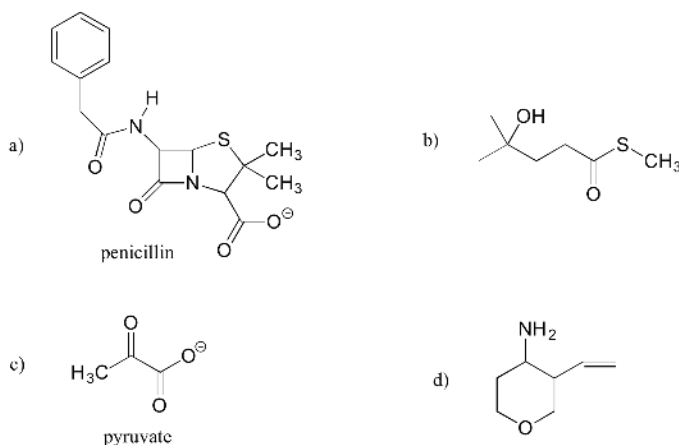
While not in any way a complete list, this section has covered most of the important functional groups that we will encounter in biological organic chemistry.

The table here summarizes the structures discussed in this chapter:

This table provides compound names, structures with functional groups in red, and examples that include formulas, structural formulas, ball-and-stick models, and names. Compound names include alkene, alkyne, alcohol, ether, aldehyde, ketone, carboxylic acid, ester, amine, and amide. Alkenes have a double bond. A formula is C_2H_4 which is named ethene. The ball-and-stick model shows two black balls forming a double bond and each is bonded to two white balls. Alkynes have a triple bond. A formula is C_2H_2 which is named ethyne. The ball-and-stick model shows two black balls with a triple bond between them each bonded to one white ball. Alcohols have an OH group. The O has two pairs of electron dots. A formula is $\text{C}_2\text{H}_5\text{OH}$ which is named ethanol. The ball-and-stick model shows two black balls and one red ball bonded to each other with a single bond. There are four white balls visible. Ethers have an O atom in the structure between two R groups. The O atom has two sets of electron dots. A formula is $\text{C}_2\text{H}_5\text{OC}_2\text{H}_5$ which is named ethyl ether. The ball-and-stick model shows two black balls bonded to a red ball which is bonded to two more black balls. All bonds are single. There are five white balls visible. Aldehydes have a C atom to which a double bonded O and an H and an R are included in the structure. The O atom has two sets of electron dots. A formula is $\text{C}_2\text{H}_4\text{O}$ which is named ethanal. The ball-and-stick model shows two black balls bonded to a red ball which is bonded to two more black balls. All bonds are single. There are three white balls visible. Ketones show a C atom to which a double bonded O is attached. The left side of the C atom is bonded to R and the right side is bonded to R prime. The O atom has two sets of electron dots. The formula is $\text{C}_4\text{H}_{10}\text{O}$ which is named methyl ethyl ketone. The ball-and-stick model shows four black balls all forming single bonds with each other. The second black ball forms a double bond with a red ball. There are five white balls visible. Carboxylic acids have a C to which a double bonded O and an OH are included in the structure. Each O atom has two sets of electron dots. A formula is $\text{C}_2\text{H}_4\text{O}_2$ which is named ethanoic or acetic acid. The ball-and-stick model shows two black balls and one red ball forming single bonds with each other. The second black ball also forms a double bond with another red ball. Three white balls are visible. Esters have a C atom which forms a double bond with an O atom and single bond with another O atom which has an attached hydrocarbon group in the structure. Each O atom has two sets of electron dots. A formula is $\text{C}_4\text{H}_8\text{O}_2$ which is named ethyl acetate. The ball-and-stick model shows two black balls, a red ball, and two more black balls forming single bonds with each other. The second black ball forms a double bond with another red ball. There are five white balls visible. Amines have an N atom in the structure to which a double bonded O and single N incorporated in a structure between two hydrocarbon groups. One hydrocarbon group is bonded to the C, the other to the N. Amides can also have a H atom bonded to the N. The O atom has two sets of electron dots, and the N atom has one set. A formula is $\text{C}_2\text{H}_5\text{NH}_2$ which is named ethylamine. The ball-and-stick model shows two black balls and one blue ball forming single bonds with each other. There are five white balls visible. Amides have a C to which a double bonded O and single N incorporated in a structure between two hydrocarbon groups. One hydrocarbon group is bonded to the C, the other to the N. Amides can also have a H atom bonded to the N. The O atom has two sets of electron dots, and the N atom has one set. A formula is $\text{C}_2\text{H}_5\text{CONH}_2$ which is named ethanamide or acetamide. The ball-and-stick model shows two black balls and one blue ball forming single bonds with each other. The second black ball forms a double bond with one red ball. There are four white balls visible.

Examples

- Identify the functional groups (other than alkanes) in the following organic compounds. State whether alcohols and amines are primary, secondary, or tertiary.



Solutions to exercises

2. Draw one example each of compounds fitting the descriptions below, using line structures. Be sure to designate the location of all non-zero formal charges. All atoms should have complete octets (phosphorus may exceed the octet rule). There are many possible correct answers for these, so be sure to check your structures with your instructor or tutor.

- a) a compound with molecular formula $\text{C}_6\text{H}_{11}\text{NO}$ that includes alkene, secondary amine, and primary alcohol functional groups
- b) an ion with molecular formula $\text{C}_3\text{H}_5\text{O}_6\text{P}^{2-}$ that includes aldehyde, secondary alcohol, and phosphate functional groups.
- c) A compound with molecular formula $\text{C}_6\text{H}_9\text{NO}$ that has an amide functional group and does *not* have an alkene group.

10.3: Functional groups is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

10.4: Organic Reactions

35

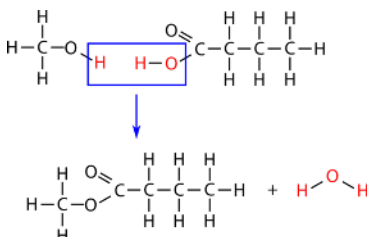
Harper College Chemistry Department

Derived from [Foundations of Introductory Chemistry-1](#) by LibreTexts and [Chemistry: Atoms First](#) by OpenStax

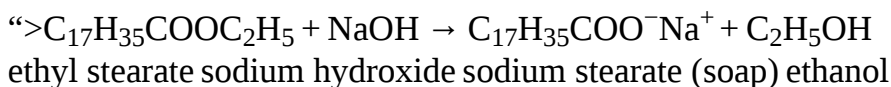
Condensation Reactions

A **condensation reaction** is a reaction in which two molecules combine to form a single molecule. A small molecule, often water, is usually removed during a condensation reaction.

An **esterification** is a condensation reaction in which an ester is formed from an alcohol and a carboxylic acid. Esterification is a subcategory of condensation reactions because a water molecule is produced in the reaction. The reaction is catalyzed by a strong acid, usually sulfuric acid. When the carboxylic acid butanoic acid is heated with an excess of methanol and a few drops of sulfuric acid, the ester methyl butanoate is produced. Methyl butanoate has the scent of pineapples. The reaction is shown below with both molecular and structural formulas.



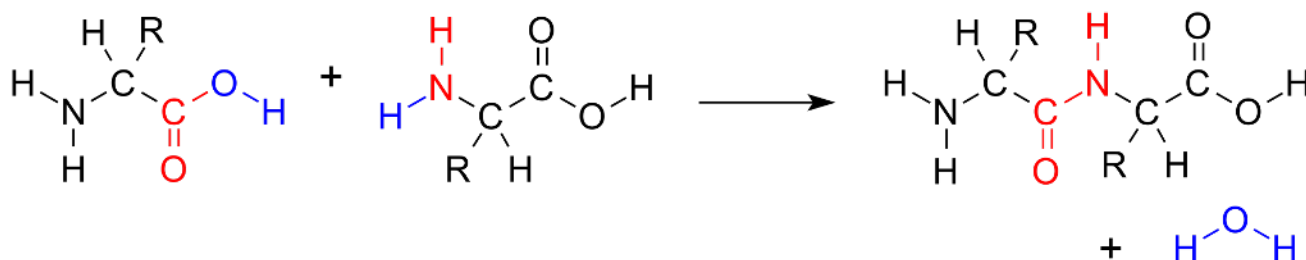
Saponification describes the alkaline hydrolysis reaction of an ester. The term saponification originally described the hydrolysis of long-chain esters called fatty acid esters to produce soap molecules, which are the salts of fatty acids. One such soap molecule is sodium stearate, formed from the hydrolysis of ethyl stearate.



The sodium hydroxide is not acting as a catalyst, but is consumed in the reaction.

Amidation reactions

Amino acids are important biological molecules that have an amine functional group on one end of the molecule and a carboxylic acid functional group on the other end. When two amino acids combine in a condensation reaction, a covalent bond forms between the amine nitrogen of one amino acid and the carboxyl carbon of the second amino acid. A molecule of water is then removed as a secondary product.



Amino acids join together to form a molecule called a dipeptide. The

“ —OH ”

from the carboxylic acid group of one amino acid combines with a hydrogen atom from the amine group of the other amino acid to produce water (blue).

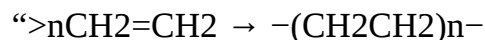
This reaction forms a molecule called a dipeptide and the carbon-nitrogen covalent bond is called a peptide bond. When repeated numerous times, a lengthy molecule called a protein is eventually produced.

Polymerization – Addition Polymers

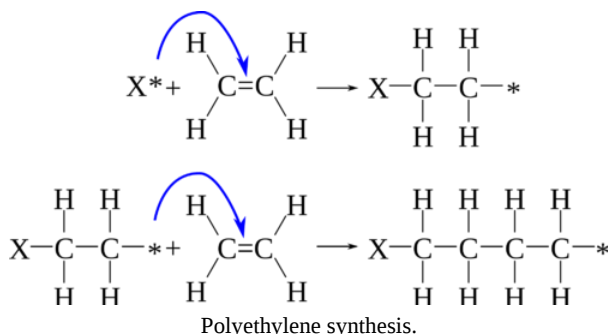
We enjoy the benefits of Styrofoam containers, but don't often think about where they end up. Styrofoam materials do not break down quickly under exposure to the elements. When buried in a landfill, styrofoam will remain intact for a long time. The good news is that there is not a lot of this pollutant found in landfills (maybe about ">0.5% by weight of the total mass of garbage). There is no good way to recycle Styrofoam at present, but someday a creative scientist may come up with one.

Polymers are very different from the other kinds of organic molecules that you have seen so far. Whereas other compounds are of relatively low molar mass, polymers are giant molecules of very high molar mass. Polymers are the primary components of all sorts of plastics and related compounds. A **polymer** is a large molecule formed of many smaller molecules covalently bonded in a repeating pattern. The small molecules which make up the polymer are called **monomers**. Polymers generally form either from an addition reaction or a condensation reaction.

An **addition polymer** is a polymer formed by chain addition reactions between monomers that contain a double bond. Molecules of ethene can polymerize with each other under the right conditions to form the polymer called polyethylene.



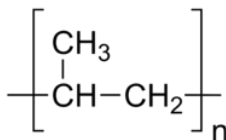
The letter ">n stands for the number of monomers that are joined in repeated fashion to make the polymer and can have a value in the hundreds or even thousands.



Polyethylene synthesis.

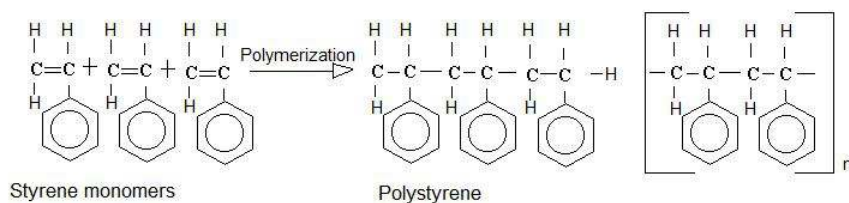
Polyethylene can have different properties depending on the length of the polymer chains and on how efficiently they pack together. Some common products made from different forms of polyethylene include plastic bottles, plastic bags, and harder plastic objects such as milk crates.

Several other kinds of unsaturated monomers can be polymerized and are components in common household products. Polypropylene is stiffer than polyethylene and is in plastic utensils and some other types of containers.



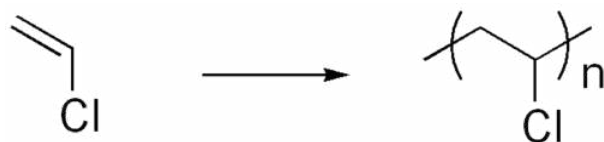
Polypropylene structure.

Polystyrene is used in insulation and in molded items such as coffee cups.



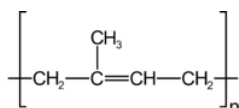
Polystyrene synthesis and structure.

Polyvinyl chloride (PVC) is extensively used for plumbing pipes.



Polyvinyl chloride.

Polyisoprene is a polymer of isoprene and is better known as rubber. It is produced naturally by rubber trees, but several variants have been developed which demonstrate improvements on the properties of natural rubber.



For your information: KEVLAR

Kevlar ([Figure 2](#)) is a synthetic polymer made from two monomers 1,4-phenylene-diamine and terephthaloyl chloride (Kevlar is a registered trademark of DuPont). Kevlar's first commercial use was as a replacement for steel in racing tires. Kevlar is typically spun into ropes or fibers. The material has a high tensile strength-to-weight ratio (it is about 5 times stronger than an equal weight of steel), making it useful for many applications from bicycle tires to sails to body armor.

A structural formula is shown for the polymer Kevlar. The structure appears inside brackets which have single dashes extending from them at the left and right ends. Outside the lower right corner of the brackets, an italicized *n* appears. The structure inside the brackets includes a C atom forming a double bond with an O atom and a bond with a benzene ring. The benzene ring forms a bond with another C atom which has a double bond with an O atom. The C atom is bonded to an N atom. The N atom is bonded to an H atom and a benzene ring. The benzene ring bonds with another N atom which is also bonded to an H atom.

This illustration shows the formula for polymeric Kevlar.

The material owes much of its strength to hydrogen bonds between polymer chains (refer back to the chapter on intermolecular interactions). These bonds form between the carbonyl group oxygen atom (which has a partial negative charge due to oxygen's electronegativity) on one monomer and the partially positively charged hydrogen atom in the N–H bond of an adjacent monomer in the polymer structure (see dashed line in [Figure 3](#)). There is additional strength derived from the interaction between the unhybridized *p* orbitals in the six-membered rings, called aromatic stacking.

This diagram shows the repeating, interlinked units that exist in Kevlar, taking on a sheet-like appearance. Dashed line segments are indicated between units. Individual units are composed of nitrogen, hydrogen, oxygen and carbon atoms. The repeating structural units include benzene rings and double bonds.

The diagram shows the polymer structure of Kevlar, with hydrogen bonds between polymer chains represented by dotted lines.

Kevlar may be best known as a component of body armor, combat helmets, and face masks. Since the 1980s, the US military has used Kevlar as a component of the PASGT (personal armor system for ground troops) helmet and vest. Kevlar is also used to protect armored fighting vehicles and aircraft carriers. Civilian applications include protective gear for emergency service personnel such as body armor for police officers and heat-resistant clothing for fire fighters. Kevlar based clothing is considerably lighter and thinner than equivalent gear made from other materials ([Figure 4](#)).

Three photos are shown. In the first, two male soldiers are shown sorting through green brown material on a table. In the second, two people are shown paddling a canoe. In the third, heavy white rope is being manipulated with a hand tool.

- (a) These soldiers are sorting through pieces of a Kevlar helmet that helped absorb a grenade blast. Kevlar is also used to make (b) canoes and (c) marine mooring lines. (credit a: modification of work by "Cla68"/Wikimedia Commons; credit b: modification of work by "OakleyOriginals"/Flickr; credit c: modification of work by Casey H. Kyhl)

In addition to its better-known uses, Kevlar is also often used in cryogenics for its very low thermal conductivity (along with its high strength). Kevlar maintains its high strength when cooled to the temperature of liquid nitrogen (–196 °C).

10.4: Organic Reactions is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

10.5: Carbonyl substitution reactions

36

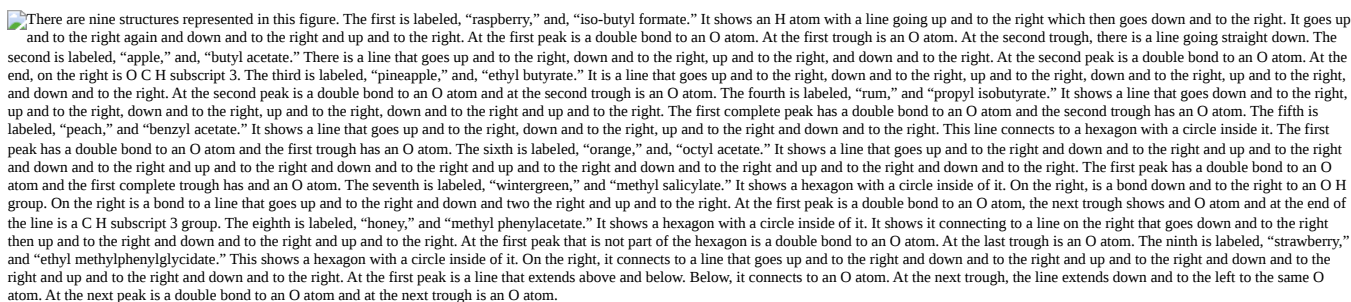
Harper College Chemistry Department

Chemistry: Atoms First

Derived from [Chemistry](#) by OpenStax

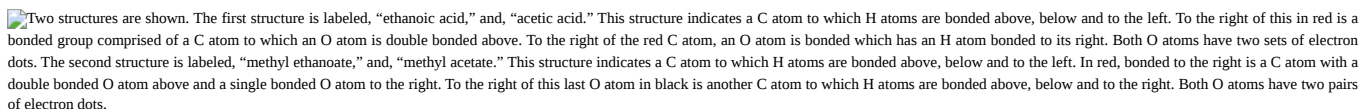
Carboxylic Acids and Esters

The odor of vinegar is caused by the presence of acetic acid, a carboxylic acid, in the vinegar. The odor of ripe bananas and many other fruits is due to the presence of esters, compounds that can be prepared by the reaction of a carboxylic acid with an alcohol. Because esters do not have hydrogen bonds between molecules, they have lower vapor pressures than the alcohols and carboxylic acids from which they are derived (see [Figure 1](#)).

There are nine structures represented in this figure. The first is labeled, "raspberry," and, "iso-butyl formate." It shows an H atom with a line going up and to the right which then goes down and to the right. It goes up and to the right again and down and to the right and up and to the right. At the first peak is a double bond to an O atom. At the first trough is an O atom. At the second trough, there is a line going straight down. The second is labeled, "apple," and, "butyl acetate." There is a line that goes up and to the right, down and to the right, up and to the right, and down and to the right. At the second peak is a double bond to an O atom. At the end, on the right is O C H subscript 3. The third is labeled, "pineapple," and, "ethyl butyrate." It is a line that goes up and to the right, down and to the right, up and to the right, down and to the right, up and to the right, and down and to the right. At the second peak is a double bond to an O atom and at the second trough is an O atom. The fourth is labeled, "rum," and, "propyl isobutyrate." It shows a line that goes down and to the right, up and to the right, down and to the right, up and to the right, down and to the right, up and to the right, and down and to the right. The first complete peak has a double bond to an O atom and the second trough has an O atom. The fifth is labeled, "peach," and, "benzyl acetate." It shows a line that goes up and to the right, down and to the right, up and to the right, down and to the right, up and to the right, and down and to the right. This line connects to a hexagon with a circle inside it. The first peak has a double bond to an O atom and the first trough has an O atom. The sixth is labeled, "orange," and, "octyl acetate." It shows a line that goes up and to the right and down and to the right and up and to the right and down and to the right and up and to the right and down and to the right and up and to the right and down and to the right. The first peak has a double bond to an O atom and the first complete trough has an O atom. The seventh is labeled, "wintergreen," and, "methyl salicylate." It shows a hexagon with a circle inside of it. On the right, is a bond down and to the right to an O H group. On the right is a bond to a line that goes up and to the right and down and to the right and up and to the right. At the first peak is a double bond to an O atom, the next trough shows an O atom and at the end of the line is a C H subscript 3 group. The eighth is labeled, "honey," and, "methyl phenylacetate." It shows a hexagon with a circle inside of it. It shows it connecting to a line on the right that goes down and to the right then up and to the right and down and to the right and up and to the right. At the first peak that is not part of the hexagon is a double bond to an O atom. At the last trough is an O atom. The ninth is labeled, "strawberry," and, "ethyl methylphenylglycidate." This shows a hexagon with a circle inside of it. On the right, it connects to a line that goes up and to the right and down and to the right and up and to the right and down and to the right and up and to the right and down and to the right. At the first peak is a line that extends above and below. Below, it connects to an O atom. At the next trough, the line extends down and to the left to the same O atom. At the next peak is a double bond to an O atom and at the next trough is an O atom.

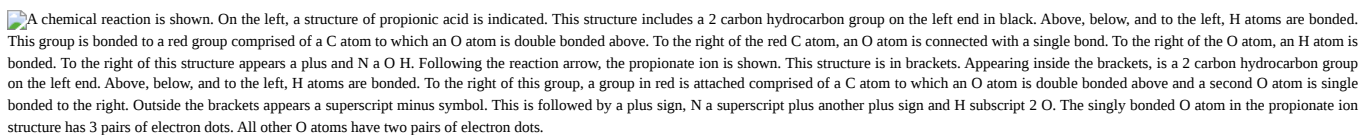
Esters are responsible for the odors associated with various plants and their fruits.

Both carboxylic acids and esters contain a carbonyl group with a second oxygen atom bonded to the carbon atom in the carbonyl group by a single bond. In a carboxylic acid, the second oxygen atom also bonds to a hydrogen atom. In an ester, the second oxygen atom bonds to another carbon atom. The names for carboxylic acids and esters include prefixes that denote the lengths of the carbon chains in the molecules and are derived following nomenclature rules similar to those for inorganic acids and salts (see these examples):

Two structures are shown. The first structure is labeled, "ethanoic acid," and, "acetic acid." This structure indicates a C atom to which H atoms are bonded above, below and to the left. To the right of this in red is a bonded group comprised of a C atom to which an O atom is double bonded above. To the right of the red C atom, an O atom is bonded which has an H atom bonded to its right. Both O atoms have two sets of electron dots. The second structure is labeled, "methyl ethanoate," and, "methyl acetate." This structure indicates a C atom to which H atoms are bonded above, below and to the left. In red, bonded to the right is a C atom with a double bonded O atom above and a single bonded O atom to the right. To the right of this last O atom in black is another C atom to which H atoms are bonded above, below and to the right. Both O atoms have two pairs of electron dots.

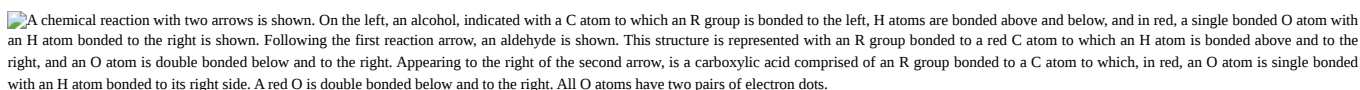
The functional groups of carboxylic acid and ester are shown in red in these formulas.

The hydrogen atom in the functional group of a carboxylic acid will react with a base to form an ionic salt:

A chemical reaction is shown. On the left, a structure of propionic acid is indicated. This structure includes a 2 carbon hydrocarbon group on the left end in black. Above, below, and to the left, H atoms are bonded. This group is bonded to a red group comprised of a C atom to which an O atom is double bonded above. To the right of the red C atom, an O atom is connected with a single bond. To the right of the O atom, an H atom is bonded. To the right of this structure appears a plus and N a O H. Following the reaction arrow, the propionate ion is shown. This structure is in brackets. Appearing inside the brackets, is a 2 carbon hydrocarbon group on the left end. Above, below, and to the left, H atoms are bonded. To the right of this group, a group in red is attached comprised of a C atom to which an O atom is double bonded above and a second O atom is single bonded to the right. Outside the brackets appears a superscript minus symbol. This is followed by a plus sign, N a superscript plus another plus sign and H subscript 2 O. The singly bonded O atom in the propionate ion structure has 3 pairs of electron dots. All other O atoms have two pairs of electron dots.

Carboxylic acids are weak acids (see the chapter on acids and bases), meaning they are not 100% ionized in water. Generally, only about 1% of the molecules of a carboxylic acid dissolved in water are ionized at any given time. The remaining molecules are undissociated in solution.

We prepare carboxylic acids by the oxidation of aldehydes or alcohols whose -OH functional group is located on the carbon atom at the end of the chain of carbon atoms in the alcohol:

A chemical reaction with two arrows is shown. On the left, an alcohol, indicated with a C atom to which an R group is bonded to the left, H atoms are bonded above and below, and in red, a single bonded O atom with an H atom bonded to the right is shown. Following the first reaction arrow, an aldehyde is shown. This structure is represented with an R group bonded to a red C atom to which an H atom is bonded above and to the right, and an O atom is double bonded below and to the right. Appearing to the right of the second arrow, is a carboxylic acid comprised of an R group bonded to a C atom to which, in red, an O atom is single bonded with an H atom bonded to its right side. A red O is double bonded below and to the right. All O atoms have two pairs of electron dots.

Esters are produced by the reaction of acids with alcohols. For example, the ester ethyl acetate, $\text{CH}_3\text{CO}_2\text{CH}_2\text{CH}_3$, is formed when acetic acid reacts with ethanol:

A chemical reaction is shown. On the left, a C H subscript 3 group bonded to a red C atom. The C atom forms a double bond with an O atom which is also in red. The C atom is also bonded to an O atom which is bonded to an H atom, also in red. A plus sign is shown, which is followed by H O C H subscript 2 C H subscript 3. The H O group is in red. Following a reaction arrow, a C H subscript 3 group is shown which is bonded to a red C atom with a double bonded O atom and a single bonded O. To the right of this single bonded O atom, a C H subscript 2 C H subscript 3 group is attached and shown in black. This structure is followed by a plus sign and H subscript 2 O. The O atoms in the first structure on the left and the structure following the reaction arrow have two pairs of electron dots.

The simplest carboxylic acid is formic acid, HCO_2H , known since 1670. Its name comes from the Latin word *formicus*, which means “ant”; it was first isolated by the distillation of red ants. It is partially responsible for the pain and irritation of ant and wasp stings and is responsible for a characteristic odor of ants that can be sometimes detected in their nests.

Acetic acid, $\text{CH}_3\text{CO}_2\text{H}$, constitutes 3–6% vinegar. Cider vinegar is produced by allowing apple juice to ferment without oxygen present. Yeast cells present in the juice carry out the fermentation reactions. The fermentation reactions change the sugar present in the juice to ethanol, then to acetic acid. Pure acetic acid has a penetrating odor and produces painful burns. It is an excellent solvent for many organic and some inorganic compounds, and it is essential in the production of cellulose acetate, a component of many synthetic fibers such as rayon.

The distinctive and attractive odors and flavors of many flowers, perfumes, and ripe fruits are due to the presence of one or more esters (Figure 2). Among the most important of the natural esters are fats (such as lard, tallow, and butter) and oils (such as linseed, cottonseed, and olive oils), which are esters of the trihydroxyl alcohol, glycerine, $\text{C}_3\text{H}_5(\text{OH})_3$, with large carboxylic acids, such as palmitic acid, $\text{CH}_3(\text{CH}_2)_{14}\text{CO}_2\text{H}$, stearic acid, $\text{CH}_3(\text{CH}_2)_{16}\text{CO}_2\text{H}$, and oleic acid, $\text{CH}_3(\text{CH}_2)_7\text{CH}=\text{CH}(\text{CH}_2)_7\text{CO}_2\text{H}$.

Oleic acid is an unsaturated acid; it contains a $\text{C}=\text{C}$ double bond. Palmitic and stearic acids are saturated acids that contain no double or triple bonds.

This is a photo of a bright red strawberry being held in a human hand.

Over 350 different volatile molecules (many members of the ester family) have been identified in strawberries. (credit: Rebecca

Siegel)**Amides** Amides are molecules that contain nitrogen atoms connected to the carbon atom of a carbonyl group. Like amines, various nomenclature rules may be used to name amides, but all use the class-specific suffix *-amide*:

This figure shows three structures. Two examples are provided. The basic structure has an H atom or R group bonded to a C atom which is double bonded to an O atom. The O atom has two sets of electron dots. The C atom is bonded to an N atom which in turn is bonded to two R groups or two H atoms. The N atom has one set of electron dots. The next structure includes acetamide, which has C H subscript 3 bonded to a C atom with a doubly bonded O atom. The second C atom is also bonded to N H subscript 2. Hexanamide has a hydrocarbon chain of length 6 involving all single bonds. The condensed structure is shown here. To the sixth C atom at the right end of the chain, an O atom is double bonded and an N H subscript 2 group is single bonded.

Amides can be produced when carboxylic acids react with amines or ammonia in a process called amidation. A water molecule is eliminated from the reaction, and the amide is formed from the remaining pieces of the carboxylic acid and the amine (note the similarity to the formation of an ester from a carboxylic acid and an alcohol discussed in the previous section):

A chemical reaction is shown between a carboxylic acid and amine to form an amide and water. Structures are shown. The carboxylic acid is shown as a C H subscript 3 group bonded to a C H subscript 2 group bonded to a C atom with a double bonded O atom above and an O H group bonded to the right. There is a plus sign. The amine is shown as an N atom with two H atoms bonded to the bottom and left sides. A C H subscript 3 group is bonded to the right side of the N atom. To the right of an arrow, an amide is shown as a C H subscript 3 group bonded to a C H subscript 2 group bonded to a C atom which is double bonded to an O atom above and an N with an H atom bonded below. The N atom is bonded to a C H subscript 3 group. The final product indicated after a plus sign is water, H subscript 2 O.

The reaction between amines and carboxylic acids to form amides is biologically important. It is through this reaction that amino acids (molecules containing both amine and carboxylic acid substituents) link together in a polymer to form proteins.

10.5: Carbonyl substitution reactions is shared under a CC BY-SA license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

11: Carbohydrates

[11.1: Carbohydrates](#)

[11.2: Fischer and Haworth projections](#)

[11: Carbohydrates](#) is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

11.1: Carbohydrates

37

Harper College Chemistry Department

Derived from [Introduction to carbohydrate structure](#) by LibreTexts and [Biology](#) by Open Stax

By the end of this section, you will be able to do the following:

- Discuss the role of carbohydrates in cells and in the extracellular materials of animals and plants
- Explain carbohydrate classifications
- List common monosaccharides, disaccharides, and polysaccharides

Most people are familiar with carbohydrates, one type of macromolecule, especially when it comes to what we eat. To lose weight, some individuals adhere to “low-carb” diets. Athletes, in contrast, often “carb-load” before important competitions to ensure that they have enough energy to compete at a high level. Carbohydrates are, in fact, an essential part of our diet. Grains, fruits, and vegetables are all natural carbohydrate sources that provide energy to the body, particularly through glucose, a simple sugar that is a component of starch and an ingredient in many staple foods. Carbohydrates also have other important functions in humans, animals, and plants.

Molecular Structures

The stoichiometric formula $(CH_2O)_n$, where n is the number of carbons in the molecule represents carbohydrates. In other words, the ratio of carbon to hydrogen to oxygen is 1:2:1 in carbohydrate molecules. The term ‘carbohydrate’, which literally means ‘hydrated carbons’, broadly refers to monosaccharides, disaccharides, oligosaccharides (shorter polymers) and polysaccharides (longer polymers).

Monosaccharides

Monosaccharides (mono- = “one”; sacchar- = “sweet”) are simple sugars, the most common of which is glucose. In monosaccharides, the number of carbons usually ranges from three to seven. Most monosaccharide names end with the suffix -ose. If the sugar has an aldehyde group (the functional group with the structure $R-CHO$), it is an aldose, and if it has a ketone group (the functional group with the structure $RC(=O)R'$), it is a ketose. Depending on the number of carbons in the sugar, they can be trioses (three carbons), pentoses (five carbons), and hexoses (six carbons). [Figure 1](#) illustrates monosaccharides.

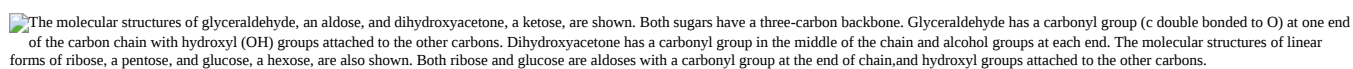


Figure 1 Scientists classify monosaccharides based on the position of their carbonyl group and the number of carbons in the backbone. Aldoses have a carbonyl group at the end of the carbon chain, and ketoses have a carbonyl group in the middle of the carbon chain. Trioses, pentoses, and hexoses have three-, five-, and six- carbon backbones, respectively.

The chemical formula for glucose is $C_6H_{12}O_6$. In humans, glucose is an important source of energy. During cellular respiration, energy releases from glucose, and that energy helps make adenosine triphosphate (ATP). Plants synthesize glucose using carbon dioxide and water, and glucose in turn provides energy requirements for the plant. Humans and other animals that feed on plants often store excess glucose as catabolized (cell breakdown of larger molecules) starch.

Galactose (part of lactose, or milk sugar) and fructose (found in sucrose, in fruit) are other common monosaccharides. Although glucose, galactose, and fructose all have the same chemical formula ($C_6H_{12}O_6$), they differ structurally and chemically (and are isomers) because of the different arrangement of functional groups around the asymmetric carbon. All these monosaccharides have more than one asymmetric carbon ([Figure 2](#)).

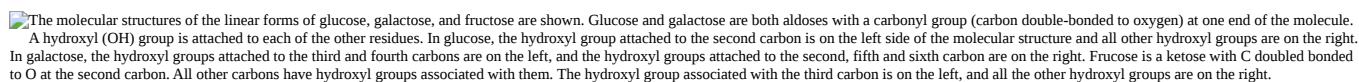


Figure 2 Glucose, galactose, and fructose are all hexoses. They are structural isomers, meaning they have the same chemical formula ($C_6H_{12}O_6$) but a different atom arrangement.

What kind of sugars are these, aldose or ketose?

Glucose, galactose, and fructose are isomeric monosaccharides (hexoses), meaning they have the same chemical formula but have slightly different structures. Glucose and galactose are aldoses, and fructose is a ketose.

Monosaccharides can exist as a linear chain or as ring-shaped molecules. In aqueous solutions they are usually in ring forms (Figure 3). Glucose in a ring form can have two different hydroxyl group arrangements (OH) around the anomeric carbon (carbon 1 that becomes asymmetric in the ring formation process). If the hydroxyl group is below carbon number 1 in the sugar, it is in the alpha (α) position, and if it is above the plane, it is in the beta (β) position.

The conversion of glucose between linear and ring forms is shown. The glucose ring has five carbons and an oxygen. In alpha glucose, the first hydroxyl group is locked in a down position, and in beta glucose, the ring is locked in an up position. Structures for ring forms of ribose and fructose are also shown. Both sugars have a ring with four carbons and an oxygen.

Figure 3 Five and six carbon monosaccharides exist in equilibrium between linear and ring forms. When the ring forms, the side chain it closes on locks into an α or β position. Fructose and ribose also form rings, although they form five-membered rings as opposed to the six-membered ring of glucose.

Disaccharides

Disaccharides (di- = “two”) form when two monosaccharides undergo a dehydration reaction (or a condensation reaction or dehydration synthesis). During this process, one monosaccharide’s hydroxyl group combines with another monosaccharide’s hydrogen, releasing a water molecule and forming a covalent bond. A covalent bond forms between a carbohydrate molecule and another molecule (in this case, between two monosaccharides). Scientists call this a glycosidic bond (Figure 4). Glycosidic bonds (or glycosidic linkages) can be an alpha or beta type. An alpha bond is formed when the OH group on the carbon-1 of the first glucose is below the ring plane, and a beta bond is formed when the OH group on the carbon-1 is above the ring plane.

The formation of sucrose from glucose and fructose is shown. In sucrose, the number one carbon of the glucose ring is connected to the number two carbon of fructose via an oxygen.

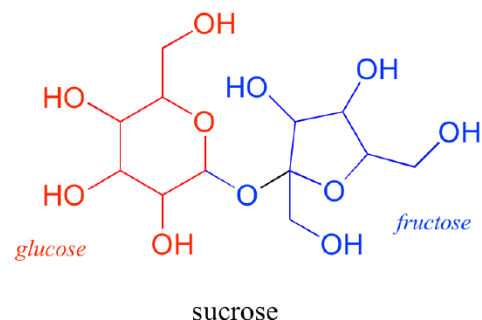


Figure 4 Sucrose forms when a glucose monomer and a fructose monomer join in a dehydration reaction to form a glycosidic bond. In the process, a water molecule is lost. By convention, the carbon atoms in a monosaccharide are numbered from the terminal carbon closest to the carbonyl group. In sucrose, a glycosidic linkage forms between carbon 1 in glucose and carbon 2 in fructose.

Common disaccharides include lactose, maltose, and sucrose (Figure 5). Lactose is a disaccharide consisting of the monomers glucose and galactose. It is naturally in milk. Maltose, or malt sugar, is a disaccharide formed by a dehydration reaction between two glucose molecules. The most common disaccharide is sucrose, or table sugar, which is comprised of glucose and fructose monomers.

The chemical structures of maltose, lactose, and sucrose are shown. Both maltose and lactose are made from two glucose monomers joined together in ring form. In maltose, the oxygen in the glycosidic bond points downward. In lactose, the oxygen in the glycosidic bond points upward. Sucrose is made from glucose and fructose monomers. The oxygen in the glycosidic bond points downward.

Figure 5 Common disaccharides include maltose (grain sugar), lactose (milk sugar), and sucrose (table sugar).

Polysaccharides

A long chain of monosaccharides linked by glycosidic bonds is a polysaccharide (poly- = “many”). The chain may be branched or unbranched, and it may contain different types of monosaccharides. The molecular weight may be 100,000 daltons or more depending on the number of joined monomers. Starch, glycogen, cellulose, and chitin are primary examples of polysaccharides.

Plants store sugars in the form of starch. In plants, an amylose and amylopectin mixture (both glucose polymers) comprise these starches. Plants are able to synthesize glucose, and they store the excess glucose, beyond their immediate energy needs, as starch in different plant parts, including roots and seeds. The starch in the seeds provides food for the embryo as it germinates and can also act as a food source for humans and animals. Enzymes break down the starch that humans consume. For example, an

amylase present in saliva catalyzes the breakdown of amylose into smaller molecules, such as maltose and glucose. The cells can then absorb the glucose.

Glucose starch comprises monomers that are joined by α 1-4 or α 1-6 glycosidic bonds. The numbers 1-4 and 1-6 refer to the carbon number of the two residues that have joined to form the bond. As [Figure 9](#) illustrates, unbranched glucose monomer chains (only α 1-4 linkages) form the starch; whereas, amylopectin is a branched polysaccharide (α 1-6 linkages at the branch points).

 The chemical structures of amylose and amylopectin are shown. Amylose consists of unbranched chains of glucose subunits, and amylopectin consists of branched chains of glucose subunits.

Figure 9 Amylose and amylopectin are two different starch forms. Unbranched glucose monomer chains comprise amylose by α 1-4 glycosidic linkages. Branched glucose monomer chains comprise amylopectin by α 1-4 and α 1-6 glycosidic linkages. Because of the way the subunits are joined, the glucose chains have a helical structure. Glycogen (not shown) is similar in structure to amylopectin but more highly branched.

Glycogen is the storage form of glucose in humans and other vertebrates and is comprised of monomers of glucose. Glycogen is the animal equivalent of starch and is a highly branched molecule usually stored in liver and muscle cells. Whenever blood glucose levels decrease, glycogen breaks down to release glucose in a process scientists call glycogenolysis.

Cellulose is the most abundant natural biopolymer. Cellulose mostly comprises a plant's cell wall. This provides the cell structural support. Wood and paper are mostly cellulosic in nature. Glucose monomers comprise cellulose that β 1-4 glycosidic bonds link ([Figure 10](#)).

 The chemical structure of cellulose is shown. Cellulose consists of unbranched chains of glucose subunits. The cellulose fibers are long, tubular, and have a slight wave shape.

Figure 10 In cellulose, glucose monomers are linked in unbranched chains by β 1-4 glycosidic linkages. Because of the way the glucose subunits are joined, every glucose monomer is flipped relative to the next one resulting in a linear, fibrous structure.

As [Figure 10](#) shows, every other glucose monomer in cellulose is flipped over, and the monomers are packed tightly as extended long chains. This gives cellulose its rigidity and high tensile strength—which is so important to plant cells. While human digestive enzymes cannot break down the β 1-4 linkage, herbivores such as cows, koalas, and buffalos are able, with the help of the specialized flora in their stomach, to digest plant material that is rich in cellulose and use it as a food source. In some of these animals, certain species of bacteria and protists reside in the rumen (part of the herbivore's digestive system) and secrete the enzyme cellulase. The appendix of grazing animals also contains bacteria that digest cellulose, giving it an important role in ruminants' digestive systems. Cellulases can break down cellulose into glucose monomers that animals use as an energy source. Termites are also able to break down cellulose because of the presence of other organisms in their bodies that secrete cellulases.

Carbohydrates serve various functions in different animals. Arthropods (insects, crustaceans, and others) have an outer skeleton, the exoskeleton, which protects their internal body parts (as we see in the bee in [Figure 11](#)). This exoskeleton is made of the biological macromolecule chitin, which is a polysaccharide-containing nitrogen. It is made of repeating N-acetyl- β -d-glucosamine units, which are a modified sugar. Chitin is also a major component of fungal cell walls. Fungi are neither animals nor plants and form a kingdom of their own in the domain Eukarya.

 A photograph shows a bee in flight, getting nectar from a flower.

Figure 11 Insects have a hard outer exoskeleton made of chitin, a type of polysaccharide. (credit: Louise Docker)

CAREER CONNECTION

Registered Dietitian

Obesity is a worldwide health concern, and many diseases such as diabetes and heart disease are becoming more prevalent because of obesity. This is one of the reasons why people increasingly seek out registered dietitians for advice. Registered dietitians help plan nutrition programs for individuals in various settings. They often work with patients in health care facilities, designing nutrition plans to treat and prevent diseases. For example, dietitians may teach a patient with diabetes how to manage blood sugar levels by eating the correct types and amounts of carbohydrates. Dietitians may also work in nursing homes, schools, and private practices.

To become a registered dietitian, one needs to earn at least a bachelor's degree in dietetics, nutrition, food technology, or a related field. In addition, registered dietitians must complete a supervised internship program and pass a national exam. Those who pursue careers in dietetics take courses in nutrition, chemistry, biochemistry, biology, microbiology, and human physiology. Dietitians must become experts in the chemistry and physiology (biological functions) of food (proteins, carbohydrates, and fats).

Benefits of Carbohydrates

Are carbohydrates good for you? Some often tell people who wish to lose weight that carbohydrates are bad and they should avoid them. Some diets completely forbid carbohydrate consumption, claiming that a low-carbohydrate diet helps people to lose weight faster. However, carbohydrates have been an important part of the human diet for thousands of years. Artifacts from ancient civilizations show the presence of wheat, rice, and corn in our ancestors' storage areas.

As part of a well balanced diet, we should supplement carbohydrates with proteins, vitamins, and fats. Calorie-wise, a gram of carbohydrate provides 4.3 Kcal. For comparison, fats provide 9 Kcal/g, a less desirable ratio. Carbohydrates contain soluble and insoluble elements. The insoluble part, fiber, is mostly cellulose. Fiber has many uses. It promotes regular bowel movement by adding bulk, and it regulates the blood glucose consumption rate. Fiber also helps to remove excess cholesterol from the body. Fiber binds to the cholesterol in the small intestine, then attaches to the cholesterol and prevents the cholesterol particles from entering the bloodstream. Cholesterol then exits the body via the feces. Fiber-rich diets also have a protective role in reducing the occurrence of colon cancer. In addition, a meal containing whole grains and vegetables gives a feeling of fullness. As an immediate source of energy, glucose breaks down during the cellular respiration process, which produces ATP, the cell's energy currency. Without consuming carbohydrates, we reduce the availability of "instant energy". Eliminating carbohydrates from the diet is not the best way to lose weight. A low-calorie diet that is rich in whole grains, fruits, vegetables, and lean meat, together with plenty of exercise and plenty of water, is the more sensible way to lose weight.

11.1: Carbohydrates is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

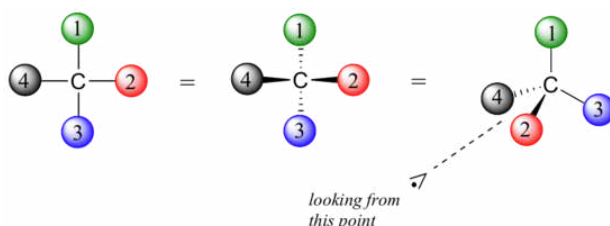
11.2: Fischer and Haworth projections

38

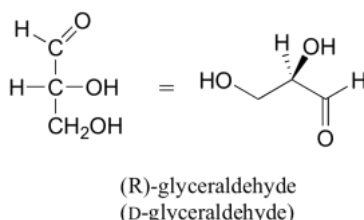
Harper College Chemistry Department

When reading the chemical and biochemical literature, you are likely to encounter several different conventions for drawing molecules in three dimensions, depending on the context of the discussion. While organic chemists prefer to use the dashed/solid wedge convention to show stereochemistry, biochemists often use drawings called **Fischer projections** and **Haworth projections** to discuss and compare the structure of sugar molecules.

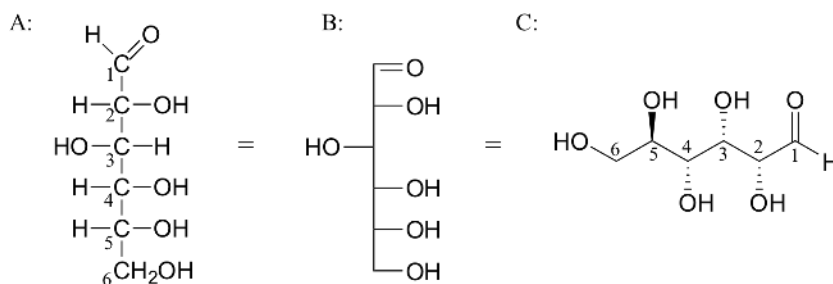
Fisher projections show sugars in their open chain form. In a Fischer projection, the carbon atoms of a sugar molecule are connected vertically by solid lines, while carbon-oxygen and carbon-hydrogen bonds are shown horizontally. Stereochemical information is conveyed by a simple rule: vertical bonds point into the plane of the page, while horizontal bonds point out of the page.



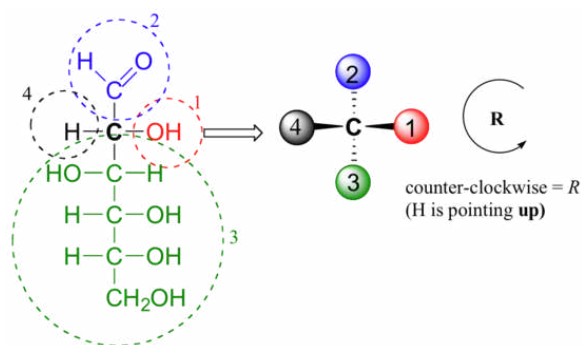
Below are two different representations of (R)-glyceraldehyde, the smallest sugar molecule (also called D -glyceraldehyde in the stereochemical nomenclature used for sugars):



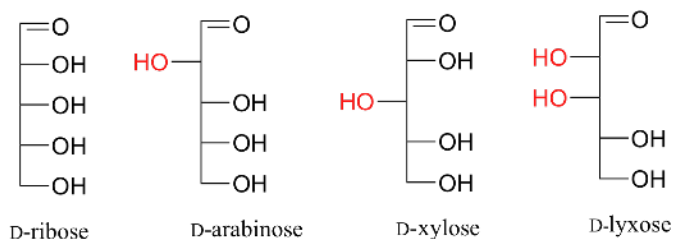
Below are three representations of the open chain form of D -glucose: in the conventional Fischer projection (A), in the “line structure” variation of the Fischer projection in which carbons and hydrogens are not shown (B), and finally in the ‘zigzag’ style (C) that is preferred by organic chemists.



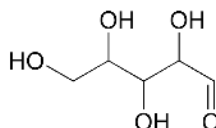
Care must be taken when ‘translating’ Fischer projection structures into ‘zigzag’ format – it is easy to get the stereochemistry wrong. Probably the best way to make a translation is to simply assign *R/S* configurations to each stereocenter, and proceed from there. When deciding whether a stereocenter in a Fischer projection is *R* or *S*, realize that the hydrogen, in a horizontal bond, is pointing *towards* you – therefore, a counterclockwise circle means *R*, and a clockwise circle means *S* (the opposite of when the hydrogen is pointing away from you).



Fischer projections are useful when looking at many different diastereomeric sugar structures, because the eye can quickly pick out stereochemical differences according to whether a hydroxyl group is on the left or right side of the structure.

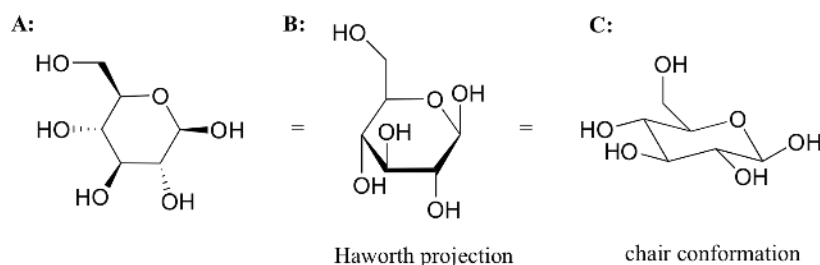


Exercise 3.24: Draw 'zigzag' structures (using the solid/dash wedge convention to show stereochemistry) for the four sugars in the figure above. Label all stereocenters *R* or *S*. To make it easy to check your answers, draw your structures using the framework below.



Solutions to exercises

While Fischer projections are used for sugars in their open-chain form, Haworth projections are often used to depict sugars in their cyclic forms. The beta diastereomer of the cyclic form of glucose is shown below in three different depictions, with the Haworth projection in the middle.



Notice that although a Haworth projection is a convenient way to show stereochemistry, it does not provide a realistic depiction of conformation. To show *both* conformation and stereochemistry, you must draw the ring in the chair form, as in structure C above.

11.2: Fischer and Haworth projections is shared under a CC BY-SA license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

12: Lipids

12.1: Lipids

12: Lipids is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

12.1: Lipids

39

Harper College Chemistry Department

Page by: OpenStax

By the end of this section, you will be able to do the following:

- Describe the four major types of lipids
- Explain the role of fats in storing energy
- Differentiate between saturated and unsaturated fatty acids
- Describe phospholipids and their role in cells
- Define the basic structure of a steroid and some steroid functions
- Explain how cholesterol helps maintain the plasma membrane's fluid nature

Lipids include a diverse group of compounds that are largely nonpolar in nature. This is because they are hydrocarbons that include mostly nonpolar carbon–carbon or carbon–hydrogen bonds. Non-polar molecules are hydrophobic (“water fearing”), or insoluble in water. Lipids perform many different functions in a cell. Cells store energy for long-term use in the form of fats. Lipids also provide insulation from the environment for plants and animals (Figure 3.12). For example, they help keep aquatic birds and mammals dry when forming a protective layer over fur or feathers because of their water-repellent hydrophobic nature. Lipids are also the building blocks of many hormones and are an important constituent of all cellular membranes. Lipids include fats, oils, waxes, phospholipids, and steroids.

 Photo shows a river otter swimming. The river otter has thick fur that is smooth and wet.

Figure 3.12 *Hydrophobic lipids in aquatic mammals' fur, such as this river otter, protect them from the elements. (credit: Ken Bosma)*

Fats and Oils

A fat molecule consists of two main moieties—glycerol and fatty acids. Glycerol is an organic compound (alcohol) with three carbons, five hydrogens, and three hydroxyl (OH) groups. Fatty acids have a long chain of hydrocarbons to which a carboxyl group is attached, hence the name “fatty acid.” The number of carbons in the fatty acid may range from 4 to 36. The most common are those containing 12–18 carbons. In a fat molecule, the fatty acids attach to each of the glycerol molecule's three carbons with an ester bond through an oxygen atom (Figure 3.13).

 The structures of glycerol, a fatty acid, and a triacylglycerol are shown. Glycerol is a chain of three carbons, with a hydroxyl (upper O upper H) group attached to each carbon. A fatty acid has an acetyl (upper C upper O upper O upper H) group attached to a long carbon chain. In triacylglycerol, a fatty acid is attached to each of glycerol's three hydroxyl groups via the carboxyl group. A water molecule is lost in the reaction so the structure of the linkage is C dash O dash C, with an oxygen double bonded to the second carbon.

Figure 3.13 *Joining three fatty acids to a glycerol backbone in a dehydration reaction forms triacylglycerol. Three water molecules release in the process.*

During this ester bond formation, three water molecules are released. The three fatty acids in the triacylglycerol may be similar or dissimilar. We also call a fat a triacylglycerol or triglyceride because of its chemical structure. Some fatty acids have common names that specify their origin. For example, palmitic acid, a saturated fatty acid, is derived from the palm tree. Arachidic acid is derived from *Arachis hypogea*, the scientific name for groundnuts or peanuts.

Fatty acids may be saturated or unsaturated. In a fatty acid chain, if there are only single bonds between neighboring carbons in the hydrocarbon chain, the fatty acid is saturated. Saturated fatty acids are saturated with hydrogen. In other words, the number of hydrogen atoms attached to the carbon skeleton is maximized. Stearic acid is an example of a saturated fatty acid (Figure 3.14).

 The structure of stearic acid is shown. This fatty acid has a hydrocarbon chain seventeen residues long attached to an acetyl group. All bonds between the carbons are single bonds.

Figure 3.14 *Stearic acid is a common saturated fatty acid.*

When the hydrocarbon chain contains a double bond, the fatty acid is unsaturated. Oleic acid is an example of an unsaturated fatty acid (Figure 3.15).

 The structure of oleic acid is shown. This fatty acid has a hydrocarbon chain seventeen residues long attached to an acetyl group. The bond between carbon eight and carbon nine is a double bond.

Figure 3.15 *Oleic acid is a common unsaturated fatty acid.*

Most unsaturated fats are liquid at room temperature. We call these oils. If there is one double bond in the molecule, then it is a monounsaturated fat (e.g., olive oil), and if there is more than one double bond, then it is a polyunsaturated fat (e.g., canola oil).

When a fatty acid has no double bonds, it is a saturated fatty acid because it is not possible to add more hydrogen to the chain's carbon atoms. A fat may contain similar or different fatty acids attached to glycerol. Long straight fatty acids with single bonds generally pack tightly and are solid at room temperature. Animal fats with stearic acid and palmitic acid (common in meat) and the fat with butyric acid (common in butter) are examples of saturated fats. Mammals store fats in specialized cells, or adipocytes, where fat globules occupy most of the cell's volume. Plants store fat or oil in many seeds and use them as a source of energy during seedling development. Unsaturated fats or oils are usually of plant origin and contain *cis* unsaturated fatty acids. *Cis* and *trans* indicate the configuration of the molecule around the double bond. If hydrogens are present on the same side of the double bond, it is a *cis* fat. If the hydrogen atoms are on two different sides, it is a *trans* fat. The *cis* double bond causes a bend or a "kink" that prevents the fatty acids from packing tightly, keeping them liquid at room temperature (Figure 3.16). Olive oil, corn oil, canola oil, and cod liver oil are examples of unsaturated fats. Unsaturated fats help to lower blood cholesterol levels; whereas, saturated fats contribute to plaque formation in the arteries. A *trans* double bond removes the 'kink' from the fatty acid chain, so the fat behaves like a saturated fat.

 A comparison of saturated and unsaturated fatty acids is shown. Stearic acid, a saturated fatty acid, has a hydrocarbon chain seventeen residues long attached to an acetyl group. Oleic acid also has a seventeen-residue hydrocarbon chain, but a double bond exists between the eighth and ninth carbon in the chain. In *cis* oleic acid, the hydrogens are on the same side of the double bond. In the *cis* oleic acid, the 2 hydrogens on the same side cause the chain to bend. In *trans* oleic acid, they are on opposite sides.

Figure 3.16 Saturated fatty acids have hydrocarbon chains connected by single bonds only. Unsaturated fatty acids have one or more double bonds. Each double bond may be in a *cis* or *trans* configuration. In the *cis* configuration, both hydrogens are on the same side of the hydrocarbon chain. In the *trans* configuration, the hydrogens are on opposite sides. A *cis* double bond causes a kink in the chain.

Trans Fats

The food industry artificially hydrogenates oils to make them semi-solid and of a consistency desirable for many processed food products. Simply speaking, hydrogen gas is bubbled through oils to solidify them. During this hydrogenation process, double bonds of the *cis*-conformation in the hydrocarbon chain may convert to double bonds in the *trans*-conformation.

Margarine, some types of peanut butter, and shortening are examples of artificially hydrogenated trans fats. Recent studies have shown that an increase in trans fats in the human diet may lead to higher levels of low-density lipoproteins (LDL), or "bad" cholesterol, which in turn may lead to plaque deposition in the arteries, resulting in heart disease. Many fast food restaurants have recently banned using trans fats, and food labels are required to display the trans fat content.

Omega Fatty Acids

Essential fatty acids are those that the human body requires but does not synthesize. Consequently, they have to be supplemented through ingestion via the diet. Omega-3 fatty acids (like those in Figure 3.17) fall into this category and are one of only two known for humans (the other is omega-6 fatty acid). These are polyunsaturated fatty acids and are omega-3 because a double bond connects the third carbon from the hydrocarbon chain's end to the fourth carbon.

 The molecular structures of alpha-linolenic acid, an omega-3 fatty acid is shown. Alpha-linolenic acid has three double bonds located eight, eleven, and fourteen residues from the acetyl group. It has a hooked shape.

Figure 3.17 Alpha-linolenic acid is an example of an omega-3 fatty acid. It has three *cis* double bonds and, as a result, a curved shape. It's also an omega-6 and an omega-9 fatty acid because it has double bonds at the 6th and 9th carbon. For clarity, the diagram does not show the carbons. Each singly bonded carbon has two hydrogens associated with it, which the diagram also does not show.

The farthest carbon away from the carboxyl group is numbered as the omega (ω) carbon, and if the double bond is between the third and fourth carbon from that end, it is an omega-3 fatty acid. These are nutritionally important because the body does not make them; omega-3 fatty acids include alpha-linoleic acid (ALA), eicosapentaenoic acid (EPA), and docosahexaenoic acid (DHA), all of which are polyunsaturated. Salmon, trout, and tuna are good sources of omega-3 fatty acids. Research indicates that omega-3 fatty acids reduce the risk of sudden death from heart attacks, lower triglycerides in the blood, decrease blood pressure, and prevent thrombosis by inhibiting blood clotting. They also reduce inflammation, and may help lower the risk of some cancers in animals.

Like carbohydrates, fats have received considerable bad publicity. It is true that eating an excess of fried foods and other "fatty" foods leads to weight gain. However, fats do have important functions. Many vitamins are fat soluble, and fats serve as a long-term

storage form of fatty acids: a source of energy. They also provide insulation for the body. Therefore, we should consume “healthy” fats in moderate amounts on a regular basis.

Waxes

Wax covers some aquatic birds’ feathers and some plants’ leaf surfaces. Because of waxes’ hydrophobic nature, they prevent water from sticking on the surface (Figure 3.18). Long fatty acid chains esterified to long-chain alcohols comprise waxes.

 The photo shows leaves on a plant; the leaves appear thick, shiny, and waxy.

Figure 3.18 Lipids comprise waxy coverings on some leaves. (credit: Roger Griffith)

Phospholipids

Phospholipids are major plasma membrane constituents that comprise cells’ outermost layer. Like fats, they are comprised of fatty acid chains attached to a glycerol or sphingosine backbone. However, instead of three fatty acids attached as in triglycerides, there are two fatty acids forming diacylglycerol, and a modified phosphate group occupies the glycerol backbone’s third carbon (Figure 3.19). A phosphate group alone attached to a diacylglycerol does not qualify as a phospholipid. It is a phosphatidate (diacylglycerol 3-phosphate), the precursor of phospholipids. An alcohol modifies the phosphate group. The modifying group contains a nitrogen and has a positive charge. Two such groups are choline and serine. When added to the phosphatidate, a phospholipid forms. Phosphatidylcholine and phosphatidylserine are two important phospholipids that are in plasma membranes.

 The molecular structure of a phospholipid is shown. It consists of two fatty acids attached to the first and second carbons in glycerol, and a phosphate group attached to the third position. The phosphate group may be further modified by addition of another molecule to one of its oxygens. Two molecules that may modify the phosphate group, choline and serine, are shown. Choline consists of a two-carbon chain with a hydroxyl group attached to one end and a nitrogen attached to the other. The nitrogen, in turn, has three methyl groups attached to it and has a charge of plus one. Serine consists of a two-carbon chain with a hydroxyl group attached to one end. An amino group and a carboxyl group are attached to the other end.

Figure 3.19 A phospholipid is a molecule with two fatty acids and a modified phosphate group attached to a glycerol backbone. Adding a charged or polar chemical group may modify the phosphate.

A phospholipid is an amphipathic molecule, meaning it has a hydrophobic and a hydrophilic part. The fatty acid chains are hydrophobic and cannot interact with water; whereas, the phosphate-containing group is hydrophilic and interacts with water (Figure 3.20).

 An illustration of a phospholipids bilayer is shown. The phospholipids bilayer consists of two layers of phospholipids. The hydrophobic tails of the phospholipids face one another while the hydrophilic head groups face outward.

Figure 3.20 The phospholipid bilayer is the major component of all cellular membranes. The hydrophilic head groups of the phospholipids face the aqueous solution. The hydrophobic tails are sequestered in the middle of the bilayer.

The head is the hydrophilic part, and the tail contains the hydrophobic fatty acids. In a membrane, a bilayer of phospholipids forms the structure’s matrix, phospholipids’ fatty acid tails face inside, away from water; whereas, the phosphate group faces the outside, aqueous side (Figure 3.20).

Phospholipids are responsible for the plasma membrane’s dynamic nature. If a drop of phospholipids is placed in water, it spontaneously forms a structure that scientists call a micelle, where the hydrophilic phosphate heads face the outside and the fatty acids face the structure’s interior.

Steroids

Unlike the phospholipids and fats that we discussed earlier, steroids have a fused ring structure. Although they do not resemble the other lipids, scientists group them with them because they are also hydrophobic and insoluble in water. All steroids have four linked carbon rings and several of them, like cholesterol, have a short tail (Figure 3.21). Many steroids also have the –OH functional group, which puts them in the alcohol classification (sterols).

 The structures of cholesterol and cortisol are shown. Each of these molecules is composed of three six-carbon rings fused to a five-carbon ring. Cholesterol has a branched hydrocarbon attached to the five-carbon ring, and a hydroxyl group attached to the terminal six-carbon ring. Cortisol has a two-carbon chain modified with a double-bonded oxygen, a hydroxyl group attached to the five-carbon ring, and an oxygen double-bonded to the terminal six-carbon ring.

Figure 3.21 Four fused hydrocarbon rings comprise steroids such as cholesterol and cortisol.

Cholesterol is the most common steroid. The liver synthesizes cholesterol and is the precursor to many steroid hormones such as testosterone and estradiol, which gonads and endocrine glands secrete. It is also the precursor to Vitamin D. Cholesterol is also the precursor of bile salts, which help emulsifying fats and their subsequent absorption by cells. Although lay people often speak negatively about cholesterol, it is necessary for the body’s proper functioning. Sterols (cholesterol in animal cells, phytosterol in plants) are components of the plasma membrane of cells and are found within the phospholipid bilayer.

For an additional perspective on lipids, explore the interactive animation “[Biomolecules: The Lipids](#)”.

12.1: [Lipids](#) is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

13: Proteins

[13.1: Properties of proteins](#)

[13.2: Amino acids](#)

[13.3: Primary structure- peptides](#)

[13.4: Secondary, tertiary, and quaternary structure of proteins](#)

[13: Proteins](#) is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

13.1: Properties of proteins

40

Harper College Chemistry Department

Proteins are one of the most abundant organic molecules in living systems and have the most diverse range of functions of all macromolecules. Proteins may be structural, regulatory, contractile, or protective. They may serve in transport, storage, or membranes; or they may be toxins or enzymes. Each cell in a living system may contain thousands of proteins, each with a unique function. Their structures, like their functions, vary greatly. They are all, however, amino acid polymers arranged in a linear sequence.

Types and Functions of Proteins

Enzymes, which living cells produce, are catalysts in biochemical reactions (like digestion) and are usually complex or conjugated proteins. Each enzyme is specific for the substrate (a reactant that binds to an enzyme) upon which it acts. The enzyme may help in breakdown, rearrangement, or synthesis reactions. We call enzymes that break down their substrates catabolic enzymes. Those that build more complex molecules from their substrates are anabolic enzymes, and enzymes that affect the rate of reaction are catalytic enzymes. Note that all enzymes increase the reaction rate and, therefore, are organic catalysts. An example of an enzyme is salivary amylase, which hydrolyzes its substrate amylose, a component of starch.

Hormones are chemical-signaling molecules, usually small proteins or steroids, secreted by endocrine cells that act to control or regulate specific physiological processes, including growth, development, metabolism, and reproduction. For example, insulin is a protein hormone that helps regulate the blood glucose level. [Table 1](#) lists the primary types and functions of proteins.

Protein Types and Functions		
Type	Examples	Functions
Digestive Enzymes	Amylase, lipase, pepsin, trypsin	Help in food by catabolizing nutrients into monomeric units
Transport	Hemoglobin, albumin	Carry substances in the blood or lymph throughout the body
Structural	Actin, tubulin, keratin	Construct different structures, like the cytoskeleton
Hormones	Insulin, thyroxine	Coordinate different body systems' activity
Defense	Immunoglobulins	Protect the body from foreign pathogens
Contractile	Actin, myosin	Effect muscle contraction
Storage	Legume storage proteins, egg white (albumin)	Provide nourishment in early embryo development and the seedling

Proteins have different shapes and molecular weights. Some proteins are globular in shape; whereas, others are fibrous in nature. For example, hemoglobin is a globular protein, but collagen, located in our skin, is a fibrous protein. Protein shape is critical to its function, and many different types of chemical bonds maintain this shape. Changes in temperature, pH, and exposure to chemicals may lead to permanent changes in the protein's shape, leading to loss of function, or denaturation. Different arrangements of the same 20 types of amino acids comprise all proteins. Two rare new amino acids were discovered recently (selenocysteine and pyrrolysine), and additional new discoveries may be added to the list.

Denaturation and Protein Folding

Each protein has its own unique sequence and shape that chemical interactions hold together. If the protein is subject to changes in temperature, pH, or exposure to chemicals, the protein structure may change, losing its shape without losing its primary sequence in what scientists call denaturation. Denaturation is often reversible because the polypeptide's primary structure is conserved in the

process if the denaturing agent is removed, allowing the protein to resume its function. Sometimes denaturation is irreversible, leading to loss of function. One example of irreversible protein denaturation is frying an egg. The albumin protein in the liquid egg white denatures when placed in a hot pan. Not all proteins denature at high temperatures. For instance, bacteria that survive in hot springs have proteins that function at temperatures close to boiling. The stomach is also very acidic, has a low pH, and denatures proteins as part of the digestion process; however, the stomach's digestive enzymes retain their activity under these conditions.

Protein folding is critical to its function. Scientists originally thought that the proteins themselves were responsible for the folding process. Only recently researchers discovered that often they receive assistance in the folding process from protein helpers, or chaperones (or chaperonins) that associate with the target protein during the folding process. They act by preventing polypeptide aggregation that comprise the complete protein structure, and they disassociate from the protein once the target protein is folded.

For an additional perspective on proteins, view [this animation](#) called “Biomolecules: The Proteins.”

13.1: Properties of proteins is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

13.2: Amino acids

41

Harper College Chemistry Department

Amino acids are the monomers that comprise proteins. Each amino acid has the same fundamental structure, which consists of a central carbon atom, or the alpha (α) carbon, bonded to an amino group (NH_2), a carboxyl group (COOH), and to a hydrogen atom. Every amino acid also has another atom or group of atoms bonded to the central atom known as the R group ([Figure 1](#)).

The molecular structure of an amino acid is given. An amino acid has an alpha carbon to which an amino group, a carboxyl group, a hydrogen, and a side chain are attached. The side chain varies for different amino acids, and is designated as the R - group.

Amino acids have a central asymmetric carbon to which an amino group, a carboxyl group, a hydrogen atom, and a side chain (R group) are attached.

Scientists use the name “amino acid” because these acids contain both amino group and carboxylic acid group in their basic structure. As we mentioned, there are 20 common amino acids present in proteins. Nine of these are essential amino acids in humans because the human body cannot produce them and we obtain them from our diet. For each amino acid, the R group (or side chain) is different ([Figure 2](#)).

The molecular structures of the twenty amino acids commonly found in proteins are given. These are divided into five categories: nonpolar aliphatic, polar uncharged, positively charged, negatively charged, and aromatic. Nonpolar aliphatic amino acids include glycine, alanine, valine, leucine, methionine, and isoleucine. Polar uncharged amino acids include serine, threonine, cysteine, proline, asparagine, and glutamine. Positively charged amino acids include lysine, arginine, and histidine. Negatively charged amino acids include aspartate and glutamate. Aromatic amino acids include phenylalanine, tyrosine, and tryptophan. For example, in the amino acid glycine, the R group is a single hydrogen; but in alanine the R group is upper C upper H subscript 3 baseline.

There are 20 common amino acids commonly found in proteins, each with a different R group (variant group) that determines its chemical nature.

Which categories of amino acid would you expect to find on a soluble protein’s surface and which would you expect to find in the interior? What distribution of amino acids would you expect to find in a protein embedded in a lipid bilayer?

The chemical nature of the side chain determines the amino acid’s nature (that is, whether it is acidic, basic, polar, or nonpolar). For example, the amino acid glycine has a hydrogen atom as the R group. Amino acids such as valine, methionine, and alanine are nonpolar or hydrophobic in nature, while amino acids such as serine, threonine, and cysteine are polar and have hydrophilic side chains. The side chains of lysine and arginine are positively charged, and therefore these amino acids are also basic amino acids. Proline has an R group that is linked to the amino group, forming a ring-like structure. Proline is an exception to the amino acid’s standard structure since its amino group is not separate from the side chain ([Figure 3](#)).

A single upper case letter or a three-letter abbreviation represents amino acids. For example, the letter V or the three-letter symbol val represent valine. Just as some fatty acids are essential to a diet, some amino acids also are necessary. These essential amino acids in humans include isoleucine, leucine, and cysteine. Essential amino acids refer to those necessary to build proteins in the body, but not those that the body produces. Which amino acids are essential varies from organism to organism.

The sequence and the number of amino acids ultimately determine the protein’s shape, size, and function. A covalent bond, or peptide bond, attaches to each amino acid, which a dehydration reaction forms. One amino acid’s carboxyl group and the incoming amino acid’s amino group combine, releasing a water molecule. The resulting bond is the peptide bond ([Figure 3.24](#)).

The formation of a peptide bond between two amino acids is shown. When the peptide bond forms, the carbon from the carbonyl group becomes attached to the nitrogen from the amino group. The upper O upper H from the carboxyl group and an upper H from the amino group form a molecule of water.

Peptide bond formation is a dehydration synthesis reaction. The carboxyl group of one amino acid is linked to the incoming amino acid’s amino group. In the process, it releases a water molecule.

The products that such linkages form are peptides. As more amino acids join to this growing chain, the resulting chain is a polypeptide. Each polypeptide has a free amino group at one end. This end the N terminal, or the amino terminal, and the other end has a free carboxyl group, also the C or carboxyl terminal. While the terms polypeptide and protein are sometimes used interchangeably, a polypeptide is technically a polymer of amino acids, whereas the term protein is used for a polypeptide or polypeptides that have combined together, often have bound non-peptide prosthetic groups, have a distinct shape, and have a unique function. After protein synthesis (translation), most proteins are modified. These are known as post-translational modifications. They may undergo cleavage, phosphorylation, or may require adding other chemical groups. Only after these modifications is the protein completely functional.

Click through the steps of protein synthesis in this [interactive tutorial](#).

EVOLUTION CONNECTION

The Evolutionary Significance of Cytochrome c

Cytochrome c is an important component of the electron transport chain, a part of cellular respiration, and it is normally located in the cellular organelle, the mitochondrion. This protein has a heme prosthetic group, and the heme's central ion alternately reduces and oxidizes during electron transfer. Because this essential protein's role in producing cellular energy is crucial, it has changed very little over millions of years. Protein sequencing has shown that there is a considerable amount of cytochrome c amino acid sequence homology among different species. In other words, we can assess evolutionary kinship by measuring the similarities or differences among various species' DNA or protein sequences.

Scientists have determined that human cytochrome c contains 104 amino acids. For each cytochrome c molecule from different organisms that scientists have sequenced to date, 37 of these amino acids appear in the same position in all cytochrome c samples. This indicates that there may have been a common ancestor. On comparing the human and chimpanzee protein sequences, scientists did not find a sequence difference. When researchers compared human and rhesus monkey sequences, the single difference was in one amino acid. In another comparison, human to yeast sequencing shows a difference in the 44th position.

13.2: Amino acids is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

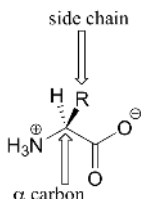
13.3: Primary structure- peptides

42

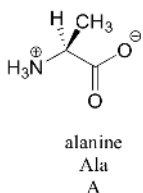
Harper College Chemistry Department

Amino acids and peptide bonds

Proteins are long chains of small ‘building block’ molecules called amino acids. There are twenty different amino acids generally found in proteins, all of which are based on the common structural framework shown below.

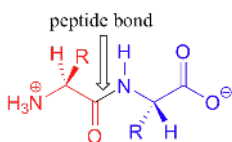


Amino acids contain an amino group on one side and a carboxylate group on the other, with a substituted carbon, referred to as the ‘alpha-carbon’, in the middle. It is the substituent on the alpha-carbon – designated ‘R’ in the general structure here, and often referred to as the amino acid ‘side chain’ – that distinguishes the 20 different amino acids. In all but one of the twenty, the alpha-carbon is a chiral center (in glycine, the ‘R’ group is simply a hydrogen, and thus the alpha-carbon is achiral). Virtually all of the amino acids found in proteins have the stereochemistry shown in the general figure above, and are referred to as L-amino acids (‘L’ stand for *levorotary*, the property of rotating plane polarized light in a counter-clockwise direction). Each amino acid is designated by both a three-letter and a single-letter abbreviation: the amino acid alanine, for example, is designated ‘Ala’ or just ‘A’.



The side chains of amino acids contain a variety of functional groups (see [table 5](#)). Some like alanine have only alkyl functionality, while others have alcohol, thiol, sulfide, or amide groups. Four amino acids have aromatic groups on their side chains: phenylalanine has a phenyl group, tyrosine a phenol, tryptophan an indole, and histidine an imidazole. Many of the side chains are ‘ionizable’, capable of acting as acids or bases depending on their protonation state. This property makes them extremely important in enzyme catalysis. Aspartate and glutamate, for example, both have carboxylate groups and are deprotonated at pH 7. Lysine and arginine have amine and imino groups, respectively, and are protonated (and positively charged) at pH 7. Histidine, with its imidazole group, is mainly protonated and positively charged at pH 7. The phenol group of tyrosine is also capable of donating a proton. We will have more to say later about the pK_a values of the different side chains in chapter 7. As you progress in your study of organic chemistry, it will become very helpful to familiarize yourself with the structures of all 20 amino acids.

Proteins are formed when amino acids are linked together by ‘peptide bonds’ to form long chains (a peptide bond is really just a specific type of amide functional group). When they are part of a protein chain, individual amino acids are often referred to as ‘**residues**’. The figure below shows a ‘dipeptide’ – two amino acids linked together by a peptide bond.



Many protein chains are several hundred amino acids long. The ‘backbone’, or ‘main chain’ of a protein refers to the repeating nitrogen – alpha-carbon – carbonyl pattern: in other words, everything but the side chain. As we have discussed previously, the

peptide bond has significant double-bond character due to resonance, and therefore there is a considerable barrier to rotation. Other sigma bonds in proteins, however, have more rotational flexibility.

13.3: Primary structure- peptides is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

13.4: Secondary, tertiary, and quaternary structure of proteins

43

Harper College Chemistry Department

Protein Structure

As we discussed earlier, a protein's shape is critical to its function. For example, an enzyme can bind to a specific substrate at an active site. If this active site is altered because of local changes or changes in overall protein structure, the enzyme may be unable to bind to the substrate. To understand how the protein gets its final shape or conformation, we need to understand the four levels of protein structure: primary, secondary, tertiary, and quaternary.

Primary Structure

Amino acids' unique sequence in a polypeptide chain is its primary structure. For example, the pancreatic hormone insulin has two polypeptide chains, A and B, and they are linked together by disulfide bonds. The N terminal amino acid of the A chain is glycine; whereas, the C terminal amino acid is asparagine (Figure 1). The amino acid sequences in the A and B chains are unique to insulin.

The amino acid sequences for the A chain and B chain of bovine insulin are shown. The A chain is 21 amino acids in length, and the B chain is 30 amino acids in length. One disulfide, or S-S bond, connects two cysteine residues in the A chain. Two other disulfide linkages connect the A chain to the B chain.

Bovine serum insulin is a protein hormone comprised of two peptide chains, A (21 amino acids long) and B (30 amino acids long). In each chain, three-letter abbreviations that represent the amino acids' names in the order they are present indicate primary structure. The amino acid cysteine (cys) has a sulfhydryl (SH) group as a side chain. Two sulfhydryl groups can react in the presence of oxygen to form a disulfide (S-S) bond. Two disulfide bonds connect the A and B chains together, and a third helps the A chain fold into the correct shape. Note that all disulfide bonds are the same length, but we have drawn them different sizes for clarity.

The gene encoding the protein ultimately determines the unique sequence for every protein. A change in nucleotide sequence of the gene's coding region may lead to adding a different amino acid to the growing polypeptide chain, causing a change in protein structure and function. In sickle cell anemia, the hemoglobin β chain (a small portion of which we show in Figure 2) has a single amino acid substitution, causing a change in protein structure and function. Specifically, valine in the β chain substitutes the amino acid glutamic. What is most remarkable to consider is that a hemoglobin molecule is comprised of two alpha and two beta chains that each consist of about 150 amino acids. The molecule, therefore, has about 600 amino acids. The structural difference between a normal hemoglobin molecule and a sickle cell molecule—which dramatically decreases life expectancy—is a single amino acid of the 600. What is even more remarkable is that three nucleotides each encode those 600 amino acids, and a single base change (point mutation), 1 in 1800 bases causes the mutation.

A portion of the hemoglobin amino acid sequence is shown. The normal hemoglobin beta chain has a glutamate at position six. The sickle cell beta chain has a valine at this position.

The beta chain of hemoglobin is 147 residues in length, yet a single amino acid substitution leads to sickle cell anemia. In normal hemoglobin, the amino acid at position seven is glutamate. In sickle cell hemoglobin, a valine replaces glutamate.

Because of this change of one amino acid in the chain, hemoglobin molecules form long fibers that distort the biconcave, or disc-shaped, red blood cells and causes them to assume a crescent or "sickle" shape, which clogs blood vessels (Figure 3). This can lead to myriad serious health problems such as breathlessness, dizziness, headaches, and abdominal pain for those affected by this disease.

This electron micrograph shows red blood cells from a patient with sickle cell anemia. Most of the cells have a normal, disk shape, but about one in five has a sickle shape. A normal blood cell is eight microns across.

In this blood smear, visualized at 535x magnification using bright field microscopy, sickle cells are crescent shaped, while normal cells are disc-shaped. (credit: modification of work by Ed Uthman; scale-bar data from Matt Russell)

Secondary Structure

The local folding of the polypeptide in some regions gives rise to the secondary structure of the protein. The most common are the α -helix and β -pleated sheet structures (Figure 4). Both structures are held in shape by hydrogen bonds. The hydrogen bonds form between the oxygen atom in the carbonyl group in one amino acid and another amino acid that is four amino acids farther along the chain.

 The illustration shows an alpha helix protein structure, which coils like a spring, and a beta-pleated sheet structure, which forms flat sheets stacked together. In an alpha-helix, hydrogen bonding occurs between the carbonyl group of one amino acid and the amino group of the amino acid that occurs four residues later. In a beta-pleated sheet, hydrogen bonding occurs between two different lengths of peptide that are antiparallel to one another.

The α -helix and β -pleated sheet are secondary structures of proteins that form because of hydrogen bonding between carbonyl and amino groups in the peptide backbone. Certain amino acids have a propensity to form an α -helix, while others have a propensity to form a β -pleated sheet.

Every helical turn in an alpha helix has 3.6 amino acid residues. The polypeptide's R groups (the variant groups) protrude out from the α -helix chain. In the β -pleated sheet, hydrogen bonding between atoms on the polypeptide chain's backbone form the "pleats". The R groups are attached to the carbons and extend above and below the pleat's folds. The pleated segments align parallel or antiparallel to each other, and hydrogen bonds form between the partially positive nitrogen atom in the amino group and the partially negative oxygen atom in the peptide backbone's carbonyl group. The α -helix and β -pleated sheet structures are in most globular and fibrous proteins and they play an important structural role.

Tertiary Structure

The polypeptide's unique three-dimensional structure is its tertiary structure (Figure 5). This structure is in part due to chemical interactions at work on the polypeptide chain. Primarily, the interactions among R groups create the protein's complex three-dimensional tertiary structure. The nature of the R groups in the amino acids involved can counteract forming the hydrogen bonds we described for standard secondary structures. For example, R groups with like charges repel each other and those with unlike charges are attracted to each other (ionic bonds). When protein folding takes place, the nonpolar amino acids' hydrophobic R groups lie in the protein's interior; whereas, the hydrophilic R groups lie on the outside. Scientists also call the former interaction types hydrophobic interactions. Interaction between cysteine side chains forms disulfide linkages in the presence of oxygen, the only covalent bond that forms during protein folding.

 This illustration shows a polypeptide backbone folded into a three-dimensional structure. Chemical interactions between amino acid side chains maintain its shape. These include an ionic bond between an amino group and a carboxyl group, hydrophobic interactions between two hydrophobic side chains, a hydrogen bond between a hydroxyl group and a carbonyl group, and a disulfide linkage.

Figure 5 A variety of chemical interactions determine the proteins' tertiary structure. These include hydrophobic interactions, ionic bonding, hydrogen bonding, and disulfide linkages.

All of these interactions, weak and strong, determine the protein's final three-dimensional shape. When a protein loses its three-dimensional shape, it may no longer be functional.

Quaternary Structure

In nature, some proteins form from several polypeptides, or subunits, and the interaction of these subunits forms the quaternary structure. Weak interactions between the subunits help to stabilize the overall structure. For example, insulin (a globular protein) has a combination of hydrogen and disulfide bonds that cause it to mostly clump into a ball shape. Insulin starts out as a single polypeptide and loses some internal sequences in the presence of post-translational modification after forming the disulfide linkages that hold the remaining chains together. Silk (a fibrous protein), however, has a β -pleated sheet structure that is the result of hydrogen bonding between different chains.

Figure 6 illustrates the four levels of protein structure (primary, secondary, tertiary, and quaternary).

 Shown are the four levels of protein structure. The primary structure is the amino acid sequence. Secondary structure is a regular folding pattern due to hydrogen bonding. Two types of secondary structure are shown: a beta pleated sheet, which is flat with regular ripples, and an alpha helix, which coils like a spring. Tertiary structure is the three-dimensional folding pattern of the protein due to interactions between amino acid side chains. Quaternary structure is the interaction of two or more polypeptide chains.

Figure 6 Observe the four levels of protein structure in these illustrations. (credit: modification of work by National Human Genome Research Institute)

Denaturation and Protein Folding

Each protein has its own unique sequence and shape that chemical interactions hold together. If the protein is subject to changes in temperature, pH, or exposure to chemicals, the protein structure may change, losing its shape without losing its primary sequence in what scientists call denaturation. Denaturation is often reversible because the polypeptide's primary structure is conserved in the process if the denaturing agent is removed, allowing the protein to resume its function. Sometimes denaturation is irreversible, leading to loss of function. One example of irreversible protein denaturation is frying an egg. The albumin protein in the liquid egg white denatures when placed in a hot pan. Not all proteins denature at high temperatures. For instance, bacteria that survive in hot springs have proteins that function at temperatures close to boiling. The stomach is also very acidic, has a low pH, and denatures proteins as part of the digestion process; however, the stomach's digestive enzymes retain their activity under these conditions.

Protein folding is critical to its function. Scientists originally thought that the proteins themselves were responsible for the folding process. Only recently researchers discovered that often they receive assistance in the folding process from protein helpers, or chaperones (or chaperonins) that associate with the target protein during the folding process. They act by preventing polypeptide aggregation that comprise the complete protein structure, and they disassociate from the protein once the target protein is folded.

LINK TO LEARNING

For an additional perspective on proteins, view [this animation](#) called “Biomolecules: The Proteins.”

13.4: Secondary, tertiary, and quaternary structure of proteins is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

CHAPTER OVERVIEW

14: Molecular Genetics

14.1: DNA structure

14: Molecular Genetics is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

14.1: DNA structure

44

Harper College Chemistry Department

Adapted from OpenStax: <https://openstax.org/books/microbiol...unction-of-dna>

LEARNING OBJECTIVES

- Describe the biochemical structure of deoxyribonucleotides
- Identify the base pairs used in the synthesis of deoxyribonucleotides
- Explain why the double helix of DNA is described as antiparallel

In previous chapters, we discussed the three classes of macromolecules: proteins, lipids and carbohydrates. In this chapter, we will discuss the genetic role of a fourth class of molecules: nucleic acids. Like other macromolecules, nucleic acids are composed of monomers, called nucleotides, which are polymerized to form large strands. Each nucleic acid strand contains certain nucleotides that appear in a certain order within the strand, called its base sequence. The base sequence of deoxyribonucleic acid (DNA) is responsible for carrying and retaining the hereditary information in a cell. In this section, we will discuss the basic structure and function of DNA.

DNA Nucleotides

The building blocks of nucleic acids are nucleotides. Nucleotides that compose DNA are called deoxyribonucleotides. The three components of a deoxyribonucleotide are a five-carbon sugar called deoxyribose, a phosphate group, and a nitrogenous base, a nitrogen-containing ring structure that is responsible for complementary base pairing between nucleic acid strands (Figure 1). The carbon atoms of the five-carbon deoxyribose are numbered 1', 2', 3', 4', and 5' (1' is read as "one prime"). A nucleoside comprises the five-carbon sugar and nitrogenous base.

a) At the center of a deoxyribonucleotide is a deoxyribose sugar. This is a pentagon shape with O at the top and H attached to the bottom right Carbon and OH attached to the bottom right Carbon. Attached to the upper left carbon is a phosphate group which consists of a Phosphate attached to 4 oxygens. Attached to the upper right carbon of the sugar is a base which consists of 1 or 2 rings that contain both carbon and nitrogen. b) A more detailed drawing of deoxyribose. This is a pentagon shaped structure with oxygen at the top corner. Moving clockwise, the upper right corner has a carbon labeled 1-prime. There is an OH attached to this carbon. The bottom right carbon is labeled 2-prime and has an H attached to it. The bottom left carbon is labeled 3-prime and has an OH group attached to it. The upper left carbon is labeled 4-prime and has CH₂OH attached. This last carbon is labeled 5-prime.

Figure 1 (a) Each deoxyribonucleotide is made up of a sugar called deoxyribose, a phosphate group, and a nitrogenous base—in this case, adenine. (b) The five carbons within deoxyribose are designated as 1', 2', 3', 4', and 5'.

The deoxyribonucleotide is named according to the nitrogenous bases (Figure 2). The nitrogenous bases adenine (A) and guanine (G) are the purines; they have a double-ring structure with a six-carbon ring fused to a five-carbon ring. The pyrimidines, cytosine (C) and thymine (T), are smaller nitrogenous bases that have only a six-carbon ring structure.

Pyrimidines have 1 ring containing both carbon and nitrogen in the ring. Cytosine and thymine are both pyrimidines. Their rings are the same but have different functional groups attached. Purines have 2 rings containing carbon and nitrogen. Adenine and Guanine are both purines but have different arrangement of atoms as part of and attached to their rings.

Figure 2 Nitrogenous bases within DNA are categorized into the two-ringed purines adenine and guanine and the single-ringed pyrimidines cytosine and thymine. Thymine is unique to DNA.

Individual nucleoside triphosphates combine with each other by covalent bonds known as 5'-3' phosphodiester bonds, or linkages whereby the phosphate group attached to the 5' carbon of the sugar of one nucleotide bonds to the hydroxyl group of the 3' carbon of the sugar of the next nucleotide. Phosphodiester bonding between nucleotides forms the sugar-phosphate backbone, the alternating sugar-phosphate structure composing the framework of a nucleic acid strand (Figure 3). During the polymerization process, deoxynucleotide triphosphates (dNTP) are used. To construct the sugar-phosphate backbone, the two terminal phosphates are released from the dNTP as a pyrophosphate. The resulting strand of nucleic acid has a free phosphate group at the 5' carbon end and a free hydroxyl group at the 3' carbon end. The two unused phosphate groups from the nucleotide triphosphate are released as pyrophosphate during phosphodiester bond formation. Pyrophosphate is subsequently hydrolyzed, releasing the energy used to drive nucleotide polymerization.

A nucleotide with a sugar in the center, the sugar is a pentagon with oxygen at the top point. Moving clockwise the carbons are numbered 1 (upper right) 2, (bottom right), 3 (bottom left), 4 (upper left) and 5 (projecting from carbon 4). Attached to carbon 1 is a base (thymine). Attached to carbon 5 is a phosphate group. Another nucleotide below has the same structure (other than that the base is C rather than T). The phosphate group attached to carbon 5 of the lower nucleotide is also attached to carbon 3 of the upper nucleotide. The lower nucleotide has an OH attached to its carbon 3. Another nucleotide has the OH group of its phosphate highlighted. A phosphodiester bond forms when water is removed from these two OH groups. This results in a bond forming between carbon 3 of the nucleotide in the chain and the phosphate group attached to carbon 5 of the new nucleotide. This is called a phosphodiester bond.

Figure 3 Phosphodiester bonds form between the phosphate group attached to the 5' carbon of one nucleotide and the hydroxyl group of the 3' carbon in the next nucleotide, bringing about polymerization of nucleotides into nucleic acid strands. Note the 5' and 3' ends of this nucleic acid strand.

CHECK YOUR UNDERSTANDING

- What is meant by the 5' and 3' ends of a nucleic acid strand?

Discovering the Double Helix

By the early 1950s, considerable evidence had accumulated indicating that DNA was the genetic material of cells and now the race was on to discover its three-dimensional structure. Around this time, Austrian biochemist Erwin Chargaff⁵ (1905–2002) examined the content of DNA in different species and discovered that adenine, thymine, guanine, and cytosine were not found in equal quantities, and that it varied from species to species, but not between individuals of the same species. He found that the amount of adenine was very close to equaling the amount of thymine, and the amount of cytosine was very close to equaling the amount of guanine, or $A = T$ and $G = C$. These relationships are also known as Chargaff's rules.

Other scientists were also actively exploring this field during the mid-20th century. In 1952, American scientist Linus Pauling (1901–1994) was the world's leading structural chemist and odds-on favorite to solve the structure of DNA. Pauling had earlier discovered the structure of protein α helices, using X-ray diffraction, and based upon X-ray diffraction images of DNA made in his laboratory, he proposed a triple-stranded model of DNA.⁶ At the same time, British researchers Rosalind Franklin (1920–1958) and her graduate student R.G. Gosling were also using X-ray diffraction to understand the structure of DNA (Figure 4). It was Franklin's scientific expertise that resulted in the production of more well-defined X-ray diffraction images of DNA that would clearly show the overall double-helix structure of DNA.

The X-ray diffraction pattern of DNA shows its helical nature. A photograph of a fuzzy spiral with fuzzy black dots forming a fuzzy figure 8.

Figure 4 The X-ray diffraction pattern of DNA shows its helical nature. (credit: National Institutes of Health)

James Watson (1928–), an American scientist, and Francis Crick (1916–2004), a British scientist, were working together in the 1950s to discover DNA's structure. They used Chargaff's rules and Franklin and Wilkins' X-ray diffraction images of DNA fibers to piece together the purine-pyrimidine pairing of the double helical DNA molecule (Figure 5). In April 1953, Watson and Crick published their model of the DNA double helix in *Nature*.⁷ The same issue additionally included papers by Wilkins and colleagues,⁸ as well as by Franklin and Gosling,⁹ each describing different aspects of the molecular structure of DNA. In 1962, James Watson, Francis Crick, and Maurice Wilkins were awarded the Nobel Prize in Physiology and Medicine. Unfortunately, by then Franklin had died, and Nobel prizes at the time were not awarded posthumously. Work continued, however, on learning about the structure of DNA. In 1973, Alexander Rich (1924–2015) and colleagues were able to analyze DNA crystals to confirm and further elucidate DNA structure.¹⁰

A photograph of a wire model in a museum.

Figure 5 In 1953, James Watson and Francis Crick built this model of the structure of DNA, shown here on display at the Science Museum in London.

CHECK YOUR UNDERSTANDING

- Which scientists are given most of the credit for describing the molecular structure of DNA?

DNA Structure

Watson and Crick proposed that DNA is made up of two strands that are twisted around each other to form a right-handed helix. The two DNA strands are antiparallel, such that the 3' end of one strand faces the 5' end of the other (Figure 6). The 3' end of each strand has a free hydroxyl group, while the 5' end of each strand has a free phosphate group. The sugar and phosphate of the polymerized nucleotides form the backbone of the structure, whereas the nitrogenous bases are stacked inside. These nitrogenous bases on the interior of the molecule interact with each other, base pairing.

Analysis of the diffraction patterns of DNA has determined that there are approximately 10 bases per turn in DNA. The asymmetrical spacing of the sugar-phosphate backbones generates major grooves (where the backbone is far apart) and minor grooves (where the backbone is close together) (Figure 6). These grooves are locations where proteins can bind to DNA. The binding of these proteins can alter the structure of DNA, regulate replication, or regulate transcription of DNA into RNA.

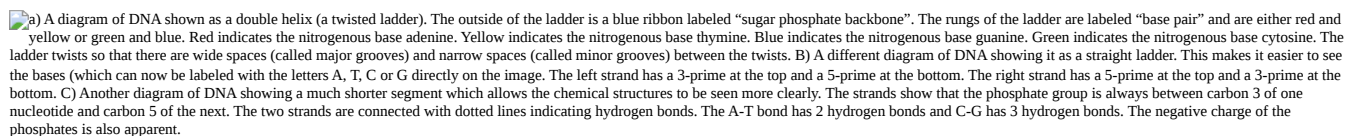
 (a) A diagram of DNA shown as a double helix (a twisted ladder). The outside of the ladder is a blue ribbon labeled “sugar phosphate backbone”. The rungs of the ladder are labeled “base pair” and are either red and yellow or green and blue. Red indicates the nitrogenous base adenine. Yellow indicates the nitrogenous base thymine. Blue indicates the nitrogenous base guanine. Green indicates the nitrogenous base cytosine. The ladder twists so that there are wide spaces (called major grooves) and narrow spaces (called minor grooves) between the twists. B) A different diagram of DNA showing it as a straight ladder. This makes it easier to see the bases (which can now be labeled with the letters A, T, C or G directly on the image). The left strand has a 3-prime at the top and a 5-prime at the bottom. The right strand has a 5-prime at the top and a 3-prime at the bottom. C) Another diagram of DNA showing a much shorter segment which allows the chemical structures to be seen more clearly. The strands show that the phosphate group is always between carbon 3 of one nucleotide and carbon 5 of the next. The two strands are connected with dotted lines indicating hydrogen bonds. The A-T bond has 2 hydrogen bonds and C-G has 3 hydrogen bonds. The negative charge of the phosphates is also apparent.

Figure 6 Watson and Crick proposed the double helix model for DNA. (a) The sugar-phosphate backbones are on the outside of the double helix and purines and pyrimidines form the “rungs” of the DNA helix ladder. (b) The two DNA strands are antiparallel to each other. (c) The direction of each strand is identified by numbering the carbons (1 through 5) in each sugar molecule. The 5' end is the one where carbon #5 is not bound to another nucleotide; the 3' end is the one where carbon #3 is not bound to another nucleotide.

Base pairing takes place between a purine and pyrimidine. In DNA, adenine (A) and thymine (T) are complementary base pairs, and cytosine (C) and guanine (G) are also complementary base pairs, explaining Chargaff's rules (Figure 7). The base pairs are stabilized by hydrogen bonds; adenine and thymine form two hydrogen bonds between them, whereas cytosine and guanine form three hydrogen bonds between them.

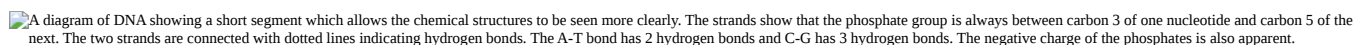
 A diagram of DNA showing a short segment which allows the chemical structures to be seen more clearly. The strands show that the phosphate group is always between carbon 3 of one nucleotide and carbon 5 of the next. The two strands are connected with dotted lines indicating hydrogen bonds. The A-T bond has 2 hydrogen bonds and C-G has 3 hydrogen bonds. The negative charge of the phosphates is also apparent.

Figure 7 Hydrogen bonds form between complementary nitrogenous bases on the interior of DNA.

In the laboratory, exposing the two DNA strands of the double helix to high temperatures or to certain chemicals can break the hydrogen bonds between complementary bases, thus separating the strands into two separate single strands of DNA (single-stranded DNA [ssDNA]). This process is called DNA denaturation and is analogous to protein denaturation, as described in Proteins. The ssDNA strands can also be put back together as double-stranded DNA (dsDNA), through reannealing or renaturing by cooling or removing the chemical denaturants, allowing these hydrogen bonds to reform. The ability to artificially manipulate DNA in this way is the basis for several important techniques in biotechnology (Figure 8). Because of the additional hydrogen bonding between the C = G base pair, DNA with a high GC content is more difficult to denature than DNA with a lower GC content.

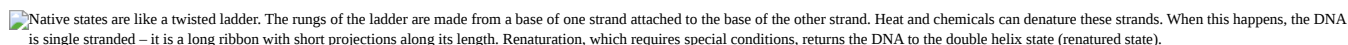
 Native states are like a twisted ladder. The rungs of the ladder are made from a base of one strand attached to the base of the other strand. Heat and chemicals can denature these strands. When this happens, the DNA is single stranded – it is a long ribbon with short projections along its length. Renaturation, which requires special conditions, returns the DNA to the double helix state (renatured state).

Figure 8 In the laboratory, the double helix can be denatured to single-stranded DNA through exposure to heat or chemicals, and then renatured through cooling or removal of chemical denaturants to allow the DNA strands to reanneal. (credit: modification of work by Hernández-Lemus E, Nicasio-Collazo LA, Castañeda-Priego R)

LINK TO LEARNING

View an [animation](#) on DNA structure from the DNA Learning Center to learn more.

CHECK YOUR UNDERSTANDING

- What are the two complementary base pairs of DNA and how are they bonded together?

DNA Function

DNA stores the information needed to build and control the cell. The transmission of this information from mother to daughter cells is called vertical gene transfer and it occurs through the process of DNA replication. DNA is replicated when a cell makes a duplicate copy of its DNA, then the cell divides, resulting in the correct distribution of one DNA copy to each resulting cell. DNA can also be enzymatically degraded and used as a source of nucleosides and nucleotides for the cell. Unlike other macromolecules, DNA does not serve a structural role in cells.

CHECK YOUR UNDERSTANDING

- How does DNA transmit genetic information to offspring?

EYE ON ETHICS

Paving the Way for Women in Science and Health Professions

Historically, women have been underrepresented in the sciences and in medicine, and often their pioneering contributions have gone relatively unnoticed. For example, although Rosalind Franklin performed the X-ray diffraction studies demonstrating the double helical structure of DNA, it is Watson and Crick who became famous for this discovery, building on her data. There still remains great controversy over whether their acquisition of her data was appropriate and whether personality conflicts and gender bias contributed to the delayed recognition of her significant contributions. Similarly, Barbara McClintock did pioneering work in

maize (corn) genetics from the 1930s through 1950s, discovering transposons (jumping genes), but she was not recognized until much later, receiving a Nobel Prize in Physiology or Medicine in 1983 (Figure 10.19).

Today, women still remain underrepresented in many fields of science and medicine. While more than half of the undergraduate degrees in science are awarded to women, only 46% of doctoral degrees in science are awarded to women. In academia, the number of women at each level of career advancement continues to decrease, with women holding less than one-third of the positions of Ph.D.-level scientists in tenure-track positions, and less than one-quarter of the full professorships at 4-year colleges and universities.¹¹ Even in the health professions, like nearly all other fields, women are often underrepresented in many medical careers and earn significantly less than their male counterparts, as shown in a 2013 study published by the *Journal of the American Medical Association*.¹²

Why do such disparities continue to exist and how do we break these cycles? The situation is complex and likely results from the combination of various factors, including how society conditions the behaviors of girls from a young age and supports their interests, both professionally and personally. Some have suggested that women do not belong in the laboratory, including Nobel Prize winner Tim Hunt, whose 2015 public comments suggesting that women are too emotional for science¹³ were met with widespread condemnation.

Perhaps girls should be supported more from a young age in the areas of science and math (Figure 10.19). Science, technology, engineering, and mathematics (STEM) programs sponsored by the American Association of University Women (AAUW)¹⁴ and National Aeronautics and Space Administration (NASA)¹⁵ are excellent examples of programs that offer such support. Contributions by women in science should be made known more widely to the public, and marketing targeted to young girls should include more images of historically and professionally successful female scientists and medical professionals, encouraging all bright young minds, including girls and women, to pursue careers in science and medicine.

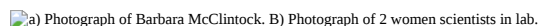
 a) Photograph of Barbara McClintock. b) Photograph of 2 women scientists in lab.

Figure 10.19 (a) Barbara McClintock's work on maize genetics in the 1930s through 1950s resulted in the discovery of transposons, but its significance was not recognized at the time. (b) Efforts to appropriately mentor and to provide continued societal support for women in science and medicine may someday help alleviate some of the issues preventing gender equality at all levels in science and medicine. (credit a: modification of work by Smithsonian Institution; credit b: modification of work by Haynie SL, Hinkle AS, Jones NL, Martin CA, Olsiewski PJ, Roberts MF)

CLINICAL FOCUS

Part 2

Based upon his symptoms, Alex's physician suspects that he is suffering from a foodborne illness that he acquired during his travels. Possibilities include bacterial infection (e.g., enterotoxigenic *E. coli*, *Vibrio cholerae*, *Campylobacter jejuni*, *Salmonella*), viral infection (rotavirus or norovirus), or protozoan infection (*Giardia lamblia*, *Cryptosporidium parvum*, or *Entamoeba histolytica*).

His physician orders a stool sample to identify possible causative agents (e.g., bacteria, cysts) and to look for the presence of blood because certain types of infectious agents (like *C. jejuni*, *Salmonella*, and *E. histolytica*) are associated with the production of bloody stools.

Alex's stool sample showed neither blood nor cysts. Following analysis of his stool sample and based upon his recent travel history, the hospital physician suspected that Alex was suffering from traveler's diarrhea caused by enterotoxigenic *E. coli* (ETEC), the causative agent of most traveler's diarrhea. To verify the diagnosis and rule out other possibilities, Alex's physician ordered a diagnostic lab test of his stool sample to look for DNA sequences encoding specific virulence factors of ETEC. The physician instructed Alex to drink lots of fluids to replace what he was losing and discharged him from the hospital.

ETEC produces several plasmid-encoded virulence factors that make it pathogenic compared with typical *E. coli*. These include the secreted toxins heat-labile enterotoxin (LT) and heat-stable enterotoxin (ST), as well as colonization factor (CF). Both LT and ST cause the excretion of chloride ions from intestinal cells to the intestinal lumen, causing a consequent loss of water from intestinal cells, resulting in diarrhea. CF encodes a bacterial protein that aids in allowing the bacterium to adhere to the lining of the small intestine.

- Why did Alex's physician use genetic analysis instead of either isolation of bacteria from the stool sample or direct Gram stain of the stool sample alone?

Jump to the [next](#) Clinical Focus box. Go back to the [previous](#) Clinical Focus box.

Footnotes

- [5](#) N. Kresge et al. "Chargaff's Rules: The Work of Erwin Chargaff." *Journal of Biological Chemistry* 280 (2005):e21.
- [6](#) L. Pauling, "A Proposed Structure for the Nucleic Acids." *Proceedings of the National Academy of Science of the United States of America* 39 no. 2 (1953):84–97.
- [7](#) J.D. Watson, F.H.C. Crick. "A Structure for Deoxyribose Nucleic Acid." *Nature* 171 no. 4356 (1953):737–738.
- [8](#) M.H.F. Wilkins et al. "Molecular Structure of Deoxypentose Nucleic Acids." *Nature* 171 no. 4356 (1953):738–740.
- [9](#) R. Franklin, R.G. Gosling. "Molecular Configuration in Sodium Thymonucleate." *Nature* 171 no. 4356 (1953):740–741.
- [10](#) R.O. Day et al. "A Crystalline Fragment of the Double Helix: The Structure of the Dinucleoside Phosphate Guanylyl-3',5'-Cytidine." *Proceedings of the National Academy of Sciences of the United States of America* 70 no. 3 (1973):849–853.
- [11](#) N.H. Wolfinger "For Female Scientists, There's No Good Time to Have Children." *The Atlantic* July 29, 2013. <http://www.theatlantic.com/sexes/arc...ildren/278165/>.
- [12](#) S.A. Seabury et al. "Trends in the Earnings of Male and Female Health Care Professionals in the United States, 1987 to 2010." *Journal of the American Medical Association Internal Medicine* 173 no. 18 (2013):1748–1750.
- [13](#) E. Chung. "Tim Hunt, Sexism and Science: The Real 'Trouble With Girls' in Labs." *CBC News Technology and Science*, June 12, 2015. <http://www.cbc.ca/news/technology/ti...labs-1.3110133>. Accessed 8/4/2016.
- [14](#) American Association of University Women. "Building a STEM Pipeline for Girls and Women." <http://www.aauw.org/what-we-do/stem-education/>. Accessed June 10, 2016.
- [15](#) National Aeronautics and Space Administration. "Outreach Programs: Women and Girls Initiative." <http://women.nasa.gov/outreach-programs/>. Accessed June 10, 2016.

14.1: DNA structure is shared under a [CC BY-SA](#) license and was authored, remixed, and/or curated by LibreTexts.

Index

D

dire

Detailed Licensing

Overview

Title: General, Organic, and Biochemistry with Problems, Case Studies, and Activities

Webpages: 68

All licenses found:

- [CC BY-SA 4.0](#): 98.5% (67 pages)
- [Undeclared](#): 1.5% (1 page)

By Page

- General, Organic, and Biochemistry with Problems, Case Studies, and Activities - [CC BY-SA 4.0](#)
 - Front Matter - [CC BY-SA 4.0](#)
 - [TitlePage](#) - [CC BY-SA 4.0](#)
 - [InfoPage](#) - [CC BY-SA 4.0](#)
 - [Table of Contents](#) - [Undeclared](#)
 - [Licensing](#) - [CC BY-SA 4.0](#)
 - 1: Scientific Method, Matter, and Measurements - [CC BY-SA 4.0](#)
 - [1.1: The Scientific Method](#) - [CC BY-SA 4.0](#)
 - [1.2: Units](#) - [CC BY-SA 4.0](#)
 - [1.3: Uncertainty in Measurement](#) - [CC BY-SA 4.0](#)
 - [1.4: Matter](#) - [CC BY-SA 4.0](#)
 - 2: Atoms and Elements - [CC BY-SA 4.0](#)
 - [2.1: Electrons and electronic structure](#) - [CC BY-SA 4.0](#)
 - 3: Ionic and Covalent Compounds - [CC BY-SA 4.0](#)
 - [3.1: Ionic and Covalent Compounds](#) - [CC BY-SA 4.0](#)
 - [3.2: Nomenclature](#) - [CC BY-SA 4.0](#)
 - 4: Chemical Bonding - [CC BY-SA 4.0](#)
 - [4.1: Lewis structures](#) - [CC BY-SA 4.0](#)
 - [4.2: Covalent bonding and electronegativity](#) - [CC BY-SA 4.0](#)
 - [4.3: Valence Bond Theory](#) - [CC BY-SA 4.0](#)
 - [4.4: Molecular Structure and Polarity](#) - [CC BY-SA 4.0](#)
 - [4.5: Atomic Structure](#) - [CC BY-SA 4.0](#)
 - 5: Unit Conversions - [CC BY-SA 4.0](#)
 - [5.1: Unit conversions](#) - [CC BY-SA 4.0](#)
 - [5.2: Density](#) - [CC BY-SA 4.0](#)
 - 6: Moles - [CC BY-SA 4.0](#)
 - [6.1: Moles](#) - [CC BY-SA 4.0](#)
 - [6.2: Formula mass](#) - [CC BY-SA 4.0](#)
 - [6.3: Converting between grams and moles](#) - [CC BY-SA 4.0](#)
 - 7: Chemical Equations and Reactions - [CC BY-SA 4.0](#)
 - [7.1: Balancing Chemical Equations](#) - [CC BY-SA 4.0](#)
 - [7.2: Energy in chemical reactions](#) - [CC BY-SA 4.0](#)
 - [7.3: Equilibrium](#) - [CC BY-SA 4.0](#)
 - [7.4: Catalysis](#) - [CC BY-SA 4.0](#)
 - 8: Gases, Liquids, and Solids - [CC BY-SA 4.0](#)
 - [8.1: Kinetic Molecular Theory](#) - [CC BY-SA 4.0](#)
 - [8.2: Pressure](#) - [CC BY-SA 4.0](#)
 - [8.3: The Gas Laws](#) - [CC BY-SA 4.0](#)
 - [8.4: Intermolecular forces](#) - [CC BY-SA 4.0](#)
 - [8.5: Types of solutions](#) - [CC BY-SA 4.0](#)
 - [8.6: Electrolytes](#) - [CC BY-SA 4.0](#)
 - [8.7: Concentration](#) - [CC BY-SA 4.0](#)
 - [8.8: Osmosis, diffusion, and dialysis](#) - [CC BY-SA 4.0](#)
 - 9: Acids and Bases - [CC BY-SA 4.0](#)
 - [9.1: What are acids and bases?](#) - [CC BY-SA 4.0](#)
 - [9.2: pH and pOH](#) - [CC BY-SA 4.0](#)
 - 10: Organic Chemistry - [CC BY-SA 4.0](#)
 - [10.1: Drawing organic structures](#) - [CC BY-SA 4.0](#)
 - [10.2: Alkanes and isomers](#) - [CC BY-SA 4.0](#)
 - [10.3: Functional groups](#) - [CC BY-SA 4.0](#)
 - [10.4: Organic Reactions](#) - [CC BY-SA 4.0](#)
 - [10.5: Carbonyl substitution reactions](#) - [CC BY-SA 4.0](#)
 - 11: Carbohydrates - [CC BY-SA 4.0](#)
 - [11.1: Carbohydrates](#) - [CC BY-SA 4.0](#)
 - [11.2: Fischer and Haworth projections](#) - [CC BY-SA 4.0](#)
 - 12: Lipids - [CC BY-SA 4.0](#)
 - [12.1: Lipids](#) - [CC BY-SA 4.0](#)
 - 13: Proteins - [CC BY-SA 4.0](#)
 - [13.1: Properties of proteins](#) - [CC BY-SA 4.0](#)
 - [13.2: Amino acids](#) - [CC BY-SA 4.0](#)
 - [13.3: Primary structure- peptides](#) - [CC BY-SA 4.0](#)
 - [13.4: Secondary, tertiary, and quaternary structure of proteins](#) - [CC BY-SA 4.0](#)
 - 14: Molecular Genetics - [CC BY-SA 4.0](#)
 - [14.1: DNA structure](#) - [CC BY-SA 4.0](#)

- [Back Matter - CC BY-SA 4.0](#)
 - [Index - CC BY-SA 4.0](#)

- [Glossary - CC BY-SA 4.0](#)
- [Detailed Licensing - CC BY-SA 4.0](#)