CHEMISTRY FOR ALLIED HEALTH

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Chemistry for Allied Health

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TABLE OF CONTENTS

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CHAPTER OVERVIEW

1: Measurements and Problem-Solving

1.1: Measurements Matter

Learning Outcomes

- Demonstrate use of scientific notation.
- Identify base and derived units for measurements.
- Memorize the metric prefixes, abbreviations, and values (Tables 1.1.1 and 1.1.3).
- Convert between temperatures in Celsius and Fahrenheit.

Many aspects of chemistry use quantitative measurements to describe behavior. In this section, we will look at how we deal with very large or very small numbers, units of measure and prefixes used with these units, and how numbers are reported.

Scientific Notation

Scientific notation is a way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power. It is a very useful tool for working with numbers that are either very large or very small. As an example, the distance from Earth to the Sun is about 150,000,000,000 meters - a very large distance indeed. In scientific notation, the distance is written as $1.5 \times 10^{11}\ \mathrm{m}$. The coefficient is the 1.5 and must be a number greater than or equal to 1 and less than 10. The power of 10, or exponent, is 11 because you would have to multiply 1.5 by 10^{11} to get the correct number. Scientific notation is sometimes referred to as exponential notation.

When working with small numbers, we use a negative exponent. So 0.1 meters is 1×10^{-1} meters, 0.01 is 1×10^{-2} and so forth. A red blood cell has a diameter of 0.000008 meters which is neither convenient to write nor say. It is much easier to report a diameter of $8 \times 10^{-6}\ \mathrm{m}$. Using scientific notation is one way to make writing very large or very small numbers more convenient. Note the use of a **leading zero** (the zero to the left of the decimal point) when writing very small numbers. The leading zero is there to help you see the decimal point more clearly. The figure 0.01 is less likely to be misunderstood that .01 where you may not see the decimal point.

Units

SI Base Units

All measurements depend on the use of units that are all well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units, five of which are shown in the table below. While some of the base units have a calculation as their standard, the kilogram has a physical standard but there is a movement to change that. Read or listen to more about changes to the standard at http://www.npr.org/templates/story/s...ryId=112003322.

Table 1.1.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol

These five units are frequently encountered in chemistry and other units of measurement, such as volume, force, and energy, can be derived from the SI base units.

Figure 1.1.1: *The standard for the unit kilogram is stored in France.*

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

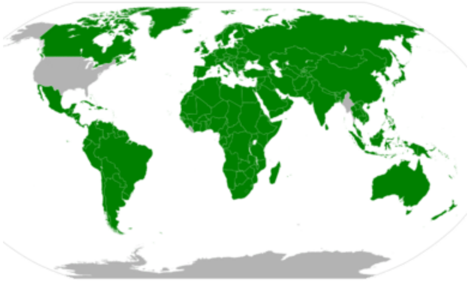

Figure 1.1.2: Countries which use the metric system.

Derived Units

Some units are combinations of SI base units. A **derived unit** is a unit that results from a mathematical combination of SI base units. These derived units are the standard for these quantities but they are not always the units used in normal practice. For example, the density of a solid or a liquid is often reported as $\mathrm{g/cm^3}$, or $\mathrm{g/mL}$ rather than $\mathrm{kg/m^3}$ (the derived SI unit). Some common examples of derived units are listed in the table below.

Table 1.1.2: Derived SI Units

Quantity	Symbol	Unit	Unit Abbreviation	Derivation
Area	A	square meter	$\mathrm{m^2}$	$\text{length} \times \text{width}$
Volume	V	cubic meter	$\mathrm{m^3}$	$\text{length} \times \text{width} \times \text{height}$
Density	D	kilograms/cubic meter	$\mathrm{kg/m^3}$	$\frac{\text{mass}}{\text{volume}}$
Concentration	c	moles/liter	$\mathrm{mol/L}$	$\frac{\text{amount}}{\text{volume}}$
Speed (velocity)	v	meters/second	$\mathrm{m/s}$	$\frac{\text{length}}{\text{time}}$
Acceleration	a	meters/second/second	$\mathrm{m/s^2}$	$\frac{\text{speed}}{\text{time}}$
Force	F	newton	N	$\text{mass} \times \text{acceleration}$
Energy	E	joule	J	$\text{force} \times \text{length}$

Metric Prefixes

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word $\mu\epsilon\gamma\alpha\varsigma$, meaning "great".

The table below lists the most common metric prefixes and their relationship to the base unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit. However, these prefixes can be used with *any* base unit.

Table 1.1.3: SI Prefixes

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

*Micro is often abbreviated "mc" in healthcare to avoid confusion between the μ and m which often look similar when written by hand.

There are more prefixes, some of them rarely used, that go far beyond what is given here. Have you ever heard of a zeptometer? You can learn more about prefixes at http://www.bipm.org/en/measurement-units/

Since the metric system is not the usual unit system we use, it can be challenging to understand the relative size of each unit. Go to http://learn.genetics.utah.edu/content/cells/scale and use the slider below the image to zoom in and see how the sizes of objects compare.

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

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs $6.5\ \mathrm{kg}$, this is easier than saying it weighs $6500\ \mathrm{g}$. Both are correct, bu the kg units in this case make for a small and easily managed number.

Temperature Scales

There are three temperature scales that are commonly used in measurement. Their units are $^\circ\mathrm{F}$ (degrees Fahrenheit), $^\circ\mathrm{C}$ (degrees Celsius), and K (Kelvin). The **Fahrenheit scale**, which is the most commonly used scale in the United States, defines the normal freezing point and boiling point of water as $32^\mathrm{o}\mathrm{F}$ and $212^\mathrm{o}\mathrm{F}$, respectively. The **Celsius scale** defines the normal freezing point and boiling point of water as $0^\mathrm{o}\mathrm{C}$ and $100^\mathrm{o}\mathrm{C}$, respectively. The Celsius scale is commonly used in most countries across the globe. The **Kelvin scale**, which is also referred to as the absolute temperature scale, defines absolute zero as the lowest theoretically possible temperature, which means that temperatures expressed in Kelvin cannot be negative numbers.

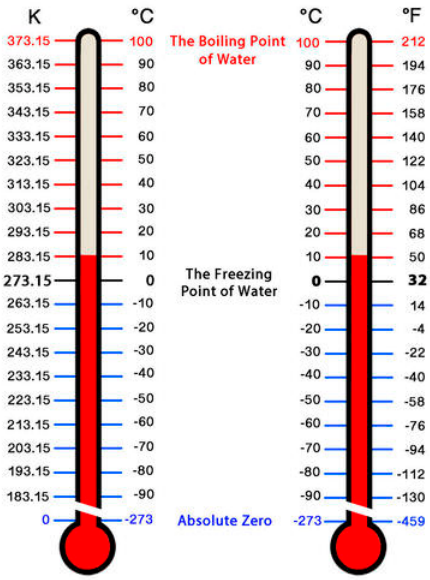

Figure 1.1.3: Comparison of Kelvin, Celsius on a Fahrenheit scale.

Converting Temperature Scales

Regardless of the temperature scale used, it is important to be able to convert from one scale to another. Here are the conversions we use.

$^{\mathrm{o}}$F to $^{\mathrm{o}}$C

$$T_{^\circ\mathrm{C}} = (T_{^\circ\mathrm{F}} - 32) \times \frac{5}{9} \tag{1.1.1}$$

$^{\mathrm{o}}$C to $^{\mathrm{o}}$F

$$T_{^\circ\mathrm{F}} = \frac{9}{5} \times (T_{^\circ\mathrm{C}}) + 32 \tag{1.1.2}$$

$^{\mathrm{o}}$C to K

$$T_{\mathrm{K}} = T_{^\circ\mathrm{C}} + 273.15 \tag{1.1.3}$$

K to $^{\mathrm{o}}$C

$$T_{^\circ\mathrm{C}} = T_{\mathrm{K}} - 273.15 \tag{1.1.4}$$

Example 1.1.1

The melting point of mercury is $-38.84^{\mathrm{o}}\mathrm{C}$. Convert this value to degrees Fahrenheit and Kelvin.

Solution

The formulas above can be used to convert among temperature units. First, the given value can be used to convert from $^{\mathrm{o}}$C to $^{\mathrm{o}}$F.

$$\begin{aligned} T_{^\circ\mathrm{F}} &= \frac{9}{5} \times (-38.84^{\mathrm{o}}\mathrm{C}) + 32 \\ T_{^\circ\mathrm{F}} &= -37.91^{\mathrm{o}}\mathrm{F} \end{aligned}$$

Then, the same initial temperature can be used to find the temperature in Kelvin.

$$\begin{aligned} T_{\mathrm{K}} &= -38.84^{\mathrm{o}}\mathrm{C} + 273.15 \\ T_{\mathrm{K}} &= 234.75\ \mathrm{K} \end{aligned}$$

Supplemental Resources

- Scientific notation practice: http://www.mathisfun.com/numbers/sci...-notation.html
- History of SI Units: http://physics.nist.gov/cuu/Units/history.html

Contributors and Attributions

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

1.2: Significant Figures

Learning Outcomes

- Distinguish between accuracy and precision.
- Explain the concept of significant figures.
- Report answers to calculations with the correct number of significant figures.

In chemistry, we are looking at measurements and not just numbers. Therefore, it is necessary to be able to count the number of significant figures and report measurements to the correct level of precision. The reported value should be as precise as possible without adding digits to a value that cannot be measured.

Accuracy and Precision

In everyday speech, the terms accuracy and precision are frequently used interchangeably. However, their scientific meanings are quite different. **Accuracy** *is a measure of how close a measurement is to the correct or accepted value of the quantity being measured.* **Precision** *is a measure of how close a series of measurements are to one another.* Precise measurements are highly reproducible, even if the measurements are not near the correct value. Darts thrown at a dartboard are helpful in illustrating accuracy and precision (see figure below).

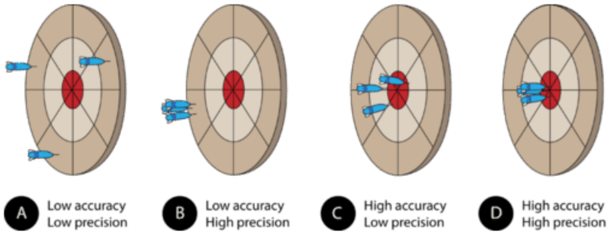

Figure 1.2.3: Distribution of darts on a dartboard to represent precision and accuracy.

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. Pictured above are the four possible outcomes.

(A) The darts have landed far from each other and far from the bulls-eye. This grouping demonstrates measurements that are neither accurate nor precise.

(B) The darts are close to one another, but far from the bulls-eye. This grouping demonstrates measurements that are precise, but not accurate. In a laboratory situation, high precision with low accuracy often results from a systematic error. Either the measurer makes the same mistake repeatedly or the measuring tool is somehow flawed. A poorly calibrated balance may give the same mass reading every time, but it will be far from the true mass of the object.

(C) The darts are not grouped very near to each other, but they are generally centered around the bulls-eye. This demonstrates poor precision but fairly high accuracy. This situation is not desirable because in a lab situation, we do not know where the "bulls-eye" actually is. Continuing with this analogy, measurements are taken in order to find the bulls-eye. If we could only see the locations of the darts and not the bulls-eye, the large spread would make it difficult to be confident about where the exact center was, even if we knew that the darts were thrown accurately (which would correspond to having equipment that is calibrated and operated correctly).

(D) The darts are grouped together and have hit the bulls-eye. This demonstrates high precision and high accuracy. Scientists always strive to maximize both in their measurements. Turning back to our laboratory situation, where we can see the darts but not the bulls-eye, we have a much narrower range of possibilities for the exact center than in the less precise situation depicted in part C.

Significant Figures in Measurements

Some error or uncertainty always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Pictured below is an object (indicated by the blue arrow) whose length is being measured by two different rulers.

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

The top ruler contains marks for tenths of a centimeter (millimeters). Now, the same object may be measured as 2.55 cm. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5.

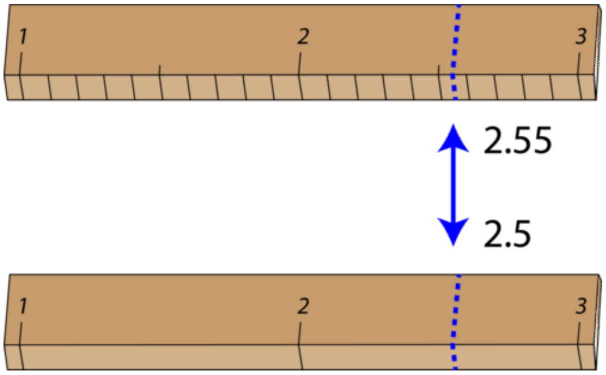

Figure $1.2.4$: Measurements can be more or less precise depending on the markings on the ruler.

Again, another measurer may report the length to be 2.54 cm or 2.66 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain (estimated) digit. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported.

Rules for Determination of Significant Figures

1. All nonzero digits are significant.

- 237 has three significant figures
- 1.897 has four significant figures

2. Zeroes between significant figures are significant.

- 39004 has five significant figures
- 2.03 has three significant figures

3. Zeroes that appear to the left of all of the nonzero digits are called placeholder zeroes and are not significant.

- 0.008 has one significant figure
- 0.0000416 has three significant figures

4. Zeroes that appear after the last nonzero digit may be significant.

a. If the number is greater than 1 and the zero(es) are not followed by a decimal point, then the number of significant figures is ambiguous.

- 140 is ambiguous

b. If the zero is followed by a decimal point, then the zero is significant.

- 140. has three significant figures (note the decimal point after the zero)

c. If the zero is after a decimal point, then the zero is significant.

- 141.0 has four significant figures

It needs to be emphasized that just because a certain digit is not significant does not mean that it is not important or that it can be left out. Though the zero in a measurement of 140 may not be significant, the value cannot simply be reported as 14. An insignificant zero functions as a placeholder for the decimal point. When numbers are written in scientific notation, this becomes more apparent. The measurement 140 can be written as 1.4×10^2, with two significant figures in the coefficient or as 1.40×10^3, with three significant figures. A number less than one, such as 0.000416, can be written in scientific notation as 4.16×10^{-4}, which has 3 significant figures. In some cases, scientific notation is the only way to correctly indicate the correct number of significant figures. In order to report a value of 15,000,00 with four significant figures, it would need to be written as 1.500×10^7.

Exact Quantities

When numbers are known exactly, the significant figure rules do not apply. This occurs when objects are counted rather than measured. For example, a carton of eggs has 12 eggs. The actual value cannot be 11.8 eggs, since we count eggs in whole number quantities. So the 12 is an exact quantity. Exact quantities are considered to have an infinite number of significant figures; the importance of this concept will be seen later when we begin looking at how significant figures are dealt with during calculations. Numbers in many conversion factors, especially for simple unit conversions, are also exact quantities and have infinite significant figures. There are exactly 100 centimeters in 1 meter and exactly 60 seconds in 1 minute. Those values are definitions and are not the result of a measurement.

Calculations

Adding and Subtracting Significant Figures

The sum or difference is determined by the smallest number of digits to the right of the decimal point in any of the original numbers.

Example 1.2.1

What is the result of $89.332 + 1.1$ when answered to the correct number of significant figures?

Solution

The mathematical result is $89.332 + 1.1 = 90.432$ However, we need to round the answer to the correct number of significant figures. For addition and subtraction, the answer should have the same number of decimal places as the starting value with the *least* number of decimal places. Since 1.1 only has one digit after the decimal place, the answer will be rounded to one decimal place. The first digit to be dropped is less than five so we round down.

The correct answer is 90.4.

Multiplying and Dividing Significant Figures

The number of significant figures in the final product or quotient is equal to the number of significant figures in the starting value that has the fewest significant figures.

Example 1.2.2

What is the product of 2.8 and 4.5039 reported to the correct number of significant figures?

Solution

First, we find the mathematical answer to the calculation which is $2.8 \times 4.5039 = 12.61092$ Next, we need to round the answer to the correct number of significant figures. For multiplication and division, the answer should be rounded to that it has the same number of significant figures as the starting value with the least number of significant figures. The value 2.8 has two significant figures while 4.5039 has five. Therefore, the answer should have two significant figures so we should round the answer to 13 because the first digit to be dropped is greater than 5, so we need to round up.

Example 1.2.3

A bag of chocolate candies has a mass of 8.25 ounces. The candies were to be divided among 4 people. How many ounces should each person get? Report your answer to the correct number of significant figures.

Solution

First, we'll find the mathematical answer to this calculation: $8.25 \div 4 = 2.0625$. Now, we need to round to the correct number of significant figures so we look back at the starting values. There are three significant figures in 8.25 but 4 is an exact number

> so it is not considered when determining the number of significant figures in the answer. Therefore, the answer should be rounded to three significant figures or 2.06 ounces.

Supplemental Resources

Contributors and Attributions

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

1.3: Scientific Dimensional Analysis

Learning Outcomes

- Convert values among units.
- Use dimensional analysis to solve problems.

Dimensional analysis (also called factor label method or unit analysis) is used to convert from one set of units to another. This method is used for both simple (feet to inches) and complex ($\mathrm{g/cm^3}$ to kg/gallon) conversions and uses relationships or conversion factors between different sets of units. While the terms are frequently used interchangeably, conversion factors and relationships are different. Conversion factors are quantities that are equal to one another, such as $100\ \mathrm{cm} = 1\ \mathrm{m}$, in which both values describe a length. Relationships are between two values that are not necessarily a measure of the same quantity. For example, the density of water is $1.00\ \mathrm{g/mL}$ Grams are a measure of mass while milliliters measure volume so this is considered a relationship rather than a conversion factor. Either way, we depend on units to help set up and solve the calculation. We will see additional examples of relationships as we explore other details about chemical substance.

Importance of using Correct Conversions

In healthcare professions, a calculation error can quite literally have a life or death consequence. Read the short article, Med-Math Errors and the Nursing Student at http://www.alysion.org/dimensional/matherrors.htm, to better understand the importance of units and correct calculations.

Conversion Factors

Many quantities can be expressed in several different ways. The English system measurement of 4 cups is also equal to 2 pints, 1 quart, and $\frac{1}{4}$ of a gallon. $4\ \mathrm{cups} = 2\ \mathrm{pints} = 1\ \mathrm{quart} = 0.25\ \mathrm{gallon}$. Notice that the numerical component of each quantity is different, while the actual amount of material that it represents is the same. That is because the units are different. We can establish the same set of equalities for the metric system: $1\ \mathrm{meter} = 10\ \mathrm{decimeters} = 100\ \mathrm{centimeters} = 1000\ \mathrm{millimeters}$ The metric system's use of powers of 10 for all conversions makes this quite simple. We can write conversion factors between any pair of equivalent quantities. In each conversion factor, the numerator and denominator represent equal quantities so they are all valid conversion factors. Additionally, these conversion factors can be inverted or used in combination with other conversion factors in a dimensional analysis problem.

$$\frac{4\ \mathrm{cups}}{2\ \mathrm{pints}} = \frac{1\ \mathrm{quart}}{4\ \mathrm{cups}} = \frac{0.25\ \mathrm{gallen}}{1\ \mathrm{quarter}} = 1$$

$$\frac{1\ \mathrm{meter}}{10\ \mathrm{decimeters}} = \frac{100\ \mathrm{centimeters}}{1000\ \mathrm{milimeters}} = \frac{1000\ \mathrm{milimeters}}{1\ \mathrm{meter}} = 1$$

Example 1.3.1

How many centimeters are in $3.4\ \mathrm{m}$?

Solution

This problem requires the conversion from one unit to another so we can use dimensional analysis to solve the problem. We need to identify the units that are given (m), the units for the answer (cm), and any relationships that relate the units of the known and unknown values. In this case, we will use the relationship of $1\ \mathrm{m} = 100\ \mathrm{cm}$. Start with the known value and its unit.

$$3.4\ \mathrm{m} \times \frac{?}{?}$$

Then, we look at the units of our relationship to see which value goes in the numerator and which value goes in the denominator. Remember, we are trying to find the value in centimeters. Since our known value is in units of meters, we need meters to be in the denominator so that it will cancel. As a result, centimeters will be in the numerator.

$$3.4\ \mathrm{m} \times \frac{100\ \mathrm{cm}}{1\ \mathrm{m}}$$

Note that the numbers stay with the appropriate unit (100 with centimeters and 1 with meters). Now, the meters will cancel out and we are left with units of centimeters. Always check that your problem is set up completely and that your units cancel correctly before you do the actual calculation.

$$\begin{aligned} 3.4 \cancel{\text{ m}} \times \frac{100 \text{ cm}}{1 \cancel{\text{ m}}} &= 340 \text{ cm} \\ &= 3.4 \times 10^2 \text{ cm} \end{aligned}$$

We find the answer to be 340 cm or $3.4 \times 10^2 \text{ cm}$.

Derived Units

Using dimensional analysis with derived units requires special care. When units are squared or cubed as with area or volume, the conversion factors themselves must also be squared or cubed. Two convenient volume units are the liter, which is equal to a cubic decimeter, and the milliliters, which is equal to a cubic centimeter. There are thus 1000 cm^3 in 1 dm^3, which is the same thing as saying there are 1000 mL in 1 L. *The conversion factor of* $1 \text{ cm}^3 = 1 \text{ mL}$ *is a very useful conversion.*

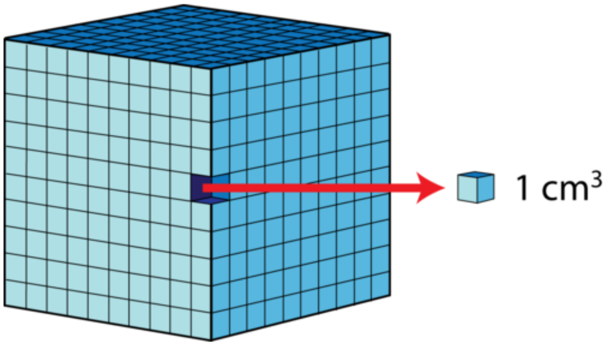

Figure 1.3.1: One centimeter cubed is the volume occupied by a cube with an edge length of 1 cm.

There are 1000 cm^3 in 1 dm^3. Since a cm^3 is equal to a mL, and a dm^3 is equal to a L, we can say that there are 1000 mL in 1 L.

Example 1.3.2

Convert 3.6 mm^3 to mL.

Solution

Determine the units of the known value $\left(\text{mm}^3\right)$ and the units of the unknown value (mL). The starting and ending units will help guide the setup of the problem. Next, list any known conversion factors that might be helpful.

- $1 \text{ m} = 1000 \text{ mm}$
- $1 \text{ mL} = 1 \text{ cm}^3$
- $1 \text{ m} = 100 \text{ cm}$

Now, we can set up the problem to find the value in units of mL. Once we know the starting units, we can then use the conversion factors to find the answer.

$$3.6 \text{ mm}^3 \times \left(\frac{?}{?}\right)$$

Continue to use the conversion factors between the units to set up the rest of the problem. Note that all of the units cancel except mL, which are the requested units for the answer. Since the values in these conversion factors are exact numbers, they will not affect the number of significant figures in the answer. Only the original value (3.6) will be considered in determining significant figures.

$$3.6 \text{ mm}^3 \times \left(\frac{1 \text{ m}}{1000 \text{ mm}}\right)^3 \times \left(\frac{?}{?}\right)$$

> Once you have solved the problem, always ask if the answer seems reasonable. Remember, a millimeter is very small and a cubic millimeter is also very small. Therefore, we would expect a small volume which means $0.0036\ \text{mL}$is reasonable.
>
> If you find that you forgot to cube numbers as well as units, you can setup the problem in an expanded form which is the equivalent to the previous method to cube the numerical values.

Contributors and Attributions

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

1.4: Percentages

Learning Outcomes

- Calculate percentages and use in calculations.

In chemistry and healthcare, values are reported with a variety of units. Percentages are often used in the concentration of IV solutions and medications. There are multiple ways to report percentages so it is important to understand the basic meaning of a percent, which will be the focus here, as well as the type of percentage, such as mass or volume percent.

Percent

Percents are used to report many different values, such as a grade on an assignment as 85%, the shoes are 20% off their original price, or the concentration of the IV solution is $5.5\%\ \mathrm{NaCl}$ by mass. While all of these numbers use percent very differently, the values are all found the same way. For now, we will focus on some basic calculations using percents. Percents are all calculated with the same basic formula using the amount of the "part" and the amount of the "whole". Note that the percent will always be less than or equal to 100% because the "part" must be less than or equal to the "whole". The basic percent equation is shown here.

$$\% = \frac{\text{part}}{\text{whole}} \times 100 \tag{1.4.1}$$

Example 1.4.1

What percent of 70 is 14?

Solution

Since this is a percent problem, first look at the equation for percent and see which values are given in the problem. In this case, we know the part (14) and the whole (70). With this information and the formula, we can solve for the percent by inserting the known values into the equation and solving for the unknown.

$$\begin{aligned} \% &= \frac{\text{part}}{\text{whole}} \times 100 \\ \% &= \frac{14}{70} \times 100 \\ \% &= 0.20 \times 100 \\ \% &= 20 \end{aligned}$$

Therefore, 14 is 20% of 70.

Example 1.4.2

What is 35% of 80?

Solution

This problem also involves percent so we will use the same equation as the previous example. However, in this example, we know the percentage (35) and the whole (80) and we are trying to find the value of the part. As before, we will insert the known values into the equation and solve for our unknown.

$$\begin{aligned} \% &= \frac{\text{part}}{\text{whole}} \times 100 \\ 35 &= \frac{\text{part}}{80} \times 100 \\ 0.35 &= \frac{\text{part}}{80} \\ \text{part} &= 28 \end{aligned}$$

So 28 is 35% of 80.

If you need to review how to rearrange equations and solve for an unknown, check out the worked examples and practice problems at https://www.mathisfun.com/algebra/in...-multiply.html.

Example 1.4.3

A patient has a fasting blood sugar level of $150\ \mathrm{mg/dL}$. The doctor has recommended some dietary changes and is hoping to see at least a 20% reduction in the patient's blood sugar level at their next appointment. If the patient meets this goal, what will their blood sugar level be?

Solution

We can start this problem like we do other percent problems. The problem gives use the percent (20) and the whole (150) and we need to calculate the part which will be the amount of decrease in the patient's blood sugar.

$$\begin{aligned}\% &= \frac{\text{part}}{\text{whole}} \times 100 \\ 20 &= \frac{\text{part}}{150} \times 100 \\ 0.20 &= \frac{\text{part}}{150} \\ \text{part} &= 30\end{aligned}$$

The part that was calculated is the *decrease* that should be seen in the patient's blood sugar level. However, it's not the actual blood sugar level that the doctor wants to see at the next visit. To find that, we need to subtract from the original value.

$$\text{original} - \text{loss} = \text{final} \tag{1.4.2}$$

$$150\ \frac{\text{mg}}{\text{dL}} - 30\ \frac{\text{mg}}{\text{dL}} = 120\ \frac{\text{mg}}{\text{dL}} \tag{1.4.3}$$

At the next visit, the patient's blood sugar level should be $120\ \mathrm{mg/dL}$ (or less!) if they are making sufficient progress towards improving their health and reducing their risk of diabetes.

Supplemental Resources

- Percents: http://www.mathisfun.com/percentage.html

Contributors and Attributions

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1.5: Measurements and Problem-Solving (Exercises)

These are homework exercises to accompany Chapter 1 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

1.1: Measurements Matter

(click here for solutions)

Q1.1.1

Express the following values in scientific notation.

a. 150,000,000
b. 0.000043
c. 332000
d. 0.0293
e. 932
f. 0.1873
g. 78,000
h. 0.0001
i. 4500
j. 0.00290
k. 6281
l. 0.00700

Q1.1.2

Express the following values in decimal notation.

a. 3.8×10^{-3}
b. 9.21×10^{5}
c. 7.91×10^{-2}
d. 2.5×10^{6}
e. 3.42×10^{-8}
f. 5.4×10^{5}
g. 3×10^{-3}
h. 7.34×10^{2}
i. 9.8×10^{-4}
j. 6×10^{7}
k. 4.20×10^{-6}
l. 4.20×10^{6}

Q1.1.3

What SI base unit would be appropriate for each measurement?

a. the length of a room
b. the amount of carbon in a diamond
c. the mass of NaCl in a bottle

Q1.1.4

List the meaning of each abbreviation of the base units.

a. m
b. K
c. kg
d. s

e. mol

Q1.1.5

What is the the derived unit from the SI base units for the relationship of each pair of quantities?

a. mass and volume
b. distance and time
c. amount of substance and volume
d. area

Q1.1.6

Give the meaning and name of each metrix prefix abbreviation.

a. M
b. m
c. n
d. d

Q1.1.7

Give the abbreviation and meaning of each metrix prefix.

a. kilo
b. centi
c. micro
d. giga

Q1.1.8

Name the prefix with the following numerical meaning.

a. 1/10
b. 1,000,000
c. 1/1,000,000
d. 1/100
e. 1

Q1.1.9

Convert each temperature to the missing one between Celsius and Fahrenheit.

a. 77°F
b. 212°F
c. 37°C
d. 22°C
e. 95°F
f. 15°C
g. 0°F
h. 0°C
i. –10°C
j. –10°F

1.2 Significant Figures

(click here for solutions)

Q1.2.1

Explain the similarities and differences between accuracy and precision.

Q1.2.2

The density of a copper sample was determined by three different students (shown below). Each performed the measurement three times and is reported below (all values in units of g/cm^3). The accepted value for the density of copper is 8.92 g/cm^3.

a. Determine if each student's data is accurate, precise, neither or both.
b. What is the average density based on Justin's data?
c. What is the average density based on Jane's data?
   - Jane: 8.94, 8.89, 8.91
   - Justin: 8.32, 8.31, 8.34
   - Julia: 8.64, 9.71, and 9.10

Q1.2.3

Determine the number of significant figures in each of the following values.

a. 406
b. 3.00
c. 3.20
d. 0.25
e. 0.0689
f. 0.002910
g. 3941
h. 46.250
i. 30.21
j. 0.10300

Q1.2.4

Write each value with three significant figures, use scientific notation if necessary.

a. 34500
b. 24
c. 0.0345
d. 0.012
e. 612.8
f. 98.22
g. 0.14928
h. 300

Q1.2.5

Give three examples of exact numbers.

Q1.2.6

Find the result of each of the following calculations and report the value with the correct number of significant figures.

a. 0.23 + 12.2 =
b. 13 - 1.03 =
c. 0.839 + 0.28925 =
d. 28 + 34.4 =
e. 0.8 + 2.3 =
f. 34.9 - 0.583 =
g. 21 - 0.132 =
h. 0.840 + 0.9334

Q1.2.7

Find the result of each of the following calculations and report the value with the correct number of significant figures.

a. 34 x 0.12 =
b. 68.2 / 0.78 =

c. $3.29 \times 10^4 \times 16.2 =$
d. 0.8449 x 29.7 =
e. $5.92 \times 10^3 / 0.628 =$
f. 3.00 x 2.6 =
g. 2.50 x 9.331 =
h. 3.20 / 12.75 =

1.3 Scientific Dimensional Analysis

(click here for solutions)

Q1.3.1

What is a conversion factor?

Q1.3.2

What is the conversion factor between each pair of units?

a. feet and inches
b. mL and cm^3
c. kg and g
d. cm and m
e. mm and cm
f. inches and centimeters
g. grams and pounds
h. g and µg (mcg)

Q1.3.3

Complete each of the following conversions.

a. 34 cm to m
b. 3.7 ft to in
c. 345 mg to Mg
d. 5.3 km to mm
e. 4.0 L to mL
f. 3.45×10^3 mm to km
g. 78 cm^3 to mL
h. 0.85 kg to dg
i. 13 pints to gallon
j. 0.35 L to cm^3

Q1.3.4

Complete each of the following conversions.

a. 342 cm^3 to dm^3
b. 2.70 g/cm^3 to kg/L
c. 34 mi/hr to km/min
d. 0.00722 km^2 to m^2
e. 4.9×10^5 mcm^3 to mm^3
f. 80. km/hr to mi/hr

1.4 Percentages

(click here for solutions)

Q1.4.1

Solve each of the following.

a. What percent of 35 is 8.2?

b. What percent of 56 is 12?
c. What percent of 923 is 38?
d. What percent of 342 is 118?

Q1.4.2

Solve each of the following?

a. What is 42% of 94?
b. What is 83% of 239?
c. What is 16% of 45?
d. What is 38% of 872?

Q1.4.3

Solve each of the following?

a. 42 is 34% of what number?
b. 73 is 82% of what number?
c. 13 is 57% of what number?
d. 75 is 25% of what number?
e. 25 is 15% of what number?
f. 98 is 76% of what number?

Q1.4.4

A patient originally weighs 182 pounds and loses 15.0% of their body weight. What is their final weight?

Q1.4.5

A patient's original weight was 135 pounds and they lose 12 pounds. What percent of their body weight did they lose?

Q1.4.6

A patient needs to increase their calcium supplement by 25% a week. If they are currently taking a 300. mg supplement, how much more will they need to take?

Q1.4.7

An infant's birth weight is 7 pounds, 1 ounce. Her discharge weight is 6 pounds, 13 ounces. What percent of her birth weight did she lose?

Q1.4.8

A patient needs a 20.% decrease in their medication dosage from 125 mg. What will his dosage be after the decrease?

Answers

1.1: Measurements Matter

Q1.1.1

a. 1.5×10^{8}
b. 4.3×10^{-5}
c. 3.32×10^{5}
d. 2.93×10^{-2}
e. 9.32×10^{2}
f. 1.873×10^{-1}
g. 7.8×10^{4}
h. 1×10^{-4}
i. 4.5×10^{3}
j. 2.9×10^{-3}
k. 6.281×10^{3}
l. 7×10^{-3}

Q1.1.2

a. 0.0038
b. 921000
c. 0.0791
d. 2500000
e. 0.0000000342
f. 540000
g. 0.003
h. 734
i. 0.00098
j. 60000000
k. 0.00000420
l. 4200000

Q1.1.3

a. meter
b. mole
c. kilogram

Q1.1.4

a. meter
b. Kelvin
c. kilogram
d. second
e. mole

Q1.1.5

a. kg/m^3
b. m/s
c. mol/m^3 is based on SI base units, but mol/L is also acceptable
d. m^2

Q1.1.6

a. Mega, 10^6
b. milli, 10^{-3}
c. nano, 10^{-9}
d. deci, 10^{-1}

Q1.1.7

a. k, 10^3
b. c, 10^{-2}
c. μ (or mc), 10^{-6}
d. G, 10^9

Q1.1.8

a. deci
b. mega
c. micro
d. centi
e. none (base unit)

Q1.1.9

a. 77°F = 25°C
b. 212°F = 100°C

c. 37°C = 98.6°F
d. 22°C = 72°F
e. 95°F = 35°C
f. 15°C = 59°F
g. 0°F = –18°C
h. 0°C = 32°F
i. –10°C = 14°F
j. –10°F = –23°C

1.2 Significant Figures

Q1.2.1

Accuracy is a measure of how close the values are close to the correct value while precision is a measure of how close values are to each other.

Q1.2.2

a.

- Jane: 8.94, 8.89, 8.91 - accurate and precise
- Justin: 8.32, 8.31, 8.34 - precise
- Julia: 8.64, 9.71, and 9.10 - neither accurate nor precise

a. 8.32 g/cm^3
b. 8.91 g/cm^3

Q1.2.3

a. 3
b. 3
c. 3
d. 2
e. 3
f. 4
g. 4
h. 5
i. 4
j. 5

Q1.2.4

a. 3.45×10^4
b. 2.40×10^1
c. 3.45×10^{-2}
d. 1.20×10^{-2}
e. 613 or 6.13×10^2
f. 9.82×10^1
g. 0.149 or 1.49×10^{-1}
h. 300. or 3.00×10^2

Q1.2.5

Answers will vary. 12 eggs, 100 cm = 1 m, 1 inch = 2.54 cm, 4 people

Q1.2.6

a. 0.23 + 12.2 = 12.43 = **12.4**
b. 13 - 1.03 = 11.97 = **12**
c. 0.839 + 0.28925 = 1.12825 = **1.128**
d. 28 + 34.4 = 62.4 = **62**
e. 0.8 + 2.3 = **3.1**

f. 34.9 - 0.583 = 34.317 = **34.3**
g. 21 - 0.132 = 20.868 = **21**
h. 0.840 + 0.9334 = 1.7734 = **1.773**

Q1.2.7

a. 34 x 0.12 = 4.08 = **4.1**
b. 68.2 / 0.78 = 87.4358974 = **87**
c. 3.29 x 10^4 x 16.2 = 5.32980 × 10^5 = 5.33 × 10^5
d. 0.8449 x 29.7 = 25.09353 = **25.1**
e. 5.92 x 10^3 / 0.628 = 9.4267515 × 10^3 = **9.43 × 10^3**
f. 3.00 x 2.6 = **7.8**
g. 2.50 x 9.331 = 23.3275 = **23.3**
h. 3.20 / 12.75 = 0.25098 = **0.251**

1.3 Scientific Dimensional Analysis

Q1.3.1

A conversion factor is a relationship between two units. The value in the numerator has some equivalence to the value in the denominator.

Q1.3.2

a. 1 foot = 12 inches
b. 1 mL = 1 cm^3
c. 1 kg = 1000 g or 1 × 10^{-3} kg = 1 g
d. 100 cm = 1 m or 1 cm = 1 × 10^{-2} m
e. 10 mm = 1 cm
f. 1 inch = 2.54 cm
g. 454 grams = 1 pound
h. 1 g = 1 × 10^6 µg (mcg) or 1 × 10^{-6} g = 1 µg (mcg)

Q1.3.3

a. $34\ cm \times \frac{1\ m}{100\ cm} = 0.34\ m$
b. $3.7\ ft \times \frac{12\ in}{1\ ft} = 44.4\ in = 44\ in$
c. $345\ mg \times \frac{1\ g}{1000\ mg} \times \frac{1\ Mg}{1 \times 10^6\ g} = 3.45 \times 10^{-7}\ Mg$
d. $5.3\ km \times \frac{1000\ m}{1\ km} \times \frac{1000\ mm}{1\ m} = 5.3 \times 10^6\ mm$
e. $4.0\ L \times \frac{1000\ mL}{1\ L} = 4.0 \times 10^3\ mL$
f. $3.45 \times 10^3\ mm \times \frac{1\ m}{1000\ mm} \times \frac{1\ km}{1000\ m} = 3.45 \times 10^{-3}\ km$
g. $78\ cm^3 \times \frac{1\ mL}{cm^3} = 78\ mL$
h. $0.85\ kg \times \frac{1000\ g}{1\ kg} \times \frac{10\ dg}{1\ g} = 8.5 \times 10^3\ dg$
i. $13\ pints \times \frac{1\ quart}{2\ pints} \times \frac{1\ gallon}{4\ quarts} = 1.6\ gallons$
j. $0.35\ L \times \frac{1000\ mL}{1\ L} \times \frac{1\ mL}{1\ cm^3} = 3.5 \times 10^2\ cm^3$

Q1.3.4

a. $342\ cm^3 \times \frac{1\ dm}{10\ cm} \times \frac{1\ dm}{10\ cm} \times \frac{1\ dm}{10\ cm} = 0.342\ dm^3$ **or** $342\ cm^3 \times \left(\frac{1\ dm}{10\ cm}\right)^3 = 342\ cm^3 \times \frac{1^3\ dm^3}{10^3\ cm^3} = 0.342\ dm^3$
b. $\frac{2.70\ g}{cm^3} \times \frac{1\ kg}{1000\ g} \times \frac{1\ cm^3}{1\ mL} \times \frac{1000\ mL}{1\ L} = \frac{2.70\ kg}{L}$
c. $\frac{34\ mi}{hr} \times \frac{5280\ ft}{1\ mi} \times \frac{12\ in}{1\ ft} \times \frac{2.54\ cm}{1\ in} \times \frac{1\ m}{100\ cm} \times \frac{1\ km}{1000\ m} \times \frac{1\ hr}{60\ min} = \frac{0.91\ km}{min}$
d. $0.00722\ km^2 \times \frac{1000\ m}{1\ km} \times \frac{1000\ m}{1\ km} = 7.22 \times 10^3\ m^2$
e. $4.95 \times 10^5\ mcm^3 \times \frac{1\ mm}{1000\ mcm} \times \frac{1\ mm}{1000\ mcm} \times \frac{1\ mm}{1000\ mcm} = 4.95 \times 10^{-4}\ mm^3$
f. $\frac{80.\ km}{hr} \times \frac{1000\ m}{1\ km} \times \frac{100\ cm}{1\ m} \times \frac{1\ in}{2.54\ cm} \times \frac{1\ ft}{12\ in} \times \frac{1\ mi}{5280\ ft} = \frac{50.\ mi}{hr}$

1.4 Percentages

Q1.4.1

a. $\% = \frac{part}{whole} \times 100 = \frac{8.2}{35} \times 100 = 23\%$
b. $\% = \frac{part}{whole} \times 100 = \frac{12}{56} \times 100 = 21\%$
c. $\% = \frac{part}{whole} \times 100 = \frac{38}{923} \times 100 = 4.1\%$
d. $\% = \frac{part}{whole} \times 100 = \frac{118}{342} \times 100 = 34.5\%$

Q1.4.2

a.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 42\% &= \frac{part}{94} \times 100 \\ part &= 39 \end{aligned}$$

b.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 83\% &= \frac{part}{239} \times 100 \\ part &= 198 = 2.0 \times 10^2 \end{aligned}$$

c.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 16\% &= \frac{part}{45} \times 100 \\ part &= 7.2 \end{aligned}$$

d.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 38\% &= \frac{part}{872} \times 100 \\ part &= 3.3 \times 10^2 \end{aligned}$$

Q1.4.3

a.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 34\% &= \frac{42}{whole} \times 100 \\ whole &= 1.2 \times 10^2 \end{aligned}$$

b.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 82\% &= \frac{73}{whole} \times 100 \\ whole &= 89 \end{aligned}$$

c.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 57\% &= \frac{13}{whole} \times 100 \\ whole &= 23 \end{aligned}$$

d.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 25\% &= \frac{75}{whole} \times 100 \\ whole &= 3.0 \times 10^2 \end{aligned}$$

e.
$$\begin{aligned} \% &= \frac{part}{whole} \times 100 \\ 15\% &= \frac{25}{whole} \times 100 \\ whole &= 1.7 \times 10^2 \end{aligned}$$

f. $$\begin{aligned}\% &= \frac{part}{whole} \times 100 \\ 76\% &= \frac{98}{whole} \times 100 \\ whole &= 129\end{aligned}$$

Q1.4.4

$$\% = \frac{part}{whole} \times 100$$
$$15.0\% = \frac{part}{182\ pounds} \times 100$$
$$part = 27.3\ pounds\ lost$$

$$182\ pounds - 27.3\ pounds = 154.7\ pounds = 155\ pounds$$

Q1.4.5

$$\% = \frac{part}{whole} \times 100$$
$$\% = \frac{12\ pounds}{135\ pounds} \times 100$$
$$\% = 8.9\%\ lost$$

Q1.4.6

$$\% = \frac{part}{whole} \times 100$$
$$25\% = \frac{part}{300.\ mg} \times 100$$
$$part = 75\ mg\ more$$

Q1.4.7

Convert both weights to ounces, find the ounces lost, and then find the percent lost.

Birth weight: $(7\ pounds \times 16) + 1\ ounce = 113\ ounces$

Discharge weight: $(6\ pounds \times 16) + 13\ ounces = 109\ ounces$

Weight lost: $113\ ounces - 109\ ounces = 4\ ounces$

Percent lost from original brith weight.

$$\% = \frac{part}{whole} \times 100$$
$$\% = \frac{4\ ounces}{113\ ounces} \times 100$$
$$\% = 3.5\% = 4\%$$

Q1.4.8

$$\% = \frac{part}{whole} \times 100$$
$$20\% = \frac{part}{125\ mg} \times 100$$
$$part = 25\ mg\ lost$$

$$125\ mg - 25\ mg = 100.\ mg$$

CHAPTER OVERVIEW

2: Elements and Ions

Chemistry is the study of matter and the changes it undergoes. Atoms are the building blocks of matter and are the smallest unit of an element.

2.1: Isotopes and Atomic Mass

Learning Outcomes

- Define atomic and mass numbers.
- Determine the number of protons, neutrons, and electrons in an atom.
- Identify the charge and relative mass of subatomic particles.
- Label the location of subatomic particles in the atom.
- Define isotope.
- Write the isotopic symbol of an atom.
- Explain the concept of average atomic mass.

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

Atomic Number

The **atomic number (Z)** *of an element is the number of protons in the nucleus of each atom of that element*. An atom can be classified as a particular element based solely on its atomic number. For example, any atom with an atomic number of 8 (its nucleus contains 8 protons) is an oxygen atom, and any atom with a different number of protons would be a different element. The periodic table (see figure below) displays all of the known elements and is arranged in order of increasing atomic number. In this table, an element's atomic number is indicated above the elemental symbol. Hydrogen, at the upper left of the table, has an atomic number of 1. Every hydrogen atom has one proton in its nucleus. Next on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, beryllium atoms have four, and so on.

Since atoms are neutral, the number of electrons in an atom is equal to the number of protons. Hydrogen atoms all have one electron occupying the space outside of the nucleus. Helium, with two protons, will have two electrons.

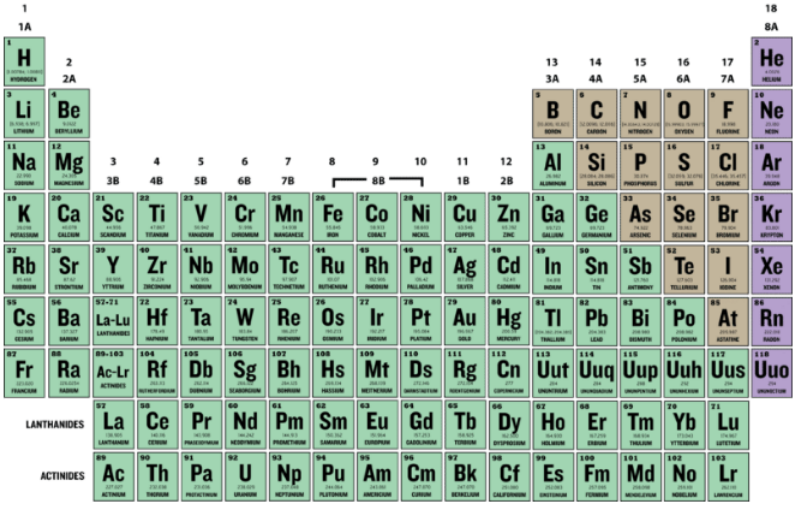

Figure 2.1.1: The periodic table of the elements.

Mass Number

Experimental data showed that the vast majority of the mass of an atom is concentrated in its nucleus, which is composed of protons and neutrons. The **mass number** *is defined as the total number of protons and neutrons in an atom.* Consider the table below, which shows data from the first six elements of the periodic table.

Table 2.1.1: Atoms of the First Six Elements

Name	Symbol	Atomic Number	Protons	Neutrons	Electrons	Mass Number

Name	Symbol	Atomic Number	Protons	Neutrons	Electrons	Mass Number
hydrogen	H	1	1	0	1	1
helium	He	2	2	2	2	4
lithium	Li	3	3	4	3	7
beryllium	Be	4	4	5	4	9
boron	B	5	5	6	5	11
carbon	C	6	6	6	6	12

View animations showing the atomic structure of the first 11 elements on the periodic table at http://web.visionlearning.com/custom...imations.shtml

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

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

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number} \tag{2.1.1}$$

Atoms of the element chromium (Cr) have an atomic number of 24 and a mass number of 52. How many neutrons are in the nucleus of a chromium atom? To determine this, you would subtract as shown:

$$52-24=28 \text{ neutrons in a chromium atom} \tag{2.1.2}$$

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

$$^{52}_{24}\text{Cr} \tag{2.1.3}$$

Another way to refer to a specific atom is to write the mass number of the atom after the name, separated by a hyphen. The above atom would be written as chromium-52, with the mass number written after the name. The atomic number does not have to be included because all atoms of chromium have the same number of protons but can vary in the atomic mass.

Isotopes

As stated earlier, not all atoms of a given element are identical. Specifically, the number of neutrons in the nucleus can vary for many elements. As an example, naturally occurring carbon exists in three forms, which are illustrated in the figure below.

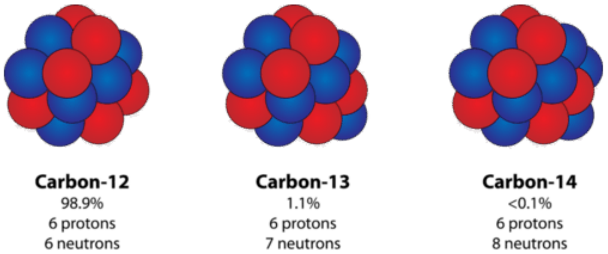

Figure 2.1.2: Nuclei of the three isotopes of carbon. Almost 99% of naturally occurring carbon is carbon-12, whose nucleus consists of six protons and six neutrons. Carbon-13 and carbon-14, with seven or eight neutrons, respectively, have a much lower nuclei abundance.

Each carbon atom has the same number of protons (6), which is equal to its atomic number. Each carbon atom also contains six electrons, allowing the atom to remain electrically neutral. However, the number of neutrons varies from six to eight. **Isotopes** *are atoms that have the same atomic number but different mass numbers due to a change in the number of neutrons.* The three isotopes of carbon can be referred to as carbon-12 $\left({}^{12}_{6}\mathrm{C}\right)$, carbon-13 $\left({}^{13}_{6}\mathrm{C}\right)$, and carbon-14 $\left({}^{14}_{6}\mathrm{C}\right)$. Naturally occurring samples of most elements are mixtures of isotopes. Carbon has only three natural isotopes, but some heavier elements have many more. Tin has ten stable isotopes, which is the most of any known element. The nucleus of a given carbon atom will be one of the three possible isotopes discussed above.

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

Size of Atoms

The graphite in your pencil is composed of the element carbon, a nonmetal. Imagine taking a small piece of carbon and grinding it until it is a fine dust. Each speck of carbon would still have all of the physical and chemical properties of carbon. Now imagine that you could somehow keep dividing the speck of carbon into smaller and smaller pieces. Eventually, you would reach a point where your carbon sample is as small as it could possibly be. This final particle is called an atom.

Atoms, as you probably know, are extremely small. In fact, the graphite in an ordinary pencil contains about 5×10^{20} atoms of carbon. This is an almost incomprehensibly large number. The population of the entire Earth is about 7×10^{9} people, meaning that there are about 7×10^{10} times as many carbon atoms in your pencil as there are people on Earth! For this to be true, atoms must be extremely small. We can only see atoms with a modern instrument called a scanning tunneling microscope. (www.nobelprize.org/educationa...opes/scanning/)

Atomic Mass

The masses of individual atoms are very, very small. However, using a modern device called a mass spectrometer, it is possible to measure such miniscule masses. An atom of oxygen-16, for example, has a mass of $2.66 \times 10^{-23}\ \mathrm{g}$. While comparisons of masses measured in grams would have some usefulness, it is far more practical to have a system that will allow us to more easily compare relative atomic masses. Scientists decided on using the carbon-12 nuclide as the reference standard by which all other masses would be compared. By definition, one atom of carbon-12 is assigned a mass of exactly 12 atomic mass units (amu). An **atomic mass unit** *is defined as a mass equal to one twelfth of an atom of carbon-12.* The mass of any isotope of any element is expressed in relation to the carbon-12 standard. For example, one atom of helium-4 has a mass of $4.0026\ \mathrm{amu}$ An atom of sulfur-32 has a mass of $31.972\ \mathrm{amu}$

The carbon-12 atom has six protons and six neutrons in its nucleus for a mass number of 12. Since the nucleus accounts for nearly all of the mass of the atom, a single proton or single neutron has a mass of approximately $1\ \mathrm{amu}$. However, as seen by the helium and sulfur examples, the masses of individual atoms are not quite whole numbers. This is because an atom's mass is affected very slightly by the interactions of the various particles within the nucleus and also includes the small mass added by each electron.

As stated in the section on isotopes, most elements occur naturally as a mixture of two or more isotopes. Listed below (see table below) are the naturally occurring isotopes of several elements along with the percent natural abundance of each.

Table 2.1.2: Atomic Masses and Percent Abundances of Some Natural Isotopes

Element	**Isotope (Symbol)**	**Percent Natural Abundance**	**Atomic mass** (amu)	**Average atomic mass** (amu)
Hydrogen	${}^{1}_{1}\mathrm{H}$	99.985	1.0078	1.0079
	${}^{2}_{1}\mathrm{H}$	0.015	2.0141	
	${}^{3}_{1}\mathrm{H}$	negligible	3.0160	
Carbon	${}^{12}_{6}\mathrm{C}$	98.89	12.000	12.011
	${}^{13}_{6}\mathrm{C}$	1.11	13.003	
	${}^{14}_{6}\mathrm{C}$	trace	14.003	
Oxygen	${}^{16}_{8}\mathrm{O}$	99.759	15.995	15.999
	${}^{17}_{8}\mathrm{O}$	0.037	16.995	

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic mass (amu)	Average atomic mass (amu)
	$^{18}_{8}\mathrm{O}$	0.204	17.999	
Chlorine	$^{35}_{17}\mathrm{Cl}$	75.77	34.969	35.453
	$^{37}_{17}\mathrm{Cl}$	24.23	36.966	
Copper	$^{63}_{29}\mathrm{Cu}$	69.17	62.930	63.546
	$^{65}_{29}\mathrm{Cu}$	30.83	64.928	

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

$$\frac{34.969+36.966}{2}=35.968 \tag{2.1.4}$$

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

$$0.7577\,(34.969)+0.2423\,(36.966)=35.453 \tag{2.1.5}$$

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

Supplemental Resources

- Elements and Atoms: http://www.youtube.com/watch?v=IFKnq9QM6_A
- Introduction to the Atom: http://www.youtube.com/watch?v=1xSQIwWGT8M
- Atomic History - A Brief Discovery: http://www.pbs.org/wgbh/nova/diamond...dehistory.html
- All About Atoms: http://education.jlab.org/atomtour/index.html
- Build and Atom: http://phet.colorado.edu/en/simulation/build-an-atom
- Molecular Workbench - Atomic Structure: http://workbench.concord.org/databas...vities/47.html
- See Inside a Diamond: http://www.pbs.org/wgbh/nova/diamond/insidewave.html
- Isotopes and Atomic Mass: http://phet.colorado.edu/en/simulati...nd-atomic-mass
- Atomic Structure: freezeray.com/flashFiles/atomcStructure.htm
- Atom Builder: freezeray.com/flashFiles/atomBuilder.htm
- Tennis Ball Isotopes: http://www.youtube.com/watch?v=oLnuXpf4hsA
- Element Math Game: http://education.jlab.org/elementmath/index.html
- Atoms and Matter Crossword Puzzle: http://education.jlab.org/sciencecro.../atoms_01.html
- Atomic Number Review #1: www.sciencegeek.net/Chemistry...micNumbers.htm
- Atomic Number Review #2: www.sciencegeek.net/Chemistry...t1Numbers2.htm
- Atomic Structure: www.sciencegeek.net/Chemistry...omicStructure/

Contributors and Attributions

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2.2: Matter

Learning Outcomes

- Define matter.
- Classify pure substances as elements or compounds.
- Classify mixtures as homogeneous or heterogeneous.
- Know the names and symbols of elements indicated in "Things to Memorize" on Canvas.
- Distinguish among metals, nonmetals, and metalloids.

Living things are made of **matter**. In fact, matter is the "stuff" of which all things are made (see figure below. Anything that occupies space and has mass is known as matter. Matter, in turn, consists of chemical substances. **Chemistry** is the study of matter and the changes it undergoes.

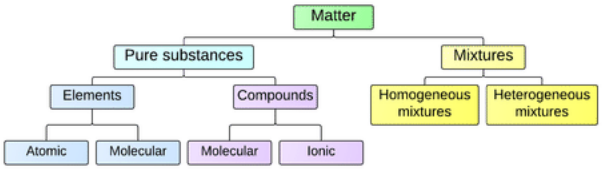

Figure 2.2.1: Matter can be categorized as pure substances or mixtures.

Pure Substances

A pure substance is a material that has a definite chemical composition. It is also homogenous, so the same chemical composition is found uniformly throughout the substance. A pure substance may be an element or a chemical compound.

Elements

An **element** is a pure substance that cannot be broken down into different types of substances. Examples of elements include carbon, oxygen, hydrogen, and iron. Each element is made up of just one type of atom. An atom is the smallest particle of an element that still characterizes the element. As shown in the figure below, at the center of an atom is a nucleus. The nucleus contains positively charged particles called protons and electrically neutral particles called neutrons. Surrounding the nucleus is a much larger electron cloud consisting of negatively charged electrons. An atom is electrically neutral if it has the same number of protons as electrons. Each element has atoms with a characteristic number of protons. For example, all carbon atoms have six protons, and all oxygen atoms have eight protons.

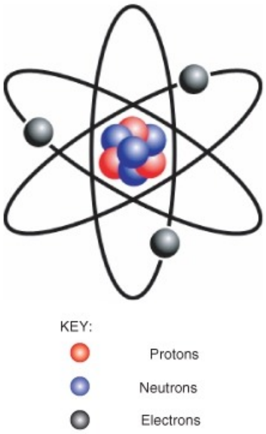

Figure 2.2.2: Model of an atom. The protons and neutrons of this atom make up its nucleus. Electrons surround the nucleus. KEY: Red = protons, Blue = neutrons, Black = electrons.

There are almost 120 known elements (see figure below) and each is represented in the periodic table by a one or two letter symbol. The majority of known elements are classified as metals. Metals are elements that are lustrous, or shiny. They are also good conductors of electricity and heat. Examples of metals include iron, gold, and copper. Fewer than 20 elements are classified as nonmetals. Nonmetals lack the properties of metals. Examples of nonmetals include oxygen, hydrogen, and sulfur. Certain other elements have properties of both metals and nonmetals. They are known as metalloids. Examples of metalloids include silicon and boron.

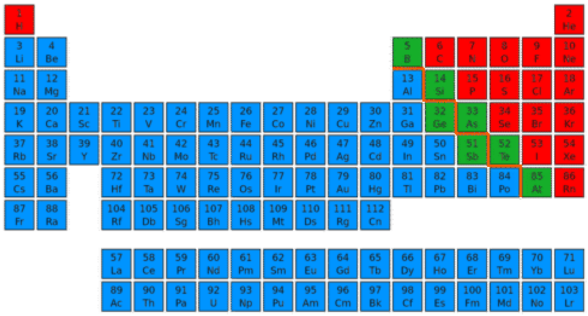

Figure 2.2.3: Periodic table of the elements showing metals (blue), nonmetals (red), and metalloids (green).

The New Periodic Table Song can be heard at https://youtu.be/zUDDiWtFtEM

Tom Lehrer performed the original *The Element Song* in 1967. You can hear it at https://youtu.be/DYW50F42ss8.

Chemical Compounds

A **chemical compound** is a pure substance that forms when atoms of two or more elements react with one another. A compound always has a unique and fixed chemical composition and the atoms of a compound are held together by chemical bonds. There are different types of chemical bonds, and they vary in how strongly they hold together the atoms of a compound. Two types of bonds are covalent and ionic bonds. Covalent bonds form when atoms *share* electrons and occur between two or more nonmetals. Ionic bonds form when electrons are *transferred* from one atom to another and usually form between a metal and a nonmetal.

An example of a chemical compound is water. A water molecule forms when oxygen (O) and hydrogen (H) atoms react and are held together by covalent bonds. Like other compounds, water always has the same chemical composition: a 2:1 ratio of hydrogen atoms to oxygen atoms. This is expressed in the chemical formula for water, $\mathrm{H_2O}$. The ratio of elements in a compound is given by the chemical formula. For example, NaCl has a 1:1 ratio between sodium and chlorine atoms. The absence of a subscripted number within the formula indicates that there is one of that element.

Mixtures

Like a chemical compound, a **mixture** consists of more than one chemical substance. Unlike a compound, a mixture does not have a fixed chemical composition. The substances in a mixture can be combined in any proportions. One characteristic of mixtures is that they can be separated into their components by physical methods. Since each part of the mixture has not reacted with another part of the mixture, the identities of the different substances remain unchanged.

The following examples illustrate these differences between mixtures and compounds. Both examples involve the same two elements: the metal iron (Fe) and the nonmetal sulfur (S).

- When iron fillings and sulfur powder are mixed together in any ratio, they form a mixture (see figure below). No chemical reaction occurs, and both elements retain their individual properties. A magnet can be used to physically separate the two elements by attracting the iron fillings out of the mixture and leaving the sulfur behind.
- When iron and sulfur are mixed together in a certain ratio and *heated*, a chemical reaction occurs. This results in the formation of a unique new compound, called iron (II) sulfide (FeS) (see figure below). A magnet cannot be used to mechanically separate the iron from the iron (II) sulfide because neither iron nor sulfur exist in the compound. Instead, another chemical reaction is required to separate the iron and sulfur.

Figure 2.2.4: Iron and sulfur as a mixture (left) and separated by physical means (right).

Figure 2.2.5: Iron (II) sulfide, FeS, is a chemical compound.

Homogeneous Mixtures

A **homogenous mixture** is a mixture in which the composition is uniform throughout the mixture. A mixture of salt and water is homogenous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same, while the composition of a mixture may vary. The amount of salt in the salt water can vary from one sample to another, while water, for example, always has the same composition.

Heterogeneous Mixtures

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. One example of a heterogeneous mixture is vegetable soup. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup. Another example of a heterogeneous mixture is soil which is composed of a variety of substances and is often of different composition depending on the sample taken. One shovel may come up with dirt and grass while the next shovel could contain an earthworm.

See more examples of heterogeneous mixtures at www.buzzle.com/articles/homog...-examples.html

Supplemental Resources

Learn More

- David Bodanis, *$E = mc^2$: A Biography of the World's Most Famous Question*. Walker and Co., 2005
- John Emsley, *Nature's Building Blocks: An A-Z Guide to the Elements*. Oxford University Press, 2003.
- Nevin Katz, *Elements, Compounds, and Mixtures: Middle and High School (Mr. Birdley Teaches Science)*. Incentive Publications, 2007.
- The Science of Macaroni Salad: What's in a Mixture? http://youtu.be/Vt7IN4QPU0k
- Heterogeneous mixtures antoine.frostburg.edu/chem/se...ogeneous.shtml
- Element Flash Cards http://education.jlab.org/elementflashcards/

Contributors and Attributions

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2.3: Mole and Molar Mass

Learning Outcomes

- Define mole.
- Determine the ratio of elements in a compound from a formula.
- Determine molar mass of an element and compound.
- Convert among mass, moles, and number of particles of a substance.

Figure $2.3.1$: Dozen is a counting number that means 12 regardless of what is being counted.

When you go to the bakery and order a dozen doughnuts, both you and the baker know that means 12. Dozen is a counting number that is defined as 12. There are other counting numbers such as pair (2), ream (500), gross (144), and score (20). The last one was famously used by Abraham Lincoln in the Gettysburg Address when he said "Four score and seven years ago". All of these words define a specific number of things regardless of what it is we are counting. A mole is also a counting number.

Avogadro's Number

It certainly is easy to count bananas or to count elephants (as long as you stay out of their way). However, you would be counting grains of sugar from your sugar canister for a long. long time. Atoms and molecules are extremely small - far, far smaller than grains of sugar. Counting atoms and molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years without stopping you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists needed a name that can stand for a very large number of items. Amedeo Avogadro (1776 - 1856), an Italian scientist, provided just such a number. He is responsible for the counting unit of measure called the mole. A **mole** (abbreviated as mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like dozen or gross, it is a name that stands for a number. There are 6.02×10^{23} atoms in a mole of carbon. There are 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.

Figure 2.3.2: Italian scientist Amedeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry.

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and the are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as H_2O and CO_2 exist as molecules and so the molecule is their representative particle. For ionic compounds such as $NaCl$ and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.

Figure 2.3.3: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year.

Conversions Between Moles and Number of Particles

Using our unit conversion techniques, we can use the mole label to convert back and forth between the number of particles and moles.

Example 2.3.1

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

In this problem, we are given the number of atoms and asked to provide the number of moles. We need to determine what factor we will use to convert from atoms to moles. Avogadro's number gives us the relationship we need, $1 \text{ mol} = 6.02 \times 10^{23}$ atoms.

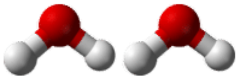
$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C} \tag{2.3.1}$$

Notice that atoms is on the bottom of the conversion factor because we need atoms to cancel so we are left with moles of carbon. Our final check is to make sure our answer is reasonable. Notice that we have 10^{24} atoms of carbon so it is reasonable that the value is greater than 1 mole.

Suppose that you wanted to know how many hydrogen atoms were in a mole of water molecules. First, you would need to know the chemical formula for water, which is H_2O. Based on the subscript for hydrogen, there are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen would there be in two water molecules? (see figure below) There would be $2 \times 2 = 4$ hydrogen atoms. How about in a dozen? In that case a dozen is 12 so $12 \times 2 = 24$ hydrogen atoms in a dozen water molecules. To get the answers (4 and 24), you had to multiply the given number of molecules by two atoms of hydrogen per molecule. So to find the number of hydrogen atoms in a mole of water molecules, the problem could be solved in a similar manner.

$$1 \text{ mol } H_2O \times \frac{6.02 \times 10^{23} \text{ molecules } H_2O}{1 \text{ mol } H_2O} \times \frac{2 \text{ atoms H}}{1 \text{ molecule } H_2O} = 1.20 \times 10^{24} \text{ atoms H} \tag{2.3.2}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.



Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms.

Example 2.3.2

Sulfuric acid has the chemical formula H_2SO_4. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid is in the sample?

Solution

In this problem, the number of atoms of oxygen is given. However, the number of molecules of H_2SO_4 is not equal to the number of atoms. Therefore, an additional step will be need to solve the problem.

We can write two relationships that will help solve this problem.

- $1 \text{ mol} = 6.02 \times 10^{23} \text{ molecules}$
- $1 \text{ molecule } H_2SO_4 = 4 \text{ O atoms}$

While 1 mole will always equal 6.02×10^{23}, the second relationship will vary depending on the identity of the compound and the element. For H_2SO_4, we could also write that 1 molecule of H_2SO_4 = 2 atoms H or 1 molecule H_2SO_4 = 1 atom S. Both of these are true, but they just aren't helpful for this problem.

$$4.89 \times 10^{25} \text{ atoms O} \times \frac{1 \text{ molecule } H_2SO_4}{4 \text{ atoms O}} \times \frac{1 \text{ mol } H_2SO_4}{6.02 \times 10^{23} \text{ molecules } H_2SO_4} = 20.3 \text{ mol } H_2SO_4 \tag{2.3.3}$$

Note that atoms O and molecules H_2SO_4 both cancel and we are left with units of mol H_2SO_4.

An alternative way to set up this problem is to use the relationship $1 \text{ mol } H_2SO_4 = 4 \text{ mol O atoms}$. Since a mole is just a counting number, we are simply multiplying both values (1 molecule and 1 atoms) by Avogadro's number to get a relationship in terms of moles. In this example, we convert from atoms to moles of O, then convert from moles of O to moles of H_2SO_4.

$$4.89 \times 10^{25} \text{ atoms O} \times \frac{1 \text{ mol O}}{6.02 \times 10^{23} \text{ atoms O}} \times \frac{1 \text{ mol } H_2SO_4}{4 \text{ mol O}} = 20.3 \text{ mol } H_2SO_4 \tag{2.3.4}$$

Both methods are correct and will give you the same answer. Use the method that makes the most sense to you.

Molar Mass

Molar mass is defined as the mass (in grams) of one mole of representative particles of a substance. By looking at a periodic table, we can conclude that the molar mass of lithium is 6.94 g, the molar mass of zinc is 65.38 g, and the molar mass of gold is 196.97 g. Each of these quantities contains 6.02×10^{23} atoms of that particular element. The units for molar mass are grams per mole or g/mol. Notice that these are the same numbers as the atomic mass but with different units. For example 1 atom of lithium has a mass of 6.94 amu while 1 mole $\left(6.02 \times 10^{23} \text{ atoms}\right)$ has a mass of 6.94 g. We can use this relationship to determine the moles of an element from its mass or vice versa.

Example 2.3.3

How many moles of carbon are in a 29.3 g sample?

Solution

We are given the mass of the sample and look up the molar mass of carbon in the periodic table which is 12.01 g/mol. Now, we can set up the calculation to solve for moles.

$$29.3 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.44 \text{ mol C} \tag{2.3.5}$$

Note that grams will cancel and we will be left with units of moles of carbon for our answer which is what is being asked for in this problem. Since the mass of the carbon sample is greater than the molar mass, it is reasonable that we have more than one mole of the carbon.

Example 2.3.4

What is the mass of aluminum in a 1.95 molsample?

Solution

As in the previous example, we will need the molar mass to solve the problem. For aluminum, the molar mass is 27.0 g/mol Now, we can set up the calculation to solve for mass.

$$1.95 \text{ mol Al} \times \frac{27.0 \text{ g Al}}{1 \text{ mol Al}} = 52.7 \text{ g Al} \tag{2.3.6}$$

In this example, we put gras on top and moles on the bottom so that the units would cancel correctly. Note that mol Al will cancel and we are left with units of grams of aluminum which is what is being asked for in the problem. The answer looks reasonable because we have almost two moles of aluminum and the mass is almost twice the molar mass.

Molar Masses of Compounds

The molecular formula of the compound carbon dioxide is CO_2. One molecule of carbon dioxide consists of 1 atom of carbon and 2 atoms of oxygen. We can calculate the mass of one molecule of carbon dioxide by adding together the masses of 1 atom of carbon and 2 atoms of oxygen.

$$12.01 \text{ amu} + 2\,(16.00 \text{ amu}) = 44.01 \text{ amu} \tag{2.3.7}$$

The **molecular mass** of a compound is the mass of one molecule of that compound. The molecular mass of carbon dioxide is 44.01 amu

The **molar mass** of any compound is the mass in grams of one mole of that compound. One mole of carbon dioxide molecules has a mass of 44.01 g The molar mass is $44.01 \text{ g/mol}$for CO_2. For water, the molar mass is 18.02 g/mol In both cases, it is the mass of 6.02×10^{23} molecules.

Example 2.3.5

Calcium nitrate, $Ca(NO_3)_2$, is used as a component in fertilizer. Determine the molar mass of calcium nitrate.

Solution

The molar mass of a compound is found from the molar masses of the elements in the compound.

- $Ca = 40.08 \text{ g/mol}$
- $N = 14.01 \text{ g/mol}$
- $O = 16.00 \text{ g/mol}$

First we need to analyze the formula. Since the Ca lacks a subscript, there is one Ca atom per formula unit. The 2 outside the parentheses means that there are two nitrate ions per formula unit and each nitrate ion consist of one nitrogen atom and three oxygen atoms. Therefore, there are a total of $1 \times 2 = 2$ nitrogen atoms and $3 \times 2 = 6$ oxygen atoms per formula unit. Thus, 1 mol of calcium nitrate contains 1 mol of Ca atoms, 2 mol of N atoms, and 6 mol of O atoms.

Use the molar masses of each atom together with the number of atoms in the formula and add together.

$$1 \text{ mol Ca} \times \frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} = 40.08 \text{ g Ca} \tag{2.3.8}$$

$$2 \text{ mol N} \times \frac{14.01 \text{ g N}}{1 \text{ mol N}} = 28.02 \text{ g N} \tag{2.3.9}$$

$$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 96.00 \text{ g O} \tag{2.3.10}$$

$$\text{molar mass of} \text{Ca(NO}_3)_2 = 40.08 \text{ g} + 28.02 \text{ g} + 96.00 \text{ g} = 164.10 \text{ g/mol} \tag{2.3.11}$$

Conversions Between Moles and Mass

Like we converted between moles and mass of an element, we can also convert between moles and mass of a compound using the molar mass of the compound. This relationship is frequently used in the laboratory. Suppose that for a certain experiment you need 3.00 moles of calcium chloride (CaCl_2). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. The molar mass of CaCl_2 is 110.98 g/mol The relationship that can be used is then based on the equality that $1 \text{ mol} = 110.98 \text{ g CaCl}_2$. Dimensional analysis will allow you to calculate the mass of CaCl_2 that you should measure.

$$3.00 \text{ mol CaCl}_2 \times \frac{110.98 \text{ g CaCl}_2}{1 \text{ mol CaCl}_2} = 333 \text{ g CaCl}_2 \tag{2.3.12}$$

When you measure the mass of 333 g of CaCl_2, you are measuring 3.00 moles of CaCl_2.

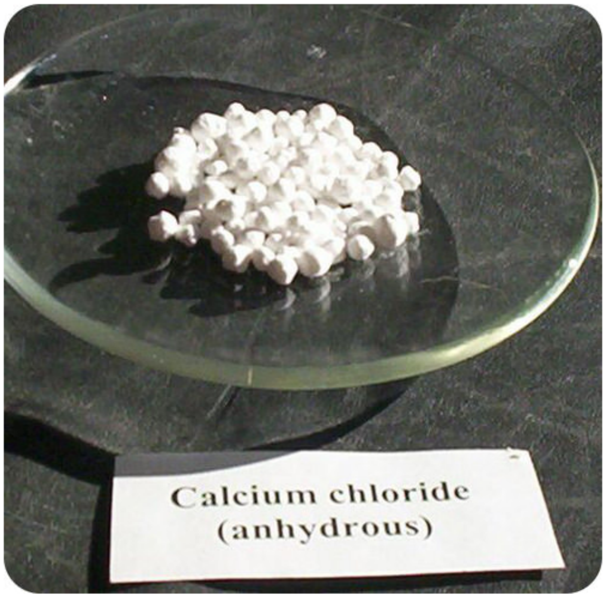


Figure 2.3.4: The mass of a substance, such as calcium chloride, is easy to measure.

Calcium chloride is used as a drying agent and as a road deicer.

Example 2.3.6

A certain reaction produces 2.81 g of copper (II) hydroxide, Cu(OH)_2. Determine the number of moles produced in the reaction.

Solution

The mass of the substance is known and we can determine the moles using the molar mass of the compound, Cu(OH)_2. First we have to find the molar mass of Cu(OH)_2 using the molar masses of the elements.

$$\text{molar mass Cu(OH)}_2 = 63.55 \text{ g/mol} + 2\,(16.00 \text{ g/mol}) + 2\,(1.008 \text{ g/mol}) = 97.57 \text{ g/mol} \tag{2.3.13}$$

Now, we can use the molar mass to convert from grams to moles.

$$2.81 \text{ g Cu(OH)}_2 \times \frac{1 \text{ mol Cu(OH)}_2}{97.57 \text{ g Cu(OH)}_2} = 0.0288 \text{ mol Cu(OH)}_2 \tag{2.3.14}$$

The units of grams will cancel and we are left with units of moles which is what is being requested in the question.

Example 2.3.7

What is the mass of water in a 3.50 mol sample?

Solution

In this example, the moles of water is known and we need to find the mass. Again, we will use the molar mass which was previously given as 18.02 g/mol

$$3.50 \text{ mol } H_2O \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 63.1 \text{ g } H_2O \tag{2.3.15}$$

The units of mol H_2O will cancel and we are left with units of grams which is what is being requested in the question.

Supplemental Resources

- Molar Mass Problems: http://misterguch.brinkster.net/molarmass.html

Contributors and Attributions

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

2.4: Electron Arrangements

Learning Outcomes

- Determine the energy levels of electrons for the first 20 elements.
- Explain the relevance of valence electrons in chemical processes.
- Identify the number of valence electrons in an element.
- Describe the stability of an atom as a result of following the octet rule.

The structure of the atom was discussed in the previous unit and now we will focus on the role that electrons play in the formation of compounds. Regardless of the type of compound or the number of atoms or electrons involved, it is the electrons of those atoms that interact to form a compound.

Electron Arrangement

Electrons are not randomly arranged in an atom and their position within the atom can be described using **electron arrangements**, which are a simplified version of electron configurations. For each element of interest, we look at the number of electrons in a single atom and then determine how those electrons are arranged based on the atomic model. The main idea behind electron arrangements is that electrons can only exist at certain energy levels. By understanding the energy levels of electrons in an atom, we can predict properties and understand behavior of the atom.

As shown in the figure below, there are multiple energy levels where electrons can be found. As the energy level increases, the energy difference between them decreases. A maximum of two electrons can be found in the $n = 1$ level; eight electrons can be in the $n = 2$ level. Although the $n = 3$ and $n = 4$ levels show only eight electrons in this diagram, those energy levels can hold more but not until we start looking at the transition metals. We will only be concerned with the electron arrangements of elements through calcium $(Z = 20)$ so we will put a maximum of eight electrons in the $n = 3$ level and two in the $n = 4$ level.

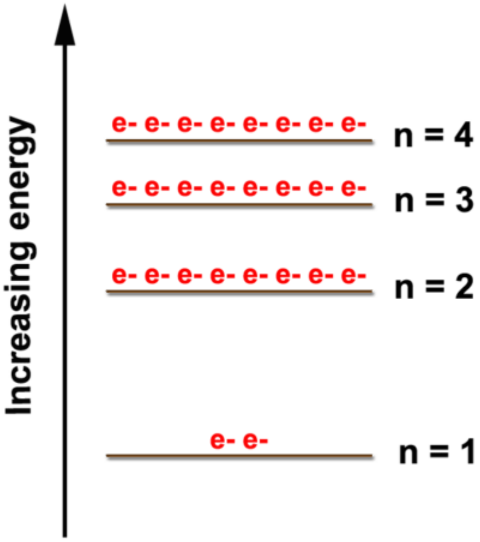

Figure 2.4.1: Energy levels of electrons.

Example 2.4.1

What is the electron arrangement of oxygen?

Solution

Oxygen has eight electrons. The first two electrons will go in the $n = 1$ level. Two is the maximum number of electrons for the level so the other electrons will have to go in a higher energy level. The $n = 2$ level can hold up to eight electrons so the remaining six electrons will go in the $n = 2$ level. The electron arrangement of oxygen is (2, 6).

Example 2.4.2

What is the electron arrangement of chlorine?

Solution

Chlorine has 17 electrons. The first two electrons will go in the $n = 1$ level. Two is the maximum number of electrons for the level so the other electrons will have to go in higher energy levels. The $n = 2$ level can hold up to eight electrons so the next 8 electrons will go in the $n = 2$ level. The remaining 7 electrons can go in the $n = 3$ level since it holds a maximum of 8 electrons. The electron arrangement of chlorine is (2, 8, 7).

The electron arrangement also provides information about the number of **valence electrons**. The valence electrons are the electrons in the highest energy level and the ones involved in ion and bond formation. Knowing the number of valence electrons will allow us to predict how a particular element will interact with other elements. Electrons in lower energy levels are called the **core electrons**.

Let's look at the figure below which shows the electron diagram for magnesium and its 12 electrons. The first two electrons are found in the $n = 1$ energy level, the next eight electrons are found in the $n = 2$ level, and the remaining two electrons are found in the $n = 3$ level. The electrons always fill the lowest energy levels available until that level is filled, then electrons fill the next energy level until it is filled. This continues for all of the electrons in an atom. We can show the electron arrangement as (2, 8, 2) representing the electrons in the $n = 1$, $n = 2$, and $n = 3$ levels, respectively.

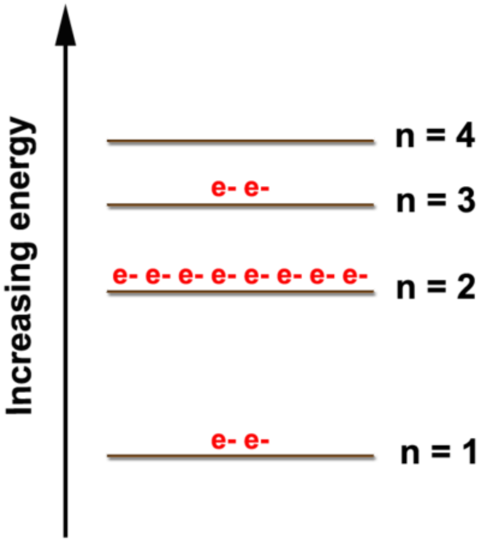

Figure 2.4.2: Electron diagram for magnesium.

The electron arrangement also shows the number of valence electrons which is two for magnesium because there are two electrons in the $n = 3$ energy level which is the highest occupied energy level for magnesium. This corresponds to the $2+$ charge formed when magnesium forms an ion. It is willing to lose 2 electrons so that it has the same electron arrangements as the nearest noble gas, which is neon (2, 8). Atoms will gain or lose electrons to look like the nearest noble gas because the noble gases are unreactive due to the stability of having eight electrons in the highest energy level. This desire of atoms to have eight electrons in their outermost shell is known as the **octet rule**.

Example 2.4.3

What is the electron arrangement of aluminum? How many valence electrons does it have?

Solution

Aluminum has 13 electrons so it will have the electron arrangement (2, 8, 3) which represents two electrons in the $n = 1$ energy level, eight electrons in the $n = 2$ level, and three electrons in the $n = 3$ level. Aluminum has three valence electrons (indicated by the three electrons in the $n = 3$ level).

Example 2.4.4

How many valence electrons does chlorine have? How many electrons will chlorine gain or lose to form an ion?

Solution

Chlorine has 7 electrons in its valence shell. To meet the octet rule, it must either gain one electron or lose seven electrons. Gaining one is easier than losing seven so it will gain one electron to have a total of eight electrons when it forms an ion (i.e. charged particle).

Contributors and Attributions

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2.5: Ion Formation

Learning Outcomes

- Define ion, cation, and anion.
- Recognize characteristics of monatomic and polyatomic ions.
- Explain how and why cations and anions are formed.
- Determine the number of subatomic particles in an ion.
- Predict the charge on ions formed from representative elements.

In many parts of the country, including parts of Kentucky, the water contains high concentrations of minerals that stain clothes, build up deposits on bathtubs and water heaters, and create problems with soap foaming properly. This problem is caused by what is called "hard water". The water contains excessive amounts of cations such as iron and calcium.

Ions

Ions are charged substances that have formed through the gain or loss of electrons. Cations form from the loss of electrons and have a positive charge while anions form through the gain of electrons and have a negative charge.

Cation Formation

Cations are the positive ions formed by the loss of one or more electrons. The most commonly formed cations of the representative elements are those that involve the loss of all of the valence electrons. Consider the alkali metal sodium (Na). It has one valence electron in the $n=3$ energy level. Upon losing that electron, the sodiu ion now has an octet of electrons from the second energy level and a charge of $1+$.

The electron arrangement of the sodium ion is now the same as that of the noble gas neon. Consider a similar process with magnesium and aluminum. In this case, the magnesium atom loses its two valence electrons in order to achieve the same arrangement as the noble gas neon and a charge of $2+$. The aluminum atom loses its three valence electrons to have the same electron arrangement as neon and a charge of $3+$. For representative elements under typical conditions, three electrons is usually the maximum number that will be los. Representative elements will not lose electrons beyond their valence because they would have to "break" the octet of the previous energy level which provides stability to the ion.

Anions

Anions are the negative ions formed from the gain of one or more electrons. When nonmetal atoms gain elections, they often do so until their outermost principal energy level achieves an octet. For fluorine, which has an electron arrangement of (2, 7), it only needs to gain one electron to have the same electron arrangement as neon. Forming an octet (eight electrons in the outer shell) provides stability to the atom. Fluorine will gain one electron and have a charge of $1-$. The electron arrangement of the fluoride ion (2, 8) will also change to reflect the gain of an electron.

Oxygen has an electron arrangement of (2, 6) and needs to gain two electrons to fill the $n=2$ energy level and achieve an octet of electrons in the outermost shell. The oxide ion will have a charge of $2-$ as a result of gaining two electrons. Under typical conditions, three electrons is the maximum that will be gained in the formation of anions.

Subatomic Particles in an Ion

Since ions form from the gain or loss of electrons, we can also look at the number of subatomic particles (protons, neutrons, and electrons) found in an ion. Remember that the number of protons determines the identity of the element and will not change in a chemical process.

Example 2.5.1

How many protons, neutrons, and electrons in a single oxide $\left(\mathrm{O}^{2-}\right)$ ion?

Solution

Oxygen has the atomic number 8 so both the atom and the ion will have 8 protons. The average atomic mass of oxygen is 16. Therefore, there will be 8 neutrons $(\text{atomic mass} - \text{atomic number} = \text{neutrons})$. A neutral oxygen atom would have 8 electrons. However, the anion has gained two electrons so O^{2-} has 10 electrons.

We can also use information about the subatomic particles to determine the identity of an ion.

Example 2.5.2

An ion with a $2+$ charge has 18 electrons. Determine the identity of the ion.

Solution

If an ion has a $2+$ charge then it must have lost electrons to form the cation. If the ion has 18 electrons and the atom lost 2 to form the ion, then the neutral atom contained 20 electrons. Since it was neutral, it must also have had 20 protons. Therefore the element is calcium.

Polyatomic Ions

A **polyatomic ion** is an ion composed of two or more atoms that have a charge as a group (poly = many). The ammonium ion (see figure below) consists of one nitrogen atom and four hydrogen atoms. Together, they comprise a single ion with a $1+$ charge and a formula of $\mathrm{NH_4^+}$. The hydroxide ion (see figure below) contains one hydrogen atom and one oxygen atom with an overall charge of $1-$. The carbonate ion (see figure below) consists of one carbon atom and three oxygen atoms and carries an overall charge of $2-$. The formula of the carbonate ion is $\mathrm{CO_3^{2-}}$. The atoms of a polyatomic ion are tightly bonded together and so the entire ion behaves as a single unit. The figures below show several examples.

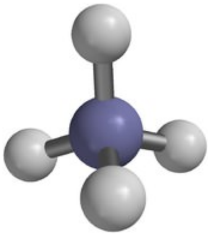

Figure $2.5.1$: The ammonium ion $\left(\mathrm{NH_4^+}\right)$ is a nitrogen atom (blue) bonded to four hydrogen atoms (white).

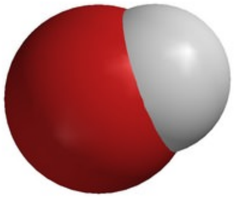

Figure $2.5.2$: The hydroxide ion $\left(\mathrm{OH^-}\right)$ is an oxygen atom (red) bonded to a hydrogen atom.

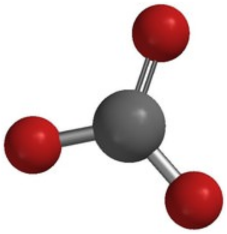

Figure $2.5.3$: The carbonate ion $\left(\mathrm{CO_3^{2-}}\right)$ is a carbon atom (black) bonded to three oxygen atoms.

The table below lists a number of polyatomic ions by name and by structure. The heading for each column indicates the charge on the polyatomic ions in that group. Note that the vast majority of the ions listed are anions - there are very few polyatomic cations.

Table $2.5.1$: Common Polyatomic Ions

$1-$	$2-$	$3-$	$1+$
acetate, $\mathrm{CH_3COO^-}$	carbonate, $\mathrm{CO_3^{2-}}$	arsenate, $\mathrm{AsO_3^{3-}}$	ammonium, $\mathrm{NH_4^+}$
bromate, $\mathrm{BrO_3^-}$	chromate, $\mathrm{CrO_4^{2-}}$	phosphite, $\mathrm{PO_3^{3-}}$	
chlorate, $\mathrm{ClO_3^-}$	dichromate, $\mathrm{Cr_2O_7^{2-}}$	phosphate, $\mathrm{PO_4^{3-}}$	

1−	2−	3−	1+
chlorite, ClO_2^-	hydrogen phosphate, HPO_4^{2-}		
cyanide, CN^-	oxalate, $C_2O_4^{2-}$		
dihydrogen phosphate, $H_2PO_4^-$	peroxide, O_2^{2-}		
hydrogen carbonate, HCO_3^-	silicate, SiO_3^{2-}		
hydrogen sulfate, HSO_4^-	sulfate, SO_4^{2-}		
hydrogen sulfide, HS^-	sulfite, SO_3^{2-}		
hydroxide, OH^-			
hypochlorite, ClO^-			
nitrate, NO_3^-			
nitrite, NO_2^-			
perchlorate, ClO_4^-			
permanganate, MnO_4^-			

The vast majority of polyatomic ions are anions, many of which end in *-ate* or *-ite*. Notice that in some cases such as nitrate $\left(NO_3^-\right)$ and nitrite $\left(NO_2^-\right)$, there are multiple anions that consist of the same two elements. In these cases, the difference between the ions is the number of oxygen atoms present, while the overall charge is the same. As a class, these are called oxyanions. When there are two oxyanions for a particular element, the one with the greater number of oxygen atoms gets the *-ate* suffix, while the one with the fewer number of oxygen atoms gets the *-ite* suffix. The four oxyanions of chlorine are shown below, which also includes the use of the prefixes *hypo-* and *per-*.

- ClO^-, hypochlorite
- ClO_2^-, chlorite
- ClO_3^-, chlorate
- ClO_4^-, perchlorate

Not your usual ion

"Drink you milk. It's good for your bones." We're told this from early childhood, and with good reason. Milk contains a good supply of calcium, part of the structure of bone. However, there are two other ionic components of hydroxyapatite, the mineral component. Phosphate ion and hydroxide ion make up the remainder of the inorganic material in bone.

News You Can Use

- Bone is a very complex structure. It is composed of protein (mainly collagen), hydroxyapatite (a calcium-phosphate-hydroxide mixture), some other minerals, and contains 10-20% water. The calcium/phosphate ratios are not stoichiometric, but vary somewhat from one portion of bone to the next.
- Bones are very strong but will break under enough stress. Regular exercise and proper nutrition help to increase bone strength. Watch a video about bone structure at http://www.youtube.com/watch?v=d9owEvYdouk

- Nitrate is an anion with a complex bonding structure. Major sources for this ion in drinking water are runoff from fertilizer, septic tank leakage, sewage, and natural deposits. High concentrations of nitrates represent a significant health hazard, especially to infants. The nitrate in the body is converted to nitrite, which then binds to hemoglobin. This binding decreases the ability of hemoglobin to transport oxygen, thus depriving the cells of the O_2 needed for proper functioning.
- Cyanide production is widespread throughout nature. Forest fires will produce significant amounts of cyanide. Many plants contain cyanide, and it is produced by a number of bacteria, algae, and fungi. Cyanide is used industrially in metal finishing, iron and steel mills, and in organic synthesis processes. This material is also an important component for the refining of precious metals. Formation of a complex between cyanide and gold allows extraction of this metal from a mixture.

Supplemental Resources

- Predict Ionic Charges www.800mainstreet.com/4/0004-001-ions.html
- Review Quiz of Protons, Neutrons, and Electrons in Ions https://www.proprofs.com/quiz-school/story.php?title=mtkyotk0na7nvu

Contributors and Attributions

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

2.6: Ionic Compounds

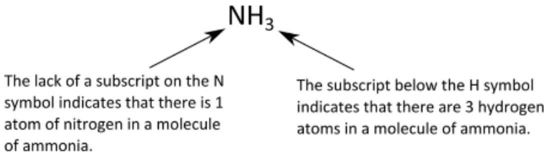
Learning Outcomes

- Describe the types of elements that form an ionic bond.
- Explain how an ionic bond is formed.
- Predict the formula of an ionic compound.

Chemistry can be broadly divided into tow main classes based on the identity of the elements present in the chemical compounds. **Organic chemistry** *is the branch of chemistry that deals with compounds containing carbon.* In this section and the following one, we will be discussing **inorganic chemistry**, *which is the branch of chemistry dealing with compounds that do not contain carbon.* Ionic compounds are examples of inorganic compounds.

Types of Formulas

Recall that a molecules includes two or more atoms that have been chemically combined. *A chemical formula that indicates how many of each type of atom are present in a single molecule* is referred to as a **molecular formula**. For example, a molecule of ammonia contains one nitrogen atom and three hydrogen atoms, so it has the following molecular formula.



Another type of chemical formula, the **empirical formula**, *shows the elements in a compound in their lowest whole-number ratio*. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O. When chemists analyze an unknown compound, often the first step is to determine its empirical formula. There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.

Unlike molecular compounds, ionic compounds are quite different. Water and other molecules exist as collections of individual molecules (see figure below) while ionic compounds do not exist as discrete molecular units. Instead,an ionic compound consists of a large three-dimensional array of alternating cations and anions. For example, sodium chloride (NaCl) is composed of many Na^+ and Cl^- ions arranged into a structure like the one pictured (see figure below). The formula for ionic compounds is always an empirical formula because it shows the smallest, whole-number ratio between the cations and anions. The actual number of ions in a sample of an ionic compound will be very large and will vary from sample to sample.

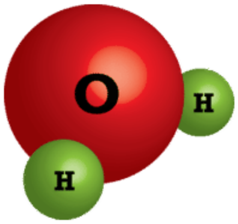


Figure 2.6.1: A water molecule consists of one atom of oxygen bonded to two atoms of hydrogen.

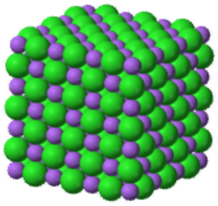


Figure 2.6.2: A crystal of table salt, sodium chloride, is a large array of alternating positive and negative ions. The purple spheres represent the Na^+ ions, while the green spheres represent the Cl^- ions.

The most straightforward way to describe this structure with a chemical formula is to give the lowest whole-number ratio between the two ions. In the case of NaCl, there are equal numbers of sodium ions and chloride ions in the salt crystal. In contrast, a crystal of magnesium chloride has twice as many chloride ions as magnesium ions, so it has a formula of $\mathrm{MgCl_2}$.

Writing Formulas for Binary Ionic Compounds

If you know the elements that form a binary ionic compound, you can write its formula. Start by writing the metal ion and its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many ions of each type are needed in order for the positive and negative charges to cancel each other out. Consider the compound aluminum nitride. The charges on each of these ions can be determined by looking at the groups in which aluminum and nitrogen are found on the periodic table. The ions are:

$$\mathrm{Al^{3\,+}} \qquad \mathrm{N^{3\,-}} \tag{2.6.1}$$

Since the ions have charges that are equal in magnitude (3, but different signs), 1:1 is the lowest ratio of ions that will produce a neutral compound. Since the charge on aluminum is 3+ and the charge on nitrogen is $3-$, the sum of their charges is zero ($+3\ +\ -3 = 0$). As a result, the formula of aluminum nitride is AlN. Another compound, lithium oxide, contains the following ions:

$$\mathrm{Li^{+}} \qquad \mathrm{O^{2\,-}} \tag{2.6.2}$$

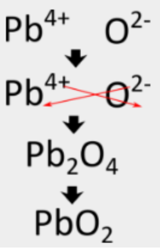
In this case, two lithium ions, each with a 1+ charge, are required to balance out the charge of each oxide ion, which has a $2-$ charge. The formula of lithium oxide is $\mathrm{Li_2O}$ because the compound must be neutral. Therefore, $(2\times +1)\ +(1\times -2)=0$.

For compounds in which the ratio of ions is not as obvious, an alternative way to determine the correct formula is to use the "crisscross" method. In this method, the numerical value of each charge crosses over to become the subscript of the opposite ion. The signs of the charges are dropped. The crisscross method is demonstrated below for aluminum oxide.

The red arrows indicate that the 3 from the $3+$ charge will cross over to become the subscript for O, while the 2 from the $2-$ charge will cross over to become the subscript for Al. The formula for aluminum oxide is $\mathrm{Al_2O_3}$.

For aluminum oxide, the crisscross method directly produces the correct formula, but in some cases, another step is required. Because ionic compounds are always described by their empirical formulas, they must be written as the lowest whole-number ratio of the ions. In the case of aluminum nitride, the crisscross method would yield a formula of $\mathrm{Al_3N_3}$, which is not correct. A second step must be performed in which the subscripts are reduced but the ratio is kept the same. $\mathrm{Al_3N_3}$ can be reduced to AlN, because both formulas describe a 1:1 ratio of aluminum ions to nitride ions. Following the crisscross method to write the formula for lead (IV) oxide would involve the following steps:

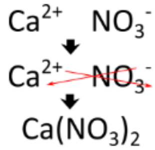
Some transition metals can have more than one possible charge. When this happens, the charge on the transition metal cation is included in parentheses in the name. For example, lead(IV) oxide has $\mathrm{Pb^{4\,+}}$ as its metal cation.



The crisscross method first yields $\mathrm{Pb_2O_4}$ for the formula, but that must be reduced to $\mathrm{PbO_2}$, which is the correct formula.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

Writing a formula for an ionic compound with polyatomic ions involves the same steps as for a binary (two element) ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion. Use the crisscross method to ensure that the final formula is neutral. Calcium nitrate is composed of calcium cations and nitrate anions.



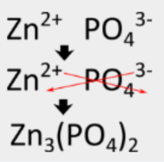
The charge is balanced by the presence of two nitrate ions and one calcium ion. Parentheses are used around the nitrate ion because more than one polyatomic ion is needed. If only one polyatomic ion is present in a formula, parentheses are not used. For example, the formula for calcium carbonate is $\mathrm{CaCO_3}$. The carbonate ion carries a $2-$ charge, so it exactly balances the $2+$ charge of the calcium ion and parentheses are not needed around the polyatomic ion.

Example 2.6.1

Write the formula for zinc phosphate which is composed of Zn^{2+} and PO_4^{3-} ions.

Solution

Write the metal cation followed by the nonmetal anion. Crisscross the ion charges in order to make the ionic compound neutral. Use parentheses around the polyatomic ion if more than one is present in the final formula. Reduce to the lowest ratio if necessary.



The formula for zinc phosphate is $\mathrm{Zn_3(PO_4)_2}$. Three zinc cations with $2+$ charges balance out two phosphate anions with $3-$ charges.

Contributors and Attributions

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2.7: Elements and Ions (Exercises)

These are homework exercises to accompany Chapter 2 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

2.1: Isotopes and Atomic Mass

(click here for solutions)

Q2.1.1

Give the names and symbols of three elements.

Q2.1.2

Describe where protons, neutrons, and electrons are located in an atom.

Q2.1.3

Why are all atoms electrically neutral?

Q2.1.4

How many protons are in the nucleus of each of the following atoms?

a. neon (Ne)
b. gold (Au)
c. strontium (Sr)
d. uranium (U)

Q2.1.5

A certain atom contains 28 protons, 28 electrons, and 31 neutrons. Provide the following:

a. atomic number
b. mass number
c. name of element

Q2.1.6

How many protons, neutrons, and electrons are in an atom of cesium-133?

Q2.1.7

How many protons, neutrons, and electrons are there in the atom $^{19}_{9}\text{F}$?

Q2.1.8

How many protons, neutrons, and electrons are there in an atom of lead-207?

Q2.1.9

A certain atom has an atomic number of 36 and a mass number of 84. Write out the designation for this isotope in both nuclide symbol form and in hyphenated form.

Q2.1.10

An atom has a mass number of 59 and contains 32 neutrons in its nucleus. What element is it?

Q2.1.11

Complete the **Table** below:

Isotope	Isotope Symbol	Atomic Number	Mass Number
sodium-23			

Isotope	Isotope Symbol	Atomic Number	Mass Number
	${}^{75}_{33}\mathrm{As}$		
silver-108			

Q2.1.12

Which one is an isotope of ${}^{40}_{18}\mathrm{Ar}$? Explain.

a. ${}^{40}_{20}\mathrm{Ca}$
b. ${}^{39}_{18}\mathrm{Ar}$
c. ${}^{40}_{18}\mathrm{Ar}$

Q2.1.13

Fill in **Table** below:

Isotope	Number of Protons	Number of Electrons	Number of Neutrons	Isotope Symbol
hydrogen-1				
hydrogen-2				
beryllium-9				
aluminum-27				

Q2.1.14

Fill in **Table** below:

Element	Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Isotope Symbol
nitrogen			14				
	B		11				
		30				35	
					77	116	
							${}^{56}_{26}\mathrm{Fe}$

2.2: Matter

(click here for solutions)

Q2.2.1

Define matter and give three examples of matter.

Q2.2.2

Explain the differences between compounds and mixtures.

Q2.2.3

Explain the differences between pure substances and mixtures.

Q2.2.4

Identify each of the following as a pure substance or a mixture.

a. salt water

b. NaCl
c. brewed coffee
d. air

Q2.2.5

Label each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.

a. silicon
b. sulfuric acid
c. air
d. soda
e. sugar
f. muddy water
g. chicken noodle soup
h. scoop of sand from the beach

Q2.2.6

Identify each of the following elements as a metal, nonmetal, or metalloid.

a. iron (Fe)
b. gallium (Ga)
c. phosphorus (P)
d. potassium (K)
e. neon (Ne)
f. germanium (Ge)

Q2.2.7

Identify each of the following elements as a metal, nonmetal, or metalloid.

a. chlorine (Cl)
b. hydrogen (H)
c. antimony (Sb)
d. titanium (Ti)
e. nitrogen (N)
f. selenium (Se)

Q2.2.8

Describe how you identify molecular and ionic compounds.

Q2.2.9

Label each as an ionic or molecular compound.

a. H_2O_2
b. N_2O_5
c. PF_3
d. $MgCl_2$
e. KBr
f. $AlCl_3$

Q2.2.10

Label each as an ionic or molecular compound.

a. CaO
b. Na_2S
c. NH_3
d. CO_2
e. N_2H_4

f. N_2O

2.3: Mole and Molar Mass

(click here for solutions)

Q2.3.1

Define "counting number".

Q2.3.2

What is the value of Avogadro's number?

Q2.3.3

What seven elements exist as diatomic molecules in nature?

Q2.3.4

How many atoms of helium are present in each of the following samples?

a. 1 mole
b. 2 moles
c. 2.5 moles
d. 0.5 moles
e. 0.35 moles

Q2.3.5

How many molecules of water are present in each of the following samples?

a. 1 mole
b. 2 moles
c. 2.5 moles
d. 0.5 moles
e. 0.35 moles

Q2.3.6

Compare the answers for each part of questions 4 and 5 to one another. How do they compare? Explain why.

Q2.3.7

How many moles of silicon is 6.73×10^{25} atoms of silicon?

Q2.3.8

How many moles of sodium is 4.29×10^{22} atoms of sodium?

Q2.3.9

How many atoms of each element are in one unit of each compound?

a. H_2O_2
b. N_2O_5
c. PF_3
d. $MgCl_2$
e. KBr
f. $AlCl_3$
g. CaO
h. Na_2S
i. NH_3
j. CO_2
k. N_2H_4
l. N_2O

Q2.3.10

How many moles of each element are in one mole of each compound?

a. H_2O_2
b. N_2O_5
c. PF_3
d. $MgCl_2$
e. KBr
f. $AlCl_3$
g. CaO
h. Na_2S
i. NH_3
j. CO_2
k. N_2H_4
l. N_2O

Q2.3.11

How do the answers in questions 9 and 10 compare to one another? Explain the similarities or differences.

Q2.3.12

How many moles of carbon are in 0.75 moles of CCl_4? How many moles of chlorine?

Q2.3.13

How many atoms of carbon are in 0.75 moles of CCl_4? How many atoms of chlorine?

Q2.3.14

How many moles of hydrogen are in 2.5 moles of H_2O? How many moles of oxygen?

Q2.3.15

How many atoms of hydrogen are in 2.5 moles of H_2O? How many atoms of oxygen?

Q2.3.16

A sample of $CaNO_3$ contains 3.87 x 10^{25} atoms of oxygen. How many molecules of $CaNO_3$ are in the sample?

Q2.3.17

A sample of propane gas (C_3H_8) contains 5.39 x 10^{24} atoms of carbon. How many atoms of hydrogen are in the sample?

Q2.3.18

What is the molar mass of each of the following elements (in atomic form)?

a. carbon
b. nitrogen
c. sodium
d. hydrogen
e. potassium
f. phosphorus

Q2.3.19

How many moles of each element listed in the previous question are present in a 25.0 g sample of the element?

Q2.3.20

For question 19, all of the samples have the same mass. Are the moles the same? Why or why not?

Q2.3.21

What is the mass of each of the following samples?

a. 0.35 moles sodium

b. 0.75 moles carbon
c. 1.34 moles potassium
d. 1.21 moles silicon
e. 0.95 moles calcium
f. 2.85 moles helium

Q2.3.22

Determine the molar mass of each of the following compounds?

a. CO_2
b. N_2H_4
c. CaF_2
d. $C_6H_{12}O_6$
e. CH_4
f. C_6H_6
g. Na_2SO_4
h. K_3PO_4
i. $Al(NO_3)_3$
j. $Mg_3(PO_4)_2$

Q2.3.23

Calculate the moles of each of the following samples.

a. 25.0 g CO_2
b. 10.0 g N_2H_4
c. 85.0 g CaF_2
d. 15.5 g $C_6H_{12}O_6$
e. 20.0 g CH_4
f. 100.0 g C_6H_6
g. 30.0 g Na_2SO_4
h. 75.0 g K_3PO_4
i. 50.0 g $Al(NO_3)_3$
j. 47.2 g $Mg_3(PO_4)_2$

Q2.3.24

Calculate the mass of each of the following samples.

a. 3.5 mol CO_2
b. 0.45 mol N_2H_4
c. 2.25 mol CaF_2
d. 1.75 mol $C_6H_{12}O_6$
e. 4.9 mol CH_4
f. 8.75 mol C_6H_6
g. 2.35 mol Na_2SO_4
h. 0.672 mol K_3PO_4
i. 0.95 mol $Al(NO_3)_3$
j. 1.15 mol $Mg_3(PO_4)_2$

2.4: Electron Arrangements

(click here for solutions)

Q2.4.1

What is the electron arrangement for each of the elements?

a. Na
b. Ne

c. Be
d. N
e. S
f. Cl

Q2.4.2

How many valence electrons are in each element?

a. K
b. P
c. F
d. S
e. Li
f. B

Q2.4.3

What is the octet rule?

2.5: Ion Formation

(click here for solutions)

Q2.5.1

Define ion.

Q2.5.2

How are anions and cation the same? Different?

Q2.5.3

What is the most common ion formed from each element?

a. Li
b. Na
c. Ca
d. B
e. P
f. S
g. Cl
h. Br

Q2.5.4

How many protons, neutrons, and electrons are present in the ions indicated in the previous question?

Q2.5.5

Identify the following elements.

a. An ion with a 3+ charge and two electrons.
b. An ion with a 1− charge and 18 electrons.
c. An ion with a 1+ charge and 18 electrons.
d. An ion with a 3− charge and 10 electrons.

Q2.5.6

Describe a polyatomic ion.

Q2.5.7

Which are polyatomic ions?

a. NO_3^-

b. O^{2-}
c. NH_4^+
d. Mg^{2+}
e. Na^+
f. O_2^{2-}

2.6: Ionic Compounds

(click here for solutions)

Q2.6.1

What element is present in all organic compounds?

Q2.6.2

Give three examples of metallic elements.

Q2.6.3

Give three examples of nonmetallic elements.

Q2.6.4

What types of elements form an ionic compound?

Q2.6.5

How do the electrons behave in the formation of an ionic bond?

Q2.6.6

What is the overall charge of an ionic compound?

Q2.6.7

What is the formula for the ionic compound formed from each of the following pairs?

a. potassium and sulfur
b. silver and chlorine (silver has a 1+ charge)
c. calcium and oxygen
d. aluminum and iodine
e. barium and nitrogen
f. sodium and phosphorus
g. lithium and fluorine
h. magnesium and nitrogen
i. calcium and sulfur
j. beryllium and bromine
k. zinc and nitrogen (zinc has a 2+ charge)
l. tin and iodine (tin has a 4+ charge)

Q2.6.8

Write the formula for the compound formed between sodium and each of these polyatomic ions. You can look up the formula and charge for each polyatomic ion.

a. carbonate
b. chlorate
c. chlorite
d. phosphate
e. nitrate
f. sulfate
g. chromate
h. dichromate

Q2.6.9

Write the formula for the compound formed between magnesium and each of the polyatomic ions listed in the previous question.

Q2.6.10

Explain when parentheses should and should not be used in the formulas of ionic compounds.

Answers

2.1: Isotopes and Atomic Mass

Q2.1.1

Answers will vary.

Q2.1.2

Protons and neutrons are in the nucleus and electrons are located outside the nucleus.

Q2.1.3

The sum of the charges on ions in an ionic compound must equal zero.

Q2.1.4

a. 10
b. 79
c. 38
d. 92

Q2.1.5

A certain atom contains 28 protons, 28 electrons, and 31 neutrons. Provide the following:

a. 28
b. 59
c. nickel

Q2.1.6

55 protons, 78 neutrons, 55 electrons

Q2.1.7

9 protons, 10 neutrons, 9 electrons

Q2.1.8

82 protons, 125 neutrons, 82 electrons

Q2.1.9

$^{84}_{36}\text{Kr}$, krypton-84

Q2.1.10

cobalt

Q2.1.11

Isotope	Isotope Symbol	Atomic Number	Mass Number
sodium-23	$^{23}_{11}\text{Na}$	11	23
aresenic-75	$^{75}_{33}\text{As}$	33	75
silver-108	$^{108}_{47}\text{Ag}$	47	108

Q2.1.12

a. $^{40}_{20}\text{Ca}$ - not an isotope because it is a different element
b. $^{39}_{18}\text{Ar}$ - isotope because it has the same atomic number but a different atomic mass
c. $^{40}_{18}\text{Ar}$ - not an isotope because it has the same atomic number and the same atomic mass

Q2.1.13

Isotope	Number of Protons	Number of Electrons	Number of Neutrons	Isotope Symbol
hydrogen-1	1	1	0	$^{1}_{1}\text{H}$
hydrogen-2	1	1	1	$^{2}_{1}\text{H}$
beryllium-9	4	4	5	$^{9}_{4}\text{Be}$
aluminum-27	13	13	14	$^{27}_{13}\text{Al}$

Q2.1.14

Element	Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Isotope Symbol
nitrogen	N	7	14	7	7	7	$^{14}_{7}\text{N}$
boron	B	5	11	5	5	6	$^{11}_{5}\text{B}$
zinc	Zn	30	65	30	30	35	$^{65}_{30}\text{Zn}$
iridium	Ir	77	193	77	77	116	$^{193}_{77}\text{Ir}$
iron	Fe	26	56	26	26	30	$^{56}_{26}\text{Fe}$

2.2: Matter

Q2.2.1

Matter is anything that has mass and occupies space. Examples of matter will vary and can be any object from an atom to a macroscopic object.

Q2.2.2

A compound is a combination of elements with a fixed composition. The elements in the compound do not retain their individual identity by have the properties of the compound. A mixture does not have a fixed composition and each component of the mixture retains its identity and properties.

Q2.2.3

A pure substance contains only one component, either an element or compound, while a mixture contains multiple pure substances.

Q2.2.4

a. mixture
b. pure substance
c. mixture
d. mixture

Q2.2.5

a. element
b. compound
c. homogeneous mixture
d. heterogeneous mixture
e. compound
f. heterogeneous mixture
g. heterogeneous mixture

h. heterogeneous mixture

Q2.2.6

a. metal
b. metal
c. nonmetal
d. metal
e. nonmetal
f. metalloid

Q2.2.7

a. nonmetal
b. nonmetal
c. metalloid
d. metal
e. nonmetal
f. nonmetal

Q2.2.8

Ionic compounds are generally formed between a metal and nonmetal or between a polyatomic ion and another ion. Molecular compounds are composed of two ore more nonmetals.

Q2.2.9

a. molecular
b. molecular
c. molecular
d. ionic
e. ionic
f. ionic

Q2.2.10

a. ionic
b. ionic
c. molecular
d. molecular
e. molecular
f. molecular

2.3: Mole and Molar Mass

Q2.3.1

A counting number is a word that is associated with a specific number.

Q2.3.2

6.022×10^{23}

Q2.3.3

H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

Q2.3.4

a. 6.022×10^{23} atoms
b. 1.204×10^{24} atoms
c. 1.506×10^{24} atoms
d. 3.011×10^{23} atoms
e. 2.108×10^{23} atoms

Q2.3.5

a. 6.022×10^{23} molecules
b. 1.204×10^{24} molecules
c. 1.506×10^{24} molecules
d. 3.011×10^{23} molecules
e. 2.108×10^{23} molecules

Q2.3.6

The numbers are the same for the same number of moles because moles are a counting number. Regardless of what is being counted, a mole will have the same number of items.

Q2.3.7

$$6.73 \times 10^{25}\ atoms\ \mathrm{Si} \left(\frac{1\ mole}{6.022 \times 10^{23}\ atoms}\right) = 112\ moles\ \mathrm{Si}$$

Q2.3.8

$$4.29 \times 10^{22}\ atoms\ \mathrm{Na} \left(\frac{1\ mole}{6.022 \times 10^{23}\ atoms}\right) = 0.0712\ moles\ \mathrm{Na}$$

Q2.3.9

a. 2 hydrogen atoms, 2 oxygen atoms
b. 2 nitrogen atoms, 5 oxygen atoms
c. 1 phosphorus atom, 3 fluorine atoms
d. 1 magnesium atom, 2 chlorine atoms
e. 1 potassium atom, 1 bromine atom
f. 1 aluminum atom, 3 chlorine atoms
g. 1 calcium atom, 1 oxygen atom
h. 2 sodium atoms, 1 oxygen atom
i. 1 nitrogen atom, 3 hydrogen atoms
j. 1 carbon atom, 2 oxygen atoms
k. 2 nitrogen atoms, 4 hydrogen atoms
l. 2 nitrogen atoms, 1 oxygen atom

Q2.3.10

a. 2 moles hydrogen, 2 moles oxygen
b. 2 moles nitrogen, 5 moles oxygen
c. 1 mole phosphorus, 3 moles fluorine
d. 1 mole magnesium, 2 moles chlorine
e. 1 mole potassium, 1 mole bromine
f. 1 mole aluminum, 3 moles chlorine
g. 1 mole calcium, 1 mole oxygen
h. 2 moles sodium, 1 mole oxygen
i. 1 mole nitrogen, 3 moles hydrogen
j. 1 mole carbon, 2 moles oxygen
k. 2 moles nitrogen, 4 moles hydrogen
l. 2 moles nitrogen, 1 mole oxygen

Q2.3.11

The numbers are the same because the ratios are the same between atoms and moles. Moles are a counting number so they are a multiple of the number of atoms.

Q2.3.12

$$0.75\ mol\ \mathrm{CCl_4} \left(\frac{1\ mol\ \mathrm{C}}{1\ mol\ \mathrm{CCl_4}}\right) = 0.75\ mol\ \mathrm{C}$$

$$0.75\ mol\ \text{CCl}_4 \left(\frac{4\ mol\ \text{Cl}}{1\ mol\ \text{CCl}_4}\right) = 3.0\ mol\ \text{Cl}$$

Q2.3.13

$$0.75\ mol\ \text{CCl}_4 \left(\frac{1\ mol\ \text{C}}{1\ mol\ \text{CCl}_4}\right)\left(\frac{6.02\times10^{23}\ atoms}{1\ mol\ \text{C}}\right) = 4.52\times10^{23}\ atoms\ \text{C}$$

$$0.75\ mol\ \text{CCl}_4 \left(\frac{4\ mol\ \text{Cl}}{1\ mol\ \text{CCl}_4}\right)\left(\frac{6.02\times10^{23}\ atoms}{1\ mol\ \text{Cl}}\right) = 1.81\times10^{24}\ atoms\ \text{Cl}$$

Q2.3.14

$$2.5\ mol\ \text{H}_2\text{O} \left(\frac{2\ mol\ \text{H}}{1\ mol\ \text{H}_2\text{O}}\right) = 5.0\ mol\ \text{H}$$

$$2.5\ mol\ \text{H}_2\text{O} \left(\frac{1\ mol\ \text{O}}{1\ mol\ \text{H}_2\text{O}}\right) = 2.5\ mol\ \text{O}$$

Q2.3.15

$$2.5\ mol\ \text{H}_2\text{O} \left(\frac{2\ mol\ \text{H}}{1\ mol\ \text{H}_2\text{O}}\right)\left(\frac{6.02\times10^{23}\ atoms}{1\ mol\ \text{H}}\right) = 3.01\times10^{24}\ atoms\ \text{H}$$

$$2.5\ mol\ \text{H}_2\text{O} \left(\frac{1\ mol\ \text{O}}{1\ mol\ \text{H}_2\text{O}}\right)\left(\frac{6.02\times10^{23}\ atoms}{1\ mol\ \text{O}}\right) = 1.51\times10^{24}\ atoms\ \text{O}$$

Q2.3.16

$$3.87\times10^{25}\ atoms\ \text{O} \left(\frac{1\ molec\ \text{CaNO}_3}{3\ atoms\ \text{O}}\right) = 1.29\times10^{25}\ molec\ \text{CaNO}_3$$

Q2.3.17

$$5.39\times10^{24}\ atoms\ \text{C} \left(\frac{8\ atoms\ \text{H}}{3\ atoms\ \text{C}}\right) = 1.44\times10^{25}\ atoms\ \text{H}$$

Q2.3.18

a. 12.01 g/mol
b. 14.01 g/mol
c. 22.99 g/mol
d. 1.008 g/mol
e. 39.10 g/mol
f. 30.97 g/mol

Q2.3.19

a. $25.0\ g\ \text{C} \left(\frac{1\ mol\ \text{C}}{12.01\ \frac{g}{mol}}\right) = 2.08\ mol\ \text{C}$

b. $25.0\ g\ \text{N} \left(\frac{1\ mol\ \text{N}}{14.01\ \frac{g}{mol}}\right) = 1.78\ mol\ \text{N}$

c. $25.0\ g\ \text{Na} \left(\frac{1\ mol\ \text{Na}}{22.99\ \frac{g}{mol}}\right) = 1.09\ mol\ \text{Na}$

d. $25.0\ g\ \text{H} \left(\frac{1\ mol\ \text{H}}{1.008\ \frac{g}{mol}}\right) = 24.8\ mol\ \text{H}$

e. $25.0\ g\ \text{K} \left(\frac{1\ mol\ \text{K}}{39.10\ \frac{g}{mol}}\right) = 0.639\ mol\ \text{K}$

f. $25.0\ g\ \text{P} \left(\frac{1\ mol\ \text{P}}{30.97\ \frac{g}{mol}}\right) = 0.807\ mol\ \text{P}$

Q2.3.20

The moles are different because a mole of atoms of each element has a different mass. Although they have the same mass, the number of atoms of each sample is different. Compare a ton of feathers to a ton of books. The same total mass but a different quantity of each.

Q2.3.21

What is the mass of each of the following samples?

a. $0.35\ mol\ \text{Na}\left(\frac{22.99\ g}{mol\ \text{Na}}\right)=8.0\ g\ \text{Na}$
b. $0.75\ mol\ \text{C}\left(\frac{12.01\ g}{mol\ \text{C}}\right)=9.0\ g\ \text{C}$
c. $1.34\ mol\ \text{K}\left(\frac{39.10\ g}{mol\ \text{K}}\right)=52.4\ g\ \text{K}$
d. $1.21\ mol\ \text{Si}\left(\frac{28.09\ g}{mol\ \text{Si}}\right)=34.0\ g\ \text{Si}$
e. $0.95\ mol\ \text{Ca}\left(\frac{40.08\ g}{mol\ \text{Ca}}\right)=38\ g\ \text{Ca}$
f. $2.85\ mol\ \text{He}\left(\frac{4.003\ g}{mol\ \text{He}}\right)=11.4\ g\ \text{He}$

Q2.3.22

a. CO_2 has 1 mole of carbon and 2 moles of oxygen; $\left(12.01\frac{g}{mol}\times 1\right)+\left(16.00\frac{g}{mol}\times 2\right)=44.01\frac{g}{mol}$
b. N_2H_4 has 2 moles of nitrogen and 4 moles of hydrogen; $\left(14.01\frac{g}{mol}\times 2\right)+\left(1.008\frac{g}{mol}\times 4\right)=32.05\frac{g}{mol}$
c. CaF_2 has 1 mole of calcium and 2 moles of fluorine; $\left(40.08\frac{g}{mol}\times 1\right)+\left(19.00\frac{g}{mol}\times 2\right)=78.08\frac{g}{mol}$
d. $C_6H_{12}O_6$ has 6 moles of carbon, 12 moles of hydrogen, and 6 moles of oxygen; $\left(12.01\frac{g}{mol}\times 6\right)+\left(1.008\frac{g}{mol}\times 12\right)+\left(16.00\frac{g}{mol}\times 6\right)=180.16\frac{g}{mol}$
e. CH_4 has 1 mole of carbon and 4 moles of hydrogen; $\left(12.01\frac{g}{mol}\times 1\right)+\left(1.008\frac{g}{mol}\times 4\right)=16.04\frac{g}{mol}$
f. C_6H_6 has 6 moles of carbon and 6 moles of hydrogen; $\left(12.01\frac{g}{mol}\times 6\right)+\left(1.008\frac{g}{mol}\times 6\right)=78.11\frac{g}{mol}$
g. Na_2SO_4 has 2 moles of sodium, 1 mole of sulfur, and 4 moles of oxygen; $\left(22.99\frac{g}{mol}\times 2\right)+\left(32.06\frac{g}{mol}\times 1\right)+\left(16.00\frac{g}{mol}\times 4\right)=142.04\frac{g}{mol}$
h. K_3PO_4 has 3 moles of potassium, 1 mole of phosphorus, and 4 moles of oxygen; $\left(39.10\frac{g}{mol}\times 3\right)+\left(30.97\frac{g}{mol}\times 1\right)+\left(16.00\frac{g}{mol}\times 4\right)=212.27\frac{g}{mol}$
i. $Al(NO_3)_3$ has 1 mole of aluminum, 3 moles of nitrogen, and 9 moles of oxygen; $\left(26.98\frac{g}{mol}\times 1\right)+\left(14.01\frac{g}{mol}\times 3\right)+\left(16.00\frac{g}{mol}\times 9\right)=213.01\frac{g}{mol}$
j. $Mg_3(PO_4)_2$ has 3 moles of magnesium, 2 moles of phosphorus, and 8 moles of oxygen; $\left(24.31\frac{g}{mol}\times 3\right)+\left(30.97\frac{g}{mol}\times 2\right)+\left(16.00\frac{g}{mol}\times 8\right)=262.87\frac{g}{mol}$

Q2.3.23

a. $25.0\ g\ \text{CO}_2\left(\frac{1\ mol\ \text{CO}_2}{44.01\ g}\right)=0.568\ mol\ \text{CO}_2$
b. $10.0\ g\ \text{N}_2\text{H}_4\left(\frac{1\ mol\ \text{N}_2\text{H}_4}{32.05\ g}\right)=0.312\ mol\ \text{N}_2\text{H}_4$
c. $85.0\ g\ \text{CaF}_2\left(\frac{1\ mol\ \text{CaF}_2}{78.08\ g}\right)=1.09\ mol\ \text{CaF}_2$
d. $15.5\ g\ \text{C}_6\text{H}_{12}\text{O}_6\left(\frac{1\ mol\ \text{C}_6\text{H}_{12}\text{O}_6}{180.16\ g}\right)=0.0860\ mol\ \text{C}_6\text{H}_{12}\text{O}_6$
e. $20.0\ g\ \text{CH}_4\left(\frac{1\ mol\ \text{CH}_4}{16.04\ g}\right)=1.25\ mol\ \text{CH}_4$
f. $100.0\ g\ \text{C}_6\text{H}_6\left(\frac{1\ mol\ \text{C}_6\text{H}_6}{78.11\ g}\right)=1.28\ mol\ \text{C}_6\text{H}_6$
g. $30.0\ g\ \text{Na}_2\text{SO}_4\left(\frac{1\ mol\ \text{Na}_2\text{SO}_4}{142.04\ g}\right)=0.211\ mol\ \text{Na}_2\text{SO}_4$
h. $75.0\ g\ \text{K}_3\text{PO}_4\left(\frac{1\ mol\ \text{K}_3\text{PO}_4}{212.27\ g}\right)=0.353\ mol\ \text{K}_3\text{PO}_4$
i. $50.0\ g\ \text{Al(NO}_3)_3\left(\frac{1\ mol\ \text{Al(NO}_3)_3}{213.01\ g}\right)=0.235\ mol\ \text{Al(NO}_3)_3$
j. $47.2\ g\ \text{Mg}_3(\text{SO}_4)_2\left(\frac{1\ mol\ \text{Mg}_3(\text{SO}_4)_2}{262.87\ g}\right)=0.180\ mol\ \text{Mg}_3(\text{SO}_4)_2$

Q2.3.24

a. $3.5\ mol\ \text{CO}_2\left(\frac{44.01\ g}{mol\ \text{CO}_2}\right)=1.50\times 10^2\ g\ \text{CO}_2$
b. $0.45\ mol\ \text{N}_2\text{H}_4\left(\frac{32.05\ g}{mol\ \text{N}_2\text{H}_4}\right)=14\ g\ \text{N}_2\text{H}_4$
c. $2.25\ mol\ \text{CaF}_2\left(\frac{78.08\ g}{mol\ \text{CaF}_2}\right)=176\ g\ \text{CaF}_2$
d. $1.75\ mol\ \text{C}_6\text{H}_{12}\text{O}_6\left(\frac{180.16\ g}{mol\ \text{C}_6\text{H}_{12}\text{O}_6}\right)=315\ g\ \text{C}_6\text{H}_{12}\text{O}_6$
e. $4.9\ mol\ \text{CH}_4\left(\frac{16.04\ g}{mol\ \text{CH}_4}\right)=79\ g\ \text{CH}_4$

f. $8.75\ mol\ \mathrm{C_6H_6}\left(\frac{78.11\ g}{mol\ \mathrm{C_6H_6}}\right) = 683\ g\ \mathrm{C_6H_6}$
g. $2.35\ mol\ \mathrm{Na_2SO_4}\left(\frac{142.04\ g}{mol\ \mathrm{Na_2SO_4}}\right) = 334\ g\ \mathrm{Na_2SO_4}$
h. $0.672\ mol\ \mathrm{K_3PO_4}\left(\frac{212.27\ g}{mol\ \mathrm{K_3PO_4}}\right) = 143\ g\ \mathrm{K_3PO_4}$
i. $0.95\ mol\ \mathrm{Al(NO_3)_3}\left(\frac{213.01\ g}{mol\ \mathrm{Al(NO_3)_3}}\right) = 2.0 \times 10^2\ g\ \mathrm{Al(NO_3)_3}$
j. $1.15\ mol\ \mathrm{Mg_3(PO_4)_2}\left(\frac{262.87\ g}{mol\ \mathrm{Mg_3(PO_4)_2}}\right) = 302\ g\ \mathrm{Mg_3(PO_4)_2}$

2.4: Electron Arrangements

Q2.4.1

a. 2, 8, 1
b. 2, 8
c. 2, 2
d. 2, 5
e. 2, 8, 6
f. 2, 8, 7

Q2.4.2

a. 1
b. 5
c. 7
d. 6
e. 1
f. 3

Q2.4.3

The octet rule predicts the stability of an atom based on having eight electrons in its electron shell.

2.5: Ion Formation

Q2.5.1

An ion is a charged species which results from the gain or loss of one ore more electrons.

Q2.5.2

Anions and cations are both charged species which results from the change in the number of electrons. Anions have a negative charge while cations have a positive charge.

Q2.5.3

a. $\mathrm{Li^+}$
b. $\mathrm{Na^+}$
c. $\mathrm{Ca^{2+}}$
d. $\mathrm{B^{3+}}$
e. $\mathrm{P^{3-}}$
f. $\mathrm{S^{2-}}$
g. $\mathrm{Cl^-}$
h. $\mathrm{Br^-}$

Q2.5.4

a. 3 protons, 4 neutrons, 2 electrons
b. 11 protons, 12 neutrons, 10 electrons
c. 20 protons, 20 neutrons, 18 electrons
d. 5 protons, 6 neutrons, 2 electrons
e. 15 protons, 16 neutrons, 18 electrons

f. 16 protons, 16 neutrons, 18 electrons
g. 17 protons, 18 neutrons, 18 electrons
h. 35 protons, 45 neutrons, 36 electrons

Q2.5.5 ($\therefore$ = therefore)

a. ion with 3+ charge and 2 electrons $\therefore$ neutral atom had 5 electrons $\therefore$ atom has 5 protons $\therefore$ boron (B)
b. ion with 1− charge and 18 electrons $\therefore$ neutral atom had 17 electrons $\therefore$ atom has 17 protons $\therefore$ chlorine (Cl)
c. ion with 1+ charge and 18 electrons $\therefore$ neutral atom had 19 electrons $\therefore$ atom has 19 protons $\therefore$ potassium (K)
d. ion with a 3− charge and 10 electrons $\therefore$ neutral atom had 7 electrons $\therefore$ atom has 7 protons $\therefore$ nitrogen (N)

Q2.5.6

A polyatomic ion contains multiple atoms working together as a group and has an overall charge.

Q2.5.7

a. polyatomic
b. monatomic
c. polyatomic
d. monatomic
e. monatomic
f. polyatomic

2.6: Ionic Compounds

Q2.6.1

carbon

Q2.6.2

Answers will vary. Most metals are in the first two columns of the periodic table or in the transition metal block.

Q2.6.3

Answers will vary. Nonmetallic elements are located in the upper right corner of the periodic table (examples include nitrogen, oxygen, phosphorus, chlorine, bromine, etc)

Q2.6.4

Ionic compounds are composed of ions of metals and nonmetals. Ionic compounds can also include a polyatomic ion.

Q2.6.5

To form an ionic bond, electrons are transferred from the metal to the nonmetal.

Q2.6.6

Zero

Q2.6.7

What is the formula for the ionic compound formed from each of the following pairs?

a. K_2S
b. $AgCl$
c. CaO
d. AlI_3
e. Ba_3N_2
f. Na_3P
g. LiF
h. Mg_3N_2
i. CaS
j. $BeBr_2$
k. Zn_3N_2

l. SnI_4

Q2.6.8

a. Na_2CO_3
b. $NaClO_3$
c. $NaClO_2$
d. Na_3PO_4
e. $NaNO_3$
f. Na_2SO_4
g. Na_2CrO_4
h. $Na_2Cr_2O_7$

Q2.6.9

a. $MgCO_3$
b. $Mg(ClO_3)_2$
c. $Mg(ClO_2)_2$
d. $Mg_3(PO_4)_2$
e. $Mg(NO_3)_2$
f. $MgSO_4$
g. $MgCrO_4$
h. $MgCr_2O_7$

Q2.6.10

Parentheses are used when there is more than one of a polyatomic ion in the formula of an ionic compound.

CHAPTER OVERVIEW

3: Compounds

Molecular compounds are chemical compounds that take the form of discrete molecules like water (H_2O) and carbon dioxide (CO_2). These compounds are very different from ionic compounds like sodium chloride $(NaCl)$, which are formed when metal atoms lose one or more of their electrons to nonmetal atoms and the resulting ions are mutually attracted to each other.

3.1: Molecular Compounds

Learning Outcomes

- Describe types of elements that form covalent bonds.
- Explain how a covalent bond is formed.
- Distinguish between ionic and molecular compounds.
- Know the meaning of prefixes used in naming molecular compounds.
- Distinguish between molecules and compounds.

Water and diamonds - two very different materials. Water can be found almost everywhere. It is in lakes, creeks, rivers, and oceans. We get water from the sky when it rains. Diamonds, on the other hand, are very rare. They are only found in a few locations on the earth and must be mined to become available. Major diamond mines are located in various African countries, Australia, and Russia. The United States has several underground sources of diamonds in Alaska, Colorado, Minnesota, Montana, and Wyoming, but the only "active" U.S. mine is the Crater of Diamonds mine in a state park near Murfreesboro, Arkansas. For a small fee, visitors can dig for diamonds. You won't get rich by visiting, though - only a few hundred carats of low-grade diamonds are found each year.

The two materials do have at least one thing in common. The atoms in the materials are held together by covalent bonds. These bonds consist of electrons shared between two or more atoms. Unlike ionic bonds, where electrons are either lost or gained by an atom to form charged ions, electrons in covalent compounds are shared between the two atoms, giving rise to properties that are quite different from those seen in ionic materials.

Molecular Compounds

Molecular compounds are chemical compounds that take the form of discrete molecules. Examples include such familiar substances as water (H_2O) and carbon dioxide (CO_2) (Figure $3.1.1$). These compounds are very different from ionic compounds like sodium chloride $(NaCl)$. Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their valence electrons in such a way that a bond forms between pairs of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

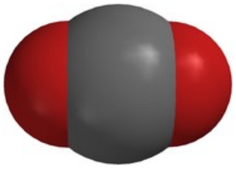

Figure $3.1.1$: Carbon dioxide molecules consist of a central carbon atom bonded to two oxygen atoms.

Binary Molecular Compound Names

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains eight atoms of carbon and eighteen atoms of hydrogen. The molecular formula of octane is C_8H_{18}.

A **binary molecular compound** *is a molecular compound that is composed of two elements*. In general, the elements that combine to form binary molecular compounds are both nonmetals. This contrasts with ionic compounds, which involve bonds between metal cations and nonmetal anions. One difference between ionic and molecular compounds is that two nonmetal atoms will frequently combine with one another in a variety of ratios. Consider the elements nitrogen and oxygen. They combine to make several binary compounds, including NO, NO_2, (see figure below), and N_2O. Obviously they can't all be called nitrogen oxide! How would someone know which one you were talking about? Each of the three compounds has very different properties and reactivity. A system to distinguish between compounds such as these is necessary.

Figure 3.1.2: NO_2 (left) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines. It is converted to N_2O (right) at low temperatures.

Prefixes are used in the names of binary molecular compounds to identify the number of atoms of each element. Although we will not cover nomenclature rules for molecular compounds you should be familiar with the prefixes used in their names. Listed below (see table below) are the prefixes up to ten.

Table 3.1.1: Numerical Prefixes

Number of Atoms	Prefix
1	*mono-*
2	*di-*
3	*tri-*
4	*tetra-*
5	*penta-*
6	*hexa-*
7	*hepta-*
8	*octa-*
9	*nona-*
10	*deca-*

Some examples of molecular compounds are listed below (see table below). Note the use of prefixes in their names.

Table 3.1.2: Examples of Molecular Compounds

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Contributors and Attributions

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

3.2: Straight-Chain Alkanes

Learning Outcomes

- Distinguish between organic and inorganic compounds.
- Describe features of alkane structure.
- Pair each prefix with its numerical definition.
- Distinguish between an alkane and a cycloalkane.

Organic Compounds

All living things on earth are formed mostly of carbon compounds. The prevalence of carbon compounds in living things has led to the epithet "carbon-based" life. The truth is we know of no other kind of life. Early chemists regarded substances isolated from *organisms* (plants and animals) as a different type of matter that could not be synthesized artificially, and these substances were thus known as *organic compounds*. The widespread belief called vitalism held that organic compounds were formed by a vital force present only in living organisms. The German chemist Friedrich Wohler was one of the early chemists to refute this aspect of vitalism, when, in 1828, he reported the synthesis of urea, a component of many body fluids, from nonliving materials. Since then, it has been recognized that organic molecules obey the same natural laws as inorganic substances, and the category of organic compounds has evolved to include both natural and synthetic compounds that contain carbon. Some carbon-containing compounds are *not* classified as organic, for example, carbonates and cyanides, and simple oxides, such as CO and CO_2. Although a single, precise definition has yet to be identified by the chemistry community, most agree that a defining trait of organic molecules is the presence of carbon as the principal element, bonded to hydrogen and other carbon atoms. Organic compounds often contain other elements such as nitrogen, oxygen, sulfur, or phosphorus. The number of carbon and hydrogen atoms in an organic compound is usually much greater than the number of other atoms.

Figure 3.2.1: *All organic compounds contain carbon and most are formed by living things, although they are also formed by geological and artificial processes. (credit left: modification of work by Jon Sullivan; credit left middle: modification of work by Deb Tremper; credit right middle: modification of work by "annszyp"/Wikimedia Commons; credit right: modification of work by George Shuklin)*

Today, organic compounds are key components of plastics, soaps, perfumes, sweeteners, fabrics, pharmaceuticals, and many other substances that we use every day. The value to us of organic compounds ensures that organic chemistry is an important discipline within the general field of chemistry. The element carbon gives rise to a vast number and variety of compounds that are found in biological and industrial settings. These compounds typically contain functional groups which include atoms other than carbon and hydrogen. A particular arrangement of atoms within a compound is known as a functional group. Types of functional groups will be discussed in the next chapter. For now, we will identify organic and inorganic (which means not organic) compounds and look at structures for the simplest of organic compounds, hydrocarbons.

Hydrocarbons

A **hydrocarbon** is an organic compound that is made up of only carbon and hydrogen. A hydrocarbon is the simplest kind of organic molecule and is the basis for all other more complex organic compounds. Hydrocarbons can have single, double, or triple bonds between the carbon atoms. Alkanes will be discussed here and other hydrocarbons will be covered later.

Alkanes

An **alkane** is a hydrocarbon in which there are only single covalent bonds. The simplest alkane is methane, with the molecular formula $\mathrm{CH_4}$. The carbon is the central atom and makes four single covalent bonds to hydrogen atoms.

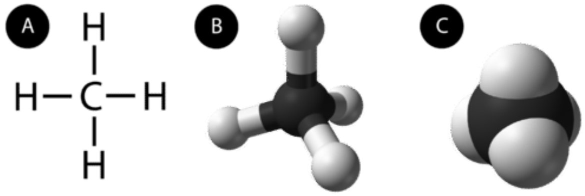
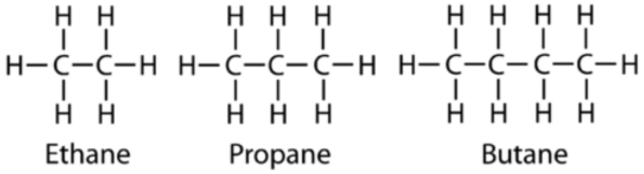


Figure 3.2.2: Methane is the simplest hydrocarbon and is shown with a structural formula, a ball-and-stick model, and a space-filling model.

The next simplest alkane is called ethane (C_2H_6) and consists of two carbon atoms with a single covalent bond between them. Each carbon is then able to bond to three hydrogen atoms. The alkane series progresses from there, increasing the length of the carbon chain by one carbon at a time. Structural formulas for ethane, propane (C_3H_8), and butane (C_4H_{10}) are shown below.



These alkanes are called straight-chain alkanes because the carbon atoms are connected in one continuous chain with no branches. Naming and writing structural and molecular formulas for the straight-chain alkanes is straightforward. The name of each alkane consists of a prefix that specifies the number of carbon atoms and the ending *-ane*. The molecular formula follows the pattern of $C_nH_{2\,n+2}$ where n is the number of carbons in the chain. The table below lists the first ten members of the alkane series.

Table 3.2.1: First Ten Members of the Alkane Series

Name	Molecular Formula	Condensed Structural Formula
Methane	CH_4	CH_4
Ethane	C_2H_6	CH_3CH_3
Propane	C_3H_8	$CH_3CH_2CH_3$
Butane	C_4H_{10}	$CH_3CH_2CH_2CH_3$
Pentane	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$
Hexane	C_6H_{14}	$CH_3CH_2CH_2CH_2CH_2CH_3$
Heptane	C_7H_{16}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$
Octane	C_8H_{18}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Nonane	C_9H_{20}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Decane	$C_{10}H_{22}$	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH$

Note that the table shows a variation of a structural formula called a condensed structural formula, which we will discuss later. In this formula, the covalent bonds are understood to exist between each carbon and the hydrogens associated with it, as well as between carbon atoms.

Cycloalkanes

Alkanes can also exist in a ring structure which is known as a cycloalkane. For example, pentane is a chain of five carbon atoms while cyclopentane is a ring of five carbons as shown in the figure below.

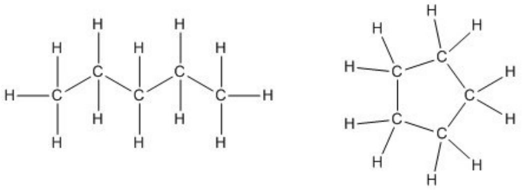


Figure 3.2.3: Structures of pentane (left) and cyclopentane (right).

Contributors and Attributions

-
- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
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3.3: Compounds (Exercises)

These are homework exercises to accompany Chapter 3 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

3.1: Molecular Compounds

(click here for solutions)

Q3.1.1

What types of elements typically form binary molecular compounds?

Q3.1.2

Describe the similarities and differences between ionic and molecular compounds.

Q3.1.3

Give the prefix that would be used in the name a molecular compound for each of the following quantities of atoms.

a. 6
b. 3
c. 9
d. 5
e. 8

Q3.1.4

Write the formula for each compound.

a. diphosphorus pentoxide
b. dinitrogen monoxide
c. trisilicon tetranitride
d. dinitrogen pentoxide
e. tetraphosphorus decasulfide
f. disulfur hexafluoride
g. triboron dicarbide
h. tetraselenium tetranitride

3.2: Straight-Chain Alkanes

(click here for solutions)

Q3.2.1

What elements are found in a hydrocarbon?

Q3.2.2

Describe an alkane.

Q3.2.3

How many carbon atoms are found in octane? propane? nonane?

Q3.2.4

What is the name for the alkane with six carbons? four carbons? two carbons?

Q3.2.5

What is the difference between an alkane and a cycloalkane?

Answers

3.1: Molecular Compounds

Q3.1.1

Binary molecular compounds are composed of two nonmetallic elements.

Q3.1.2

They both form as a result of bonding between atoms. Ionic compounds result from the transfer of electrons from one element to another while molecular compounds form bonds through the sharing of electrons.

Q3.1.3

a. hexa
b. tri
c. nona
d. penta
e. octa

Q3.1.4

a. P_2O_5
b. N_2O
c. Si_3N_4
d. N_2O_5
e. P_4S_{10}
f. S_2F_6
g. B_3C_2
h. Se_4N_4

3.2: Straight-Chain Alkanes

Q3.2.1

carbon and hydrogen

Q3.2.2

An alkane contains only carbon and hydrogen atoms with the carbons connected by single bonds.

Q3.2.3

octane, 8; propane, 3; nonane, 9

Q3.2.4

6, hexane; 4, butane; 2, ethane

Q3.2.5

An alkane contains a chain of carbon atoms while a cycloalkane contains carbons in a ring structure.

CHAPTER OVERVIEW

4: Structure and Function

The three dimensional shape or **configuration** of a molecule is an important characteristic. This shape is dependent on the preferred spatial orientation of covalent bonds to atoms having two or more bonding partners.

4.1: Lewis Electron Dot Structures

Learning Outcomes

- Determine Lewis structure of a compound.

Ionic substances are completely held together by ionic bonds. The full charges of the ions (for example, Na^+ and Cl^- in sodium chloride) cause electrostatic interactions that result in a stable crystal lattice. Ionic compounds exist as extended, orderly arrangements of ions. This is quite different from the structure of molecular substances, which take the form of collections of individual molecules.

The electrons that form a covalent bond are not fully possessed by a single atom (as the electrons in an ion would be) but are shared between the two atoms involved in the bond. The concept of the covalent bond was first proposed in 1916 by the American chemist G. N. Lewis (1875 - 1946), who suggested that sharing electrons was one way that atoms could attain a complete octet of valence electrons. This idea was expanded upon by Linus Pauling (1901 - 1994), who eventually won the Nobel Prize in Chemistry in 1954 for his work on chemical bonding.

Single Covalent Bonds

The simplest covalent bond is formed between two hydrogen atoms. Each hydrogen atom has a single electron, and each needs two electrons for a full outer shell. The hydrogen molecule, H_2, consists of two hydrogen atoms sharing their two valence electrons. Hydrogen can also form covalent bonds with other atoms. For example, hydrogen and chlorine each need one more electron to achieve a noble gas configuration. By sharing valence electrons (each atom donates one), the stable HCl molecule is formed.

We will use a simplified representation of covalent bonds known as **Lewis structures**. These drawings are also know by various other names, including Lewis dot structures or electron dot structures. Each dot in the structure represents one valence electron in the compound. For example, H_2 could be drawn as $\mathrm{H} : \mathrm{H}$. Each dot represents one valence electron, and the fact that they are placed between the two atoms means that they are being shared bas a covalent bond. For larger molecules, it can become cumbersome to draw out all of the valence electrons, so a bonding pair of electrons can also be drawn as a straight line. Thus, H_2 can also be represented as $\mathrm{H}{-}\mathrm{H}$.

If we wanted to show the Lewis structure of HCl, we would draw the following:

We can see that the covalent bond consists of two electrons between the H and the Cl. The H has a full outer shell of two electrons and the chlorine has a full outer shell of eight electrons. Covalent bonds with other halogens can be written the same way.

Similar types of Lewis structures can be written for other molecules that form covalent bonds. Many compounds that contain O, N, C, S, and P are held together by covalent bonds. The number of covalent bonds an atom will form can generally be predicted by the number of electrons an atom requires to fill its valence shell. For example, oxygen has 6 electrons in its outer shell and needs two more to fill this shell, so it will only form two covalent bonds with other atoms. If we look at the water molecule $(\mathrm{H}_2\mathrm{O})$ (see figure below), we see that the oxygen atom makes two total bonds (one with each hydrogen atom).

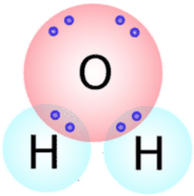
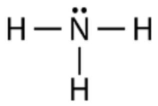


Figure 4.1.1: Electron sharing in a water molecule.

As you can see, there are two pairs of electrons not involved in covalent bonding. These unbonded pairs of electrons are known as **lone pairs** and contribute to the overall shape of the molecule. Similarly, nitrogen needs three electrons to complete its valence shell, so it tends to make three covalent bonds, with one lone pair of non-bonding electrons left over.



Again, each of the lines stands for a *pair* of bonding electrons (a single bond), and the lone pair of nitrogen is drawn as two dots.

Double and Triple Bonds

So far we have considered only single bonds, formed by the sharing of one electron from each atom. Many molecules contain **double bonds**, in which each atom shares two electrons, or **triple bonds**, in which each atom shares three electrons. These are represented by drawing two or three lines between the bonded atoms. For example, a carbon-carbon double bond can be written as $\mathrm{C::C}$ or $\mathrm{C{=}C}$. A carbon-carbon triple bond is shown as $\mathrm{C:::C}$ or with three lines between the two carbon atoms, as seen in the structure of an organic molecule called acetylene (shown in the figure below). Just as N wants to form 3 bonds, other elements tend to form the same number of bonds in different compounds.

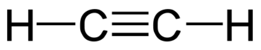
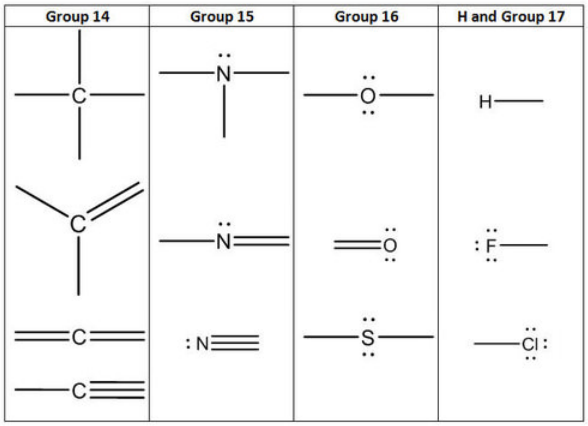


Figure 4.1.2: Acetylene molecule.

We will build Lewis structures that satisfy the octet rule to determine how atoms are attached to one another using the components shown in the table below to create structures that follow the octet rule.

Table 4.1.1: *Components to Build Lewis Structures*



Supplemental Resources

Practice drawing Lewis Dot Structures of Covalent Compounds: http://www.wisc-online.com/objects/V...spx?ID=GCH6404

Contributors and Attributions

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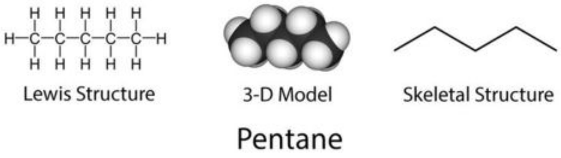
4.2: Representing Structures

Learning Outcomes

- Use condensed structural formulas to represent molecules.
- Use skeletal structures to represent molecules.

Lewis structures are great for small molecules but can be complicated when dealing with large molecules. There are a variety of ways that we can represent molecules that provide use with enough information without creating a complicated figure. The structures in the figure below all represent the molecules pentane, which is a straight-chain alkane with five carbon atoms.



Lewis Structures

The Lewis structure shows bonds to represent the sharing of electrons and include all atoms in the molecule so we have information about connectivity of atoms but not the molecular shape or geometry.

Space-Filling Model

A 3-D space-filling model shows connectivity of atoms and the shape of the molecule. These work well for smaller molecules, but can be a challenge to interpret since atoms are so close together. This is the most accurate representation of what a molecule looks like.

Skeletal Structure

These are the simplest way to represent the structure of a molecule. Hydrogen atoms are omitted (except when attached to O or S) and carbon atoms are at the intersection of each pair of lines and at the terminal end of each line. Each of the circles in the figure below indicate where a carbon is located.

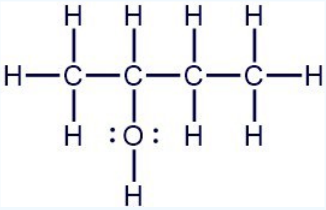
Figure 4.2.1: Pentane molecule with circles indicating locations of carbon atoms.

Condensed Structures

These provide more information than a chemical formula (like C_5H_{12} for pentane) but are condensed as given by their name, so they can be written out rather than in a drawing. The best way to write a condensed structure is to "read" a Lewis structure and see how atoms are connected.

Example 4.2.1

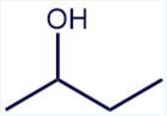
Draw the skeletal and condensed structure of 2-butanol based on the given Lewis structure.



Skeletal Structure

Count the number of carbon atoms and draw the carbon chain with the correct number of carbon atoms. This represents a chain of four carbon atoms.

Now look for any atoms other than hydrogen that are attached to the carbon chain. The second carbon has an $-\mathrm{OH}$ group. Add that to the second carbon in the chain.



The skeletal structure is complete and all of the remaining hydrogen atoms are implied because every carbon must have four bonds.

Condensed Structure

For the condensed structure, we "read" the Lewis structure from left to right.

The first carbon has three hydrogen atoms attached so we start with

$$\mathrm{CH_3} \tag{4.2.1}$$

The second carbon has one hydrogen attached and one $-\mathrm{OH}$ group. Since the $-\mathrm{OH}$ group is to the side, we put it in parentheses

$$\mathrm{CH_3CH(OH)} \tag{4.2.2}$$

The third carbon has two hydrogen atoms attached so we add $\mathrm{CH_2}$

$$\mathrm{CH_3CH(OH)CH_2} \tag{4.2.3}$$

The fourth, and last, carbon has three hydrogen atoms attached, so we add $\mathrm{CH_3}$

$$\mathrm{CH_3CH(OH)CH_2CH_3} \tag{4.2.4}$$

The result is the condensed structure of 2-butanol.

Contributors and Attributions

-
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4.3: Electron Group Geometry

Learning Outcomes

- Explain the basis of VSEPR theory.
- Predict the shapes of molecules using VSEPR theory.
- Predict bond angles in a molecule.
- Account for variations in bond angles on the relative repulsive forces exerted by lone pairs and bonding pairs of electrons.

Electron group geometry is the three-dimensional arrangement of atoms in a molecule. The geometry of a molecule is an important factor that affects the physical and chemical properties of a compound. Those properties include melting and boiling points, solubility, density, and the types of chemical reactions that a compound undergoes. In this section, you will learn a technique to predict the geometry of a molecule based on a its Lewis electron dot structure.

VSEPR Theory

The **valence shell** is the outermost occupied shell of electrons in an atom. This shell holds the valence electrons, which are the electrons that are involved in bonding and shown in a Lewis structure. **Valence-shell electron pair repulsion theory**, or VSEPR theory, states that a molecule will adjust its shape so that the valence electron pairs stay as far apart from each other as possible. This makes sense, based on the fact that negatively charged electrons repel one another. We will systematically classify molecules according to the number of bonding pairs of electrons and the number of nonbonding or lone pairs around the central atom. For the purposes of the VSEPR model, a double or triple bond is no different in terms of repulsion than a single bond. We will begin by examining molecules in which the central atom does not have any lone pairs.

Central Atom with No Lone Pairs

In order to easily understand the types of molecules possible, we will use a simple system to identify the parts of any molecule.

$$\text{A} = \text{central atom in a molecule}$$

$$\text{B} = \text{atoms surrounding the central atom}$$

Subscripts after the B will denote the number of B atoms that are bonded to the central A atom. For example, AB_4 is a molecule with a central atom surrounded by four covalently bonded atoms. Again, it does not matter if those bonds are single, double, or triple bonds.

AB_2

Beryllium hydride (BeH_2) consists of a central beryllium atom with two single bonds to hydrogen atoms. Note that it violates the octet rule, because the central atom has only 4 valence electrons. This is acceptable because beryllium only has two valence electrons to begin with, so it is not possible for it to create more than two covalent bonds with hydrogen atoms.

H—Be—H

According to the requirement that electron pairs maximize their distance from one another, the two bonding pairs in the BeH_2 molecules will arrange themselves on directly opposite sides of the central Be atom. The resulting geometry is a *linear* molecule, shown in a "ball-and-stick" model in the figure below.

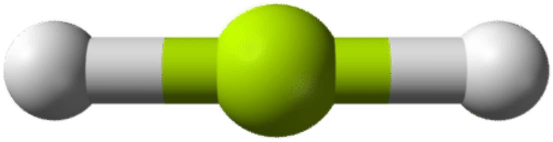
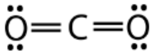


Figure 4.3.1: BeH_2

The $\text{H}-\text{Be}-\text{H}$ bond angle is 180^o because of its linear geometry.

Carbon dioxide is another example of a molecule which falls under the AB_2 category. Its Lewis structure consists of double bonds between the central carbon atom and each oxygen atom.



The repulsion between the two double bonds on either side of the carbon atom is no different than the repulsion between the two single bonds on either side of the beryllium in the previous example. Therefore carbon dioxide is also linear, as this achieves the maximum distance between the electron pair bonds.

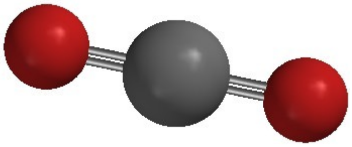
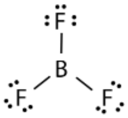


Figure 4.3.2: CO_2.

AB_3

Boron trifluoride $\left(BF_3\right)$ consists of a central boron atom with three single bonds to fluorine atoms. The boron atom is an exception to the octet rule, and generally only needs 6 electrons to be stable in a bonded molecule.



The geometry of the BF_3 molecule is called *trigonal planar*. The fluorine atoms are positioned at the vertices of an equilateral triangle. The $F-B-F$ angle is 120^o, and all four atoms lie in the same plane.

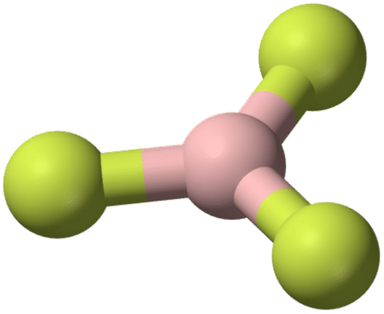
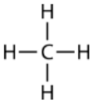


Figure 4.3.3: BF_3.

AB_4

Methane $\left(CH_4\right)$ is an organic compound that is the primary component of natural gas. Its structure consists of a central carbon atom with four single bonds to hydrogen atoms.



In order to maximize their distance from one another, the four groups of bonding electrons do not lie in the same plane. Instead, each of the hydrogen atoms lies at the corners f a geometrical shape called a tetrahedron. The carbon atom is at the center of the tetrahedron. Each face of a tetrahedron is an equilateral triangle.

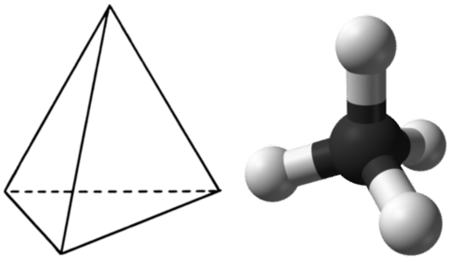


Figure 4.3.4: CH_4.

The electron group geometry of the methane molecule is referred to as *tetrahedral*. The $H-C-H$ bond angles are 109.5^o, which is larger than the 90° that they would be if the molecule was planar. This way, the bonds are as far apart as possible to minimize electron repulsion. When drawing a structural formula for a molecule such as methane, it is advantageous to be able to indicate the three-dimensional character of its shape. The structural formula in the figure below is called a *perspective drawing*. The dotted line bond should be visualized as going back into the page, while the solid triangle bond should be visualized as coming out of the page.

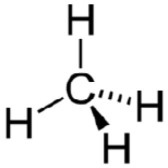


Figure 4.3.5: Perspective drawing of CH_4.

There are structures with five and six bonds that we will not explore because they are not seen in biological molecules.

Table 4.3.1: Geometries of Molecules in Which the Central Atom Has No Lone Pairs

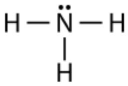
Atoms Around Central Atom	Electron Group Geometry	Example
2 AB_2	Linear	$BeCl_2$
3 AB_3	Trigonal Planar	BF_3
4 AB_4	Tetrahedral	CH_4

Central Atom with One or More Lone Pairs

The number of bonds to the central atom plus the number of lone pairs on the central atom gives us what is called the **electron group geometry**. Electron group geometries refer to the five geometries: linear, trigonal planar, tetrahedral, trigonal bipyramidal, or octahedral. If one or more of the bonding pairs of electrons is replaced with a lone pair, the electron geometry does not change but the the shape of the molecule is altered. Understanding the impact of lone pairs of electrons will be important when looking at properties of compounds that are affected by polarity of a molecule. We will focus on the tetrahedral electron group geometry since it is most relevant to organic and biological molecules. E represents a non-bonding pair of electrons.

AB_3E

The ammonia molecule contains three single bonds and one lone pair on the central nitrogen atom.



The electron group geometry for a molecule with four electron pairs is tetrahedral, as was seen with CH_4. In the ammonia molecule, one of the electron pairs is a lone pair rather than a bonding pair. Although the lone pair is not visible, it will affects the location and bond angles among other atoms in the molecule.

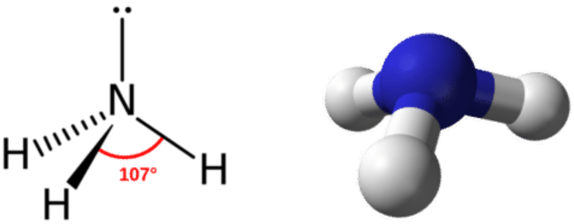


Figure 4.3.6: NH_3.

Recall that the bond angles in the tetrahedral CH_4 molecule are all equal to 109.5^o. One might expect the $H-N-H$ bond angles in ammonia to be 109.5^o as well, but slight adjustments need to be made for the presence of lone pairs. Within the context of the VSEPR model, lone pairs of electrons are considered to be slightly more repulsive than bonding pairs of electrons, due to their closer proximity to the central atom. In other words, lone pairs "take up more space". Therefore the $H-N-H$ angle is slightly less than 109.5^o. Its actual value is approximately 107^o.

AB_2E_2

A water molecule consists of two bonding pairs and two lone pairs of electrons.

The water molecule, like the ammonia and methane molecules, has a tetrahedral electron group geometry. In the water molecule, two of the electron pairs are lone pairs rather than bonding pairs. The $H-O-H$ bond angle is 104.5^o, which is smaller than the bond angle in NH_3.

Table 4.3.2: Geometries of Molecules in Which the Central Atom Has One or More Lone Pairs

Atoms Plus Lone Pairs Around Central Atom	Number of Surrounding Atoms	Number of Lone Pairs	Electron Group Geometry	Example
3 AB_2E	2	1	Trigonal Planar	O_3
4 AB_3E	3	1	Tetrahedral	NH_3
4 AB_2E_2	2	2	Tetrahedral	H_2O

Contributors and Attributions

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

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4.4: Functional Groups

Learning Outcomes

- Describe the importance and purpose of functional groups in organic reactions.
- Identify and name functional groups in compounds.

With over twenty million known organic compounds in existence, it would be very challenging to memorize chemical reactions for each one. Fortunately, molecules with similar functional groups tend to undergo similar reactions. A **functional group** is defined as an atom or group of atoms within a molecule that has similar chemical properties whenever it appears in various compounds. Even if other parts of the molecule are quite different, certain functional groups tend to react in certain ways.

We've already looked at alkanes, but they are generally unreactive. We primarily use alkanes as a source of energy when they are combusted. While the majority of functional groups involve atoms other than carbon and hydrogen, we will also look at some that include only carbon and hydrogen. Some of the most common functional groups are presented in the following sections.

Organic molecules vary greatly in size and when focusing on functional groups, we want to direct our attention to the atoms involved in the functional group. As a result, the abbreviation R is used in some examples. The letter R is used in molecular structures to represent the "**R**est of the molecule". It consists of a group of carbon and hydrogen atoms of any size. It is used as an abbreviation since a group of carbon and hydrogen atoms does not affect the functionality of the compound. In some molecules, you will see R, R', or R'' which indicates that the R groups in the molecule can be different from one another. For example, R might be $–CH_2CH_3$ while R' is $–CH_2CH_2CH_2CH_3$.

Alkenes and Alkynes

Alkenes are hydrocarbons that contain one or more double bonds between neighboring carbon atoms. Alkynes contain one or more triple bonds between neighboring carbon atoms. The presence of a double or triple bond results in reactivity not present in alkane and alkynes (see figure below) undergo characteristic reactions which will be covered later.

Figure 4.4.1: Alkenes (left) contain one or more double bonds while alkynes (right) contain one or more triple bonds.

Aromatics

The additional functional group that contains only carbon and hydrogen is an aromatic ring which is a six-carbon ring with alternative double bonds. The aromatic ring is also shown as a ring with a circle in the middle representing the double bonds. Aromatic rings (see figure below) are found in many compounds including steroids and medications.

Figure 4.4.2: Aromatic rings contain alternating double bonds.

Alcohols

The **alcohol** functional group involves an oxygen atom that is bonded to one hydrogen atom and one carbon atom. The carbon atom will be part of a larger organic structure. One way to indicate a generic alcohol would be with the formula $\mathrm{R{-}OH}$. R represents any organic fragment in which a carbon atom is directly bonded to the explicitly indicated functional group (in this case, OH). The R group is typically a chain of carbon atoms.

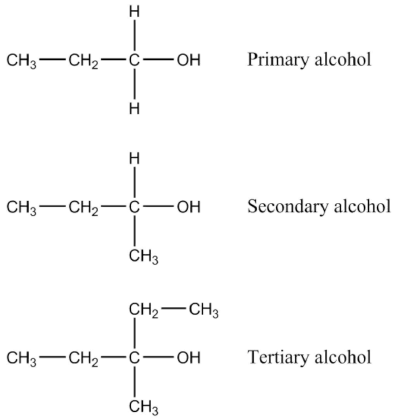


Figure 4.4.3: Primary, secondary, and tertiary alcohols.

Alcohols can be classified as primary, secondary, or tertiary based on the characteristics of the carbon to which it is attached. In a primary alcohol, the carbon bonded directly to the oxygen atom is also bonded to exactly one carbon atom, with the other bonds generally going to hydrogen atoms. In a secondary alcohol, the carbon is attached to two other carbon atoms, and in a tertiary alcohol, the carbon is bonded to three other carbon atoms. The type of alcohol being used will determine the product of certain reactions. Note the naming of alcohols as illustrated in the figure above. The location of the $-\mathrm{OH}$ group is indicated with the number of the carbon to which it is attached.

We are already familiar with several common alcohols. For example, ethanol $\left(\mathrm{CH_3CH_2OH}\right)$ is the alcohol present in alcoholic beverages. It is also widely used in the industrial manufacture of other chemicals. Methanol $\left(\mathrm{CH_3OH}\right)$ is used as a gasoline additive or alternative. Additionally, methanol can be used to manufacture formaldehyde, which is employed in the production of plastics, paints, and other useful substances. Isopropanol is commonly known as rubbing alcohol. In addition to its industrial uses, isopropanol is used to clean various surfaces, including computer monitors, whiteboards, and even skin (e.g., before getting blood drawn).

Ethers

The **ether** functional group consists of an oxygen atom that forms single bonds with two carbon atoms.

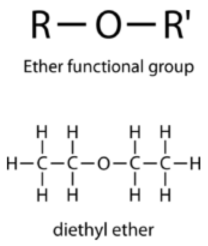


Figure 4.4.4: Ethers

Ethers are good solvents for other organic compounds because of their low reactivity. They readily dissolve nonpolar molecules. Diethyl ether is perhaps the best known ether. It is widely used as a solvent and has been used as an inhalable anesthetic.

Although ethers themselves are relatively unreactive, they can be converted to peroxides after prolonged exposure to oxygen. Peroxides are very reactive and are often explosive at elevated temperatures. Many commercially available ethers come with a small amount of a peroxide scavenger dissolved in them to help prevent this type of safety hazard.

Thiol

The **thiol** functional group contains a sulfur atom bonded to a hydrogen atom. It is very similar to an alcohol functional group with the sulfur replacing the O.

Figure 4.4.5: Thiol

Thiols are also called mercaptans which is derived from the Latin phrase for "capturing mercury" because of the strong bonds it forms with mercury-containing compounds. Some thiol compounds have a distinctive smell similar to rotten eggs. They are often added to natural gas, which itself has no odor, as a way to detect leaks since its odor can be detected by humans in very small amounts. A thiol group is also present in the amino acid cysteine which will be discussed later.

Amines

An **amine** consists of a nitrogen atom bonded to some combination of carbons and hydrogens.

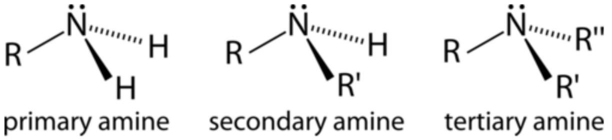


Figure 4.4.6: Amines

Like alcohols, amines can be classified as primary, secondary, or tertiary. However, the rules for assigning these categories are slightly different. In an alcohol, the oxygen atom is always bonded to exactly one carbon atom, so we look at the branching on the adjacent carbon, not the oxygen atom itself. In a neutral amine, the nitrogen can be bonded to one, two, or three carbon atoms, and this is how we decide whether it is called a primary, secondary, or tertiary amine.

Neutral amines are weak bases, because the lone pair on nitrogen can act as a proton acceptor. Many smaller amines have very strong and offensive odors. For example, the aptly-named compounds cadaverine and putrescine are foul-smelling amines, formed as a part of the decay process after death.

Amines serve a wide variety of uses. Diphenylamine acts as a stabilizer for certain types of explosives. Amines are found as components in some lubricating materials, in developers, and are a part of waterproofing textiles. Some amines, such as novocaine, are used as anesthetics. Many pharmaceutical compounds contain amines, including 8 of the 10 most prescribed medications in 2012.

Aldehydes

A very common structural component of organic structures is the **carbonyl**, which is simply a carbon atom and an oxygen atom connected by a double bond. The reactivity of carbonyls is primarily dictated by the polarization of the $\mathrm{C{=}O}$ bond, but the surrounding atoms also play a role in its specific reaction pathways. While carbonyl is a component of many functional groups, it is not itself a functional group.

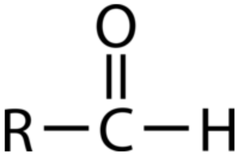


Figure 4.4.7: Aldehyde

An **aldehyde** is a carbonyl in which the carbon atom is bonded to at least one hydrogen atom. The other group attached to the carbonyl may be an R-group or a hydrogen atom. Because the hydrogen atom is so small, the partial positive charge on the carbonyl carbon is very easy for other molecules to approach, making aldehydes a particularly reactive type of carbonyl. Aldehydes are versatile reactants for a wide variety of organic syntheses. Many aldehydes also have distinctive flavors and aromas. For example, the flavor of cinnamon is primarily due to the molecule cinnamaldehyde, and vanillin is the aldehyde most responsible for the smell and taste of vanilla extract.

A special aldehyde is the molecule in which the carbonyl is bonded to two hydrogen atoms. This molecule, called formaldehyde, has a wide variety of uses. By itself, it can be used as a tissue preservative or as a very harsh disinfectant. It is also used as a precursor to various materials, including plastics, resins, and other polymers.

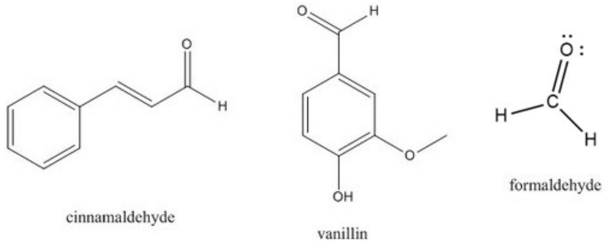
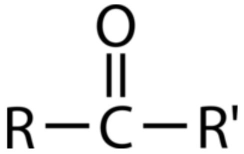


Figure 4.4.8: Examples of molecules with aldehyde functional groups.

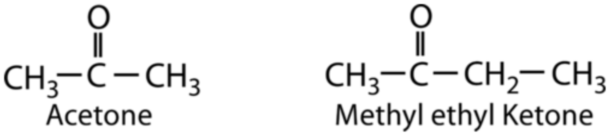
Ketones



$$R-\overset{\overset{\Large O}{\|}}{C}-R'$$

Figure 4.4.9: Ketone

A **ketone** involves a carbonyl in which the carbon atom makes single bonds with two R-groups. Ketones undergo most of the same reactions as aldehydes, but they tend to be slightly less reactive. The simplest ketone is acetone, in which the carbonyl carbon is bonded to two CH_3 groups. This ketone is commonly used to remove fingernail polish and serves as an industrial solvent. Methyl ethyl ketone is used as a paint stripper and a solvent. Ketones are also used in the production of various polymers, either as a building block or as a solvent. The R-group in a ketone can be the same or different as seen in the example.



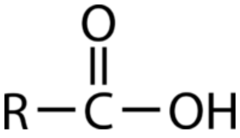
$$CH_3-\overset{\overset{O}{\|}}{C}-CH_3 \qquad CH_3-\overset{\overset{O}{\|}}{C}-CH_2-CH_3$$

Acetone Methyl ethyl Ketone

Figure 4.4.10: Two examples of ketones.

Carboxylic Acids

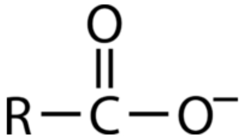
Carboxylic acids are another carbonyl-containing functional group, in which the carbon atom is bonded to an OH group on one side and either a carbon or hydrogen atom on the other.



$$R-\overset{\overset{\Large O}{\|}}{C}-OH$$

Figure 4.4.11*: Carboxylic acid*

As the name implies, carboxylic acids are weak acids. An OH group that is directly connected to a carbonyl will ionize to a small extent when dissolved in water. The reason for this is the relative stability of the resulting anion. A carboxylate ion (see figure below), in which the negative charge is spread over two different oxygen atoms through resonance structures, is more stable than an isolated oxygen-centered anion. The carboxylic acid and carboxylate ion are interchangeable. Carboxylate ions are often present in amino acids.



$$R-\overset{\overset{\Large O}{\|}}{C}-O^-$$

Figure 4.4.12: Carboxylate ion

Carboxylic acids are used in a variety of environments. Formic acid acts as a protective chemical for many stinging insects and plants. Acetic acid gives vinegar its characteristic smell and flavor and is a fundamental biological and industrial building block. Carboxylic acids with longer carbon chains (fatty acids) are used by animals as a way of storing energy and are widely used in the manufacture of soaps. Some compounds contain multiple carboxylic acids within a single molecule. For example, citric acid (three carboxyl groups) is especially abundant in citrus fruits and it used as a flavoring and preservative in many foods and beverages.

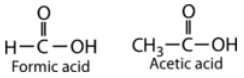


Figure 4.4.13: Carboxylic acids

Esters

An **ester** is similar to a carboxylic acid, in that it contains a carbonyl where the carbon is bonded to one additional oxygen atom and one carbon or hydrogen atom. However, the second oxygen atom is bonded to another carbon instead of to an acidic hydrogen atom. Structurally, carboxylic acids and esters are related to one another in the same way as alcohols and ethers.

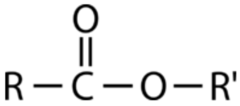


Figure 4.4.14: Ester

Esters can be formed by heating carboxylic acids and alcohols in the presence of an acid catalyst. This process is reversible, and the starting materials can be regenerated by reacting an ester with water in the presence of a weak base.

Some esters have very pleasant odors, so they are used in the manufacture of many perfumes. Propyl acetate contributes to the odor of pears, while isoamyl acetate gives bananas their smell. This ester also serves as an alarm signal for honeybees. Esters are employed in the manufacture of fabrics (polyesters) and Plexiglass. Anesthetics such as procaine and benzocaine also contain esters.

Amides

An **amide** is a carbonyl in which the carbon is attached to one nitrogen atom and one carbon or hydrogen atom. Alternatively, we could define an amide as an amine in which one of the carbon atoms attached to the nitrogen is part of a carbonyl.

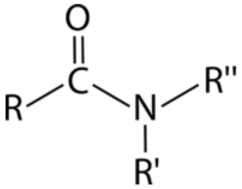


Figure 4.4.15: Amide

An amide can be formed by combining a carboxylic acid and an amine. Only primary and secondary amines can be sued to form amides, since they have a hydrogen that can be replaced with the carbonyl carbon; tertiary amines will not form amides. The amide shown in the figure above was formed from a carboxylic acid and a primary amine.

Amides are used as coloring agents in crayons, pencils, and ink. They are employed in the paper, plastic, and rubber industries. Polyacrylamide is a very widely used amide; it is involved in t he treatment of drinking water and sewage, and in plastics manufacture. The amide Kevlar is widely employed for the production of body armor, and nylon is another type of amide-based polymer.

Haloalkanes

The haloalkanes, also known as alkyl halides, are a group of chemical compounds comprised of an alkane with one or more hydrogens replaced by a halogen atom (Group 17 atom). There is a fairly large distinction between the structural and physical properties of haloalkanes and the structural and physical properties of alkanes. A

$$R-X$$

Figure 4.4.16: Amide

Haloalkanes are found in fire extinguishers, refrigerants, propellants, solvents, and medications. They are also a significant source of pollution and their use has been reduced or eliminated in some products. Chlorofluorocarbons (CFCs) were used as refrigerants in air-conditioners but were found to be a major cause of the depletion of the ozone layer. Research and development of alternatives began in the 1970s. Hydrochlorofluorocarbons (HCFCs) have been used for many years since they cause less damage to the ozone layer, but many countries agreed to eliminate HCFCs by 2020.

Contributors and Attributions

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4.5: Structure and Function (Exercises)

These are homework exercises to accompany Chapter 4 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

4.1: Lewis Electron Dot Structures

(Click here for solutions)

Q4.1.1

Describe a Lewis structure.

Q4.1.2

How many bonds are usually formed by each element?

a. sulfur
b. hydrogen
c. nitrogen
d. phosphorus

Q4.1.3

How many bonds does oxygen tend to form? What combinations of bonds (single, double, triple) can it have?

Q4.1.4

How many bonds does carbon tend to form? What combinations of bonds (single, double, triple) can it have?

Q4.1.5

Using Table 4.1.1, draw the Lewis structure for each molecule.

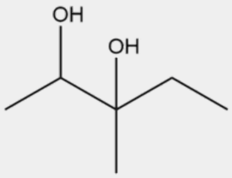
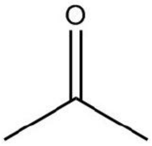
a. H_2S
b. CCl_4
c. PF_3
d. C_2H_4

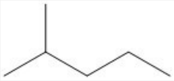
4.2: Representing Structures

(click here for solutions)

Q4.2.1

Given one type of structure (Lewis, condensed, or skeletal), draw the other two.

a.	$CH_3CH_2CH_3$
b.	
c.	
d.	$CH_3CHBrCH_2CH_2CHClCH_3$
e.	$CH_2(OH)CH_2CH_2CH_3$

f.	
g.	

4.3: Molecular Shapes

(click here for solutions)

Q4.3.1

How does the presence of a lone pair affect the shape of the molecule?

Q4.3.2

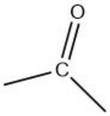
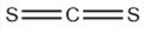
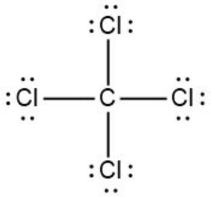
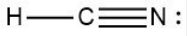
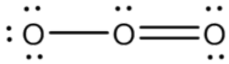
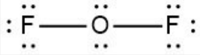
What is the "ideal" bond angle in a tetrahedral (AB_4) molecule? How does that change in trigonal pyramidal (AB_3E)? Bent (AB_2E_2)?

Q4.3.3

What is the ideal bond angle for trigonal planar (AB_3) molecules? What would be the expected bond angle for a bent (AB_2E) molecule?

Q4.3.4

Determine the electron geometry for each of the following molecules.

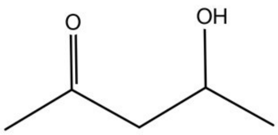
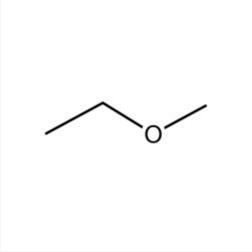
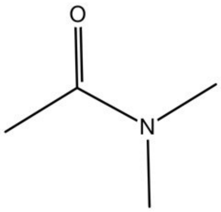
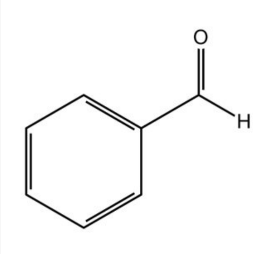
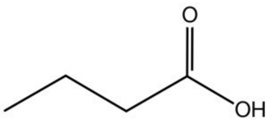
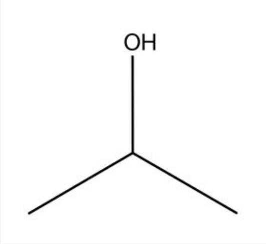
a.	
b.	
c.	
d.	
e.	
f.	

4.4: Functional Groups

(click here for solutions)

Q4.4.1

Circle and label the functional groups in each molecule. There may be more than one functional group in a molecule.

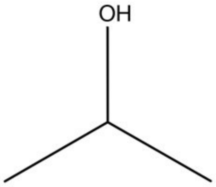
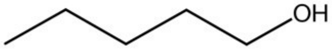
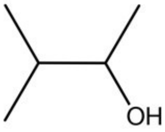
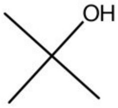
a.	
b.	
c.	
d.	
e.	
f.	

Q4.4.2

Draw a molecule which contains an ester, an alcohol, and an aldehyde.

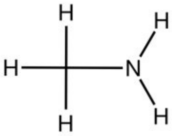
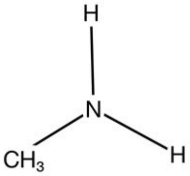
Q4.4.3

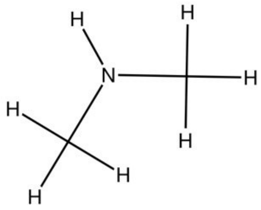
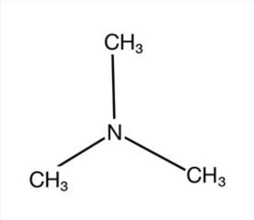
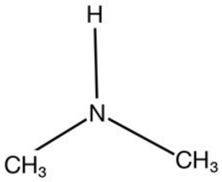
Label each alcohol as primary, secondary, or tertiary.

a.	
b.	
c.	
d.	

Q4.4.4

Label each amine as primary, secondary, or tertiary.

a.	
b.	

c.	
d.	
e.	

Answers

4.1: Lewis Electron Dot Structures

Q4.1.1

A Lewis structure shows the connections between atoms (single, double, or triple bonds) as well as any non-bonding electrons.

Q4.1.2

a. 2
b. 1
c. 3
d. 3

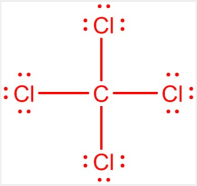
Q4.1.3

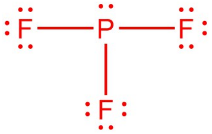
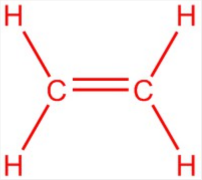
2 bonds total; 2 single bonds or 1 double bond

Q4.1.4

4 bonds total; 4 single or 2 single and 1 double or 2 double or 1 single and 1 triple

Q4.1.5

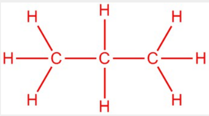
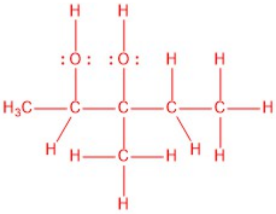
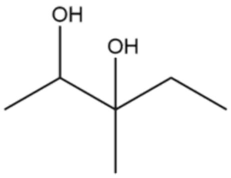
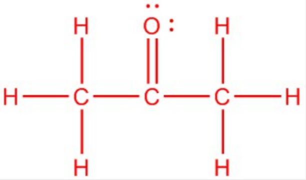
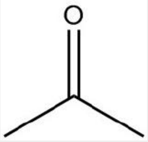
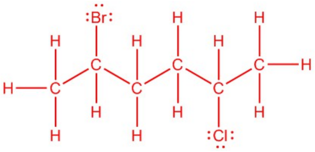
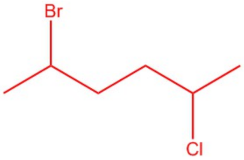
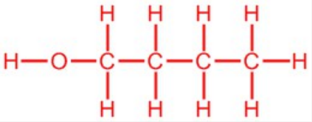
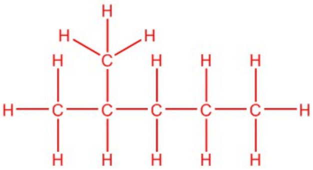
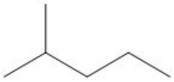
a.	
b.	

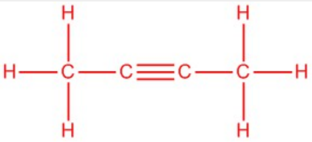
c.	
d.	

4.2: Representing Structures

Q4.2.1

Given one type of structure (Lewis, condensed, or skeletal), draw the other two.

	Lewis Structure	Condensed Structure	Skeletal Structure
a.		$CH_3CH_2CH_3$	
b.		$CH_3CH(OH)C(CH_3)(OH)CH_2CH_3$	
c.		$CH_3C(O)CH_3$	
d.		$CH_3CHBrCH_2CH_2CHClCH_3$	
e.		$CH_2(OH)CH_2CH_2CH_3$	
f.		$CH_3CH(CH_3)CH_2CH_2CH_3$	

g.		CH_3CCCH_3	

4.3: Molecular Shapes

Q4.3.1

A lone pair causes the compression of other bond angles in the molecule.

Q4.3.2

AB_4: 109.5°

AB_3E: 107° (less than AB_4)

AB_2E_2: 104.5° (less than AB_3E)

Q4.3.3

AB_3: 120°

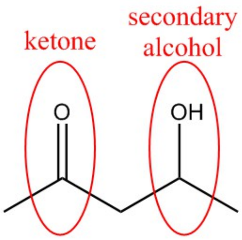
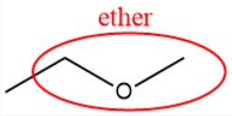
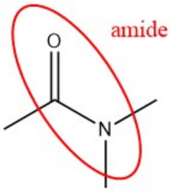
AB_2E: less than 120°

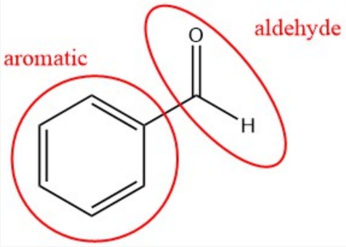
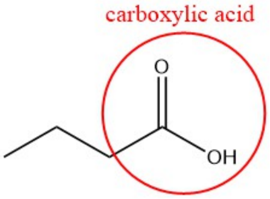
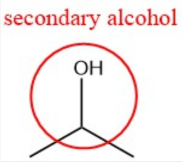
Q4.3.4

a. trigonal planar
b. linear
c. tetrahedral
d. linear
e. trigonal planar
f. tetrahedral

4.4: Functional Groups

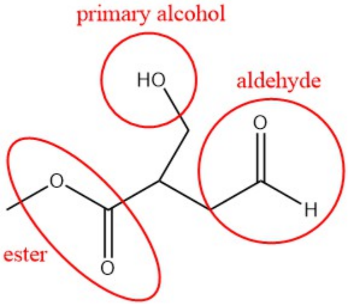
Q4.4.1

a.	
b.	
c.	

d.	
e.	
f.	

Q4.4.2

Answers will vary. One example is shown.



Q4.4.3

a. secondary
b. primary
c. secondary
d. tertiary

Q4.4.4

a. primary
b. primary
c. secondary
d. tertiary
e. secondary

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CHAPTER OVERVIEW

5: Properties of Compounds

5.1: Isomers

Learning Outcomes

- Define isomer.
- Determine the isomeric relationship between a pair of molecules.
- Identify the chiral centers in a molecule.
- Describe different types of isomers.

One of the interesting aspects of organic chemistry is that it is three-dimensional. A molecule can have a shape in space that may contribute to its properties. Molecules can differ in the way the atoms are arranged - the same combination of atoms can be assembled in more than one way. These compounds are known as **isomers**. Isomers are molecules with the same molecular formulas, but different arrangements of atoms. There are several different types of isomers which will be described and a flowchart (see figure below) can help you determine which type of isomers are present.

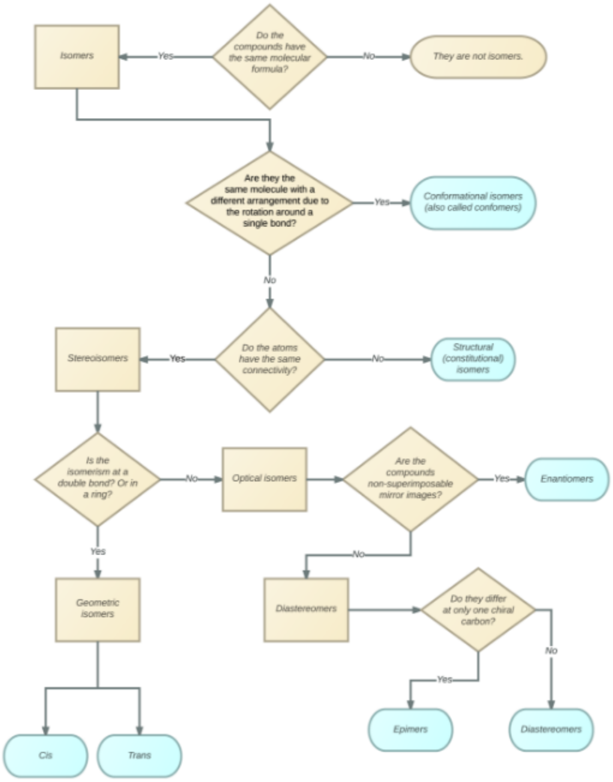


Figure 5.1.1: Isomer flowchart.

Conformational Isomers

Conformational isomers, also known as conformers, differ from one another by their rotation around a single bond. Rotations occur freely around single carbon-carbon bonds. Unlike double and triple bonds, which are "locked" in their orientation, single bonds have no such restrictions.

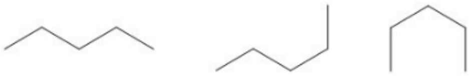
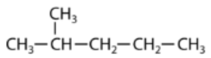


Figure 5.1.2: Conformational isomers of pentane.

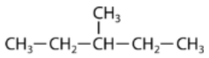
Structural Isomers

A **structural isomer**, also known as a constitutional isomer, is one in which two or more organic compounds have the same molecular formulas but different structures. The two molecules below have the same chemical formula, but are different molecules because they differ in the location of the methyl group.



$$CH_3-\overset{\overset{\large CH_3}{|}}{C}H-CH_2-CH_2-CH_3$$

2-methylpentane



$$CH_3-CH_2-\overset{\overset{\large CH_3}{|}}{C}H-CH_2-CH_3$$

3-methylpentane

Alkenes can also demonstrate structural isomerism. In alkenes, there are multiple structural isomers based on where in the chain the double bond occurs. The condensed structural formulas of 1-butene and 2-butene show this.

$CH_2=CH-CH_2-CH_3$ 1-butene

$CH_3-CH=CH-CH_3$ 2-butene

The number in the name of the alkene refers to the lowest numbered carbon in the chain that is part of the double bond.

Stereoisomers

Stereoisomers have the same connectivity in their atoms but a different arrangement in three-dimensional space. There are different classifications of stereoisomers depending on how the arrangements differ from one another. Notice that in the structural isomers, there was some difference in the connection of atoms. For example, 1-butene has a double bond followed by two single bonds while 2-butene has a single bond, then a double bond, then a single bond. A stereoisomer will have the same connectivity among all atoms in the molecule.

Geometric Isomers

With a molecule such as 2-butene, a different type of isomerism called geometric isomerism can be observed. **Geometric isomers** are isomers in which the order of atom bonding is the same but the arrangement of atoms in space is different. The double bond in an alkene is not free to rotate because of the nature of the bond. Therefore, there are two different ways to construct the 2-butene molecule (see figure below). The image below shows the two geometric isomers, called *cis*-2-butene and *trans*-2-butene.

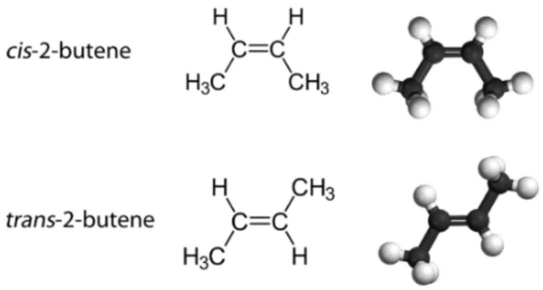


Figure 5.1.3: 2-butene

The *cis* isomer has the two single hydrogen atoms on the same side of the molecule, while the *trans* isomer has them on opposite sides of the molecule. In both molecules, the bonding order of the atoms is the same. In order for geometric isomers to exist, there must be a rigid structure in the molecule to prevent free rotation around a bond. This occurs with a double bond or a ring. In addition, the two carbon atoms must each have two different groups attached in order for there to be geometric isomers. Propene (see figure below) has no geometric isomers because one of the carbon atoms (the one on the far left) involved in the double bond has two single hydrogens bonded to it.

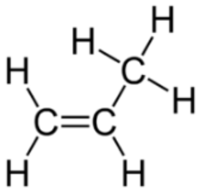


Figure 5.1.4: Propene does not have a geometric isomer.

Physical and chemical properties of geometric isomers are generally different. As with alkenes, alkynes display structural isomerism beginning with 1-butyne and 2-butyne. However, there are no geometric isomers with alkynes because there is only one other group bonded to the carbon atoms that are involved in the triple bond.

Optical Isomers

Stereoisomers that are not geometric isomers are known as optical isomers. Optical isomers differ in the placement of substituted groups around one or more atoms of the molecule. They were given their name because of their interactions with plane-polarized light. Optical isomers are labeled enantiomers or diastereomers.

Enantiomers are non-superimposable mirror images. A common example of a pair of enantiomers is your hands. Your hands are mirror images of one another but no matter how you turn, twist, or rotate your hands, they are not superimposable.

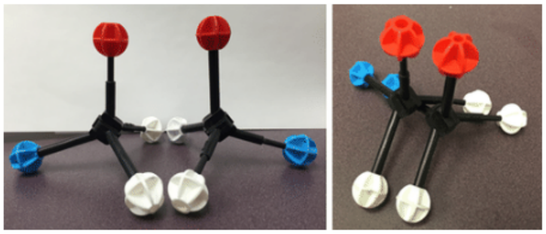


Figure 5.1.5: Two models that are mirror images and superimposable. Since they are superimposable, they are the same molecule and are not isomers.

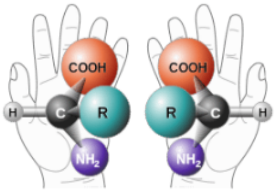


Figure 5.1.6: Your hands and some molecules are mirror images but are not superimposable. These pairs of molecules are called enantiomers.

Objects that have non-superimposable mirror images are called **chiral**. When examining a molecule, carbon atoms with four unique groups attached are considered chiral. Look at the figure below to see an example of a chiral molecule. Note that we have to look beyond the first atom attached to the central carbon atom. The four circles indicate the four unique groups attached to the central carbon atom, which is chiral.

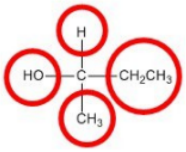


Figure 5.1.7: A chiral carbon has four unique groups attached to it.

Another type of optical isomer are **diastereomers**, which are non-mirror image optical isomers. Diastereomers have a different arrangement around one or more atoms while some of the atoms have the same arrangement. As shown in the figure below, note that the orientation of groups on the first and third carbons are different but the second one remains the same so they are not the same molecule. The solid wedge indicates a group coming out of the page/screen towards you and the dashed line indicates that a group is going away from you "behind" the page/screen.

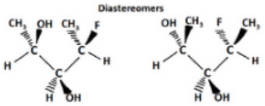


Figure 5.1.8: Diastereomers differ at one or more atom. These molecules are not mirror images and they are not superimposable. They are optical isomers because they have the same connectivity between atoms but a different arrangement of substituent groups.

Epimers are a sub-group of diastereomers that differ at only one location. All epimers are diastereomers but not all diastereomers are epimers.

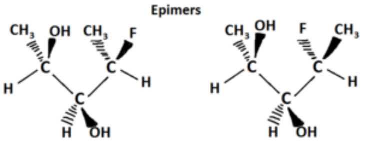


Figure 5.1.9: Epimers have a different arrangement around one atom, while arrangements around the other atoms are the same.

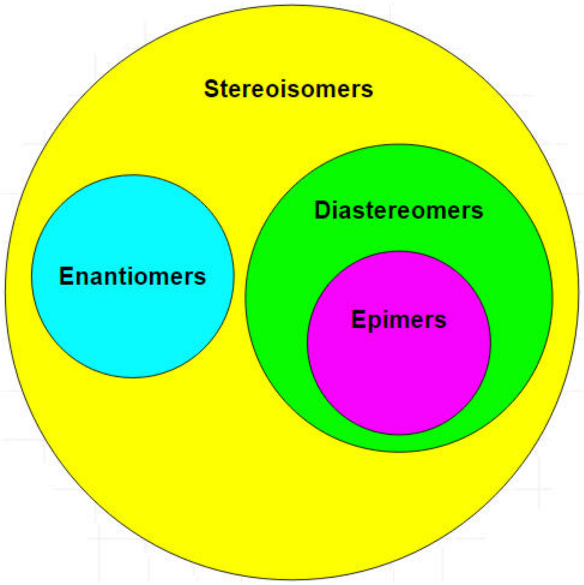


Figure 5.1.10: Diagram showing the division of stereoisomers (also known as optical isomers).

Contributors and Attributions

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5.2: Carbohydrate Structures

Learning Outcomes

- Describe the structure and function of carbohydrates.
- Identify functional groups of carbohydrates.
- Give general name for a carbohydrate molecule (i.e. aldotetrose, ketopentose, etc)
- Label carbohydrates as either D- or L-enantiomers.
- Draw the mirror image of a carbohydrate molecule.
- Distinguish between monosaccharides, disaccharides, and polysaccharides.
- Describe the structure of complex carbohydrates.
- Recognize how carbohydrates determine blood type.

The brain is a marvelous organ. And it's a hungry one, too. The major fuel for the brain is the carbohydrate glucose. The average adult brain represents about 2% of our body's weight, but uses 25% of the glucose in the body. Moreover, specific areas of the brain use glucose at different rates. If you are concentrating hard, (taking a test, for example) certain parts of the brain need a lot of extra glucose while other parts of the brain only use their normal amount. Something to think about.

As a child, you may have been told that sugar is bad for you. Well, that's not exactly true. Essentially, carbohydrates are made of sugar, from a single sugar molecule to thousands of sugar molecules all attached together. Why? One reason is to store energy. But that does not mean you should eat it by the spoonful.

Carbohydrates

Carbohydrates are organic compounds that contain only carbon (C), hydrogen (H), and oxygen (O). They contain a chain of carbons, an aldehyde or a ketone, and hydroxyl groups. Every carbon atom is attached to one oxygen atom. There are thousands of different carbohydrates, but they all consist of one or more smaller units called monosaccharides.

Monosaccharides

The general formula for a **monosaccharide** is $(\mathrm{CH_2O})_n$, where n can be any number greater than two. For example, if n is 6, then the formula can be written $\mathrm{C_6H_{12}O_6}$. This is the formula for the monosaccharide glucose. Another monosaccharide, fructose, has the same chemical formula as glucose, but the atoms are arranged differently. Carbohydrates have many isomers because of the arrangement of the $-\mathrm{OH}$ groups in their structures. Compare the glucose and fructose molecules in the figure below. Can you identify their differences? The only differences are the positions of some of the atoms. These differences affect the properties of the two monosaccharides.

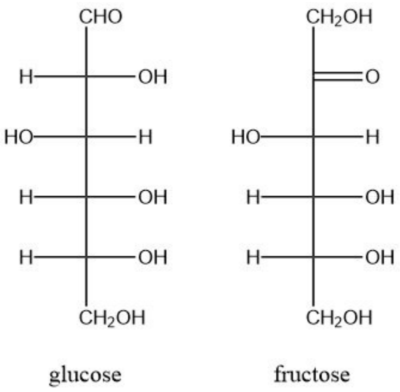


Figure 5.2.1: Structures of glucose and fructose.

Monosaccharides can be classified by the number of carbon atoms they contain: diose (2), triose (3), tetrose (4), pentose (5), hexose (6), heptose (7), and so on. They can also be classified based on whether or not they contain an aldehyde (aldose) or ketone (ketose). We can also combine these two designations to refer to classes of carbohydrates. For example, an aldohexose is a carbohydrate (indicated by the *-ose* ending) with six carbons (*hex*) and an aldehyde group (*aldo*). A ketopentose is a carbohydrate with a ketone and 5 carbons. Both glucose and fructose are hexoses because they contain six carbons but glucose is an aldohexose while fructose (also known as "fruit sugar") is a ketohexose. Other common monosaccharides include galactose (part of lactose), xylose ("wood sugar"), ribose (in RNA), and deoxyribose (in DNA).

Fischer Projections

There are several ways to draw the structure of carbohydrate molecules. The Fischer projection (straight chain) makes it appear that the molecule is flat but it is a three-dimensional molecule. Although we will not be concerned with the 3D orientation, know that the arrangement in the Fischer projection does provide information about the orientation of atoms around each carbon atom.

These projections simplify the drawing of molecules yet retain important information about the arrangement of atoms within the structure. The figure below shows the Fischer projections for the enantiomers (non-superimposable mirror images) of ephedrine and pseudoephedrine. While it may appear that the molecules are the same, they are not because the Fischer projection does not explicitly show the three-dimensional geometry of the molecule.

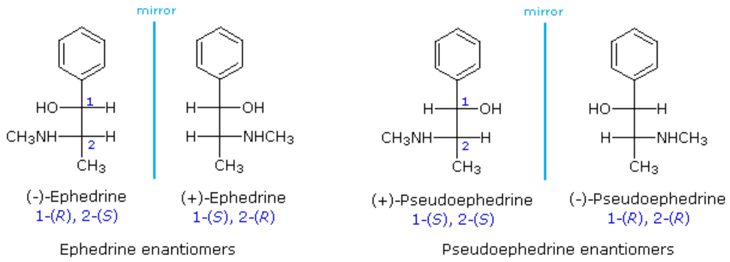


Figure 5.2.2: Fischer projections of ephedrine and pseudoephedrine.

Fischer projections provide an easy way to distinguish among the many, similar carbohydrate molecules that exist. For example, there are sixteen aldohexoses (see figure below). Note the different patterns of the $-\mathrm{OH}$ bonds on the left and right sides of the Fischer projection for each. Changing the orientation of one or more of the $-\mathrm{OH}$ groups changes the identity of the molecule.

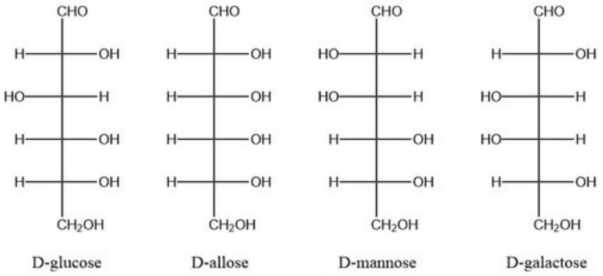


Figure 5.2.3: Four of the sixteen aldohexoses.

Each carbohydrate molecule also has an enantiomer and the two are designated as the D- and L- versions of the compound. The designation is based on the orientation of the $-\mathrm{OH}$ group on the chiral carbon farthest from the aldehyde or ketone. The structures of D-glucose and L-glucose are shown in the figure below. The orientation of all $-\mathrm{OH}$ groups are reversed but only the arrangement of at the carbon indicated by the arrow determines whether the sugar is a D-sugar with the $-\mathrm{OH}$ group on the right or an L-sugar with the $-\mathrm{OH}$ group on the left.

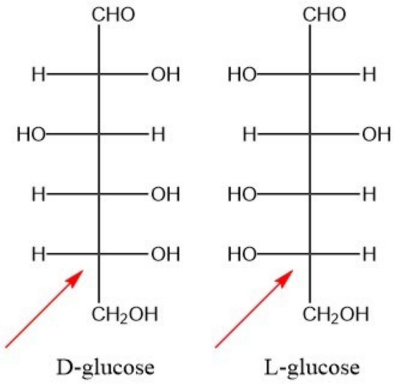
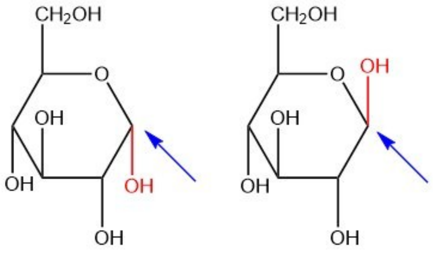


Figure 5.2.4: D-glucose and L-glucose are mirror images of one another.

Haworth Structures

Like Fischer projections, the Haworth structures provide information about a molecule's three-dimensional structure without explicitly showing it in the drawing. Carbohydrates are present in the body in both the chain and ring forms with the latter being more common. Haworth projections provide a simple way to display the ring structures and may or may not show the hydrogen atoms attached to each carbon. Remember, every carbon has four bonds so hydrogens are implied when the structure does not show all four bonds. When the cyclic monosaccharide forms, there are two versions that can form, called α (alpha\) and β (beta) (see figure below). The arrow in the figure indicates the **anomeric** carbon which it the location where the ring forms and where the orientation of the $-\mathrm{OH}$ group can change. The orientation of the other $-\mathrm{OH}$ groups are fixed because they are determined by the orientation of the $-\mathrm{OH}$ groups in the particular monosaccharide (compare to the orientation of the $-\mathrm{OH}$ groups on the left and right sides of the Fischer projections). Each monosaccharide can exist in either α or β form and the two forms will interconvert as the ring opens and closes. The α form occurs when the $-\mathrm{OH}$ group on the anomeric carbon is pointing down and the β version exists when the $-\mathrm{OH}$ group on the anomeric carbon is pointing up.



α-D-glucose β-D-glucose

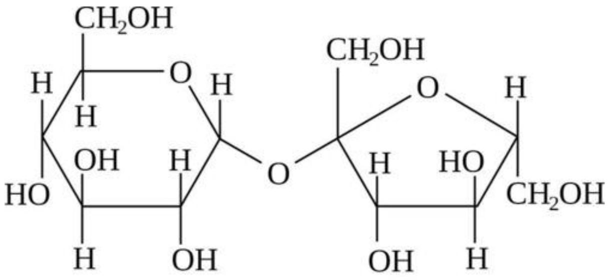
Figure 5.2.5: The cyclic forms of carbohydrates can interconvert between the alpha and beta forms.

As a result of these different orientations, we can have four forms of each monosaccharide. For example, glucose can exist as α-D-glucose, α-L-glucose, β-D-glucose, or β-L-glucose. While the α and β forms can interconvert, the same cannot be said for D and L versions. Naturally occurring monosaccharides are in the D version, called "D sugars". The arrangement within the D or L form is fixed and they cannot interconvert.

Disaccharides

If two monosaccharides bond together, they form a carbohydrate called a **disaccharide**. Two monosaccharides will bond together through a dehydration reaction, in which a water molecule is lost. A dehydration reaction is a **condensation reaction**, a chemical reaction in which two molecules combine to form one single molecule, losing a small molecule in the process. In the dehydration reaction, this small molecule is water. The bond between two monosaccharides is known as a **glycosidic bond**.

An example of a disaccharide is sucrose (table sugar), which consists of the monosaccharides glucose and fructose (see figure below). Other common disaccharides include lactose ("milk sugar") and maltose. Monosaccharides and disaccharides are also called *simple sugars*. They provide the major source of energy to living cells.



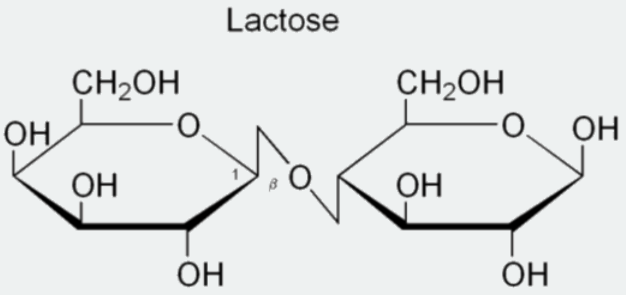
KEY: C = Carbon, H = Hydrogen, O = Oxygen

NOTE: Each unlabeled point where lines intersect represents another carbon atom.

Figure 5.2.6: Sucrose molecule. This sucrose molecule is a disaccharide. It is made up of two monosaccharides: glucose on the left and fructose on the right.

Got milk?

Milk is one of the basic foods needed for good nutrition, especially for growing children. It contains vitamins and minerals necessary for healthy development. Unfortunately, milk and other dairy products also contain lactose, a carbohydrate that can make some people very ill. Lactose intolerance is a condition in which the lactose in milk cannot be digested well in the small intestine. The undigested lactose then moves into the large intestine where bacteria attack it, forming large amounts of gas. Symptoms of lactose intolerance include bloating, cramps, nausea, and vomiting. Avoidance of foods containing lactose is recommended for people who show signs of lactose intolerance. Since dairy products can provide many vital nutrients, tablets can be taken that provide the needed digestive materials in the small intestine. Lactose-free milk is also readily available.



Oligosaccharides

An **oligosaccharide** is a saccharide polymer containing a small number (typically two to ten) of monosaccharides. Oligosaccharides can have many functions; for example, they are commonly found on the plasma membrane of animal cells where they can play a role in cell-cell recognition. In general, they are found attached to compatible amino acid side-chains in proteins or to lipids.

Oligosaccharides are often found as a component of **glycoproteins** or **glycolipids**. They are often used as chemical markers on the outside of cells, often for cell recognition. Oligosaccharides are also responsible for determining blood type.

Blood Type

Carbohydrates attached to red blood cells also determine blood type (see figure below). Of the four blood types, type O has the fewest types of saccharides attached to it while type AB has the most. As a result, type O blood is considered the universal donor because it doesn't have any saccharides present that will appear as foreign when transfused into blood of another type. The reverse is not true. For example, if type A blood is given to a patient with type O blood, it will be rejected by the body because there is an unknown species being introduced to the body. Type A blood cells contain N-acetyl-galactosamine which is not present in type O blood. A person with type O blood would undergo rejection upon receiving type A blood. The Rhesus factor (Rh) in blood also affects donor and acceptor properties but it does not depend on carbohydrates. The Rh factor is determined by the presence (Rh+) or absence (Rh-) of a specific protein on the surface of red blood cells.

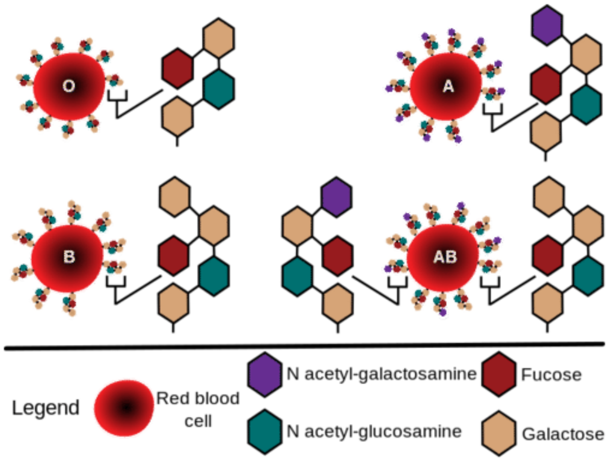


Figure 5.2.7: ABO blood types.

Polysaccharides

Polysaccharides are long carbohydrate molecules of repeated monomer units joined together by glycosidic bonds. A polysaccharide may contain anywhere from a few monosaccharides to several thousand monosaccharides. Polysaccharides are also called **complex carbohydrates**. Polysaccharides have a general formula of $\mathrm{Cx(H_2O)y}$, where x is usually a large number between 200 and 2500.

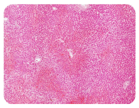
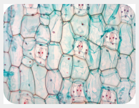
Starches are one of the more common polysaccharides. Starch is made up of a mixture of amylose (15-20%) and amylopectin (80-85%). Amylose consists of a linear chain of several hundred glucose molecules and amylopectin is a branched molecules made of several thousand glucose units. Starches can be digested by **hydrolysis reactions**, catalyzed by enzymes called **amylases**, which can break the glycosidic bonds. Humans and other animals have amylases, so they can digest starches. Potato, rice, wheat, and maize are major sources of starch in the human diet. The formations of starches are the ways that plants store glucose. **Glycogen** is sometimes referred to as *animal starch*. Glycogen is used for long-term energy storage in animal cells. Glycogen is made primarily by the liver and the muscles.

Are we there yet?

As the weather warms up, the runners come out. Not just the casual joggers, but those really serious ones who actually enjoy running all 26.2 miles of a marathon. Prior to these races (and a lot of shorter ones), you hear a lot about carbo-loading. This practice involves eating a lot of starch in the days prior to the race. The starch is converted to glucose, which is normally used for biochemical energy. Excess glucose is stored as glycogen in liver and muscle tissue to be used when needed. If there is a lot of glycogen available, the muscles will have more biochemical energy to draw on when needed for the long run. The rest of us will just sit at the sidewalk restaurant eating our spaghetti and enjoying watching other people work hard.

The main functions of polysaccharides are to store energy and form structural tissues. Examples of several other polysaccharides and their roles are listed in the table below. These complex carbohydrates play important roles in living organisms.

Table 5.2.1: Complex Carbohydrates

Complex Carbohydrate	Function	Organism
Starch	Stores energy	Plants 
Amylose	Stores energy	Plants
Glycogen	Stores energy	Animals 
Cellulose	Forms cell walls	Plants 
Chitin	Forms an exoskeleton	Some animals 

Contributors and Attributions

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5.3: Polarity and Intermolecular Forces

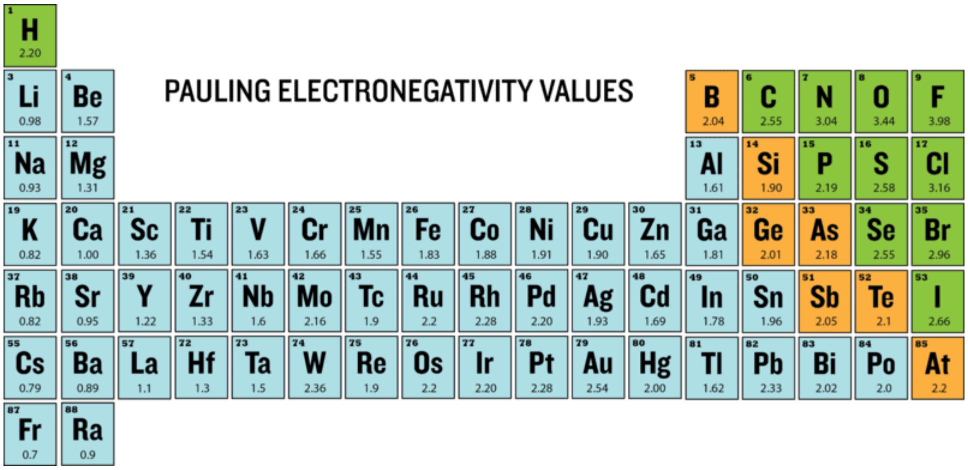
Learning Outcomes

- Define electronegativity.
- Describe how the electronegativity difference between two atoms in a covalent bond results in the formation of a nonpolar covalent, polar covalent, or ionic bond.
- Predict polarity of a molecule.
- Describe how molecular geometry plays a role in determining whether a molecule is polar or nonpolar.
- Distinguish between the following three types of intermolecular forces: dipole-dipole forces, London dispersion forces, and hydrogen bonds.
- Identify types of intermolecular forces in a molecule.
- Describe how chemical bonding and intermolecular forces influence the properties of various compounds.

In an ionic bond, one or more electrons are transferred from one atom to another. In a covalent bond, one or more pairs of electrons are shared between atoms. However, bonding between atoms of different elements is rarely purely ionic or purely covalent.

Bond Polarity

Bond polarity is determined by the difference in electronegativity and is defined as the relative ability of an atom to attract electrons when present in a compound. The electronegativities of various elements are shown below. Note that elecronegativity values increase from left to right and from bottom to top on the periodic table. The degree to which a given bond is ionic or covalent is determined by calculating the difference in electronegativity between the two atoms involved in the bond.



PAULING ELECTRONEGATIVITY VALUES

1 H 2.20																
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
55 Cs 0.79	56 Ba 0.89	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.20	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Tl 1.62	82 Pb 2.33	83 Bi 2.02	84 Po 2.0	85 At 2.2
87 Fr 0.7	88 Ra 0.9															

Figure 5.3.1: Electronegativities of the elements.

As an example, consider the bond that occurs between an atom of potassium and an atom of fluorine. Using the table, the difference in electronegativity is $4.0 - 0.8 = 3.2$. Because the difference in electronegativity is relatively large, the bond between the two atoms is primarily ionic. Since the fluorine atom has a much larger attraction for electrons than the potassium atom does, the valence electron from the potassium atom is considered to have completely transferred to the fluorine atom. The figure below shows how the difference in electronegativity relates to the ionic or covalent character of a chemical bond.

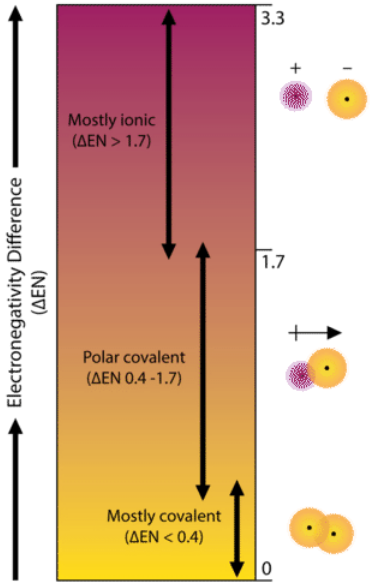


Figure 5.3.2: The difference in electronegativity between the two elements involved in a chemical bond is predictive of the type of bond made by those two atoms. A small difference (<0.4) results in a nonpolar covalent bond, an intermediate difference (0.4 - 1.7) results in a polar covalent bond, and a large difference (>1.7) results in an ionic bond.

According to the figure above, a difference in electronegativity (Δ EN) greater than 1.7 results in a bond that is mostly ionic in character.

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.7 is considered to be mostly covalent in character. However, a distinction is often made between two general types of covalent bonds. A **nonpolar covalent bond** *is a covalent ond in which the onding electrons are shared equally between the two atoms.* In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms (see figure below).

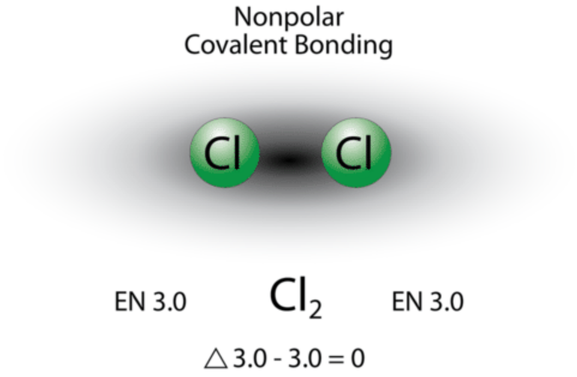


Figure 5.3.3: A nonpolar covalent bond is one in which the shared electrons are distributed equally between the two atoms.

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the $\mathrm{Cl_2}$ molecule is symmetrical. Any diatomic molecule in which the two atoms are the same element must be joined by a nonpolar covalent bond.

There are seven **diatomic elements**, which are *elements whose natural form is of a diatomic molecule.* They are hydrogen $(\mathrm{H_2})$, nitrogen $(\mathrm{N_2})$, oxygen $(\mathrm{O_2})$, fluorine $(\mathrm{F_2})$, chorine $(\mathrm{Cl_2})$, bromine $(\mathrm{Br_2})$, and iodine $(\mathrm{I_2})$. By forming a diatomic molecule, both atoms in each of these molecules satisfy the octet rule, resulting in a structure that is much more stable than the isolated atoms

Notice from the figure above that molecules in which the electronegativity difference is very small (<0.4) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine (Δ EN $= 3.16 - 2.96 = 0.20$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.4 and 1.7 is called a polar covalent bond. A **polar covalent bond** *is a covalent bond in which the atoms have an unequal attraction for electrons, so the sharing is unequal.* In a polar covalent bond, sometimes simply called a polar bond, the distribution of shared electrons within the molecule is no longer symmetrical (see figure below).

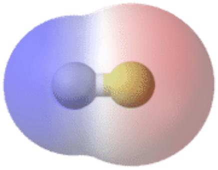
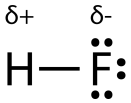


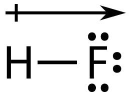
Figure 5.3.4: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

The hydrogen fluoride molecule has an electronegativity difference of 1.9, which places it in the category of being slightly ionic. However, the hydrogen ion $\left(H^+\right)$ is so very small that it is not capable of adopting the crystal lattice structure of an ionic compound. Hydrogen fluoride is a highly polar molecule. Because of its greater electronegativity, the electron density around the fluorine atom is much higher than the electron density around the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ) along with a positive or negative sign to indicate that an atom has a partial positive or negative charge.



The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.



Example 5.3.1

Which type of bond will form between each of the following pairs of atoms?

a. C and O
b. Na and N
c. B and H

Solution

Step 1: List the known quantities and plan the problem.

Known

Using the electronegativity chart:

- $\mathrm{C} = 2.5, \ \mathrm{O} = 3.5$
- $\mathrm{Na} = 0.9, \ \mathrm{N} = 3.0$
- $\mathrm{B} = 2.0, \ \mathrm{H} = 2.1$

Step 2: Solve.

Calculate the difference and use the diagram above to identify the bond type.

$$3.5 - 2.5 = 1.0 \rightarrow \mathrm{C{-}O} \text{ bond is polar covalent} \tag{5.3.1}$$

$$3.0 - 0.9 = 2.1 \rightarrow \mathrm{Na{-}N} \text{ bond is ionic} \tag{5.3.2}$$

$$2.1 - 2.0 = 0.1 \rightarrow \mathrm{B{-}H} \text{ bond is nonpolar covalent} \tag{5.3.3}$$

> *Step 3: Think about your result.*
>
> Bonds between nonmetal atoms are generally covalent in nature (A and C), while bond between a metal atom and a nonmetal atom are generally ionic.

Molecular Polarity

A **polar molecule** *is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative*. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. *A molecule with two poles is called a* **dipole**. Hydrogen fluoride is a dipole. A simplified way to depict molecules is pictured below (see figure below).

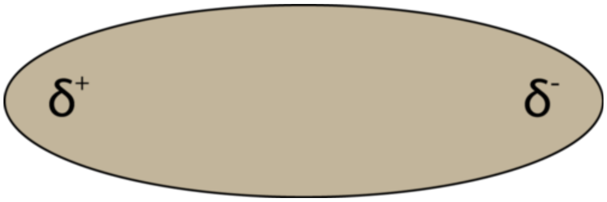


Figure 5.3.5: A molecular dipole results from the unequal distribution of electron density throughout a molecule.

When placed between oppositely charged plates, polar molecules orient themselves so that their positive ends are closer to the negative plate and their negative ends are closer to the positive plate (see figure below).

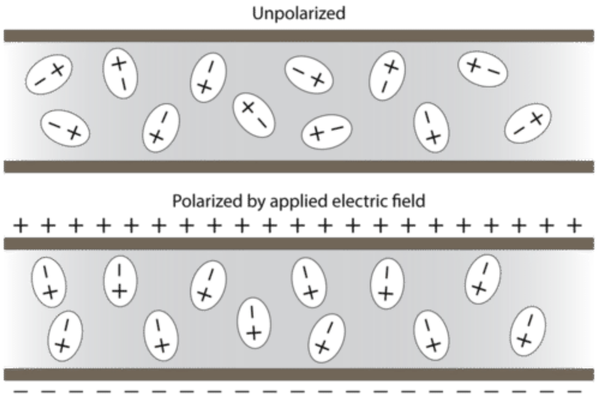


Figure 5.3.6: Polar molecules are randomly oriented in the absence of an applied electric field (top). In an electric field, the molecules orient themselves to maximize the attraction between opposite charges (bottom).

Experimental techniques involving electric fields can be used to determine if a certain substance is composed of polar molecules and to measure the degree of polarity.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. Pictured below (see figure below) is a comparison between carbon dioxide and water. Carbon dioxide (CO_2) is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented in this way, they cancel each other out, and the overall molecular polarity of CO_2 is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape the dipoles do not cancel each other out, and the water molecule is polar. In the figure below, the net dipole is shown in blue and points upward.

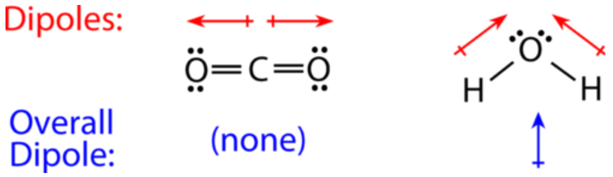


Figure 5.3.7: The molecular geometry of a molecule affects its polarity. In CO_2, the two polar bonds cancel each other out, and the result is a nonpolar molecule. Water is polar because its bent shape means that the two polar bonds do not cancel.

Some other molecules are shown below (see figure below). Notice that a tetrahedral molecule such as $\mathrm{CH_4}$ is nonpolar. However, if one of the peripheral H atoms is replaced by another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule $(\mathrm{BF_3})$ may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule $(\mathrm{NH_3})$ is polar because of the pair of electrons in the nitrogen atoms.

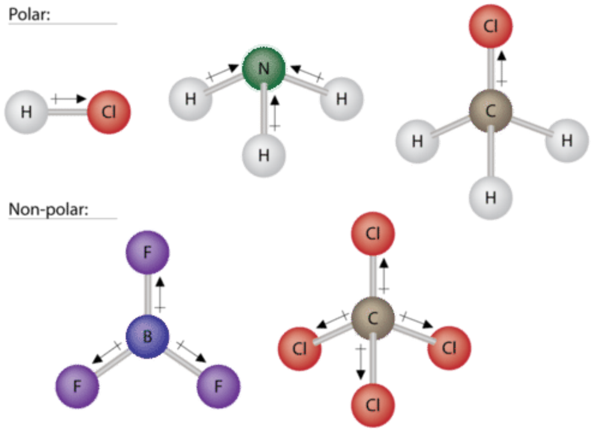


Figure 5.3.8: Some examples of polar and nonpolar molecules with various molecular geometries.

Intermolecular Forces

Covalent and ionic bonds can be called intramolecular forces: forces that act within a molecule or crystal. Molecules also attract other molecules. **Intermolecular forces** *are attractions that occur between molecules.* Intermolecular forces are weaker than either ionic or covalent bonds. However, the varying strengths of different types of intermolecular forces are responsible for physical properties of molecular compounds such as melting and boiling points and the amount of energy needed for changes in state.

London Dispersion Forces

Dispersion forces are the weakest of all intermolecular forces. They are often called London forces after Fritz London (1900 - 1954), who first proposed their existence in 1930. **London dispersion forces** *are intermolecular forces that occur between all atoms and molecules due to the random motion of electrons.*

For example, the electron cloud of a helium atom contains two electrons, and, when averaged over time, these electrons will distribute themselves evenly around the nucleus. However, at any given moment, the electron distribution may be uneven, resulting in an *instantaneous dipole*. This weak and temporary dipole can subsequently influence neighboring helium atoms through electrostatic attraction and repulsion. The formation of an induced dipole is illustrated below.

Figure 5.3.9: Random fluctuations in the electron density within the electron cloud of a helium atom results in a short-lived ("instantaneous") dipole. The attractive force between instantaneous dipoles and the resulting induced dipoles in neighboring molecules is called the London dispersion force.

The instantaneous and induced dipoles are weakly attracted to one another. The strength of dispersion forces increases as the total number of electrons in the atoms or nonpolar molecules increases. The halogen group consists of four elements that all take the form of nonpolar diatomic molecules. Listed below is a comparison of the melting and boiling points for each.

Table 5.3.1: Melting and Boiling Points of Halogens

Molecule	Total Number of Electrons	Melting Point $(^\circ\mathrm{C})$	Boiling Point $(^\circ\mathrm{C})$	Physical State at Room Temperature
$\mathrm{F_2}$	18	-220	-188	gas
$\mathrm{Cl_2}$	34	-102	-34	gas

Molecule	Total Number of Electrons	Melting Point $(^\circ\text{C})$	Boiling Point $(^\circ\text{C})$	Physical State at Room Temperature
Br_2	70	-7	59	liquid
I_2	106	114	184	solid

The dispersion forces are strongest for iodine molecules because they have the greatest number of electrons. The relatively stronger forces result in melting and boiling points which are the highest of the halogen group. These forces are strong enough to hold iodine molecules close together in the solid state at room temperature. The dispersion forces are progressively weaker for bromine, chlorine, and fluorine, as illustrated by their steadily lower melting and boiling points. Bromine is a liquid at room temperature, while chlorine and fluorine are gases. Because gaseous molecules are so far apart from one another, intermolecular forces are nearly nonexistent in the gas state, and so the dispersion forces in chlorine and fluorine only become measurable as the temperature decreases and they condense into the liquid state.

Dipole-Dipole Forces

Dipole-dipole forces *are the attractive forces that occur between polar molecules* (see figure below). A molecule of hydrogen chloride has a partially positive hydrogen atom and a partially negative chlorine atom. A collection of many hydrogen chloride molecules will align themselves so that the oppositely charged regions of neighboring molecules are near each other.

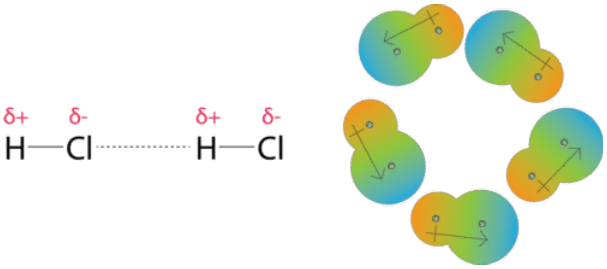


Figure 5.3.10: Dipole-dipole forces result from the attraction between the positive end of one dipole and the negative end of a neighboring dipole. Dipole-dipole forces are similar to ionic bonds, but because they involve only partial charges, they are much weaker.

Hydrogen Bonding

The attractive force between water molecules is an unusually strong type of dipole-dipole interaction. Water contains hydrogen atoms that are bound to a highly electronegative oxygen atom, making for very polar bonds. The partially positive hydrogen atom of one molecule is then attracted to the oxygen atom of a nearby water molecule (see figure below).

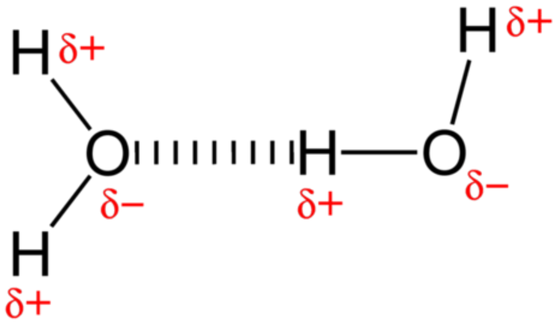


Figure 5.3.11: A hydrogen bond in water occurs between the hydrogen atom of one water molecule and the lone pair of electrons on the oxygen atom of a neighboring water molecule.

A **hydrogen bond** *is an intermolecular attractive force in which a hydrogen atom, that is covalently bonded to a small, highly electronegative atom, is attracted to a lone pair of electrons on an atom in a neighboring molecule.* Hydrogen bonds are very strong compared to other dipole-dipole interactions, but still much weaker than a covalent bond. A typical hydrogen bond is about 5% as strong as a covalent bond.

Hydrogen bonding occurs only in molecules where hydrogen is covalently bonded to one of three elements: fluorine, oxygen, or nitrogen. These three elements are so electronegative that they withdraw the majority of the electron density from the covalent bond with hydrogen, leaving the H atom very electron-deficient. Because the hydrogen atom does not have any electrons other than the ones in the covalent bond, its positively charged nucleus is almost completely exposed, allowing strong attractions to other nearby lone pairs of electrons.

The hydrogen bonding that occurs in water leads to some unusual, but very important properties. Most molecular compounds that have a mass similar to water are gases at room temperature. However, because of the strong hydrogen bonds, water molecules are able to stay condensed in the liquid state. The figure below shows how its bent shape and the presence of two hydrogen atoms per molecule allows each water molecule to hydrogen bond with several other molecules.

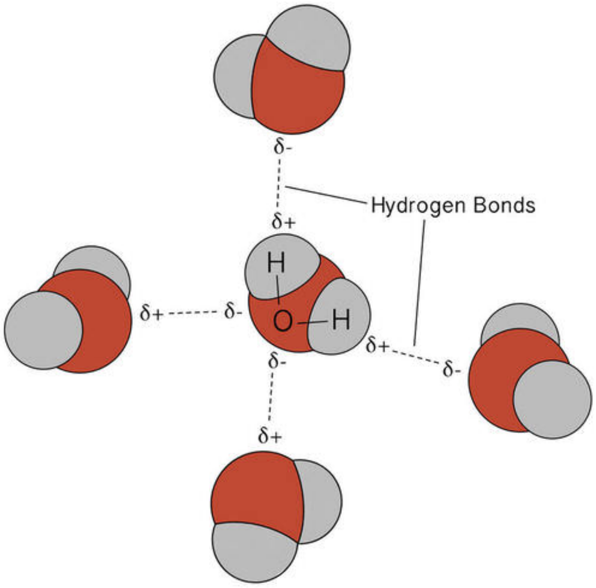


Figure 5.3.12: Multiple hydrogen bonds occur simultaneously in water because of its bent shape and the presence of two hydrogen atoms per molecule.

In the liquid state, the hydrogen bonds of water can break and reform as the molecules flow from one place to another. When water is cooled, the molecules begin to slow down. Eventually, when water is frozen to ice, the hydrogen bonds become more rigid and form a well-defined network (see figure below).

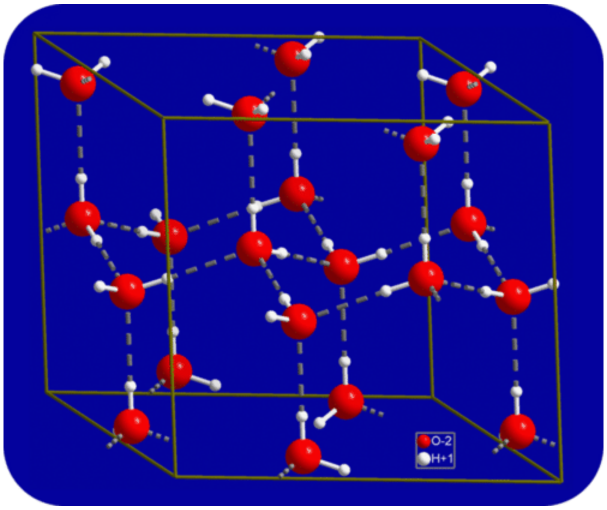


Figure 5.3.13: When water freezes to ice, the hydrogen bonding network becomes fixed until the ice melts. Each oxygen atom has an approximately tetrahedral geometry that includes two covalent bonds and two hydrogen bonds.

The bent shape of the molecules leads to gaps in the hydrogen bonding network of ice. Ice has the very unusual property that its solid state is less dense than its liquid state. As a result, ice floats in liquid water. Virtually all other substances are denser in the solid state than in the liquid state. Hydrogen bonds also play a very important biological role in the physical structures of proteins and nucleic acids.

Boiling Points and Bonding Types

In order for a substance to enter the gas phase, its particles must completely overcome the intermolecular forces holding them together. Therefore, a comparison of boiling points is essentially equivalent to comparing the strengths of the attractive intermolecular forces exhibited by the individual molecules. For small molecular compounds, London dispersion forces are the weakest intermolecular forces. Dipole-dipole forces are somewhat stronger, and hydrogen bonding is a particularly strong form of dipole-dipole interaction. However, when the mass of a nonpolar molecule is sufficiently large, its dispersion forces can be stronger than the dipole-dipole forces in a lighter polar molecule. Thus, nonpolar Cl_2 has a higher boiling point than polar HCl.

Table $5.3.2$: Intermolecular Forces and Boiling Points

Substance	Strongest Intermolecular Force	Boiling Point $(^\circ C)$
H_2	dispersion	-253
Ne	dispersion	-246
O_2	dispersion	-183
Cl_2	dispersion	-34
HCl	dipole-dipole	-85
HBr	dipole-dipole	-66
H_2S	dipole-dipole	-61
NH_3	hydrogen bonding	-33
HF	hydrogen bonding	20
H_2O	hydrogen bonding	100

Supplemental Resources

- Electronegativity: www.chemguideco.uk/atoms/bond...elecroneg.html
- Intermolecular Bonding - van der Waals Forces: www.chemguidecouk/atoms/bonding/vdw.html
- Intermolecular Bonding - Hydrogen Bonds: www.chemguide.co.uk/bonding/hbond.html
- Ionic bond formation: www.dlt.ncssm/edu/core/Chapte...icBonding.html
- Polar covalent bond formation: http://www.dlt.ncssm.edu/core/Chapte...arBonding.html
- Nonpolar covalent bond formation: www.dlt.ncssm/edu/core/Chapte...ntBonding.html

Contributors and Attributions

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

5.4: Chromatography

Learning Outcomes

- Define chromatography.
- Distinguish between the stationary and mobile phases.
- Explain how components of a substance are separated based on attraction to each phase.

When a pure substance is mixed with another pure substance in which it is soluble, the substances become completely interspersed at the molecular level. In thinking about making solutions at the molecular level an analogy to a can of marbles may be useful. In the analogy, a layer of red marbles is placed in the bottom of a can and covered with a second layer of white marbles. After shaking the can for a short time, the marbles are mixed randomly.

Now let us imagine that you want to collect all the red marbles again. If you simply shake the can, it is unlikely that you will ever divide the marbles into two layers, each with only one kind of marble. Similarly, if two miscible liquids are combined, a chemist cannot simply un-mix the liquids into pure components.

Continuing the analogy, what if a few green marbles and blue marbles are placed into the can? Given enough red and white marbles, it may be difficult to determine that the green marbles and blue marbles are actually there. Similarly, when chemists have a multi-component solution which may contain traces of important chemical species, they are faced with the challenge of detecting whether these chemicals are present in solution.

To deal with these difficulties, chemists employ different methods to separate solutions into their components. Two essential techniques are distillation and chromatography.

Another useful set of techniques for separating mixtures is called **chromatography**. Perhaps the simplest of these techniques to describe is paper chromatography, an example of which is shown in the video below.

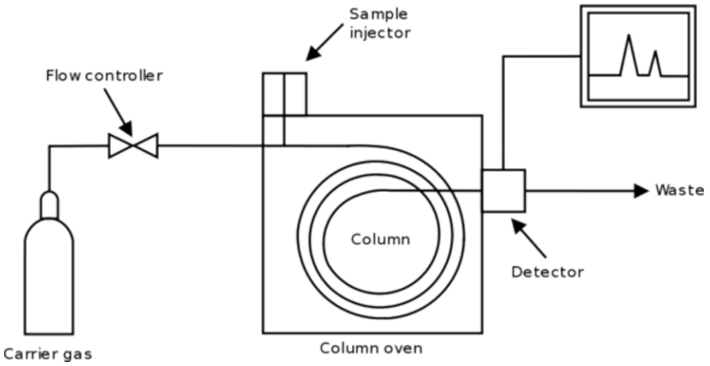
Video: Simple paper chromatography. A simple demonstration on paper chromatography using marker ink and water

Three substances are applied to a strip of chromatography-grade paper (the **stationary phase** of this experiment). As the liquid level rises and meets the spots, the sample partially dissolves in the liquid (the **mobile phase** because it is moving) and travels up the plate within the solution. Different substances will travel different distances along the plate. The distance that a substance will travel depends on how strongly it adheres to the stationary phase (a process called **adsorption**) versus how much time it spends dissolved in the mobile phase.

The more a substance adsorbs, the less it dissolves and the less it moves along the plate. The pink and blue spots at the end of the video are examples of substances highly adsorbed to the stationary phase. The less a substance adsorbs, the more it dissolves and the farther it travels, such as the yellow on the far left. The process is continued until a good separation is created. In this manner, a mixture of substances may be separated such as the middle sample, which was originally green but separated into blue and yellow. Notice that while it was not initially obvious that the middle spot contained both substances, this fact is clear after performing paper chromatography.

All forms of chromatography work on the same general principle as paper chromatography. There is always a stationary phase which does not move and a mobile phase which does. The various components in the mixture being chromatographed separate from each other because they are more strongly held by one phase or the other. Those which have the greatest affinity for the mobile phase move along the fastest.

The most important form of chromatography is **gas chromatography** or vapor-phase chromatography. A long column is packed with a finely divided solid whose surface has been coated with an inert liquid. This liquid forms the stationary phase. The mobile phase is provided by an inert **carrier gas**, such as He or N_2, which passes continuously through the column (seen below), among the solid particles. A liquid sample can be injected into the gas stream at the **sample injector** and vaporized just before it enters the tube. As this sample is carried through the **column** by the slow stream of gas, those components which are most soluble in the inert liquid are held up, while the less-soluble components move on more rapidly. The components thus emerge one by one from the end of the tube, into the detector. In this way it is possible to separate and analyze mixtures of liquids which it would be impossible to deal with by distillation or any other technique.



The development of chromatography is one of the major revolutions in technique in the history of chemistry, comparable to that which followed the development of an accurate balance. Separations which were previously considered impossible are now easily achieved, sometimes with quite simple apparatus. This technique is particular essential to the science of biochemistry, in which complex mixtures are almost always encountered. In the field of environmental chemistry, chromatography has helped us separate and detect very low concentrations of contaminants like DDT or PCB (polychlorinated biphenyls). The major drawback to chromatography is that it does not lend itself to large-scale operation. As a result it remains largely a laboratory, rather than an industrial, technique for separating mixtures.

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5.5: Properties of Compounds (Exercises)

These are homework exercises to accompany Chapter 5 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

5.1: Isomers

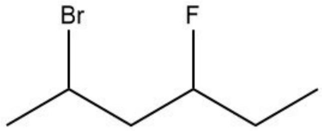
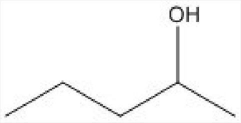
(click here for solutions)

Q5.1.1

Define isomer.

Q5.1.2

Circle the chiral carbons in each structure.

a.	
b.	
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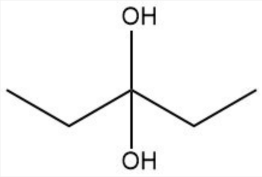
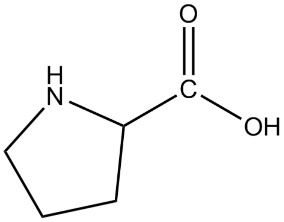
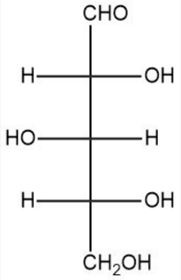
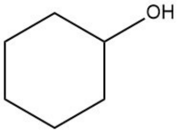
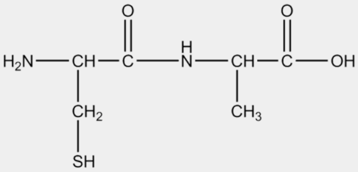
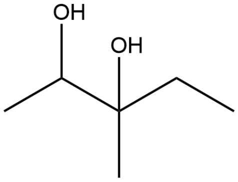
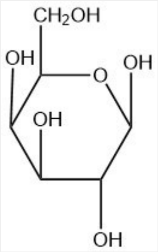
d.	
e.	
f.	

g.	
h.	 $H_2N-CH(CH_2SH)-C(=O)-NH-CH(CH_3)-C(=O)-OH$
i.	
j.	

Q5.1.3

Determine whether each molecule can have a geometric isomer.

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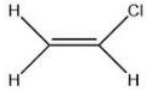

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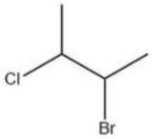

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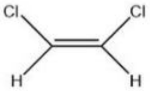

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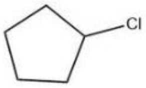

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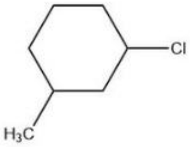

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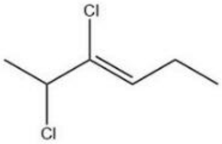
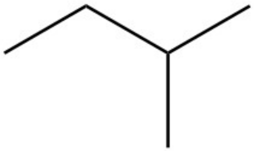
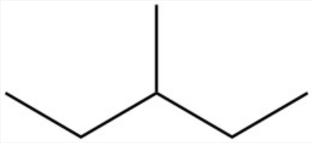
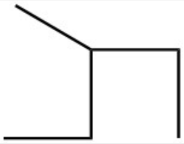
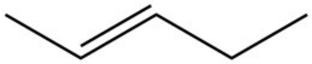
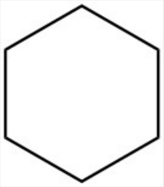
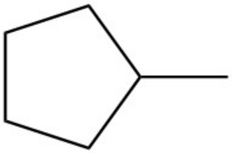


Q5.1.4

For each pair, determine if they are isomers. If they are isomers, identify the type.

a.		

b.		
c.		
d.		
e.		
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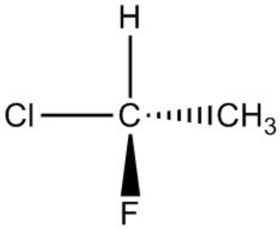
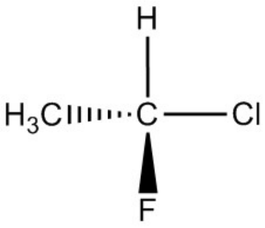
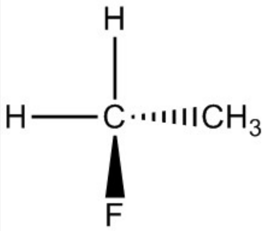
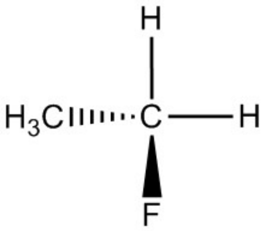
g.		
h.		

Q5.1.5

Draw two isomers with the formula C_2H_6O

5.2: Carbohydrate Structures

(click here for solutions)

Q5.2.1

A monosaccharide has 4 carbon atoms. What is its chemical formula?

Q5.2.2

What are the differences between monosaccharides, disaccharides, oligosaccharides, and polysaccharides?

Q5.2.3

Compare and contrast simple and complex carbohydrates.

Q5.2.4

Draw each of the following structures.

a. aldoheptose
b. ketopentose
c. aldotriose
d. ketotetrose

Q5.2.5

Draw the Fischer projections of two aldohexoses that are different from the four given in Figure 5.2.3.

Q5.2.6

Draw the enantiomers of D-glucose, D-allose, D-mannose, and D-galactose. (See Figure 5.2.3).

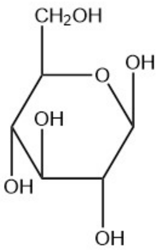
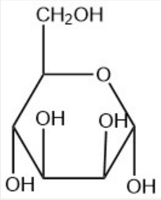
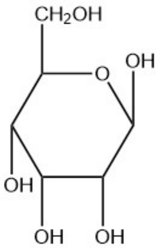
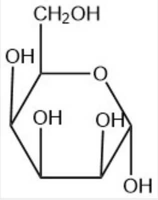
Q5.2.7

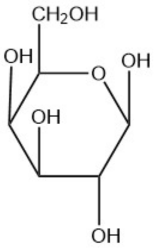
Determine if each of the following pairs are diastereomers? Epimers? (Look up the structures in the chapter or online.)

a. D-glucose and D-allose
b. D-glucose and D-galactose
c. D-allose and D-mannose
d. D-allose and D-galactose

Q5.2.8

For each of the given Haworth structures, identify the sugar as being in the α or β form.

a.	
b.	
c.	
d.	

e.	

Q5.2.9

Refer to Figure 5.2.7 when answering these questions.

a. Which blood type is the universal donor? Why?
b. Which blood type is the universal acceptor? Why?
c. Type A blood can be donated to recipients with which types of blood?
d. Type B recipients can accept which types of blood?

5.3: Polarity and Intermolecular Forces

(click here for solutions)

Q5.3.1

Define electronegativity.

Q5.3.2

Describe the periodic trends for electronegativity values.

Q5.3.3

Distinguish between nonpolar and polar covalent bonds.

Q5.3.4

Describe the bond between each pair of elements as ionic, polar covalent, or nonpolar covalent. Refer to Table 5.3.1 when answering this question. For the exam, if you **need** specific values, a table of electronegativity values will be provided. Note that many questions can be answered **without** the table by knowing the periodic trends.

a. N and O
b. C and P
c. Si and Cl
d. Al and F
e. Al and I
f. P and S
g. C and N
h. B and Cl
i. Be and Br
j. Si and P

Q5.3.5

Place the following bonds in order from least to most polar. Refer to Table 5.3.1 when answering this question.

a. Fe-N
b. H-Cl
c. Ca-O
d. C-S

Q5.3.6

Place the following bonds in order from least to most polar. Use periodic trends to determine the correct order without looking at electronegativity values.

a. PCl
b. SCl
c. PBr
d. CBr

Q5.3.7

Label each of the molecules as nonpolar or polar covalent.

a. CO_2
b. CCl_4
c. NH_3

Q5.3.8

Describe the types of molecules that have the following types of intermolecular forces.

a. London dispersion forces
b. dipole-dipole forces
c. hydrogen bonding

Q5.3.9

Why are the intermolecular forces in H_2O and H_2S so different from one another?

Q5.3.10

What type(s) of intermolecular forces are present in each of the molecules in the question 5.3.7?

Q5.3.11

What is the relationship between the strength of intermolecular forces in a molecule and its boiling point?

Q5.3.12

Rank the following in order of increasing (smallest to greatest) boiling point.

a. N_2
b. CH_3OH
c. PH_3

5.4: Chromatography

(click here for solutions)

Q5.4.1

Define chromatography.

Q5.4.2

List two types of chromatography.

Q5.4.3

Distinguish between the stationary and mobile phases.

Answers

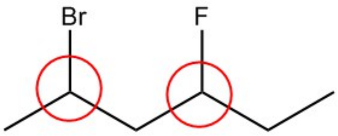
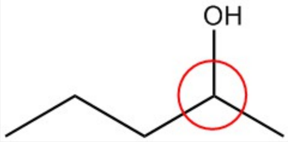
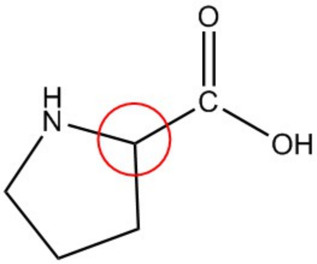
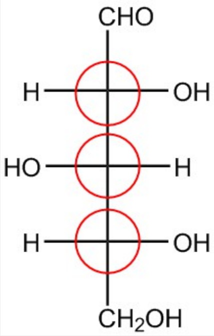
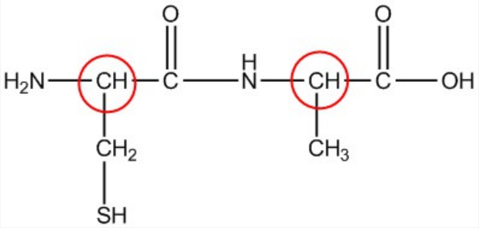
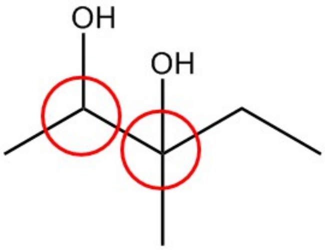
5.1: Isomers

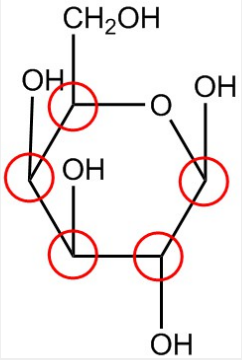
Q5.1.1

Isomers are molecules with the same chemical formula but a different structure or arrangement of atoms.

Q5.1.2

Circle the chiral carbons in each structure.

a.	
b.	
c.	No chiral carbons
d.	No chiral carbons
e.	
f.	
g.	No chiral carbons.
h.	
i.	

j.	

Q5.1.3

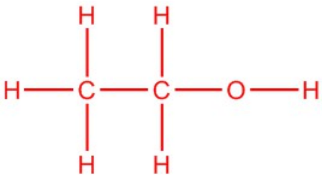
a. No, because of two H atoms on left side of double bond.
b. Yes, because the chlorine atoms are shown trans but could be drawn cis.
c. Yes, because the two groups could be on the same side of the ring or on opposite sides.
d. No, because you need a double bond or a ring to have a geometric isomer.
e. No, because there is only one non-hydrogen group on the ring.
f. Yes, because each carbon in the double bond has two different groups so it could be described as cis or trans.

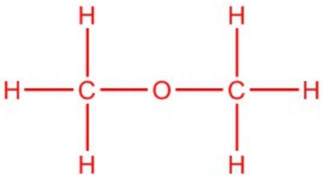
Q5.1.4

For each pair, determine if they are isomers. If they are isomers, identify the type.

a. structural
b. conformational
c. not isomers
d. structural
e. structural
f. not isomers
g. enantiomers
h. same molecule

Q5.1.5





5.2: Carbohydrate Structures

Q5.2.1

$C_4H_8O_4$

Q5.2.2

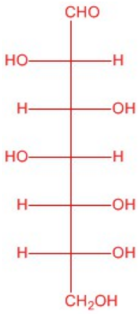
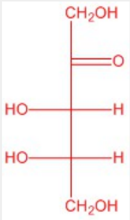
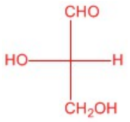
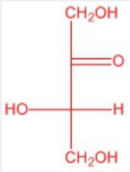
Monosaccharides are a single carbohydrate molecule. Disaccharides have two carbohydrate molecules bonded together. Polysaccharides have three or more carbohydrate molecules bonded together.

Q5.2.3

Simple carbohydrates are monosaccharides or disaccharides and are easily broken down in the body. Complex carbohydrates are polysaccharides such as starches and fiber and take longer to break down in the body.

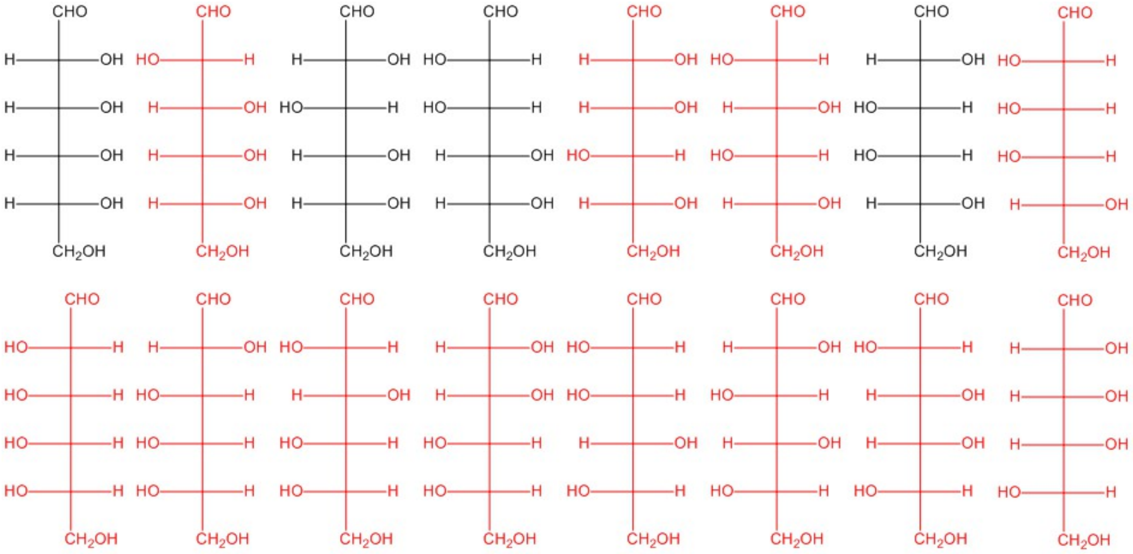
Q5.2.4

Answers will vary due to orientation of H and OH at each chiral carbon.

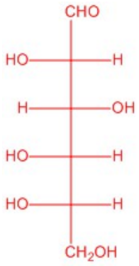
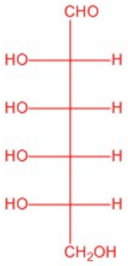
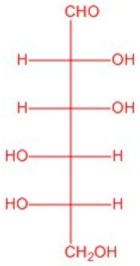
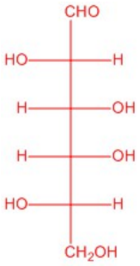
a.	
b.	
c.	
d.	

Q5.2.5

Any two of the structures shown here in red. The structures in black are from Figure 5.2.3.



Q5.2.6

			
L-glucose	L-allose	L-mannose	L-galactose

Q5.2.7

a. Epimers because the orientation differs at a single chiral carbon.
b. Diastereomers because the orientation differs at multiple (but not all) chiral carbons.
c. Diastereomers because the orientation differs at multiple (but not all) chiral carbons.
d. Diastereomers because the orientation differs at multiple (but not all) chiral carbons.

Q5.2.8

a. beta
b. alpha
c. beta
d. alpha
e. beta

Q5.2.9

a. O because it has the fewest types of carbohydrates.
b. AB because it has all carbohydrates found on red blood cells.
c. A can give to A or AB.
d. B can receive B or O.

5.3: Polarity and Intermolecular Forces

Q5.3.1

Electronegativity is an atom's attraction to an electron in a bond.

Q5.3.2

Electronegativity increases from left to right and from bottom to top with a maximum at fluorine.

Q5.3.3

Nonpolar covalent bonds have even sharing of electrons between atoms while polar electrons share electrons unevenly.

Q5.3.4

a. polar
b. nonpolar
c. polar
d. ionic
e. polar
f. nonpolar
g. polar
h. polar
i. polar
j. nonpolar

Q5.3.5

C-S < H-Cl < Fe-N < Ca-O

Calculate the difference in electronegativity for each bond. The smaller the difference, the more nonpolar; the greater the difference, the more polar. If a bond has a large enough difference, it is so polar that it is considered ionic.

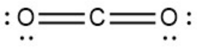
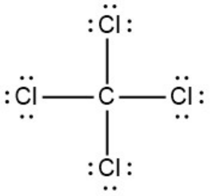
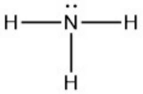
Q5.3.6

S-Cl < P-Cl < P-Br < C-Br

- S and Cl are closest together so they will have the smallest difference in electronegativity and be the least polar.
- S-Cl and P-Cl can be compared since they share a common element. P is farther left than S, so the electronegativity of P must be less than that of S. Therefore, the difference between P and Cl must be greater than the difference between S and Cl since we are comparing them both to the same element (Cl).
- Compare P-Cl and P-Br. Since Br is further down the periodic table, it has a lower electronegativity value than Cl. Therefore the difference between P and Cl is less than the difference between P and Br. P-Br is more polar than P-Cl.
- Compare P-Br and C-Br. C is further from Br than P is so the difference in electronegativity is greater for C-Br than P-Br making C-Br the more polar bond.

Q5.3.7

a. nonpolar
b. nonpolar
c. polar

Q5.3.8

a. all molecules
b. polar molecules
c. molecules with an H bonded to F, O, or N which is attracted to the F, O, or N on another molecule

Q5.3.9

H_2O has hydrogen bonding while H_2S does not.

Q5.3.10

a. dispersion
b. dispersion
c. dispersion, dipole-dipole, and hydrogen bonding

Q5.3.11

The stronger the intermolecular forces, the higher the boiling point.

Q5.3.12

N_2 < PH_3 < CH_3OH

All three of these molecules are comparable in size so the differences in dispersion forces are small. N_2 is nonpolar and has only dispersion forces. PH_3 is polar, so it has dispersion and dipole-dipole forces but no hydrogen bonding. CH_3OH has dispersion, dipole-dipole, and hydrogen bonding forces.

5.4: Chromatography

Q5.4.1

Chromatogaphy is used to separate a mixture into its components when they travel at different rates in the mobile phase.

Q5.4.2

Paper, thin-layer, liquid, and gas are the most well-known types of chromatography

Q5.4.3

The mobile phase is a fluid which moves through the stationary phase. The stationary phase holds the sample until the mobile phase moves it along the stationary phase.

CHAPTER OVERVIEW

6: Energy and Properties

6.1: Heat Flow

Learning Outcomes

- Describe how chemical potential energy relates to heat and work.
- Use the specific heat equation to perform calculations that relate mass, specific heat, change in temperature, and the amount of heat absorbed or released.
- Convert among joule, calorie, and Calorie.
- Define endothermic and exothermic.
- Explain the difference between heat and specific heat capacity.

Chemical reactions are accompanied by transfers of energy. Keeping track of heat flow and energy requirements is important for a full understanding of chemical processes. Energy is the capacity for doing work or supplying heat. When you fill your car with gasoline, you are providing it with potential energy. **Chemical potential energy** *is the energy stored in the chemical bonds of a substance*. The various chemicals in gasoline contain a large amount of chemical potential energy that is released when the gasoline is burned in a controlled way in the engine of the car. The release of that energy does two things. Some of the potential energy is transformed into work, which is used to move the car (see figure below). At the same time, some of the potential energy is converted to heat, making the car's engine very hot. The energy changes of a system occur as either heat or work, or some combination of both.

Figure 6.1.1: Combustion of gasoline produces energy which moves the car and generates heat.

Heat *is energy that is transferred from one object or substance to another because of a difference in temperature between them.* Heat always flows from an object at a higher temperature to an object at a lower temperature (see figure below). The flow of heat will continue until the two objects are at the same temperature.

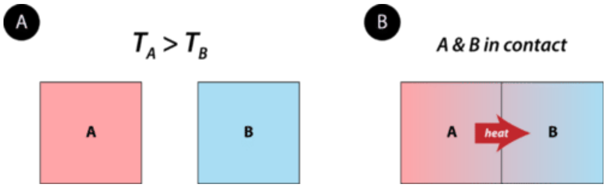


Figure 6.1.2: (A) Object A starts with a higher temperature than object B. No heat flows when the objects are isolated from each other. (B). When brought into contact, heat flows from A to B until the temperatures of the two objects are the same.

Thermochemistry *is the study of energy changes that occur during chemical reactions and during changes of state.* When chemical reactions occur, some chemical bonds are broken, while new chemical bonds form. As a result of the rearrangement of atoms, the total chemical potential energy of the system either increases or decreases.

Exothermic and Endothermic Processes

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

In the study of thermochemical processes, things are viewed from the point of view of the system. A chemical reaction or physical change is **endothermic** *if heat is absorbed by the system from the surroundings.* In the course of an endothermic process, the system gains heat from the surroundings, so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter q. The sign of q for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** *if heat is released by the system into the surroundings*. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of q for an exothermic process is negative because the system is losing heat. The difference between an endothermic reaction and an exothermic reaction is illustrated below (see figure below).

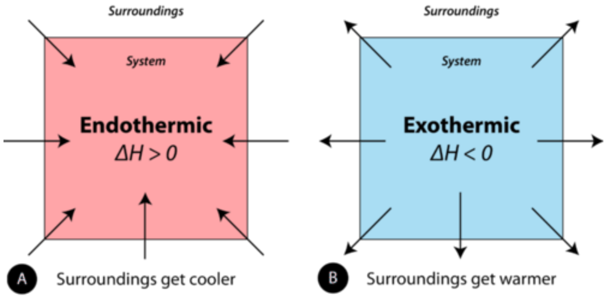


Figure 6.1.3: (A) In an endothermic reaction, heat flows from the surroundings into the system, decreasing the temperature of the surroundings. (B) In an exothermic reaction, heat flows from the system into the surroundings, increasing the temperature of the surroundings.

Units of Heat

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

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

$$1 \text{ kilocalorie} = 1 \text{ Calorie} = 1000 \text{ calories} \tag{6.1.1}$$

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

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

$$1\ \mathrm{J} = 0.2390\ \mathrm{cal} \text{ or } 1\ \mathrm{cal} = 4.184\ \mathrm{J} \tag{6.1.2}$$

Example 6.1.1

How many joules of energy are released when a 400. Calorie hamburger is digested?

Solution

Use the relationship between Calories and calories which is $1000 \text{ calories} = 1 \text{ Calorie}$ and the conversion factor for joules to calories to find the value in joules. Note that all of the units cancel except joules which is what is being asked for in the question.

$$400\ \mathrm{Cal}\left(\frac{1000\ \mathrm{cal}}{1\ \mathrm{Cal}}\right)\left(\frac{4.184\ \mathrm{J}}{1\ \mathrm{cal}}\right) = 1.67 \times 10^6\ \mathrm{J} \tag{6.1.3}$$

Heat Capacity and Specific Heat

If a swimming pool and a bucket, both full of water at the same temperature, were subjected to the same input of heat energy, the bucket of water would certainly rise in temperature more quickly than the swimming pool. **Heat capacity** *is the amount of heat required to raise the temperature of an object by* $1^\circ\mathrm{C}$. The heat capacity of an object depends both on its mass and its chemical composition. Because of its much larger mass, the swimming pool of water has a larger heat capacity than the bucket of water.

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water in the same sun will not become nearly as hot. Water is very resistant to changes in temperature, while metals in general are not. The **specific heat** *of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by* 1^oC. The table below lists the specific heat of some common substances. The symbol for specific heat is C_p, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree $(\text{J/g} \cdot^\text{o} \text{C})$, or calories per gram per degree $(\text{cal/g} \cdot^\text{o} \text{C})$. This text will use $\text{J/g} \cdot^\text{o} \text{C}$ for specific heat. Note that the specific heat of a substance depends not only on its identity but also its state. For example, ice, liquid water, and steam all have different specific heat values.

***Table 6.1.1**: Specific Heats of Some Common Substances*

Substance	**Specific Heat** $(\text{J/g} \cdot^\text{o} \text{C})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

Notice that water has a very high specific heat compared to most other substance. Water is commonly used as a coolant for machinery because it is able to absorb large quantities of heat (see figure below). Coastal climates are much more moderate than inland climates because of the presence of the ocean. Water in lakes or oceans absorbs heat from the air on hot days and releases it back into the air on cool days.

Figure 6.1.4: This power plant in West Virginia, like many others, is located next to a large lake so that the water from the lake can be used as a coolant. Cool water from the lake is pumped into the plant, while warmed water is pumped out of the plant and back into the lake.

Specific Heat Calculations

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (C_p), mass (m), and temperature change (ΔT) is shown below.

$$q = m \times C_p \times \Delta T \tag{6.1.4}$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Example 6.1.2

A 15.0 g piece of cadmium metal absorbs 134 J of heat as its temperature is increased from 24.0^oC to 62.7^oC. Calculate the specific heat of cadmium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Heat $= q = 134 \text{ J}$
- Mass $= m = 15.0 \text{ g}$
- $\Delta T = 62.7^\text{o}\text{C} - 24.0^\text{o}\text{C} = 38.7^\text{o}\text{C}$

Unknown

- C_p of cadmium $= ? \text{ J/g} \cdot {}^\text{o}\text{C}$

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$C_p = \frac{q}{m \times \Delta T} = \frac{134 \text{ J}}{15.0 \text{ g} \times 38.7^\text{o}\text{C}} = 0.231 \text{ J/g} \cdot {}^\text{o}\text{C} \tag{6.1.5}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals in the table above. The result has three significant figures.

Since most specific heats are known, they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that a 60.0 g sample of water at 23.52^oC was cooled by the removal fo 813 J of heat. The change in temperature can be calculated using the specific heat equation.

$$\Delta T = \frac{q}{C_p \times m} = \frac{-813 \text{ J}}{4.18 \text{ J/g} \cdot {}^\text{o}\text{C} \times 60.0 \text{ g}} = -3.24^\text{o}\text{C} \tag{6.1.6}$$

Since the water was being cooled, heat is removed from the system. Therefore, q is negative, and the temperature decreases. The final temperature is:

$$T_f = 23.52^\text{o}\text{C} - 3.24^\text{o}\text{C} = 20.28^\text{o}\text{C} \tag{6.1.7}$$

Supplemental Resources

- Energy in Chemical Reactions: http://www.wisc-online.com/Objects/V...spx?ID=GCH8705

Contributors and Attributions

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

6.2: Energy and Properties (Exercises)

These are homework exercises to accompany Chapter 6 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

6.1: Heat Flow

(click here for solutions)

Q6.1.1

Define potential energy and chemical potential energy.

Q6.1.2

What is one potential use for substances that have a large amount of chemical potential energy?

Q6.1.3

Describe what happens when two objects that have different temperatures come into contact with one another.

Q6.1.4

Distinguish between system and surroundings.

Q6.1.5

Distinguish between endothermic and exothermic.

Q6.1.6

Two different reactions are performed in two identical test tubes. In reaction A, the test tube becomes very warm as the reaction occurs. In reaction B, the test tube becomes cold. Which reaction is endothermic and which is exothermic? Explain.

Q6.1.7

What is the sign of q for an endothermic process? For an exothermic process?

Q6.1.8

Classify the following as endothermic or exothermic processes.

a. Boiling water
b. Sweating
c. Burning paper
d. Water freezing

Q6.1.9

Convert each value to the indicated units.

a. 150. kcal to Cal
b. 355 J to cal
c. 200. Cal to J
d. 225 kcal to cal
e. 3450. cal to kcal
f. 450. Cal to kJ
g. 175 kJ to cal

Q6.1.10

Equal amounts of heat are applied to 10.0 g samples of iron and aluminum, both originally at 25°C. Which one will be at the higher temperature?

Q6.1.11

Which sample will require more heat to increase the temperature by 10°C?

a. 25.0 g copper
b. 25.0 g lead

Q6.1.12

How much energy is required to heat 50.0 g of silver from 30°C to 50°C?

Q6.1.13

What is the final temperature when 125 J is applied to 20.0 g of lead, initially at 15°C?

Q6.1.14

How much energy is required to raise the temperature of 13.7 g of aluminum from 25.2°C to 61.9°C?

Q6.1.15

A 274 g sample of air is heated with 2250 J of heat and its temperature rises by 8.11°C. What is the specific heat of air at these conditions?

Q6.1.16

98.3 J of heat is supplied to 12.28 g of a substance, and its temperature rises by 5.42°C. What is the specific heat of the substance?

Q6.1.17

A quantity of ethanol is cooled from 47.9°C to 12.3°C and releases 3.12 kJ of heat. What is the mass of the ethanol sample?

Answers

6.1: Heat Flow

Q6.1.1

Potential energy is usually described as the energy of position. Chemical potential energy is energy stored within the chemical bonds of a substance.

Q6.1.2

Answers will vary. The most common example in every day life is the burning of fossil fuels to generate electricity or to run a vehicle.

Q6.1.3

The temperature of the hot object decreases and the temperature of the cold object increases as heat is transferred from the hot object to the cold object. The change in temperature of each depends on the identity and properties of each substance.

Q6.1.4

The system is the specific portion of matter being observed in an experiment and is designated by the experimenter. The surroundings is everything that is not the system.

Q6.1.5

Endothermic processes result in the gain of heat to the system while exothermic processes are associated with the loss of heat from the system.

Q6.1.6

Reaction A is exothermic because heat is leaving the system making the test tube feel hot. Reaction B is endothermic because heat is being absorbed by the system making the test tube feel cold.

Q6.1.7

q is positive for endothermic processes and q is negative for exothermic processes.

Q6.1.8

Classify the following as endothermic or exothermic processes.

a. Endothermic because heat is being added to the water to get it from the liquid state to the gas state.
b. Endothermic because energy is consumed to evaporate the moisture on your skin which lowers your temperature.
c. Exothermic because burning (also known as combustion) releases heat.
d. Exothermic because energy is exiting the system in order to go from liquid to solid. Another way to look at it is to consider the opposite process of melting. Energy is consumed (endothermic) to melt ice (solid to liquid) so the opposite process (liquid to solid) must be exothermic.

Q6.1.9

Convert each value to the indicated units.

a. $150\ kcal \left(\frac{1\ Cal}{1\ kcal}\right) = 150\ Cal$
b. $355\ J \left(\frac{1\ cal}{4.184\ J}\right) = 84.8\ cal$
c. $200.\ Cal \left(\frac{1000\ cal}{1\ Cal}\right) \left(\frac{4.184\ J}{1\ cal}\right) = 8.37 \times 10^5\ J$
d. $225\ kcal \left(\frac{1000\ cal}{1\ kcal}\right) = 2.25 \times 10^5\ cal$
e. $3450.\ cal \left(\frac{1 kcal\ kJ}{1000\ cal}\right) = 3.450\ kcal$
f. $450.\ Cal \left(\frac{1000\ cal}{1\ Cal}\right) \left(\frac{4.184\ J}{1\ cal}\right) \left(\frac{1\ kJ}{1000\ J}\right) = 1.88 \times 10^3\ kJ$ or $450.\ Cal \left(\frac{4.184\ kJ}{1\ Cal}\right) = 1.88 \times 10^3\ kJ$
g. $175\ kJ \left(\frac{1000\ J}{1\ kJ}\right) \left(\frac{1\ cal}{4.184\ J}\right) = 4.18 \times 10^4\ cal$

Q6.1.10

Iron has a specific heat capacity of $0.449\ \mathrm{J/g \cdot {}^{o}C}$ which means it takes 0.449 J of energy to raise 1 gram of iron by 1^{o}C. Aluminum has a specific heat capacity of $0.897\ \mathrm{J/g \cdot {}^{o}C}$ which means it takes 0.897 J of energy to raise 1 gram of aluminum by 1 $^{\text{o}}\text{C}$. When equal amounts of heat are applied, the temperature of the iron will increase more because it takes less energy (heat) to raise its temperature so iron will be at a higher temperature since they both start at 25^{o}C.

Q6.1.11

Both samples are the same mass so a comparison of the specific heat must be compared. Copper has a specific heat of 0.385 $\mathrm{J/g \cdot {}^{o}C}$ which means it takes 0.385 J of energy to raise 1 gram of copper by 1^\text{o} \text{C}\). Lead has a specific heat of 0.129 $\mathrm{J/g \cdot {}^{o}C}$ which means it takes 0.129 J of energy to raise 1 gram of copper by 1^\text{o} \text{C}\). More energy is needed to raise the temperature of copper so more heat will be needed to increase the temperature of copper by 10 ^\text{o} \text{C}\).

Q6.1.12

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ q &= 50.0\ g \cdot 0.233\ \frac{J}{g \cdot {}^{o}\mathrm{C}} \cdot (50^{o}\mathrm{C} - 30^{o}\mathrm{C}) \\ q &= 50.0\ g \cdot 0.233\ \frac{J}{g \cdot {}^{o}\mathrm{C}} \cdot (20^{o}\mathrm{C}) \\ q &= 233\ J \end{aligned}$$

Q6.1.13

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ 125\ J &= 20.0\ g \cdot 0.129\ \frac{J}{g \cdot {}^{o}\mathrm{C}} \cdot (T_f - 15^{o}\mathrm{C}) \\ 48.4^{o}\mathrm{C} &= T_f - 15^{o}\mathrm{C} \\ T_f &= 63^{o}\mathrm{C} \end{aligned}$$

Q6.1.14

$$\begin{aligned} q &= m \cdot C_p \cdot \Delta T \\ q &= 13.7\ g \cdot 0.897\ \frac{J}{g \cdot {}^{o}\mathrm{C}} \cdot (61.9^{o}\mathrm{C} - 25.2^{o}\mathrm{C}) \\ q &= 13.7\ g \cdot 0.897\ \frac{J}{g \cdot {}^{o}\mathrm{C}} \cdot (36.7^{o}\mathrm{C}) \\ q &= 451\ J \end{aligned}$$

Q6.1.15

$$
\begin{aligned}
q &= m \cdot C_p \cdot \Delta T \\
2250\ J &= 274\ g \cdot C_p\ \cdot 8.11^{\mathrm{o}}\mathrm{C} \\
C_p &= 1.01 \frac{J}{g \cdot {}^{\mathrm{o}}\mathrm{C}}
\end{aligned}
$$

Q6.1.16

$$
\begin{aligned}
q &= m \cdot C_p \cdot \Delta T \\
98.3\ J &= 12.28\ g \cdot C_p\ \cdot 5.42^{\mathrm{o}}\mathrm{C} \\
C_p &= 1.48 \frac{J}{g \cdot {}^{\mathrm{o}}\mathrm{C}}
\end{aligned}
$$

Q6.1.17

$$
\begin{aligned}
q &= m \cdot C_p \cdot \Delta T \\
-3.12\ kJ &= m \cdot 2.44 \frac{J}{g \cdot {}^{\mathrm{o}}\mathrm{C}} (12.3^{\mathrm{o}}\mathrm{C} - 47.9^{\mathrm{o}}\mathrm{C}) \\
-3.12 \times 10^3\ J &= m \cdot 2.44 \frac{J}{g \cdot {}^{\mathrm{o}}\mathrm{C}} (-35.6^{\mathrm{o}}\mathrm{C}) \\
m &= 35.9\ g
\end{aligned}
$$

CHAPTER OVERVIEW

7: Solids, Liquids, and Gases

University of Kentucky

Chemistry 103 – Chemistry for Allied Health

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7.1: States of Matter

Learning Outcomes

- List the three states of matter and give examples of each.
- Describe the properties of each state of matter.
- Identify and describe each type of change in state.
- Recognize that boiling and melting points vary with each substance.
- Recognize that a substance's boiling point depends on the pressure.

States of Matter

Another way that we can describe the properties of matter is the state (also called phase). The amount of energy in molecules of matter determines the **state of matter**. Matter can exist in one of several different states, including a gas, liquid, or solid state. These different states of matter have different properties, which are illustrated in the figure below.

- A **gas** is a state of matter in which atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide. Forces between atoms or molecules are not strong enough to hold them together.
- A **liquid** is a state of matter in which atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another. Forces between atoms or molecules are strong enough to keep the molecules relatively close together but not strong enough to prevent them from moving past one another.
- A **solid** is a state of matter in which atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another. Forces between atoms or molecules are strong enough to keep the molecules together and to prevent them from moving past one another.

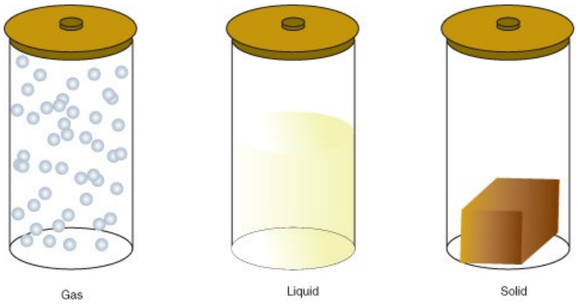

Figure $7.1.1$: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

What Determines a Substance's State?

Which state a substance is in depends partly on temperature and air pressure. For example, at the air pressure found at sea level, water exists as a liquid at temperatures between 0^oC and 100^oC. Above 100^oC, water exists as a gas (water vapor). Below 0^oC, water exists as a solid (ice). Different substances have a different range of temperatures at which they exist in each state. For example, oxygen is a gas above -183^oC, but iron is a gas only above 2861^oC. These differences explain why some substances are always solids at normal Earth temperatures, whereas others are always gases or liquids.

Water can take many forms. At low temperatures (below 0^oC), it is a solid. It is a liquid between 0^oC and 100^oC. While at temperatures above 100^oC, water is a gas (steam). The state the water is in depends upon the temperature. Each state (solid, liquid, and gas) has its own unique set of physical properties.

Table $7.1.1$: Characteristics of each Phase of Matter

	Gas	Liquid	Solid
Shape	no definite shape (takes the shape of its container)	no definite shape (takes the shape of its container)	definite shape (rigid)
Volume	particles move in random motion with little or no attraction to each other	has definite volume	definite volume

	Gas	**Liquid**	**Solid**
Mobility	particles move in random motion with little or no attraction to each other	particles are free to move over each other, but are still attracted to each other	particles vibrate around fixed axes

	Gas	Liquid	Solid
Compressibility	highly compressible	weakly compressible	weakly compressible

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

Changing States

Matter constantly goes through cycles that involve changing states (also called change phases). Water and all the elements important to organisms, including carbon and nitrogen, are constantly recycled on Earth. As matter moves through its cycles, it changes state repeatedly. For example, in the water cycle, water repeatedly changes from a gas to a liquid or solid and back to a gas again. How does this happen?

Adding energy to matter gives its atoms or molecules the ability to resist some of the forces holding them together. For example, heating ice to its **melting point** gives its molecules enough energy to move. The ice melts and becomes liquid water. Similarly, heating liquid water to its **boiling point** gives its molecules enough energy to pull apart from one another so they no longer have contact. The liquid water vaporizes and becomes water vapor.

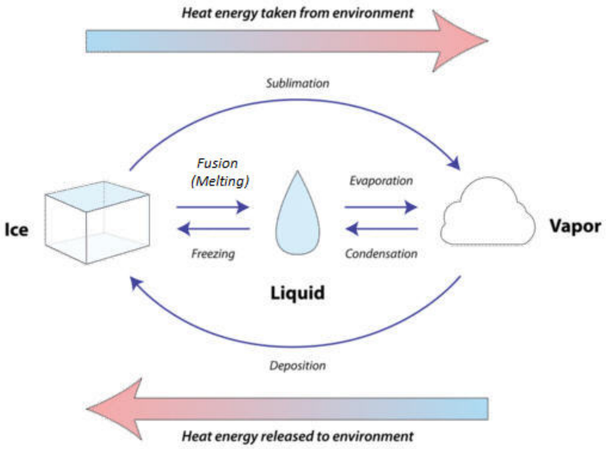

Figure 7.1.2: Matter can move between any two states (or phases) of matter depending on the pressure and temperature conditions. Although the figure shows the states for water, these names apply to all substances. *Use of the word fusion for melting is based on its Latin roots.*

The temperature of the melting and boiling points depend on the identity of the substance and the atmospheric pressure. Each substance has its own boiling and melting points that depend on the properties of the substance, including the strength of its intermolecular forces. As an example, the values for water are given in Table 7.1.1. Note how the boiling point of water varies greatly with pressure.

Table 7.1.1: Boiling point of water as a function of pressure

Altitude (ft)	**Pressure** (atm)	**Boiling Point** (°C)
-500	1.05	100.5
0	1.00	100
4000	0.892	96
7000	0.797	93

Contributors and Attributions

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7.2: State Changes and Energy

Learning Outcomes

- Identify whether energy is consumed or released during a phase change.
- Calculate the amount of energy consumed or released during a phase change.
- Recognize that the enthalpies of vaporization and condensation are equal in magnitude but opposite in sign.
- Recognize that the enthalpies of fusion and freezing are equal in magnitude but opposite in sign.

Heating Curve

When heat energy is supplied to a solid (like ice) at a steady rate by means of an electrical heating coil, we find that the temperature climbs steadily until the melting point is reached and the first signs of liquid formation become evident. Thereafter, even though we are still supplying heat energy to the system, the temperature remains constant as long as both liquid and solid are present. On the graph below, this is represented by the flat line, where energy is being added to the ice, but no change is occurring in the temperature. All energy added to the system at this stage is used to convert solid ice into liquid water.

Once all of the sample is in the liquid phase, the addition of energy now increases the temperature until the boiling point is reached and the first signs of gas formation are seen. The temperature remains constant even though energy is being added to the system. The energy is being used to convert the liquid to a gas. Once all of the sample is in the gas phase, additional energy can be added to increase the temperature of the gas.

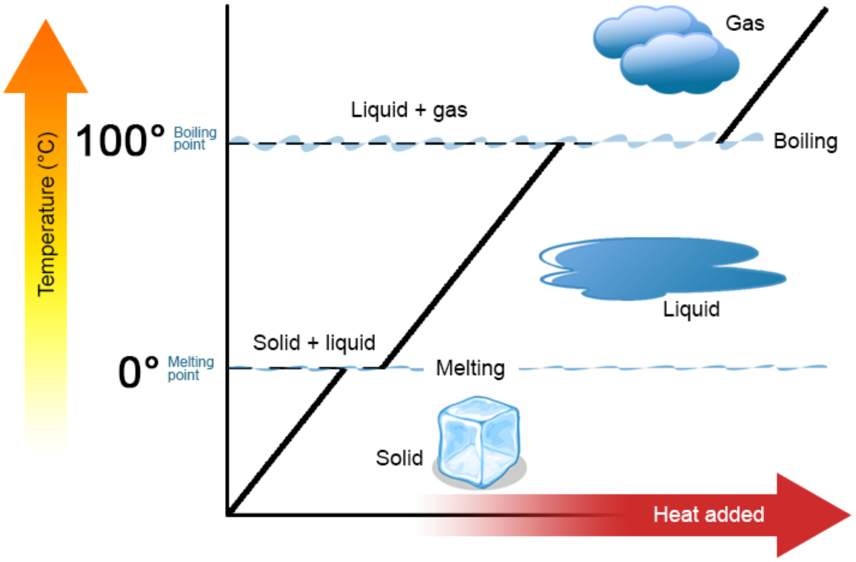

Figure 7.2.1: Heating curve of water at 1 atm. By Community College Consortium for Bioscience Credentials (Own work) [CC BY 3.0 (http://creativecommons.org/licenses/by/3.0)], via Wikimedia Commons

This macroscopic behavior demonstrates quite clearly that energy must be supplied to a solid in order to melt or vaporize it. On a microscopic level melting or vaporization involves separating molecules which are attracted to each other. The amount of energy needed to separate the molecules is proportional to the intermolecular forces between the molecules.

Enthalpy

The heat energy which a solid absorbs when it melts is called the **enthalpy of fusion (**or heat of fusion and is usually quoted on a molar basis. (The word *fusion* means the same thing as "melting.") When 1 mol of ice, for example, is melted, we find from experiment that 6.01 kJ of energy is needed. The molar enthalpy of fusion of ice is thus +6.01 kJ per mol ($\frac{6.01\ kJ}{mol}$), and we can write

$$\mathrm{H_2O}(s) \rightarrow \mathrm{H_2O}(l) \tag{7.2.1}$$

$$\triangle H_{fus} = \frac{6.01\ kJ}{mol} \tag{7.2.2}$$

If two moles of water are melted, then it would require twice as much energy (see below). The amount of energy needed to melt a substance will depend on the amount of the substance.

$$2\ mol \times \frac{6.01\ kJ}{mol} = \frac{12.0\ kJ}{mol}$$

Selected molar enthalpies of fusion are tabulated in Table 7.2.1. Solids like ice which have strong intermolecular forces have much higher values than those like CH_4 with weak ones.

When a liquid is boiled (or vaporized), energy is required to move the molecules apart to go from the liquid phase to the gas phase. The energy which a liquid absorbs when it vaporizes is known as the **enthalpy of vaporization (**. In the case of water, the molar enthalpy of vaporization is ($\frac{40.67\ kJ}{mol}$). In other words

$$H_2O(s) \rightarrow H_2O(l) \tag{7.2.3}$$

$$\triangle H_{vap} = 40.67 \frac{\text{kJ}}{\text{mol}} \tag{7.2.4}$$

Heat energy is absorbed to vaporize a liquid because molecules which are held together by intermolecular forces in the liquid are separated as the gas is formed. Such a separation requires energy. As with melting, the amount of energy needed to vaporize a substance is proportional to the amount of substance present. The more liquid there is, the more energy required to vaporize it.

In general the energy needed differs from one liquid to another depending on the magnitude of the intermolecular forces. We can expect liquids with strong intermolecular forces to have larger enthalpies of vaporization. See the examples in the table below. Note that the enthalpy of vaporization of a substance is always higher than its enthalpy of fusion. When a solid melts, the molecules are not separated from each other to nearly the same extent as when a liquid boils.

Melting and boiling points are given in units of Kelvin ($K = T_{\circ C} + 273.15$)

Table 7.2.1: Molar Enthalpies of Fusion and Vaporization of Selected Substances.

Substance	Formula	$\triangle H_{fus}$ $(\frac{kJ}{mol})$	Melting Point (K)	$\triangle H_{vap}$ $(\frac{kJ}{mol})$	Boiling Point (K)
neon	Ne	0.33	24	1.80	27
oxygen	O_2	0.44	54	6.82	90.2
methane	CH_4	0.94	90.7	8.18	112
ethane	C_2H_6	2.85	90.0	14.72	184
chlorine	Cl_2	6.40	172.2	20.41	239
carbon tetrachloride	CCl_4	2.67	250.0	30.00	350
water*	H_2O	6.01	273.1	40.7	373.1
n-nonane	C_9H_{20}	19.3	353	40.5	491
lead	Pb	4.77	601	178	2022
ammonia	NH_3	5.636	195	23.35	240
ethanol	CH_3COOH	4.9	159	38.56	351

*http://www1.lsbu.ac.uk/water/data.html

Example 7.2.1: Heat Energy

Calculate the heat energy required to (a) vaporize 100. g of lead, (b) melt 100. g of lead, (c) vaporize 100. g water, and (d) melt 100. g of water.

Solution

(a)To vaporize 100. g of lead:

$$\mathrm{Pb}(l) \rightarrow \mathrm{Pb}(g) \qquad \triangle H_{vap} = 178 \frac{kJ}{mol} \tag{7.2.5}$$

$$100.\ g \times \frac{1\ mol\ \mathrm{Pb}}{207.2\ g\ \mathrm{Pb}} \times \frac{178\ kJ}{mol} = 85.9\ kJ \tag{7.2.6}$$

(b) To melt 100. g of lead:

$$\mathrm{Pb}(s) \rightarrow \mathrm{Pb}(l) \qquad \triangle H_{fus} = 4.77 \frac{kJ}{mol} \tag{7.2.7}$$

$$100.\ g \times \frac{1\ mol\ \mathrm{Pb}}{207.2\ g\ \mathrm{Pb}} \times \frac{4.77\ kJ}{mol} = 2.30\ kJ \tag{7.2.8}$$

(c) To vaporize 100. g of water:

$$\mathrm{H_2O}(s) \rightarrow \mathrm{H_2O}(l) \qquad \triangle H_{vap} = 40.657 \frac{kJ}{mol} \tag{7.2.9}$$

$$100.\ g \times \frac{1\ mol\ \mathrm{H_2O}}{18.0\ g\ \mathrm{H_2O}} \times \frac{40.657\ kJ}{mol} = 226\ kJ \tag{7.2.10}$$

(d) To melt 100. g of water:

$$\mathrm{H_2O}(s) \rightarrow \mathrm{H_2O}(l) \qquad \triangle H_{fus} = 6.01 \frac{kJ}{mol} \tag{7.2.11}$$

$$100.\ g \times \frac{1\ mol\ \mathrm{H_2O}}{18.0\ g\ \mathrm{H_2O}} \times \frac{6.01\ kJ}{mol} = 33.4\ kJ \tag{7.2.12}$$

It might be surprising that the heat required to melt or vaporize 100 g of lead is so much less than that require to melt or vaporized water. First, the temperature at which the substance melts has nothing to do with the enthalpy of fusion. Remember, we are only looking at the energy required to change the phase, not the energy required to get the substance to the melting or boiling point.

Also note that the enthalpies of fusion and vaporization are given as kJ per mole. Although the water and the lead have the same mass, the moles of each substance is very different (5.5 moles of water vs. 0.48 moles of lead).

Freezing and Condensation

The discussion here has focused on fusion (melting) and vaporization. But what about freezing and condensation? Fusion (solid to liquid) and freezing (liquid to solid) are opposite processes. As a result, the magnitude of energy for each is the same, but the sign is different. With fusion, energy is a positive value because it is endothermic (consuming energy) while the energy for freezing is negative because it is an exothermic process (releasing energy).

Let's compare the enthalpies of fusion and freezing for methane. Both values have the same magnitude of 0.94 but they have the opposite sign.

$$\triangle H_{fusion} = 0.94\ \frac{kJ}{mol} \qquad \triangle H_{freezing} = -0.94\ \frac{kJ}{mol}$$

The same concept applies to vaporization (liquid to gas) and condensation (gas to liquid). Energy is consumed during vaporization (positive energy) and released during condensation (negative energy). The enthalpies of vaporization and condensation are equal in magnitude but opposite in sign.

$$\triangle H_{vaporization} = 8.18\ \frac{kJ}{mol} \qquad \triangle H_{condensation} = -8.18\ \frac{kJ}{mol}$$

Energy is consumed to change a substance from solid to liquid to gas. Energy is released to change a substance from gas to liquid to solid.

Contributors and Attributions

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7.3: Kinetic-Molecular Theory

Learning Outcomes

- Define pressure and describe how gases exert pressure.
- Convert between units of gas pressure (atm, mmHg, torr, and Pa).
- Relate temperature to average kinetic energy.
- Relate mass to molecular speed.
- Describe the behavior of an ideal gas.

The **kinetic-molecular theory** *is a theory that explains the states of matter and is based on the idea that matter is composed of tiny particles that are always in motion*. The theory helps explain observable properties and behaviors of solids, liquids, and gases. However, the theory is most easily understood as it applies to gases. The theory applies specifically to a model of a gas called an ideal gas. An **ideal gas** *is an imaginary gas whose behavior perfectly fits all the assumptions of the kinetic-molecular theory*. In reality, gases are not ideal, but they are very close to being so under most everyday conditions.

The kinetic-molecular theory, as it applies to gases, has five basic assumptions.

1. **Gases consist of very large numbers of tiny spherical particles that are far apart from one another compared to their size**. The particles of a gas may be either atoms or molecules. The distance between the particles of a gas is much, much greater than the distances between the particles of a liquid or a solid. Most of the volume of a gas, therefore, is composed of the empty space between the particles. In fact, the volume of the particles themselves is considered to be insignificant compared to the volume of the empty space.
2. **Gas particles are in constant rapid motion in random directions**. The fast motion of gas particles gives them a relatively large amount of kinetic energy. Recall that kinetic energy is the energy that an object possesses because of its motion. The particles of a gas move in a straight line until they collide with another particle or with one of the walls of their container (see figure below).
3. **Collisions between gas particles and between particles and the container walls are elastic collisions**. An elastic collision is one in which there is no overall loss of kinetic energy. Kinetic energy may be transferred from one particle to another during an elastic collision, but there is no change in the total energy of the colliding particles.
4. **There are no forces of attraction or repulsion between gas particles**. Attractive forces are responsible for particles of a real gas condensing together to form a liquid. It is assumed that the particles of an ideal gas have no such attractive forces. The motion of each particle is completely independent of the motion of all other particles.
5. **The average kinetic energy of gas particles is dependent upon the temperature of the gas**. As the temperature of a gas is increased, its component particles begin to move faster, resulting in an increase in their kinetic energies. Not all particles in a given sample have the same speed, so the sample will contain particles with a range of different kinetic energies. However, the average kinetic energy of the particles in a sample is proportional to its temperature.

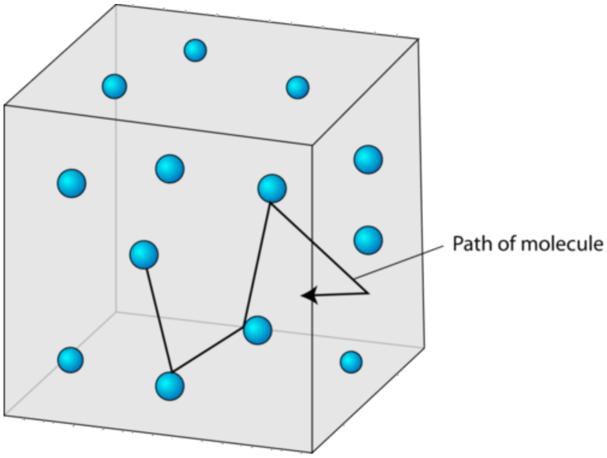

Figure 7.3.1: Gas particles move in a random, linear fashion according to the kinetic-molecular theory. The space between particles is very large compared to the size of the particles.

Gas Pressure

Pressure *is defined as the force per unit area on a surface.*

$$\text{Pressure} = \frac{\text{force}}{\text{area}} \tag{7.3.1}$$

When a person stands on the floor, his feet exert pressure on the surface. That pressure is related to both the mass of the person and the surface area of his feet. If the person were holding a heavy object, the pressure would increase because of a greater force. Alternatively, if the person stands on his toes, the pressure also increases because of a decrease in the surface area.

Gas molecules also exert pressure. Earth's atmosphere exerts pressure because gravity acts on the huge number of gas particles contained in the atmosphere, holding it in place. Pressure is also exerted by small samples of gas, such as the outward pressure exerted by the gas inside a balloon. **Gas pressure** *is the pressure that results from collisions of gas particles with an object*. Inside the balloon, the gas particles collide with the balloon's inner walls. It is those collisions that keep the balloon inflated. If the gas particles were to suddenly stop moving, the balloon would instantly deflate. The figure below is an illustration of gas particles exerting pressure inside a container.

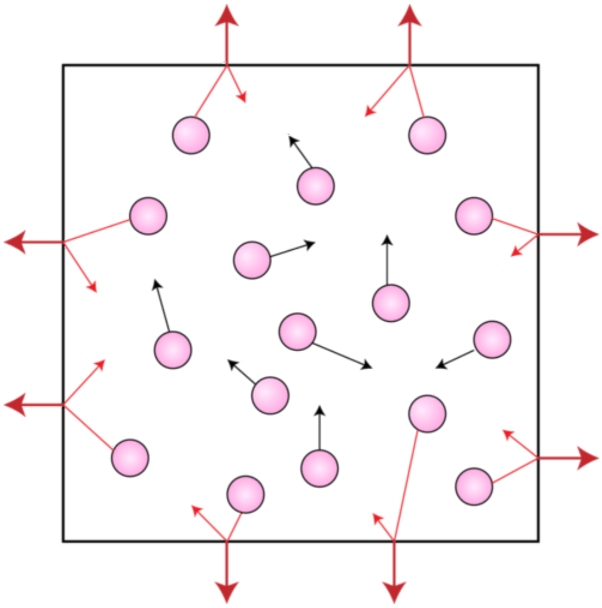

Figure $7.3.2$: Gas pressure results from collisions with the surface of its container.

Measuring Pressure

Atmospheric pressure *is the pressure exerted by the gas particles in Earth's atmosphere as those particles collide with objects*. A **barometer** *is an instrument used to measure atmospheric pressure*. A traditional mercury barometer consists of an evacuated tube immersed in a container of mercury. Air molecules from the atmosphere push down on the outer surface of the mercury, but, because the inside of the tube is a vacuum, there is no corresponding downward push on the mercury in the tube. As a result, the mercury rises inside the tube. The height to which the mercury rises is dependent on the external air pressure.

At sea level, a mercury column will rise a distance of 760 mm. This atmospheric pressure is reported as 760 mm Hg (millimeters of mercury). At higher altitudes, the atmospheric pressure is lower, so the column of mercury will not rise as high. For example, on the summit of Mt. Everest (at an elevation of 8848 m), the air pressure is 253 mm Hg. Atmospheric pressure is also slightly dependent on weather conditions.

A more convenient barometer, called an aneroid barometer, measures pressure by the expansion and contraction of a small spring within an evacuated metal capsule.

Units of Gas Pressure

A barometers measures gas pressure by the height of the column of mercury. One unit of gas pressure is the millimeter of mercury (mm Hg). An equivalent unit to the mm Hg is called the torr, in honor of the inventor of the barometer, Evangelista Torricelli. The **pascal** (Pa) *is the standard unit of pressure*. A pascal is a very small amount of pressure, so a more useful unit for everyday gas pressures is the kilopascal (kPa). A kilopascal is equal to 1000 pascals. Another commonly used unit of pressure is the atmosphere (atm). Standard atmospheric pressure is called $1 \mathrm{~atm}$ of pressure and is equal to $760 \mathrm{~mm} \mathrm{Hg}$ and $101.3 \mathrm{kPa}$ The relationships between the most common units of pressure are shown below.

$$1 \mathrm{~atm}=760 \mathrm{~mm} \mathrm{Hg}=760 \text { torr }=101.3 \mathrm{kPa} \tag{7.3.2}$$

Example 7.3.1

The atmospheric pressure in a mountainous location is measured to be $613 \mathrm{~mm} \mathrm{Hg}$. What is this pressure in atm and in kPa?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Given: $613 \mathrm{~mm} \mathrm{Hg}$
- $1 \mathrm{~atm}=760 \mathrm{~mm} \mathrm{Hg}$
- $101.3 \mathrm{kPa}=760 \mathrm{~mm} \mathrm{Hg}$

Unknown

- Pressure $=? \mathrm{~atm}$
- Pressure $=? \mathrm{kPa}$

Use conversion factors from the equivalent pressure units to convert from $\mathrm{mm} \mathrm{Hg}$ to atm and from $\mathrm{mm} \mathrm{Hg}$ to kPa.

Step 2: Solve.

$$613 \mathrm{~mm} \mathrm{Hg} \times \frac{1 \mathrm{~atm}}{760 \mathrm{~mm} \mathrm{Hg}}=0.807 \mathrm{~atm}$$

$$613 \mathrm{~mm} \mathrm{Hg} \times \frac{101.3 \mathrm{kPa}}{760 \mathrm{~mm} \mathrm{Hg}}=81.7 \mathrm{kPa}$$

Step 3: Think about your result.

The air pressure is about 80% of the standard atmospheric pressure at sea level. The standard pressure of $760 \mathrm{~mm} \mathrm{Hg}$ can be considered to have three significant figures.

Kinetic Energy and Temperature

As stated in the kinetic-molecular theory, the temperature of a substance is related to the average kinetic energy of the particles of that substance. When a substance is heated, some of the absorbed energy is stored within the particles, while some of the energy increases the speeds at which the particles are moving. This is observed as an increase in the temperature of the substance.

Average Kinetic Energy

At any given temperature, not all of the particles in a sample of matter have the same kinetic energy. Instead, the particles display a wide range of kinetic energies. Most of the particles have a kinetic energy near the middle of the range. However, some of the particles have kinetic energies a great deal lower or a great deal higher than the average (see figure below).

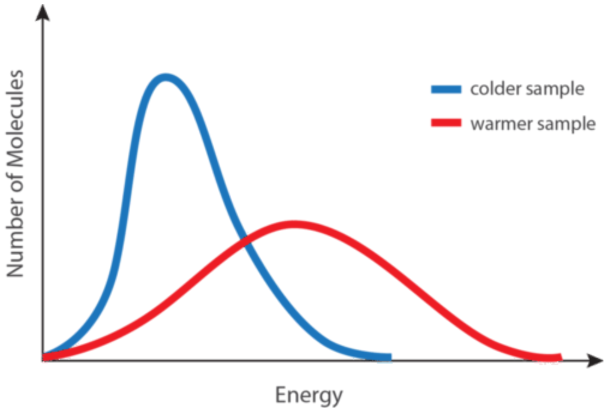

Figure 7.3.3: A distribution of molecular kinetic energies as a function of temperature. The blue curve is for a low temperature, while the red curve is for a high temperature.

The blue curve shown (see figure above) is for a sample of matter at a relatively low temperature, while the red curve is for a sample at a relatively high temperature. In both cases, most of the particles have intermediate kinetic energies, close to the average. Notice that as temperature increases, the range of kinetic energies increases and the distribution curve "flattens out".

At a given temperature, the particles of any substance have the same average kinetic energy. At room temperature, the molecules in a sample of liquid water have the same average kinetic energy as the molecules in a sample of oxygen gas or the ions in a sample of sodium chloride.

Absolute Zero

As a sample of matter is continually cooled, the average kinetic energy of its particles decreases. Eventually, one would expect the particles to stop moving completely. **Absolute zero** *is the temperature at which the motion of particles theoretically ceases.* Absolute zero has never been attained in the laboratory, but temperatures on the order of $1 \times 10^{-10}\ \text{K}$ have been achieved. The Kelvin temperature scale is based on this theoretical limit, so absolute zero is equal to $0\ \text{K}$. The Kelvin temperature of a substance is directly proportional to the average kinetic energy of the particles of the substance. For example, the particles in a sample of hydrogen gas at $200\ \text{K}$ have twice the average kinetic energy as the particles in a hydrogen sample at $100\ \text{K}$.

Contributors and Attributions

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7.4: The Ideal Gas Equation

Learning Outcomes

- State the ideal gas law and identify the variables.
- Use the ideal gas law to solve for an unknown.
- State the combined gas law.
- Simplify the combined gas law for any values that are constant.
- Use the combined gas law to solve for an unknown value.

Individual Gas Laws

Properties of gases such as pressure (P), volume (V), temperature (T), and moles(n) are relatively easy to measure. Unlike with liquids and solids, the particles (molecules or atoms) in a gas phase sample are very far apart from one another. As a result, their behavior is much more predictable because intermolecular forces become insignificant for most samples in the gas phase even over a wide range of conditions. The presence of intermolecular forces in liquid and solid samples makes their behavior harder to predict.

Experiments with gas phase samples over time showed the relationship between pairs of variables (P, V, T, and n) and individual gas laws (equations) show the quantitative relationship between those variables. Avogadro's law tells us that at constant P and T, the volume of a gas is directly proportional to the amount of gas. Boyle's law says that volume is inversely proportional to pressure at constant T and n. Charles' indicates that volume is directly proportional to temperature at constant P and n.

The video below shows a situation where 3 variables, pressure, volume, and amount of substance (moles) are all interrelated: inside our lungs.

As you can see in the video, when the pink balloon on the bottom (the "diaphragm") is pulled down, the balloon inside expands. This expansion causes a decrease in pressure (Boyle's Law). The pressure decrease causes a pressure differential, drawing air in through the straw, an increase in the amount of air (moles). So in your lungs, volume, pressure, and amount of air are all related. But none of the current laws explain the relation between 2 variables. How can this be resolved?

The Ideal Gas Law

These three laws may all be applied at once if we write ($\propto$ means "proportional to"):

$$V \propto n \times \frac{1}{P} \times T \tag{7.4.1}$$

or, introducing a constant of proportionality R,

$$V = R\,\frac{nT}{P} \tag{7.4.2}$$

This is known as the ideal gas law which results from the combination of the individual gas laws. Equation 7.4.2 applies to all gases at low pressures and high temperatures and is a very good approximation under nearly all conditions. The value of R, the **gas constant**, is independent of the kind of gas, the temperature, or the pressure and has a value of $\frac{0.08206\ L \cdot atm}{mol \cdot K}$.

Equation 7.4.2 is usually rearranged by multiplying both sides by P, so that it reads

$$PV = nRT \tag{7.4.3}$$

This is called the ideal gas equation or the ideal gas law. With the ideal gas equation we can convert from volume of a gas to amount of substance (provided that P and T are known). This is very useful since the volume, pressure, and temperature of a gas are easier to measure than mass, and because knowledge of the molar mass is unnecessary.

Note that for any gas law calculations, the temperature must be in units of Kelvin. The relationship between ^{o}C and K is K = °C + 273.15.

Example 7.4.1 : MOLES of Gas

Calculate the moles of gas in a 0.100 L sample at a temperature of 300 K and a pressure of 0.987 atm.

Solution

$$PV = nRT \tag{7.4.4}$$

$$(0.987\ atm)(0.100\ L) = n \left(\frac{0.08206\ L \cdot atm}{mol \cdot K} \right) (300\ K) \tag{7.4.5}$$

$$n = 0.00401\ mol \tag{7.4.6}$$

Example 7.4.2 : unit considerations

A sample of benzene (C_6H_6) was heated to 100.°C in an evacuated flask whose volume was 247.2 ml, a sample of benzene vaporized. When the benzene was condensed to a liquid, its mass was found to be 0.616 g. What was the pressure in the flask?

Solution

The problem gives values for temperature, volume, and mass of the sample. Since R has units of \(\left (\frac{L\cdot atm}{mol\cdot K} \right)), we need to have the temperature in units of Kelvin, the volume in liters, and the amount of sample in moles.

Temperature:

K = °C + 273.15

K = 100.°C + 273.15

K = 373 K

Volume: $247.2\ mL \left(\frac{10^{-3}\ L}{1\ mL} \right) = 0.2472\ L$

Moles: $0.616\ g \left(\frac{1\ mol}{78.11\ g} \right) = 7.89 \times 10^{-3}\ mol$

Now that all of the values are in the correct units, the value for the unknown pressure can be determined.

$$PV = nRT \tag{7.4.7}$$

$$P(0.2472\ L) = (7.89 \times 10^{-3}\ mol) \left(\frac{0.08206\ L \cdot atm}{mol \cdot K} \right) (373\ K) \tag{7.4.8}$$

$$P = .977 atm \tag{7.4.9}$$

Combined Gas law

While the ideal gas law is useful in solving for a single unknown when the other values are known, the combined gas law is useful when comparing initial and final situations. The ideal gas law can be rearranged to solve for R, the gas constant.

$$R = \frac{PV}{nT}$$

Under the initial conditions, $R = \frac{P_iV_i}{n_iT_i}$ and under final conditions, $R = \frac{P_fV_f}{n_fT_f}$. Since both expressions are equal to R, they are equal to each other.

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$

This equation is typically used when one or more of the variables is constant. As a result, that variable is canceled from the equation. For example, the equation $2x^2 = 2y$ can be simplified to $x^2 = y$ since the 2 is on both sides of the equation.

What happens to the combined gas law equation when the initial and final pressures are equal ($P_i = P_f$)? Since they are equal, P_i can replace P_f.

$$\frac{P_iV_i}{n_iT_i} = \frac{P_iV_f}{n_fT_f}$$

which simplifies to

$$\frac{V_i}{n_iT_i} = \frac{V_f}{n_fT_f}$$

If two variables are constant, the equation can be simplified even more. If temperature and volume are constant, then $T_i = T_f$ and $V_i = V_f$. Then,

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_i}{n_fT_i}$$

simplifies to

$$\frac{P_i}{n_i} = \frac{P_f}{n_f}$$

Example 7.4.3

Imagine a 1855 L balloon initially at 30°C and 745 mmHg. The balloon rises to an altitude of 23,000 ft and that the pressure and temperature at that altitude were 312 mmHg and −30°C, respectively. To what volume would the balloon have to expand to hold the same amount of hydrogen gas at the higher altitude?

Solution:

Begin by setting up a table of the two sets of conditions (note that some values will need to be converted to different units):

Initial	Final
$P_i = 745 \text{ mmHg} = 0.980 \text{ atm}$	$P_f = 312 \text{ mmHg} = 0.411 \text{ atm}$
$T_i = 30\ ^\circ\text{C} = 303 \text{ K}$	$T_f = -30\ ^\circ\text{C} = 243 \text{ K}$
$V_i = 1855 \text{ L}$	$V_f = ?$

By eliminating the constant property (n) of the gas, the combined gas law is simplified to

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f} \tag{7.4.10}$$

By solving the equation for V_f, we get:

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f} \tag{7.4.11}$$

$$\frac{0.980\ atm \cdot 1855\ L}{303\ K} = \frac{0.411\ atm \cdot V_f}{243\ K} \tag{7.4.12}$$

$$V_f = 3.55 \times 10^3\ L \tag{7.4.13}$$

Contributors and Attributions

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-

7.5: Aqueous Solutions

Learning Outcomes

- Define a solution and describe the parts of a solution.
- Describe how an aqueous solution is formed from both ionic compounds and molecular compounds.
- Recognize that some compounds are insoluble in water.
- Describe the differences among strong electrolytes, weak electrolytes, and nonelectrolytes.

Forming a Solution

When one substance dissolves into another, a solution is formed. A **solution** *is a homogenous mixture consisting of a solute dissolved into a solvent.* The **solute** *is the substance that is being dissolved,* while the **solvent** *is the dissolving medium*. Solutions can be formed with many different types and forms of solutes and solvents. In this chapter, we will focus on solution where the solvent is water. An **aqueous solution** *is water that contains one or more dissolved substance.* The dissolved substances in an aqueous solution may be solids, gases, or other liquids.

In order to be a true solution, a mixture must be stable. When sugar is fully dissolved into water, it can stand for an indefinite amount of time, and the sugar will not settle out of the solution. Further, if the sugar-water solution is passed through a filter, it will remain with the water. This is because the dissolved particles in a solution are very small, usually less than $1\ \mathrm{nm}$ in diameter. Solute particles can be atoms, ions, or molecules, depending on the type of substance that has been dissolved.

Figure $7.5.1$: When a colored solution is passed through a filter, the entire solution, both solute and solvent, pass through unchanged.

The Dissolving Process

Water typically dissolves most ionic compounds and polar molecules. Nonpolar molecules, such as those found in grease or oil, do not dissolve in water. We will first examine the process that occurs when an ionic compound, such as table salt (sodium chloride), dissolves in water.

Water molecules move about continuously due to their kinetic energy. When a crystal of sodium chloride is placed into water, the water's molecules collide with the crystal lattice. Recall that the crystal lattice is composed of alternating positive and negative ions. Water is attracted to the sodium chloride crystal because water is polar; it has both a positive and a negative end. The positively charged sodium ions in the crystal attract the oxygen end of the water molecules because they are partially negative. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules because they are partially positive. The action of the polar water molecules takes the crystal lattice apart (see figure below).

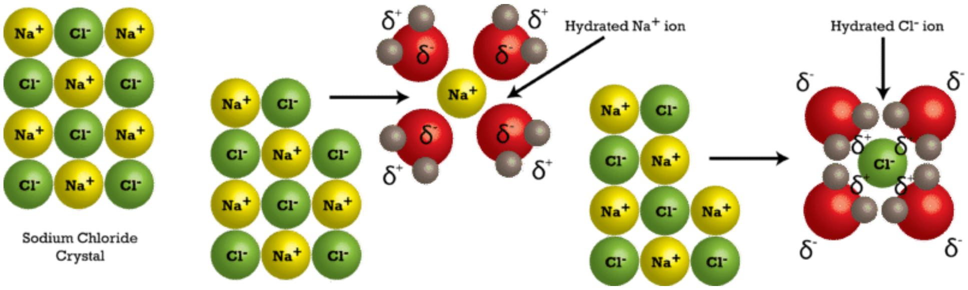


Figure $7.5.2$: The process of an ionic sodium chloride crystal being dissolved and hydrated by water.

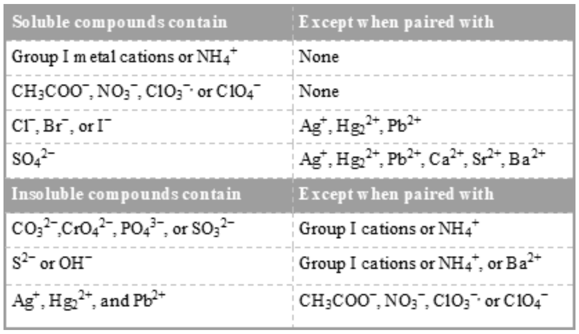
After coming apart from the crystal, *the individual ions are then surrounded by solvent particles in a process called* **solvation**. Note in the figure above that the individual $\mathrm{Na^+}$ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. **Hydration** *is the process of solute particles being surrounded by water molecules arranged in a specific manner*. Hydration helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and forming a precipitate.

Table sugar is made of the molecular compound sucrose $\left(\mathrm{C_{12}H_{22}O_{11}}\right)$. Solid sugar consists of individual sugar molecules held together by intermolecular attractive forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the attractive forces, but it *does not break the covalent bonds* between the carbon, hydrogen, and oxygen atoms. Dissolved sugar molecules are also hydrated. The hydration shell around a molecule of sucrose is arranged so that its partially negative oxygen atoms are near the partially positive hydrogen atoms in the solvent, and vice versa.

Insoluble Compounds

Not all compounds dissolve well in water. Some ionic compounds, such as calcium carbonate $\left(\mathrm{CaCO_3}\right)$ and silver chloride $\left(\mathrm{AgCl}\right)$, are nearly insoluble. This is because the attractions between the ions in the crystal lattice are stronger than the attraction that the water molecules have for the ions. As a result, the crystal remains intact. The solubility of ionic compounds can be predicted using the solubility rules as shown in Table $7.5.1$.

Table* $7.5.1$*: *Solubility rules for ionic compounds in water.*



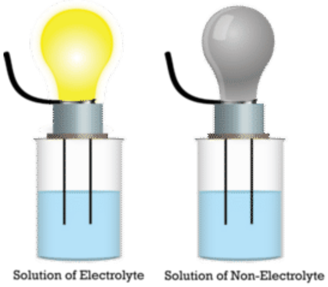
Soluble compounds contain	Except when paired with
Group I metal cations or NH_4^+	None
CH_3COO^-, NO_3^-, ClO_3^-· or ClO_4^-	None
Cl^-, Br^-, or I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
SO_4^{2-}	Ag^+, Hg_2^{2+}, Pb^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}
Insoluble compounds contain	**Except when paired with**
CO_3^{2-},CrO_4^{2-}, PO_4^{3-}, or SO_3^{2-}	Group I cations or NH_4^+
S^{2-} or OH^-	Group I cations or NH_4^+, or Ba^{2+}
Ag^+, Hg_2^{2+}, and Pb^{2+}	CH_3COO^-, NO_3^-, ClO_3^-· or ClO_4^-

Nonpolar compounds also do not dissolve in water. The attractive forces that operate between the particles in a nonpolar compound are weak dispersion forces. In order for a nonpolar molecule to dissolve in water, it would need to break up some of the hydrogen bonds between adjacent water molecules. In the case of an ionic substance, these favorable interactions are replaced by other attractive interactions between the ions and the partial charges on water. However, interactions between nonpolar molecules and water are less favorable than the interactions that water makes with itself. When a nonpolar liquid such as oil is mixed with water, two separate layers form, because the liquids will not dissolve into each other (see figure below). When a polar liquid like ethanol is mixed with water, they completely blend and dissolve into one another. *Liquids that dissolve in one another in all proportions are said to be* **miscible**. *Liquids that do not dissolve in one another are called* **immiscible**. The general rule for deciding if one substance is capable of dissolving another is "like dissolves like", where the property being compared is the overall polarity of the substance. For example, a nonpolar solid such as iodine will dissolve in nonpolar lighter fluid, but it will not dissolve in polar water.

Figure $7.5.3$: Water and oil form separate layers when they are mixed because the nonpolar oil will not dissolve into the polar water. The oil forms the top layer because it is less dense than water.

Electrolytes and Nonelectrolytes

An **electrolyte** *is a compound that conducts an electric current when it is dissolved in water or melted.* In order to conduct a current, a substance must contain mobile ions that can move from one electrode to the other. All ionic compounds are electrolytes. When ionic compounds dissolve, they break apart into ions, which are then able to conduct a current. Even insoluble ionic compounds, such as $\mathrm{CaCO_3}$, are considered electrolytes because they can conduct a current in the molten (melted) state.



A conductivity apparatus is an incomplete electrical circuit that contains a source of electricity and a light bulb or meter that will show when current is flowing through the circuit. The ends of the incomplete circuit are prongs that can be lowered into a solution. If the prongs are lowered into a solution containing a sufficient number of ions, the circuit will be completed by the solution, current will flow, and the light bulb will light up. If the prongs are lowered into a solution with no ions or an insufficient number of ions, not enough current will flow to light the bulb.

Figure $7.5.4$: An apparatus for testing the conductivity of a solution.

A **nonelectrolyte** *is a compound that does not conduct an electric current in either aqueous solution or in the molten state.* Many molecular compounds, such a sugar or ethanol, are nonelectrolytes. When these compounds dissolve in water, they do not produce ions. Illustrated below is the difference between an electrolyte and a nonelectrolyte.

Dissociation

Earlier, you saw how an ionic crystal lattice breaks apart when it is dissolved in water. **Dissociation** *is the separation of ions that occurs when a solid ionic compound dissolves*. Simply undo the crisscross method that you learned when writing chemical formulas for ionic compounds, and you are left with the components of an ionic dissociation equation. The subscripts for the ions in the chemical formulas become the coefficients of the respective ions on the product side of the equations. Shown below are dissociation equations for NaCl, $\mathrm{Ca(NO_3)_2}$, and $\mathrm{(NH_4)_3PO_4}$.

$$\mathrm{NaCl}\,(s) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) \tag{7.5.1}$$

$$\mathrm{Ca(NO_3)_2}\,(s) \rightarrow \mathrm{Ca^{2+}}\,(aq) + 2\mathrm{NO_3^-}\,(aq) \tag{7.5.2}$$

$$\mathrm{(NH_4)_3PO_4}\,(s) \rightarrow 3\mathrm{NH_4^+}\,(aq) + \mathrm{PO_4^{3-}}\,(aq) \tag{7.5.3}$$

One formula unit of sodium chloride dissociates into one sodium ion and one chloride ion. The calcium nitrate formula unit dissociates into one calcium ion and two nitrate ions, because the $2+$ charge of each calcium ion requires two nitrate ions (each with a charge of $1-$) to form an electrically neutral compound. The ammonium phosphate formula unit dissociates into three ammonium ions and one phosphate ion.

Do not confuse the subscripts of the atoms within the polyatomic ion for the subscripts that result from the crisscrossing of the charges that make the original compound neutral. The 3 subscript of the ntirate ion and the 4 subscript of the ammonium ion are part of the polyatomic ion and remain a part of the ionic formula after the compound dissociates. Notice that the compounds are solids (s) that become ions when dissolved in water, producing an aqueous solution (aq).

Figure $7.5.5$: Calcium nitrate is a typical ionic compound. In an aqueous solution, it dissociates into calcium ions and nitrate ions.

Nonelectrolytes do not dissociate when forming an aqueous solution. An equation can still be written that simply shows the solid going into solution. For example, the process of dissolving sucrose in water can be written as follows:

$$\mathrm{C_{12}H_{22}O_{11}}\,(s) \rightarrow \mathrm{C_{12}H_{22}O_{11}}\,(aq) \tag{7.5.4}$$

Strong and Weak Electrolytes

Some polar molecular compounds are nonelectrolytes when the are in their pure state but become electrolytes when they are dissolved in water. Hydrogen chloride (HCl) is a gas in its pure molecular state and is a nonelectrolyte. However, when HCl is dissolved in water, it conducts a current well because the HCl molecule ionizes into hydrogen and chloride ions.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq) \tag{7.5.5}$$

When HCl is dissolved into water, it is called hydrochloric acid. Ionic compounds and some polar compounds are completely broken apart into ions and thus conduct a current very well. A **strong electrolyte** *is a solution in which almost all of the dissolved solute exists as ions*.

Some other polar molecular compounds become electrolytes upon being dissolved into water but do not ionize to a very great extent. For example, nitrous acid $\left(\mathrm{HNO_2}\right)$ only partially ionizes into hydrogen ions and nitrite ions when dissolved in water. Aqueous nitrous acid is composed of only about 5% ions and 95% intact nitrous acid molecules A **weak electrolyte** *is a solution in which only a small fraction of the dissolved solute exists as ions*. The equation showing the ionization of a weak electrolyte utilizes an equilibrium arrow, indicating an equilibrium between the reactants and products.

$$\mathrm{HNO_2}\,(aq) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{NO_2^-}\,(aq) \tag{7.5.6}$$

Contributors and Attributions

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7.6: Colloids and Suspensions

Learning Outcomes

- Describe the properties of a suspension.
- Describe the properties of a colloid.
- Distinguish among suspensions. colloids, and solutions.

Suspensions

Take a glass of water and throw in a handful of sand or dirt. Stir it and stir it and stir it. Have you made a solution? Sand and dirt do not dissolve in water, and, though it may look homogenous for a few moments, the sand or dirt gradually sinks to the bottom of the glass (see figure below). Some medications are delivered as suspensions and must be mixed well before the doses measured to make sure the patient is receiving the correct amount of medication.

Figure 7.6.1: A mixture of sand and water forms a suspension.

A **suspension** *is a heterogeneous mixture in which some of the particles settle out of the mixture upon standing*. The particles in a suspension are far larger than those of a solution, so gravity is able to pull them down out of the dispersion medium (water). The diameter for the dispersed particles in a suspension, such as the sand in the suspension described above, is typically at least 1000 times greater than those in a solution. Unlike a solution, the dispersed particles can be separated from the dispersion medium by filtering. Suspensions are considered heterogeneous because the different substances in the mixture will not remain uniformly distributed if they are not actively being mixed.

Colloids

A colloid is a heterogeneous mixture in which the dispersed particles are intermediate in size between those of a solution and a suspension. The particles are spread evenly throughout the dispersion medium, which can be a solid, liquid, or gas. Because the dispersed particles of a colloid are not as large as those of a suspension, they do not settle out upon standing. The table below summarizes the properties and distinctions between solutions, colloids, and suspensions.

Table 7.6.1: Properties of Solutions, Colloids, and Suspensions

Solution	Colloids	Suspensions
Homogeneous	Heterogeneous	Heterogeneous
Particle size: 0.01-1 nm; atoms, ions or molecules	Particle size: 1-1000 nm, dispersed; large molecules or aggregates	Particle size: over 1000 nm, suspended: large particles or aggregates
Do not separate on standing	Do not separate on standing	Particles settle out
Cannot be separated by filtration	Cannot be separated by filtration	Can be separated by filtration
Do not scatter light	Scatter light (Tyndall effect)	May either scatter light or be opaque

Colloids are unlike solutions because their dispersed particles are much larger than those of a solution. The dispersed particles of a colloid cannot be separated by filtration, but they scatter light, a phenomenon called the Tyndall effect.

Tyndall Effect

Colloids are often confused with true homogenous solutions because the individual dispersed particles of a colloid cannot be seen. When light is passed through a true solution, the dissolved particles are too small to deflect the light. However, the dispersed particles of a colloid, being larger, do deflect light (see figure below). The **Tyndall effect** *is the scattering of visible light by colloidal particles*. You have undoubtedly "seen" a light beam as it passes through fog, smoke, or a scattering of dust particles suspended in air. All three are examples of colloids. Suspensions may scatter light, but if the number of suspended particles is sufficiently large, the suspension may simply be opaque, and the light scattering will not occur.

Figure 7.6.2: Light passes through a colorless solution and is not scattered. When it passes through a diluted milk solution, the light is scattered by colloidal particles, an observation of the Tyndall effect. The Tyndall effect allows sunlight to be seen as it passes through a fine mist.

Examples of Colloids

Listed in the table below are examples of colloidal systems, most of which are very familiar. Some of these are shown below (see figure below). The dispersed phase describes the particles, while the dispersion medium is the material in which the particles are distributed.

Table 7.6.2: Classes of Colloids

Class of Colloid	Dispersed Phase	Dispersion Medium	Examples
Sol and gel	solid	liquid	paint, jellies, blood, gelatin, mud
Solid aerosol	solid	gas	smoke, dust in air
Solid emulsion	liquid	solid	cheese, butter
Liquid emulsion	liquid	liquid	milk, mayonnaise
Liquid aerosol	liquid	gas	fog, mist, clouds, aerosol spray
Foam	gas	solid	marshmallow
Foam	gas	liquid	whipped cream, shaving cream

Figure 7.6.3: Some common colloids (A) gelatin dessert, (B) smoke (solid aerosol), (C) butter (solid emulsion), (D) mayonnaise (liquid emulsion), (E) fog (liquid aerosol), (F) marshmallows (foam), (G) whipped cream (foam)

Emulsions

Butter and mayonnaise are examples of a class of colloids called emulsions. An **emulsion** *is a colloidal dispersion of a liquid in either a liquid or a solid.* A stable emulsion requires an emulsifying agent to be present. Mayonnaise is made in part of oil and vinegar. Since oil is nonpolar, and vinegar is a polar aqueous solution, the two do not mix and would quickly separate into layers. However, the addition of egg yolk causes the mixture to become stable and not separate. Egg yolk is capable of interacting with both the polar vinegar and the nonpolar oil. The egg yolk is called the emulsifying agent. Soap acts as an emulsifying agent because one end of a soap molecule is polar, and the other end is nonpolar. This allows the grease to be removed from your hands or your clothing by washing with soapy water.

Supplemental Resources

- Types of Mixtures - Solutions, Suspensions, Colloids: www.edinformatics.com/math_sc.../mixtures.html

Contributors and Attributions

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7.7: Solubility

Learning Outcomes

- List examples of solutions made from different solute-solvent combinations.
- Distinguish between saturated and unsaturated solutions.
- Explain the effects of temperature on the solubility of solids and gases.
- Use a solubility curve to determine the solubilities of substances at various temperatures.
- Explain the effect of pressure on the solubility of gases.

Previously, we looked at the primary characteristics of a solution and how water is able to dissolve solid solutes. There are many examples of solutions that do not involve water at all, or solutions that involve solutes that are not solids. The table below summarizes the possible combinations of solute-solvent states, along with examples of each.

Table 7.7.1: Solute-Solvent Combinations

Solute State	Solvent State	Example
liquid	gas	water in air
gas	gas	oxygen in nitrogen (gas mixture)
solid	liquid	salt in water
liquid	liquid	alcohol in water
gas	liquid	carbon dioxide in water
solid	solid	zinc in copper (brass alloy)
liquid	solid	mercury in silver and tin (dental amalgam)

Air is a homogeneous mixture of many different gases and, therefore, qualifies as a solution. Solid-solid solutions, such as brass, bronze, and sterling silver, are called alloys. Fish depend on oxygen gas that is dissolved in the water found in oceans, lakes and rivers (see figure below). While solid-liquid and aqueous solutions comprise the majority of solutions encountered in the chemistry laboratory, it is important to be aware of the other possibilities.

Figure 7.7.1: Large aquariums like this salt-water tank have air continually bubbled into the water so that the fish have enough dissolved oxygen to breathe.

Rate of Dissolving

We know that the dissolving of a solid by water depends upon the collisions that occur between the solvent molecules and the particles in the solid crystal. Anything that can be done to increase the frequency of those collisions and/or to give those collisions more energy will increase the rate of dissolving. Imagine that you were trying to dissolve some sugar in a glassful of tea. A packet of granulated sugar would dissolve faster than a cube of sugar. The rate of dissolving would be increased by stirring, or agitating the solution. Finally, the sugar would dissolve faster in hot tea than it would in cold tea.

Surface Area

The rate at which a solute dissolves depends upon the size of the solute particles. Dissolving is a surface phenomenon, since it depends on solvent molecules colliding with the outer surface of the solute. A given quantity of solute dissolves faster when it is ground into small particles than if it is in the form of a large chunk, because more surface area is exposed. The packet of granulated sugar exposes far more surface area to the solvent and dissolves more quickly than the sugar cube.

Agitation of the Solution

Dissolving sugar in water will occur more quickly if the water is stirred. The stirring allows fresh solvent molecules to continually be in contact with the solute. If it is not stirred, then the water right at the surface of the solute becomes saturated with dissolved sugar molecules, meaning that it is more difficult for additional solute to dissolve. The sugar cube would eventually dissolve because random motions of the water molecules would bring enough fresh solvent into contact with the sugar, but the process would take much longer. It is important to realize that neither stirring nor breaking up a solute affect the overall amount of solute that dissolves. It only affects the rate of dissolving.

Temperature

Heating up the solvent gives the molecules more kinetic energy. The more rapid motion means that the solid or liquid solvent molecules collide with the solute with greater frequency, and the collisions occur with more force. Both factors increase the rate at which the solid or liquid solute dissolves. As we will see in the next section, a temperature change not only affects the rate of dissolving, but it also affects the amount of solute that can be dissolved.

Types of Solutions

Table salt $\left(\mathrm{NaCl} \right)$ readily dissolves in water. Suppose that you have a beaker of water to which you add some salt, stirring until it dissolves. Then you add more, and that dissolves as well. If you keep adding more and more salt, eventually you will reach a point at which no more of the salt will dissolve, no matter how long or how vigorously you stir it. Why? On the molecular level, we know that the action of the water causes the individual ions to break apart from the salt crystal and enter the solution, where they remain hydrated by water molecules. What also happens is that some of the dissolved ions collide back again with the crystal and remain there. **Recrystallization** *is the process of dissolved solute returning to the solid state*. At some point, the rate at which the solid salt is dissolving becomes equal to the rate at which the dissolved solute is recrystallizing. When that point is reached, the total amount of dissolved salt remains unchanged.

When the solution equilibrium point is reached and no more solute will dissolve, the solution is said to be saturated. A **saturated solution** *is a solution that contains the maximum amount of solute that is capable of being dissolved*. At 20^oC, the maximum amount of NaCl that will dissolve in $100. \: \text{g}$ of water is $36.0 \: \text{g}$. If any more NaCl is added past that point, it will not dissolve because the solution is saturated. What if more water is added to the solution instead? Now, more NaCl would be capable of dissolving, since there is additional solvent present. An **unsaturated solution** *is a solution that contains less than the maximum amount of solute that is capable of being dissolved*. The figure below illustrates the above process, and shows the distinction between unsaturated and saturated.

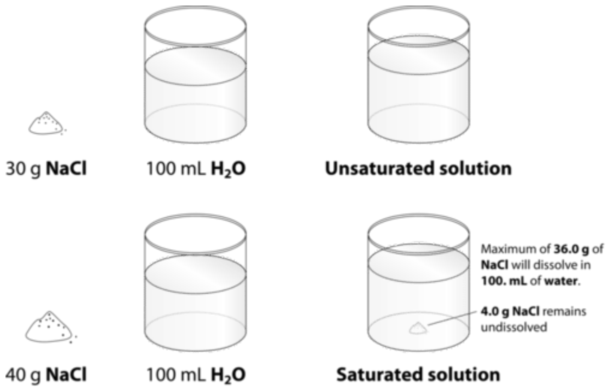

Figure $7.7.2$: When $30.0 \: \text{g}$ of NaCl is added to $100 \: \text{mL}$ of water at 20^oC, it all dissolves, forming an unsaturated solution. When $40.0 \: \text{g}$ is added, $36.0 \: \text{g}$ dissolves and $4.0 \: \text{g}$ remains undissolved, forming a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has extra undissolved solute at the bottom of the container, must be saturated (see figure below).

Figure 7.7.3: A saturated solution of salt (sodium chloride) with undissolved solute remaining at the bottom of the measuring cup.

Solubility Values

The **solubility** *of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature*. Solubility is often measured in grams of solute per 100 g of solvent. The solubility of sodium chloride in water is 36.0 g per 100 g of water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. In other words, the solubility of sodium chloride would be different in another solvent. For the purposes of this text, the solubility of a substance will refer to aqueous solubility unless otherwise specified. Solubilities for different solutes have a very wide variation, as can be seen by the data presented in the table below.

Table 7.7.2: Solubility of Solutes at Different Temperatures (g solute in 100 g H_2O)

Substance	**0°C**	**20°C**	**40°C**	**60°C**	**80°C**	**100°C**
$AgNO_3$	122	216	311	440	585	733
$Ba(OH)_2$	1.67	3.89	8.22	20.94	101.4	__
$C_{12}H_{22}O_{11}$	179	204	238	287	362	487
$Ca(OH)_2$	0.189	0.173	0.141	0.121	__	0.07
KCl	28.0	34.2	40.1	45.8	51.3	56.3
KI	128	144	162	176	192	206
KNO_3	13.9	31.6	61.3	106	167	245
$LiCl$	69.2	83.5	89.8	98.4	112	128
$NaCl$	35.7	35.9	36.4	37.1	38.0	39.2
$NaNO_3$	73	87.6	102	122	148	180
CO_2 (1 atm)	0.335	0.169	0.0973	0.058	__	__
O_2 (1 atm)	0.00694	0.00537	0.00308	0.00227	0.00138	0.00

Factors Affecting Solubility

The solubility of a solid or a liquid solute in a solvent is affected by the temperature, while the solubility of a gaseous solute is affected by both the temperature and the pressure of the gas. We will examine the effects of temperature and pressure separately.

Temperature

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a solubility curve, which is a graph of the solubility vs. temperature. Examine the solubility curves shown (see figure below).

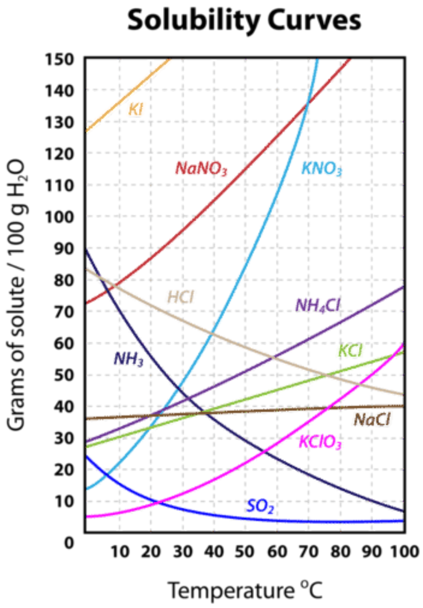

Figure 7.7.4: A solubility curve is a graph of the solubility of a substance as a function of temperature.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for $\mathrm{KNO_3}$, on the other hand, is very steep; an increase in temperature dramatically increases the solubility of $\mathrm{KNO_3}$.

Several substances listed on the graph - HCl, $\mathrm{NH_3}$, and $\mathrm{SO_2}$ - have solubilities that decrease as the temperature increases. These substances are all gases over the indicated temperature range when at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increases. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return back to the gas phase. As a result, the solubility of a gas decreases as the temperature increases. This has some profound environmental consequences. Industrial plants situated near bodies of water often use that water as a coolant, returning the warmer water back to the lake or river. This increases the overall temperature of the water, which lowers the quantity of dissolved oxygen, affecting the survival of fish and other organisms.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that 80 g of $\mathrm{KNO_3}$ is added to 100 g of water at $30^{\circ}\mathrm{C}$. According to the solubility curve, approximately 48 g of $\mathrm{KNO_3}$ will dissolve at $30^{\circ}\mathrm{C}$. This means that the solution will be saturated, since 48 g is less than 80 g. We can also determine that there will be $80 - 48 = 32$ g of undissolved $\mathrm{KNO_3}$ remaining at the bottom of the container. Now, suppose that this saturated solution is heated to $60^{\circ}\mathrm{C}$. According to the curve, the solubility of $\mathrm{KNO_3}$ at $60^{\circ}\mathrm{C}$ is about 107 g the solution is now unsaturated, since it still contains only the original 80 g of solute, all of which is now dissolved. Then, suppose the solution is cooled all the way down to $0^{\circ}\mathrm{C}$. The solubility at $0^{\circ}\mathrm{C}$ is about 14 g, meaning that $80 - 14 = 66$ g of the $\mathrm{KNO_3}$ will recrystallize.

Some solutes, such as sodium acetate, do not recrystallize easily. Suppose an exactly saturated solution of sodium acetate is prepared at $50^{\circ}\mathrm{C}$. As it cools back to room temperature, crystals do not immediately appear in the solution, even though the solubility of sodium acetate is lower at room temperature. A **supersaturated solution** *is a solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature*. The recrystallization of the excess dissolved solute is a supersaturated solution can be initiated by the addition of a tiny crystal of solute, called a seed crystal. The seed crystal provides a nucleation site on which the excess dissolved crystals can begin to grow. Recrystallization from a supersaturated solution is typically very fast.

Pressure

Pressure has very little effect on the solubility of solids or liquids, but it has a significant effect on the solubility of gases. Gas solubility increases as the partial pressure of a gas above the liquid increases. Suppose a certain volume of water is in a closed container with the space above it occupied by carbon dioxide gas at standard pressure. Some of the $\mathrm{CO_2}$ molecules come into contact with the surface of the water and dissolve into the liquid. Now suppose that more $\mathrm{CO_2}$ is added to the space above the container, causing a pressure increase. More $\mathrm{CO_2}$ molecules are now in contact with the water, so more of them dissolve. Thus the solubility increases as the pressure increases. As with a solid, the $\mathrm{CO_2}$ that is undissolved reaches an equilibrium with the dissolved $\mathrm{CO_2}$, represented by the following equation.

$$\mathrm{CO_2}\left(g\right) \rightleftharpoons \mathrm{CO_2}\left(aq\right) \tag{7.7.1}$$

At equilibrium, the rate of gaseous $\mathrm{CO_2}$ dissolving is equal to the rate of dissolved $\mathrm{CO_2}$ coming out of the solution.

When carbonated beverages are packaged, they are done so under high $\mathrm{CO_2}$ pressure so that a large amount of carbon dioxide dissolves in the liquid. When the bottle is opened, equilibrium is disrupted because the $\mathrm{CO_2}$ pressure above the liquid decreases. Immediately, bubbles of $\mathrm{CO_2}$ rapidly exit the solution and escape out of the top of the open bottle. The amount of dissolved $\mathrm{CO_2}$ decreases. If the bottle is left open for an extended period of time, the beverage becomes "flat" as more and more $\mathrm{CO_2}$ comes out of the liquid.

The relationship of gas solubility to pressure is described by Henry's law, named after English chemist William Henry (1774 - 1836). **Henry's law** *states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid*.

Supplemental Resources

- Solubility of Things: http://www.solubilityofthings.com/
- Solubility Curves: http://www.kentchemistry.com/links/K...lityCurves.htm
- Henry's Law: http://www.kentchemistry.com/links/K.../HenrysLaw.htm

Contributors and Attributions

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7.8: Solutions (Exercises)

These are homework exercises to accompany Chapter 7 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Answers are available below the solutions.

Questions

7.1: States of Matter

(click here for solutions)

Q7.1.1

In which state(s) of matter are atoms

a. closest together?
b. farthest apart?
c. filling the container?
d. fixed in position relative to one another?
e. moving past one another?
f. taking on the shape of the container?

Q7.1.2

Which of the following statements are true? Correct any false statements.

a. All substances exist as a liquid at room temperature and pressure.
b. Water changes from liquid to solid at 32°C.
c. All substances can exist as solids, liquids, or gases.

Q7.1.3

Which state of matter is most compressible?

Q7.1.4

Use online resources to find the boiling point of ethanol and dimethyl ether. Which one is higher? Why?

Q7.1.5

Describe the relationship between boiling point and altitude.

Q7.1.6

Where is the boiling point of ethanol the highest? The lowest?

a. Lexington, KY
b. New Orleans, LA
c. Salt Lake City, UT
d. Same at all locations.

7.2: Heat and Changes of State

(click here for solutions)

7.2.1

What phase change is described by each term? Is the process endothermic or exothermic?

a. sublimation
b. vaporization
c. fusion
d. deposition

Q7.2.2

List two phase changes that consume energy.

Q7.2.3

List two phase changes that release energy.

Refer to Table 7.2.1 for enthalpy values.

Q7.2.4

What is the enthalpy of fusion, vaporization, freezing, and condensation for each substance?

a. oxygen, O_2
b. ethane, C_2H_6
c. carbon tetrachloride, CCl_4
d. lead, Pb

Q7.2.5

How much energy is needed to vaporize 1.4 moles of ammonia (NH_3)?

Q7.2.6

How much energy is needed to melt 3.0 moles of ice (H_2O)?

Q7.2.7

What is the change in energy when 2.0 moles of ethanol is condensed?

Q7.2.8

What is the change in energy when 2.2 moles of oxygen is condensed?

Q7.2.9

Using the molar mass of water, convert the molar heats of fusion and vaporization for water from units of kJ/mol to kJ/g.

Q7.2.10

Calculate the quantity of heat that is absorbed or released during each process.

a. 655 g of water vapor condenses at 100°C
b. 8.20 kg of water is frozen
c. 40.0 mL of ethanol is vaporized. The density of ethanol is 0.789 g/mL.
d. 25.0 mL of ethanol condenses. The density of ethanol is 0.789 g/mL.

Q7.2.11

Various systems are each supplied with 9.25 kJ of heat. Calculate the mass of each substance that will undergo the indicated process with this input of heat.

a. melt ice at 0°C
b. vaporize water at 100°C
c. vaporize ethanol at 351 K

Q7.2.12

15.5 kJ of energy is released from each change. What mass of substance is involved?

a. condensation of NH_3
b. freezing water
c. condensation of ethanol

Q7.2.13

What is ΔH_{vap} for benzene (C_6H_6) if 7.88 kJ of energy is needed to vaporize 20.0 g of benzene?

7.3: Kinetic-Molecular Theory

(click here for solutions)

7.3.1

How are gases different from liquids and solids in terms of the distance between the particles?

Q7.3.2

Under what conditions do gases exhibit the most ideal behavior?

Q7.3.3

Which of the following are behaviors of a gas that can be explained by the kinetic-molecular theory?

a. Gases are compressible.
b. Gases exert pressure.
c. All particles of a gas sample move at the same speed.
d. Gas particles can exchange kinetic energy when they collide.
e. Gas particles move in a curved-line path.

Q7.3.4

What is an elastic collision?

Q7.3.5

Perform the indicated conversions for the following pressure measurements.

a. 1.721 atm to mmHg
b. 559 torr to kPa
c. 91.1 kPa to atm
d. 2320 mmHg to atm

Q7.3.6

a. A typical barometric pressure in Redding, California, is about 755 mmHg. Calculate this pressure in atm and kPa.
b. A typical barometric pressure in Denver, Colorado, is 615 mmHg. What is this pressure in atmospheres and kilopascals?

Q7.3.7

How does the average kinetic energy of an air sample near a campfire compare to the average kinetic energy of a sample of air that is far away from it?

7.4: The Ideal Gas Equation

(click here for solutions)

7.4.1

Complete the missing temperature values in the table.

ºC	ºF	K
25		
	99	
32		
		0
		300
	65	

Q7.4.2

What units must temperature be in for gas law calculations?

Q7.4.3

Based on R = 0.08206 $\frac{L\dot{a}tm}{mol\dot{K}}$, what units should be used in ideal gas law calculations?

Q7.4.4

A 1.00 mol sample of gas is at 300 K and 4.11 atm. What is the volume of the gas under these conditions?

Q7.4.5

What is the pressure in a 2.5 L container with 2.5 moles of gas at 293 K?

Q7.4.6

How many moles of carbon monoxide, CO, are in an 11.2-L sample at 744 torr at 55 °C?

Q7.4.7

A weather balloon contains 8.80 moles of helium at a pressure of 0.992 atm and a temperature of 25 °C at ground level. What is the volume of the balloon under these conditions?

Q7.4.8

A balloon inflated with three breaths of air has a volume of 1.7 L. At the same temperature and pressure, what is the volume of the balloon if five more same-sized breaths are added to the balloon?

Q7.4.9

The volume of an automobile air bag was 66.8 L when inflated at 25 °C with 77.8 g of nitrogen (N_2) gas. What was the pressure in the bag in kPa?

Q7.4.10

How many moles of gaseous boron trifluoride, BF_3, are contained in a 4.3410-L bulb at 788.0 K if the pressure is 1.220 atm? How many grams of BF_3?

Q7.4.11

How is the combined gas law is simplified for each set of conditions?

a. constant V and n
b. constant n
c. constant P and V
d. constant T and n
e. constant V and T
f. constant P and n
g. constant T

Q7.4.12

A nitrogen sample has a pressure of 0.56 atm with a volume of 2.0 L. What is the final pressure if the volume is compressed to a volume of 0.75 L? Assume constant moles and temperature.

Q7.4.13

A 2.50-L volume of hydrogen measured at –196 °C is warmed to 100 °C. Calculate the volume of the gas at the higher temperature, assuming no change in pressure.

Q7.4.14

A high altitude balloon is filled with 1.41×10^4 L of hydrogen at a temperature of 21 °C and a pressure of 745 torr. What is the volume of the balloon at a height of 20 km, where the temperature is –48 °C and the pressure is 63.1 torr?

Q7.4.15

A cylinder of medical oxygen has a volume of 35.4 L, and contains O_2 at a pressure of 151 atm and a temperature of 25 °C. What volume of O_2 does this correspond to at normal body conditions, that is, 1 atm and 37 °C?

Q7.4.16

A 0.50 L container of helium expands to 1.50 L. By what factor does the pressure change? Assume constant moles and temperature.

Q7.4.17

A sample of oxygen gas has an initial pressure and volume of 1.0 L and 1.0 atm. What is the final pressure if the volume is compressed to 0.50 L? Assume constant moles and temperature.

Q7.4.18

A sample of gas has a volume of 2.75 L at a temperature of 100 K. What is the volume of the gas when the temperature increases to 200 K? Assume constant pressure and moles.

Q7.4.19

What is the final volume of a gas that was originally at 0.75 L at 25°C and a final temperature of 50°C? Assume constant pressure and moles.

Q7.4.20

A sample of nitrogen is at 45°C with a volume of 2.5 L. What is the final temperature in °C if the volume is compressed to 1.4 L? Assume constant pressure and moles.

Q7.4.21

A 2.00 mole sample of gas is in a 3.50 L container. What happens to the volume when an additional 0.75 moles of gas is added? Assume pressure and temperature are constant.

Q7.4.22

A 1.85 mole sample of helium has a volume of 2.00 L. Additional helium is added at constant pressure and temperature until the volume is 3.25 L. What is the total moles of helium present in the sample? What mass of helium was added?

Q7.4.23

If the temperature of a fixed amount of a gas is doubled at constant volume, what happens to the pressure?

Q7.4.24

If the volume of a fixed amount of a gas is tripled at constant temperature, what happens to the pressure?

7.5: Aqueous Solutions

(click here for solutions)

7.5.1

Describe the solution, solvent, and solute.

Q7.5.2

How do solutions differ from compounds? Are solutions heterogeneous or homogeneous mixtures?

Q7.5.3

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?

Q7.5.4

What are the differences between strong, weak and non-electrolytes.

Q7.5.5

Write dissociation equations for the following strong electrolytes.

a. $NaCl(s)$
b. $CoCl_3(s)$
c. $Li_2S(s)$
d. $MgBr_2(s)$

e. $CaF_2(s)$

Q7.5.6

Based on the given information, identify each as a strong, weak, or non-electrolyte.

a. $C_6H_{12}O_6(s) \rightarrow C_6H_{12}O_6(aq)$
b. NaCl is added to water and the conductivity increases dramatically.
c. 1.5 moles of HCl are added to a container of water. The resulting solution has 1.5 moles of H^+ ions and 1.5 moles of Cl\- ions.
d. Acetic acid (CH_3COOH) partially dissociates in water.
e. An HCN solution contains 0.50 moles of HCN molecules and 0.05 moles of H^+ ions and 0.05 moles of Cl\- ions.
f. Acetone is added to water and the conductivity does not change.

7.6: Colloids and Suspensions

(click here for solutions)

7.6.1

How can you distinguish between a suspension and a solution?

Q7.6.2

How big are the particles in a colloid compared to those of a suspension and a solution?

Q7.6.3

What is the Tyndall effect? Why don't solutions demonstrate the Tyndall effect?

Q7.6.4

Explain the difference between the dispersed phase and the dispersing medium of a colloid.

Q7.6.5

Identify each of the following descriptions or examples as being representative of a solution, suspension, or colloid. More than one answer may apply.

a. dispersed particles can be filtered out
b. heterogeneous
c. particles are not visible to the unaided eye
d. paint
e. lemonade with no pulp
f. particle size larger than 1 nm
g. milk
h. particles do not settle upon standing
i. fog

7.7: Solubility

(click here for solutions)

7.7.1

Describe the difference between saturated and unsaturated solutions.

Q7.7.2

What are two things that you could do to change an unsaturated solution into a saturated solution?

Q7.7.3

A given solution is clear and colorless. A single crystal of solute is added to the solution. Describe what happens in each of the following situations.

a. The original solution was saturated.
b. The original solution was unsaturated.

Q7.7.4

List the original states (solid, liquid, or gas) of the solute and solvent that are combined to make each of the following solutions.

a. an alloy
b. salt water
c. carbonated water
d. oil in gasoline

Q7.7.5

Answer the following using the solubility curve diagram.

a. How many grams of NH_4Cl are required to make a saturated solution in 100 g of water at 70°C?
b. How many grams of NH_4Cl could be dissolved in 200 g of water at 70°C?
c. At what temperature is a solution of 50 grams of KNO_3 dissolved in 100 grams of water a saturated solution?
d. Which two substances in the above graph have the same solubility at 85°C?
e. How many grams of $NaNO_3$ can be dissolved in 100 grams of water to make a saturated solution at 25°C?
f. How much KI can be dissolved in 5 grams of water at 20°C to make a saturated solution?

Q7.7.6

An exactly saturated solution of $KClO_3$ is prepared at 90°C using 100 grams of water. If the solution is cooled to 20°C, how many grams of $KClO_3$ will recrystallize (i.e. come out of solution)?

Q7.7.7

Indicate whether the following solutions are unsaturated or saturated.

a. 22 grams of $KClO_3$ is dissolved in 100 g of water at 50°C.
b. 60 grams of KNO_3 is dissolved in 100 g of water at 50°C.
c. 50 grams of NaCl is dissolved in 100 g of water at 50°C.

Q.7.7.8

Under which set of conditions is the solubility of a gas in a liquid the greatest?

a. low temperature and low pressure
b. low temperature and high pressure
c. high temperature and low pressure
d. high temperature and high pressure

Answers

7.1: States of Matter

Q7.1.1

a. solid
b. gas
c. gas
d. solid
e. liquid and gas
f. liquid and gas

Q7.1.2

Which of the following statements are true? Correct any false statements.

a. All substances exist as a liquid at ~~room temperature and pressure~~. at some temperature and pressure.
b. Water changes from liquid to solid at 32~~°C~~ °F.
c. True (although some states are rarely seen for some substances).

Q7.1.3

gas

Q7.1.4

ethanol 78°C; dimethyl ether −24°C

Ethanol has stronger intermolecular forces due to having hydrogen bonding which is not seen in dimethyl ether. The stronger the intermolecular forces, the higher the boiling point.

Q7.1.5

As the altitude increases, the boiling point decreases

Q7.1.6

a. Lexington, KY (altitude = 978 feet)
b. New Orleans, LA (altitude = 2 feet) - HIGHEST
c. Salt Lake City, UT (altitude = 4226 feet) - LOWEST

7.2: Heat and Changes of State

Q7.2.1

a. solid to gas; endothermic
b. liquid to gas; endothermic
c. solid to liquid; endothermic
d. gas to solid; exothermic

Q7.2.2

Any two of fusion, vaporization, or sublimation.

Q7.2.3

Any two of freezing, condensation, deposition.

Q7.2.4

Substance	ΔH_{fus} (kJ/mol)	ΔH_{vap} (kJ/mol)	$\Delta H_{freezing}$ (kJ/mol)	$\Delta H_{condensation}$ (kJ/mol)
oxygen, O_2	0.44	6.82	-0.44	-6.82
ethane, C_2H_6	2.85	14.72	-2.85	-14.72
carbon tetrachloride, CCl_4	2.67	30.0	-2.67	-30.0
lead, Pb	4.77	178	-4.77	-178

Q7.2.5

$$1.4\ mol\ \mathrm{NH_3}\left(\frac{23.35\ kJ}{mol}\right) = 33\ kJ$$

Q7.2.6

$$3.0\ mol\ \mathrm{H_2O}\left(\frac{6.01\ kJ}{mol}\right) = 18\ kJ$$

Q7.2.7

$$2.0\ mol\ \mathrm{CH_3CH_2OH}\left(\frac{-38.56\ kJ}{mol}\right) = -77\ kJ$$

Q7.2.8

$$2.2\ mol\ \mathrm{O_2}\left(\frac{-6.82\ kJ}{mol}\right) = -15\ kJ$$

Q7.2.9

$$\frac{6.01\ kJ}{mol}\left(\frac{1\ mol}{18.02\ g}\right) = \frac{0.334\ kJ}{g}$$

$$\frac{40.7\ kJ}{mol}\left(\frac{1\ mol}{18.02\ g}\right)=\frac{2.26\ kJ}{g}$$

Q7.2.10

a. $655\ g\ \mathrm{H_2O}\left(\frac{1\ mol}{18.02\ g}\right)\left(\frac{-40.7\ kJ}{mol}\right)=-1.48\times10^{3}\ kJ$
b. $8.20\ kg\ \mathrm{H_2O}\left(\frac{1000\ g}{1\ kg}\right)\left(\frac{1\ mol}{18.02\ g}\right)\left(\frac{-6.01\ kJ}{mol}\right)=-2.73\times10^{3}\ kJ$
c. $40.0\ mL\ \mathrm{CH_3CH_2OH}\left(\frac{0.789\ g}{1\ mL}\right)\left(\frac{1\ mol}{46.07\ g}\right)\left(\frac{38.56\ kJ}{mol}\right)=26.4\ kJ$
d. $25.0\ mL\ \mathrm{CH_3CH_2OH}\left(\frac{0.789\ g}{1\ mL}\right)\left(\frac{1\ mol}{46.07\ g}\right)\left(\frac{-38.56\ kJ}{mol}\right)=-16.5\ kJ$

Q7.2.11

a. $9.25\ kJ\left(\frac{mol}{6.01\ kJ}\right)\left(\frac{18.02\ g}{mol}\right)=27.7\ g\ \mathrm{H_2O}$
b. $9.25\ kJ\left(\frac{mol}{40.7\ kJ}\right)\left(\frac{18.02\ g}{mol}\right)=4.10\ g\ \mathrm{H_2O}$
c. $9.25\ kJ\left(\frac{mol}{38.56\ kJ}\right)\left(\frac{46.07\ g}{mol}\right)=11.1\ g\ \mathrm{CH_3CH_2OH}$

Q7.2.12

a. $-15.5\ kJ\left(\frac{mol}{-23.35\ kJ}\right)\left(\frac{17.03\ g}{mol}\right)=11.3\ g\ \mathrm{NH_3}$
b. $-15.5\ kJ\left(\frac{mol}{-6.01\ kJ}\right)\left(\frac{18.02\ g}{mol}\right)=46.5\ g\ \mathrm{H_2O}$
c. $-15.5\ kJ\left(\frac{mol}{-38.56\ kJ}\right)\left(\frac{46.07\ g}{mol}\right)=18.5\ g\ \mathrm{CH_3CH_2OH}$

Q7.2.13

Find the moles of benzene.

$$20.0\ g\ \mathrm{C_6H_6}\left(\frac{1\ mol}{78.11\ g}\right)=0.256\ mol\ \mathrm{C_6H_6}$$

Combine the energy with the moles to calculate the enthalpy of vaporization.

$$\Delta H_{vap}=\frac{7.88\ kJ}{0.256\ mol}=\frac{30.8\ kJ}{mol}$$

7.3: Kinetic-Molecular Theory

Q7.3.1

Gas particles are much farther from one another than liquid or solid particles.

Q7.3.2

Gases have the most ideal behavior at high temperatures (molecules moving more quickly than at low temperatures so less time to interact) and at low pressure (molecules are farther apart from one another than at high pressure).

Q7.3.3

a. Molecules are very far apart from one another and are compressible.
b. Gases are in constant random motion so they collide with the walls of the container.
c. False. Molecules of the same substance are moving at a range of speeds.
d. Collisions are elastic. Energy is exchanged but not lost when two particles coll
e. False. Particles move in a straight line.

Q7.3.4

A collision in which no energy is lost.

Q7.3.5

a. $1.721\ atm\left(\frac{760\ mmHg}{1\ atm}\right)=1308\ mmHg$
b. $559\ torr\left(\frac{101.3\ kPa}{760\ torr}\right)=74.5\ kPa$
c. $91.1\ kPa\left(\frac{1\ atm}{101.3\ kPa}\right)=0.899\ atm$

d. $2320\ mmHg\left(\frac{1\ atm}{760\ mmHg}\right)=3.05\ atm$

Q7.3.6

a. $755\ mmHg\left(\frac{1\ atm}{760\ mmHg}\right)=0.993\ atm\,755\ mmHg\left(\frac{101.3\ kpa}{760\ mmHg}\right)=101\ kPa$
b. $615\ mmHg\left(\frac{1\ atm}{760\ mmHg}\right)=0.809\ atm\,615\ mmHg\left(\frac{101.3\ kpa}{760\ mmHg}\right)=82.0\ kPa$

Q7.3.7

Closer to the fire, it is warmer and the kinetic energy of the particles (and therefore the average speed) will be greater.

7.4: The Ideal Gas Equation

Q7.4.1

^{o}C	^{o}F	K
25	77	298
37	99	310
32	90	305
-273	-459	0
27	80	300
18	65	291

Q7.4.2

Kelvin

Q7.4.3

P (atm), V (L), n (mol), T (K)

Q7.4.4

A 1.00 mol sample of gas is at 300 K and 4.11 atm. What is the volume of the gas under these conditions?

$$PV=nRT$$
$$(4.11\ atm)\,V=(1.00\ mol)\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)(300\ K)$$
$$V=5.99\ L$$

Q7.4.5

$$PV=nRT$$
$$P\,(2.5\ L)=(2.5\ mol)\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)(293\ K)$$
$$P=24\ atm$$

Q7.4.6

$$744\ torr\left(\frac{1\ atm}{760\ mmHg}\right)=0.979\ atm$$

T = 55°C + 273.15 = 328 K

$$PV=nRT$$
$$(0.979\ atm)\,(11.2\ L)=n\left(0.08206\frac{L\cdot atm}{mol\cdot K}\right)(328\ K)$$
$$n=0.407\ mol$$

Q7.4.7

T = 25°C + 273.15 = 298 K

$$PV = nRT$$
$$(0.992\ atm)\,V = (8.80\ mol)\left(0.08206\,\frac{L\cdot atm}{mol\cdot K}\right)(298\ K)$$
$$V = 217\ L$$

Q7.4.8

x = volume of one breath of air

$3x = 1.7\ L$

$x = 0.57\ L$

Balloon will have a total of 8 breaths of air (3 original plus 5 additional)

$V = 8x = 8(0.57\ L) = 4.6L$

Q7.4.9

$$77.8\ g\ N_2\left(\frac{1\ mol}{28.02\ g}\right) = 2.78\ mol\ N_2$$

T = 25°C + 273.15 = 298 K

$$PV = nRT$$
$$P\,(66.8\ L) = (2.78\ mol)\left(0.08206\,\frac{L\cdot atm}{mol\cdot K}\right)(298\ K)$$
$$P = 1.02\ atm$$

$$1.02\ atm\left(\frac{101.3\ kPa}{1\ atm}\right) = 103\ kPa$$

Q7.4.10

$$PV = nRT$$
$$(1.220\ atm)\,(4.3410\ L) = mol\left(0.08206\,\frac{L\cdot atm}{mol\cdot K}\right)(788.0\ K)$$
$$n = 0.08190\ mol\ BF_3$$

$$0.08190\ mol\left(\frac{67.82\ g}{mol}\right) = 5.554\ g\ BF_3$$

Q7.4.11

a. $\frac{P_i}{T_i} = \frac{P_f}{T_f}$
b. $\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$
c. $\frac{1}{n_iT_i} = \frac{1}{n_fT_f}$ or $n_iT_i = n_fT_f$
d. $P_iV_i = P_fV_f$
e. $\frac{P_i}{n_i} = \frac{P_f}{n_f}$
f. $\frac{V_i}{T_i} = \frac{V_f}{T_f}$
g. $\frac{P_iV_i}{n_i} = \frac{P_fV_f}{n_f}$

Q7.4.12

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$P_iV_i = P_fV_f$$
$$0.56\ atm\cdot 2.0L = P_f\cdot 0.75\ L$$
$$P_f = 1.5\ atm$$

Q7.4.13

$$K = -196^{\circ}C + 273.15 = 77\ K$$
$$K = 100^{\circ}C + 273.15 = 373\ K$$

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$

$$\frac{2.50\ L}{77\ K} = \frac{V_f}{373\ K}$$
$$V_f = 12\ L$$

Q7.4.14

$$K = 21^{\circ}C + 273.15 = 294\ K$$
$$K = -48^{\circ}C + 273.15 = 225\ K$$

$$745\ torr\left(\frac{1\ atm}{760\ torr}\right) = 0.980\ atm$$
$$63.1\ torr\left(\frac{1\ atm}{760\ torr}\right) = 0.0830\ atm$$

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$
$$\frac{0.980\ atm \cdot 1.4 \times 10^4\ L}{294\ K} = \frac{P_i \cdot 0.0830\ atm}{225}$$
$$P_f = 1.27 \times 10^5\ atm$$

Q7.4.15

$$K = 25^{\circ}C + 273.15 = 298\ K$$
$$K = -37^{\circ}C + 273.15 = 310\ K$$

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$
$$\frac{151\ atm \cdot 35.4\ L}{298\ K} = \frac{1\ atm \cdot V_f}{310}$$
$$V_f = 5.56 \times 10^3\ L$$

Q7.4.16

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$P_iV_i = P_fV_f$$

Set the initial pressure = x to calculate the factor of change in terms of x.

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$P_iV_i = P_fV_f$$
$$x \cdot 0.50\ L = P_f \cdot 1.50\ L$$
$$P_f = \frac{1}{3}x$$

The final pressure is one third of the original pressure.

Q7.4.17

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$P_iV_i = P_fV_f$$
$$1.0\ atm \cdot 1.0\ L = P_f \cdot 0.50\ L$$
$$P_f = 2.0\ atm$$

Q7.4.18

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$\frac{V_i}{T_i} = \frac{V_f}{T_f}$$
$$\frac{2.75\ L}{100\ K} = \frac{V_f}{200\ K}$$
$$V_f = 5.50\ L$$

Q7.4.19

$$K = 25^{\circ}C + 273.15 = 298\ K$$
$$K = -50^{\circ}C + 273.15 = 323\ K$$

$$\frac{P_iV_i}{n_iT_i}=\frac{P_fV_f}{n_fT_f}$$
$$\frac{V_i}{T_i}=\frac{V_f}{T_f}$$
$$\frac{0.75\ L}{298\ K}=\frac{V_f}{323\ K}$$
$$V_f=0.813\ L$$

Q7.4.20

$$K=45^\circ C+273.15=318\ K$$

$$\frac{P_iV_i}{n_iT_i}=\frac{P_fV_f}{n_fT_f}$$
$$\frac{V_i}{T_i}=\frac{V_f}{T_f}$$
$$\frac{2.5\ L}{318\ K}=\frac{1.4\ L}{T_f}$$
$$T_f=178\ K$$

$$^\circ C=K-273.15$$
$$^\circ C=178-273.15$$
$$^\circ C=-95\ K$$

Q7.4.21

$$\frac{P_iV_i}{n_iT_i}=\frac{P_fV_f}{n_fT_f}$$
$$\frac{V_i}{n_i}=\frac{V_f}{n_f}$$
$$\frac{3.50\ L}{2.00\ mol}=\frac{V_f}{2.75\ mol}$$
$$V_f=4.81\ L$$

Note the final moles is 2.75 because the problem says that 0.75 moles of gas is **added** to the original amount of 2.00 moles.

Q7.4.22

$$\frac{P_iV_i}{n_iT_i}=\frac{P_fV_f}{n_fT_f}$$
$$\frac{V_i}{n_i}=\frac{V_f}{n_f}$$
$$\frac{2.00\ L}{1.85\ mol}=\frac{3.25\ L}{n_f}$$
$$n_f=3.01\ mol$$

$$\text{moles added}=3.01\ mol-1.85\ mol$$
$$\text{moles added}=1.16\ mol$$

$$1.16\ mol\ He\left(\frac{4.003\ g}{mol}\right)=4.64\ g\ \text{He}$$

Q7.4.23

$$\frac{P_iV_i}{n_iT_i}=\frac{P_fV_f}{n_fT_f}$$
$$\frac{P_i}{T_i}=\frac{P_f}{T_f}$$

The temperature is doubled so $T_f=2\cdot T_i$

Let $P_i=x$ to see the factor the pressure changes.

$$\frac{P_iV_i}{n_iT_i}=\frac{P_fV_f}{n_fT_f}$$
$$\frac{P_i}{T_i}=\frac{P_f}{T_f}$$
$$\frac{x}{T_i}=\frac{P_f}{2\cdot T_i}$$
$$x=\frac{P_f}{2}$$
$$P_f=2x$$

The final pressure is twice the initial pressure.

Q7.4.24

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$P_iV_i = P_fV_f$$

The volume is tripled so $V_f = 3 \cdot V_i$

Let $P_i = x$ to see the factor the pressure changes.

$$\frac{P_iV_i}{n_iT_i} = \frac{P_fV_f}{n_fT_f}$$
$$P_iV_i = P_fV_f$$
$$x \cdot V_i = P_f \cdot 3V_i$$
$$x = 3P_f$$
$$P_f = \frac{1}{3}x$$

The final pressure is one-third of the initial pressure.

7.5: Aqueous Solutions

Q7.5.1

The solute is present in the smaller amount, the solvent is present in the larger amount, and the solution is the combination of the solute and solvent.

Q7.5.2

Solutions are a homoogeneous mixture of two or more compounds.

Q7.5.3

Endothermic because heat was needed to dissolve the KNO_3. Heat present in the solution was consumed by the dissolution process.

Q7.5.4

Strong electrolytes completely dissociate into ions in aqueous solution and are conductors of electricity. Weak electrolytes partially dissociate into ions in aqueous solutions are are weak conductors of electricity. Non-electrolytes do not dissociate into ions in aqueous solution and are poor conductors of electricity.

Q7.5.5

a. $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$
b. $CoCl_3(s) \rightarrow Co^{3+}(aq) + 3Cl^-(aq)$
c. $Li_2S(s) \rightarrow 2Li^+(aq) + S^{2-}(aq)$
d. $MgBr_2(s) \rightarrow Mg^{2+}(aq) + 2Br^-(aq)$
e. $CaF_2(s) \rightarrow Ca^{2+}(aq) + 2F^-(aq)$

Q7.5.6

Based on the given information, identify each as a strong, weak, or non-electrolyte.

a. non-electrolyte
b. strong electrolyte
c. strong electrolyte
d. weak electrolyte
e. weak electrolyte
f. non-electrolyte

7.6: Colloids and Suspensions

Q7.6.1

A suspension can be separated from the solvent by filtration while a solution cannot because particles settle out of suspensions but not solutions.

Q7.6.2

Particles in a solution are less than 1 nanometer, colloids have particles from 1-1000 nm, and suspensions have particles over 1000 nm.

Q7.6.3

The Tyndall effect is the scattering of visible light by particles. The particles in colloids are large enough to scatter light while the particles in solutions are too small to scatter light. Solutions are transparent (we can see through them) because the particles are so small.

Q7.6.4

The dispersed phase is present in the smaller amount and the dispersing medium is present in a larger amount.

Q7.6.5

a. suspension
b. colloids and suspensions
c. solution
d. colloid
e. solution
f. colloids and suspensions
g. colloid
h. solutions and colloids
i. colloids

7.7: Solubility

Q7.7.1

A saturated solution has the maximum amount of solute dissolved. An unsaturated solution does not have the maximum amount dissolved; additional solute can be added and will dissolve.

Q7.7.2

1. Addition of solute to the solution until no more dissolves.
2. Removal of solvent such as through evaporation.

Q7.7.3

a. The added solute will not dissolve.
b. The added solute will dissolve.

Q7.7.4

a. The solute and solvent are both solids.
b. The solute is a solid and the solvent is a liquid.
c. The solute is a gas and the solvent is a liquid.
d. The solute and solvent are both liquids.

Q7.7.5

a. 60 g NH_4Cl
b. 120 g NH_4Cl
c. 31^oC
d. HCl and $KClO_3$
e. 90 g
f. 7 g

Q7.7.6

At 90°C, 50 g of $KClO_3$ will dissolve in 100 g of water for a saturated solution. At 20°C, only 10 g of $KClO_3$ is dissolved in 100 g of water for a saturated solution. 40 grams of $KClO_3$ will precipitate out of solution.

Q7.7.7

a. saturated
b. unsaturated
c. saturated (with additional undissolved solute)

Q.7.7.8

The solubility of a gas in a liquid is the greatest at low temperature and high pressure.

CHAPTER OVERVIEW

8: Properties of Solutions

8.1: Concentrations of Solutions

Learning Outcomes

- Define concentration.
- Use the terms concentrated and dilute to describe the relative concentration of a solution.
- Calculate the molarity of a solution.
- Calculate percentage concentration (m/m, v/v, m/v).
- Describe a solution whose concentration is in ppm or ppb.
- Use concentration units in calculations.
- Determine equivalents for an ion.
- Complete calculations relating equivalents to moles, volumes, or mass.
- Complete dilution calculations.

There are several ways to express the amount of solute present in a solution. The **concentration** *of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution*. A **concentrated solution** *is one that has a relatively large amount of dissolved solute.* A **dilute solution** *is one that has a relatively small amount of dissolved solute*. However, these terms are relative, and we need to be able to express concentration in a more exact, quantitative manner. Still, concentrated and dilute are useful as terms to compare one solution to another (see figure below). Also, be aware that the terms "concentrate" and "dilute" can be used as verbs. If you were to heat a solution, causing the solvent to evaporate, you would be concentrating it, because the ratio of solute to solvent would be increasing. If you were to add more water to an aqueous solution, you would be diluting it because the ratio of solute to solvent would be decreasing.

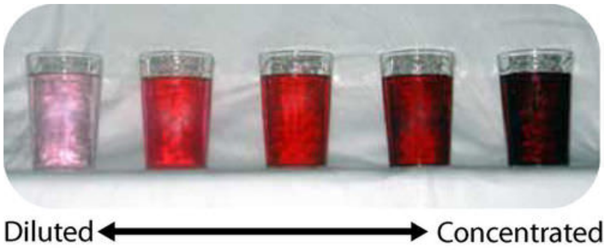

Figure 8.1.1: Solutions of a red dye in water from the most dilute (on the left) to the most concentrated (on the right).

Percent Concentration

One way to describe the concentration of a solution is by the percent of the solution that is composed of the solute. This percentage can be determined in one of three ways: (1) the mass of the solute divided by the mass of solution, (2) the volume of the solute divided by the volume of the solution, or (3) the mass of the solute divided by the volume of the solution. Because these methods generally result in slightly different vales, it is important to always indicate how a given percentage was calculated.

Mass Percent

When the solute in a solution is a solid, a convenient way to express the concentration is a mass percent (mass/mass), which is the grams of solute per 100 g of solution.

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{8.1.1}$$

Suppose that a solution was prepared by dissolving 25.0 g of sugar into 100 g of water. The percent by mass would be calculated as follows:

$$\text{Percent by mass} = \frac{25\text{ g sugar}}{125\text{ g solution}} \times 100\% = 20\%\text{ sugar} \tag{8.1.2}$$

Sometimes, you may want to make a particular amount of solution with a certain percent by mass and will need to calculate what mass of the solute is needed. For example, let's say you need to make 3.00×10^3 g of a sodium chloride solution that is 5.00% by mass. You can rearrange and solve for the mass of solute.

$$\% \text{ by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{8.1.3}$$

$$5.00\% = \frac{\text{mass of solute}}{3.00 \times 10^3 \text{ g solution}} \times 100\% \tag{8.1.4}$$

$$\text{mass of solute} = 150.\text{ g} \tag{8.1.5}$$

You would need to weigh out 150 g of NaCl and add it to 2850 g of water. Notice that it was necessary to subtract the mass of the NaCl (150 g) from the mass of solution $(3.00 \times 10^3 \text{ g})$ to calculate the mass of the water that would need to be added.

Volume Percent

The percentage of solute in a solution can more easily be determined by volume when the solute and solvent are both liquids. The volume of the solute divided by the volume of the solution expressed as a percent, yields the percent by volume (volume/volume) of the solution. If a solution is made by taking 40. mL of ethanol and adding enough water to make 240. mL of solution, the percent by volume is:

$$\text{Percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{8.1.6}$$

$$= \frac{40 \text{ mL ethanol}}{240 \text{ mL solution}} \times 100\% \tag{8.1.7}$$

$$= 16.7\% \text{ ethanol} \tag{8.1.8}$$

Frequently, ingredient labels on food products and medicines have amounts listed as percentages (see figure below).

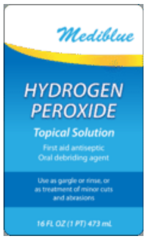

Figure 8.1.2: Hydrogen peroxide is commonly sold as a 3% by volume solution for use as a disinfectant.

It should be noted that, unlike in the case of mass, you cannot simply add together the volumes of solute and solvent to get the final solution volume. When adding a solute and solvent together, mass is conserved, but volume is not. In the example above, a solution was made by starting with 40 mL of ethanol and adding enough water to make 240 mL of solution. Simply mixing 40 mL of ethanol and 200 mL of water would not give you the same result, as the final volume would probably not be exactly 240 mL.

The mass-volume percent is also used in some cases and is calculated in a similar way to the previous two percentages. The mass/volume percent is calculated by dividing the mass of the solute by the volume of the solution and expressing the result as a percent.

For example, if a solution is prepared from 10 NaCl in enough water to make a 150 mL solution, the mass-volume concentration is

$$\text{Mass-volume concentration} \frac{\text{mass solute}}{\text{volume solution}} \times 100\% \tag{8.1.9}$$

$$= \frac{10 \text{ g NaCl}}{150 \text{ mL solution}} \times 100\% \tag{8.1.10}$$

$$= 6.7\% \tag{8.1.11}$$

Parts per Million and Parts per Billion

Two other concentration units are parts per million and parts per billion. These units are used for very small concentrations of solute such as the amount of lead in drinking water. Understanding these two units is much easier if you consider a percentage as parts per hundred. Remember that 85% is the equivalent of 85 out of a hundred. A solution that is 15 ppm is 15 parts solute per 1 million parts solution. A 22 ppb solution is 22 parts solute per billion parts solution. While there are several ways of expressing two units of ppm and ppb, we will treat them as mg or μg of solutes per L solution, respectively.

For example, 32 ppm could be written as $\frac{32 \text{ mg solute}}{1 \text{ L solution}}$ while 59 ppb can be written as $\frac{59\ \mu\text{g solute}}{1 \text{ L solution}}$.

Molarity

Chemists primarily need the concentration of solutions to be expressed in a way that accounts for the number of particles present that could react according to a particular chemical equation. Since percentage measurements are based on either mass or volume, they are generally not useful for chemical reactions. A concentration unit based on moles is preferable. The **molarity** (M) *of a solution is the number of moles of solute dissolved in one liter of solution*. To calculate the molarity of a solution, you divide the moles of solute by the volume of the solution expressed in liters.

$$\text{Molarity } (\mathrm{M}) = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\mathrm{L}} \tag{8.1.12}$$

Note that the volume is in liters of solution and not liters of solvent. When a molarity is reported, the unit is the symbol M, which is read as "molar". For example, a solution labeled as $1.5 \mathrm{~M~NH_3}$ is a "1.5 molar solution of ammonia".

Example 8.1.1

A solution is prepared by dissolving $42.23 \mathrm{~g}$ of $\mathrm{NH_4Cl}$ into enough water to make $500.0 \mathrm{~mL}$ of solution. Calculate its molarity.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass of $\mathrm{NH_4Cl} = 42.23 \mathrm{~g}$
- Molar mass of $\mathrm{NH_4Cl} = 53.50 \mathrm{~g/mol}$
- Volume of solution $= 500.0 \mathrm{~mL} = 0.5000 \mathrm{~L}$

Unknown

- Molarity $= ? \mathrm{~M}$

The mass of the ammonium chloride is first converted to moles. Then, the molarity is calculated by dividing by liters. Note that the given volume has been converted to liters.

Step 2: Solve.

$$42.23 \mathrm{~g~NH_4Cl} \times \frac{1 \mathrm{~mol~NH_4Cl}}{53.50 \mathrm{~g~NH_4Cl}} = 0.7893 \mathrm{~mol~NH_4Cl} \tag{8.1.13}$$

$$\frac{0.7893 \mathrm{~mol~NH_4Cl}}{0.5000 \mathrm{~L}} = 1.579 \mathrm{~M} \tag{8.1.14}$$

Step 3: Think about your result.

The molarity is $1.579 \mathrm{~M}$, meaning that a liter of the solution would contain 1.579 moles of $\mathrm{NH_4Cl}$. Having four significant figures is appropriate.

Figure 8.1.3: Volumetric flasks come in many sizes, each designed to prepare a different volume of solution.

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. This is because the number of moles of the solute does not change, but the total volume of the solution increases. We can set up an equality between the moles of the solute before the dilution (1) and the moles of the solute after the dilution (2).

$$\text{mol}_1 = \text{mol}_2 \tag{8.1.15}$$

Since the moles of solute in a solution is equal to the molarity multiplied by the volume in liters, we can set those equal.

$$M_1 \times L_1 = M_2 \times L_2 \tag{8.1.16}$$

Finally, because the two sides of the equation are set equal to one another, the volume can be in any units we choose, as long as that unit is the same on both sides. Our equation for calculating the molarity of a diluted solution becomes:

$$M_1 \times V_1 = M_2 \times V_2 \tag{8.1.17}$$

Additionally, the concentration can be in any other unit as long as M_1 and M_2 are in the same unit.

Suppose that you have $100.\ \text{mL}$ of a $2.0\ \text{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \text{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2.

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \text{M} \times 100.\ \text{mL}}{500.\ \text{mL}} = 0.40\ \text{M HCl} \tag{8.1.18}$$

The solution has been diluted by a factor of five, since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value. Another common dilution problem involves deciding how much a highly concentrated solution is required to make a desired quantity of solution with a lower concentration. The highly concentrated solution is typically referred to as the stock solution.

Example 8.1.2

Nitric acid $\left(\text{HNO}_3\right)$ is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \text{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \text{L}$ of a $0.50\ \text{M}$ solution?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Stock HNO_3 $(M_1) = 16\ \text{M}$

- $V_2 = 8.00 \text{ L}$
- $M_2 = 0.50 \text{ M}$

Unknown

- Volume of stock HNO_3 $(V_1) = ? \text{ L}$

The unknown in the equation is V_1, the necessary volume of the concentrated stock solution.

Step 2: Solve.

$$V_1 = \frac{M_2 \times V_2}{V_1} = \frac{0.50 \text{ M} \times 8.00 \text{ L}}{16 \text{ M}} = 0.25 \text{ L} = 250 \text{ mL} \tag{8.1.19}$$

Step 3: Think about your result.

250 mL of the stock HNO_3 solution needs to be diluted with water to a final volume of 8.00 L. The dilution from 16 M to 0.5 M is a factor of 32.

Equivalents

Concentration is important in healthcare because it is used in so many ways. It's also critical to use units with any values to ensure the correct dosage of medications or report levels of substances in blood, to name just two.

Another way of looking at concentration such as in IV solutions and blood is in terms of equivalents. One equivalent is equal to one mole of charge in an ion. The value of the equivalents is always positive regardless of the charge. For example, Na^+ and Cl^- both have 1 equivalent per mole.

Ion	**Equivalents**
Na^+	1
Mg^{2+}	2
Al^{3+}	3
Cl^-	1
NO_3^-	1
SO_4^{2-}	2

(8.1.20)

Equivalents are used because the concentration of the charges is important than the identity of the solutes. For example, a standard IV solution does not contain the same solutes as blood but the concentration of charges is the same.

Sometimes, the concentration is lower in which case milliequivalents (mEq) is a more appropriate unit. Just like metric prefixes used with base units, milli is used to modify equivalents so $1 \text{ Eq} = 1000 \text{ mEq}$.

Example 8.1.3

How many equivalents of Ca^{2+} are present in a solution that contains 3.5 moles of Ca^{2+}?

Solution

Use the relationship between moles and equivalents of Ca^{2+} to find the answer.

$$3.5 \text{ mol} \cdot \frac{2 \text{ Eq}}{1 \text{ mol Ca}^{2+}} = 7.0 \text{ Eq Ca}^{2+} \tag{8.1.21}$$

Example 8.1.4

A patient received 1.50 L of saline solution which has a concentration of 154 mEq/L Na^+. What mass of sodium did the patient receive?

Solution

Use dimensional analysis to set up the problem based on the values given in the problem, the relationship for Na^+ and equivalents and the molar mass of sodium. Note that if this problem had a different ion with a different charge, that would need to be accounted for in the calculation.

$$1.50 \text{ L} \cdot \frac{154 \text{ mEq}}{1 \text{ L}} \cdot \frac{1 \text{ Eq}}{1000 \text{ mEq}} \cdot \frac{1 \text{ mol Na}^+}{1 \text{ Eq}} \cdot \frac{22.99 \text{ g}}{1 \text{ mol Na}^+} = 5.31 \text{ g Na}^+ \tag{8.1.22}$$

Contributors and Attributions

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8.2: Chemical Equilibrium

Learning Outcomes

- Explain chemical equilibrium.
- Write expression for calculating K.
- Calculate and compare Q and K values.
- Predict relative amounts of reactants and products based on equilibrium constant K.

Hydrogen and iodine gases react to form hydrogen iodide according to the following reaction:

$$\mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightleftharpoons 2\mathrm{HI}\,(g) \tag{8.2.1}$$

$$\text{Forward reaction: } \mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightarrow 2\mathrm{HI}\,(g) \tag{8.2.2}$$

$$\text{Reverse reaction: } 2\mathrm{HI}\,(g) \rightarrow \mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \tag{8.2.3}$$

Initially, only the forward reaction occurs because no HI is present. As soon as some HI has formed, it begins to decompose back into $\mathrm{H_2}$ and $\mathrm{I_2}$. Gradually, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually the rate of combination of $\mathrm{H_2}$ and $\mathrm{I_2}$ to produce HI becomes equal to the rate of decomposition of HI into $\mathrm{H_2}$ and $\mathrm{I_2}$. When the rates of the forward and reverse reactions have become equal to one another, the reaction has achieved a state of balance. **Chemical equilibrium** is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.

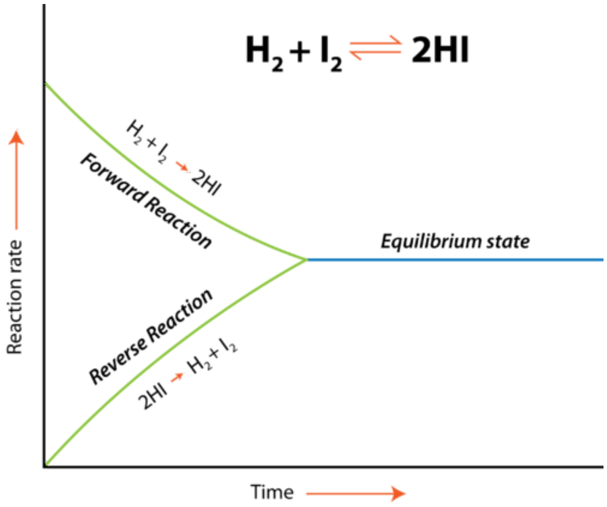

Figure 8.2.1: Equilibrium in reaction: $\mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightleftharpoons 2\mathrm{HI}\,(g)$.

Chemical equilibrium can be attained whether the reaction begins with all reactants and no products, all products and no reactants, or some of both. The figure below shows changes in concentration of $\mathrm{H_2}$, $\mathrm{I_2}$, and HI for two different reactions. In the reaction depicted by the graph on the left (A), the reaction begins with only $\mathrm{H_2}$ and $\mathrm{I_2}$ present. There is no HI initially. As the reaction proceeds towards equilibrium, the concentrations of the $\mathrm{H_2}$ and $\mathrm{I_2}$ gradually decrease, while the concentration of the HI gradually increases. When the curve levels out and the concentrations all become constant, equilibrium has been reached. At equilibrium, concentrations of all substances are constant.

In reaction B, the process begins with only HI and no $\mathrm{H_2}$ or $\mathrm{I_2}$. In this case, the concentration of HI gradually decreases while the concentrations of $\mathrm{H_2}$ and $\mathrm{I_2}$ gradually increase until equilibrium is again reached. Notice that in both cases, the relative position of equilibrium is the same, as shown by the relative concentrations of reactants and products. The concentration of HI at equilibrium is significantly higher than the concentrations of $\mathrm{H_2}$ and $\mathrm{I_2}$. This is true whether the reaction began with all reactants or all products. The position of equilibrium is a property of the particular reversible reaction and does not depend upon how equilibrium was achieved.

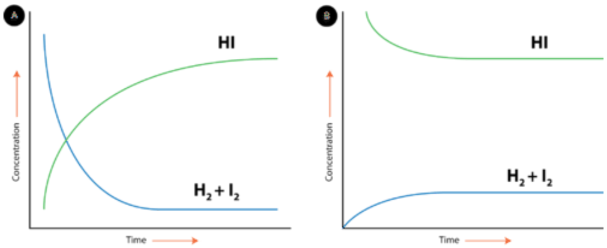

Figure 8.2.2: Equilibrium between reactants and products is achieved regardless of whether the reaction starts with the reactants or products.

Conditions for Equilibrium and Types of Equilibrium

It may be tempting to think that once equilibrium has been reached, the reaction stops. Chemical equilibrium is a dynamic process. The forward and reverse reactions continue to occur even after equilibrium has been reached. However, because the rates of the reactions are the same, there is no change in the relative concentrations of reactants and products for a reaction that is at equilibrium. The conditions and properties of a system at equilibrium are summarized below.

1. The system must be closed, meaning no substances can enter or leave the system.
2. Equilibrium is a dynamic process. Even though we don't necessarily see the reactions, both forward and reverse are taking place.
3. The rates of the forward and reverse reactions must be equal.
4. The amount of reactants and products do not have to be equal. However, after equilibrium is attained, the amounts of reactants and products will be constant.

The description of equilibrium in this concept refers primarily to equilibrium between reactants and products in a chemical reaction. Other types of equilibrium include phase equilibrium and solution equilibrium. A phase equilibrium occurs when a substance is in equilibrium between two states. For example, a stoppered flask of water attains equilibrium when the rate of evaporation is equal to the rate of condensation. A solution equilibrium occurs when a solid substance is in a saturated solution. At this point, the rate of dissolution is equal to the rate of recrystallization. Although these are all different types of transformations, most of the rules regarding equilibrium apply to any situation in which a process occurs reversibly.

Red blood cells transport oxygen to the tissues so they can function. In the absence of oxygen, cells cannot carry out their biochemical responsibilities. Oxygen moves to the cells attached to hemoglobin, a protein found in the red cells. In cases of carbon monoxide poisoning, CO binds much more strongly to the hemoglobin, blockin oxygen attachment and lowering the amount of oxygen reaching the cells. Treatment involves the patient breathing pure oxygen to displace the carbon monoxide. The equilibrium reaction shown below illustrates the shift toward the right when excess oxygen is added to the system:

$$\text{Hb(CO)}_4\,(aq) + 4\text{O}_2\,(g) \rightleftharpoons \text{Hb(O}_2)_4\,(aq) + 4\text{CO}\,(g) \tag{8.2.4}$$

Equilibrium Constant

Consider the hypothetical reversible reaction in which reactants A and B react to form products C and D. This equilibrium can be shown below, where the lowercase letters represent the coefficients of each substance.

$$a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D} \tag{8.2.5}$$

As we have established, the rates of the forward and reverse reactions are the same at equilibrium, and so the concentrations of all of the substances are constant. Since that is the case, it stands to reason that a ratio of the concentration for any given reaction at equilibrium maintains a constant value. The **equilibrium constant** $(\boldsymbol{K}_{\text{eq}})$ is the ratio of the mathematical product of the products of a reaction to the mathematical product of the concentrations of the reactants of the reaction. Each concentration is raised to the power of its coefficient in the balanced chemical equation. For the general reaction above, the equilibrium constant expression is written as follows:

$$K_{\text{eq}} = \frac{[\text{C}]^c[\text{D}]^d}{[\text{A}]^a[\text{B}]^b} \tag{8.2.6}$$

The concentrations of each substance, indicated by the square brackets around the formula, are measured in molarity units (mol/L).

The value of the equilibrium constant for any reaction is only determined by experiment. As detailed in the above section, the position of equilibrium for a given reaction does not depend on the starting concentrations and so the value of the equilibrium constant is truly constant. It does, however, depend on the temperature of the reaction. This is because equilibrium is defined as a condition resulting from the rates of forward and reverse reactions being equal. If the temperature changes, the corresponding change in those reaction rates will alter the equilibrium constant. For any reaction in which a K_{eq} is given, the temperature should be specified.

When K_{eq} is greater than 1, the numerator is larger than the denominator so the products are favored, meaning the concentration of its products are greater than that of the reactants.

If K_{eq} is less than 1, then the reactants are favored because the denominator (reactants) is larger than the numerator (products).

When K_{eq} is equal to 1, then the concentration of reactants and products are approximately equal.

Reaction Quotient

The reaction quotient, Q, is used when questioning if we are at equilibrium. The calculation for Q is **exactly** the same as for K but we can only use K when we know we are at equilibrium. Comparing Q and K allows the direction of the reaction to be predicted.

- $Q = K$ equilibrium
- $Q < K$ reaction proceeds to the right to form more products and decrease amount of reactants so value of Q will increase
- $Q > K$ reaction proceeds to the left to form more reactants and decrease amount of products so value of Q will decrease

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8.4: Osmosis and Diffusion

Learning Outcomes

- Define osmosis and diffusion.
- Distinguish among hypotonic, hypertonic, and isotonic solutions.
- Describe a semipermeable membrane.
- Predict behavior of blood cells in different solution types.
- Describe flow of solvent molecules across a membrane.
- Identify the polar and nonpolar regions of a cell membrane.
- Explain the components present in a phospholipid.

Fish cells, like all cells, have semipermeable membranes. Eventually, the concentration of "stuff" on either side of them will even out. A fish that lives in salt water will have somewhat salty water inside itself. Put it in freshwater, and the freshwater will, through osmosis, enter the fish, causing its cells to swell, and the fish will die. What will happen to a freshwater fish in the ocean?

Osmosis

Imagine you have a cup that has $100 \mathrm{~mL}$ water, and you add $15 \mathrm{~g}$ of table sugar to the water. The sugar dissolves and the mixture that is now in the cup is made up of a **solute** (the sugar) that is dissolved in the **solvent** (the water). The mixture of a solute in a solvent is called a **solution**.

Imagine now that you have a second cup with $100 \mathrm{~mL}$ of water, and you add $45 \mathrm{~g}$ of table sugar to the water. Just like the first cup, the sugar is the solute, and the water is the solvent. But now you have two mixtures of different solute concentrations. In comparing two solutions of unequal solute concentration, the solution with the higher solute concentration is **hypertonic**, and the solution with the lower solute concentration is **hypotonic**. Solutions of equal solute concentration are **isotonic**. The first sugar solution is hypotonic to the second solution. The second sugar solution is hypertonic to the first.

You now add the two solutions to a beaker that has been divided by a semipermeable membrane, with pores that are too small for the sugar molecules to pass through, but are big enough for the water molecules to pass through. The hypertonic solution is one one side of the membrane and the hypotonic solution on the other. The hypertonic solution has a lower water concentration than the hypotonic solution, so a concentration gradient of water now exists across the membrane. Water molecules will move from the side of higher water concentration to the side of lower concentration until both solutions are isotonic. At this point, **equilibrium** is reached.

Red blood cells behave the same way (see figure below). When red blood cells are in a hypertonic (higher concentration) solution, water flows out of the cell faster than it comes in. This results in *crenation* (shriveling) of the blood cell. On the other extreme, a red blood cell that is hypotonic (lower concentration outside the cell) will result in more water flowing into the cell than out. This results in swelling of the cell and potential *hemolysis* (bursting) of the cell. In an isotonic solution, the flow of water in and out of the cell is happening at the same rate.

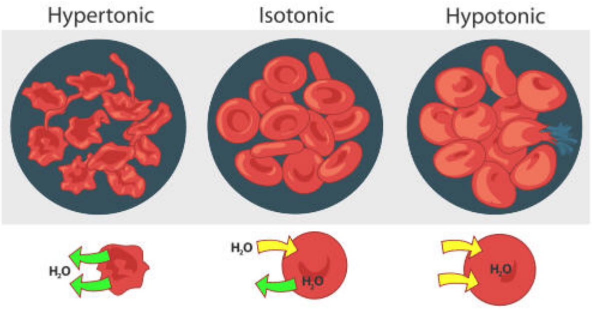

Figure 8.4.1: Red blood cells in hypertonic, isotonic, and hypotonic solutions.

Osmosis is the diffusion of water molecules across a semipermeable membrane from an area of lower concentration solution (i.e., higher concentration of water) to an area of higher concentration solution (i.e., lower concentration of water). Water moves into and out of cells by osmosis.

- If a cell is in a hypertonic solution, the solution has a lower water concentration than the cell cytosol, and water moves out of the cell until both solutions are isotonic.
- Cells placed in a hypotonic solution will take in water across their membranes until both the external solution and the cytosol are isotonic.

A red blood cell will swell and undergo hemolysis (burst) when placed in a hypotonic solution. When placed in a hypertonic solution, a red blood cell will lose water and undergo *crenation* (shrivel). Animal cells tend to do best in an isotonic environment, where the flow of water in and out of the cell is occurring at equal rates.

Diffusion

Passive transport is a way that small molecules or ions move across the cell membrane without input of energy by the cell. The three main kinds of passive transport are diffusion (or simple diffusion), osmosis, and facilitated diffusion. Simple diffusion and osmosis do not involve transport proteins. Facilitated diffusion requires the assistance of proteins.

Diffusion is the movement of molecules from an area of high concentration of the molecules to an area with a lower concentration. For cell transport, diffusion is the movement of small molecules across the cell membrane. The difference in the concentrations of the molecules in the two areas is called the **concentration gradient**. The kinetic energy of the molecules results in random motion, causing diffusion. In simple diffusion, this process proceeds without the aid of a transport protein. It is the random motion of the molecules that causes them to move from an area of high concentration to an area with a lower concentration.

Diffusion will continue until the concentration gradient has been eliminated. Since diffusion moves materials from an area of higher concentration to the lower, it is described as moving solutes "down the concentration gradient". The end result is an equal concentration, or **equilibrium**, of molecules on both sides of the membrane. At equilibrium, movement of molecules does not stop. At equilibrium, there is equal movement of materials in both directions.

Not everything can make it into your cells. Your cells have a plasma membrane that helps to guard your cells from unwanted intruders.

The Plasma Membrane and Cytosol

If the outside environment of a cell is water-based, and the inside of the cell is also mostly water, something has to make sure the cell stays intact in this environment. What would happen if a cell dissolved in water, like sugar does? Obviously, the cell could not survive in such an environment. So something must protect the cell and allow it to survive in its water-based environment. All cells have a barrier around them that separates them from the environment and from other cells. This barrier is called the **plasma membrane**, or cell membrane.

The Plasma Membrane

The plasma membrane (see figure below) is made of a double layer of special lipids, known as **phospholipids**. The phospholipid is a lipid molecule with a hydrophilic ("water-loving") head and two hydrophobic ("water-hating") tails. Because of the hydrophilic and hydrophobic nature of the phospholipid, the molecule must be arranged in a specific pattern as only certain parts of the molecule can physically be in contact with water. Remember that there is water outside the cell, and the **cytoplasm** inside the cell is mostly water as well. So the phospholipids are arranged in a double layer (a bilayer) to keep the cell separate from its environment. Lipids do not mix with water (recall that oil is a lipid), so the phospholipid bilayer of the cell membrane acts as a barrier, keeping water out of the cell, and keeping the cytoplasm inside the cell. The cell membrane allows the cell to stay structurally intact in its water-based environment.

The function of the plasma membrane is to control what goes in and out of the cell. Some molecules can go through the cell membrane to enter and leave the cell, but some cannot. The cell is therefore not completely permeable. "Permeable" means that anything can cross a barrier. An open door is completely permeable to anything that wants to enter or exit through the door. The plasma membrane is **semipermeable**, meaning that some things can enter the cell, and some things cannot.

Molecules that cannot easily pass through the bilayer include ions and small hydrophilic molecules, such as glucose, and macromolecules, including proteins and RNA. Examples of molecules that can easily diffuse across the plasma membrane include carbon dioxide and oxygen gas. These molecules diffuse freely in and out of the cell, along their concentration gradient. Though water is a polar molecule, it can also diffuse through the plasma membrane.

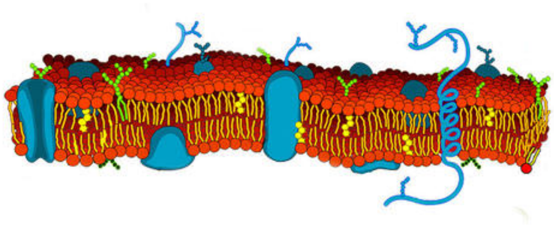

Figure 8.4.2: Plasma membranes are primarily made up of phospholipids (orange). The hydrophilic ("water-loving") head and two hydrophobic ("water-hating") tails are shown. The phospholipids form a bilayer (two layers). The middle of the bilayer is an area without water. There can be water on either side of the bilayer. There are many proteins throughout the membrane.

Cytosol

The inside of all cells also contain a jelly-like substance called **cytosol**. Cytosol is composed of water and other molecules, including **enzymes**, which are proteins that speed up the cell's chemical reactions. Everything in the cell sits in the cytosol, like fruit in a Jell-o mold. The term cytoplasm refers to the cytosol and all of the organelles, the specialized compartments of the cell. The cytoplasm does not include the nucleus. As a prokaryotic cell does not have a nucleus, the DNA is in the cytoplasm.

Supplemental Resources

- The Plasma Membrane: http://www.youtube.com/watch?v=moPJkCbKjBs

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8.5: Acid-Base Definitions

Learning Outcomes

- Describe the properties of acids and bases.
- Define an acid and a base according to the Arrhenius theory.
- Define an acid and a base according to the Brønsted-Lowry theory.
- Identify the conjugate acid-base pairs in a Brønsted-Lowry acid-base reaction.

Acid-Base Properties

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C (see figure below). Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Bases are less common as foods, but they are nonetheless present in many household products (see figure below). Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. Antacids, which combat excess stomach acid, are comprise of bases such as magnesium hydroxide or sodium hydrogen carbonate.

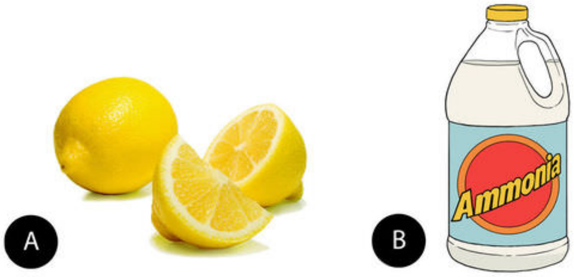


Figure 8.5.1: (A) Lemons and other citrus fruits contain citric and ascorbic acids. (B) Ammonia is a base that is present in many household cleaners.

Acids

Acids are a distinct class of compounds because of the properties of their aqueous solutions. Those properties are outlined below.

1. Aqueous solutions of acids are electrolytes, meaning that they conduct an electrical current. Some acids are strong electrolytes because they ionize completely in water. Other acids are weak electrolytes which partially ionize when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicators .Two common indicators are litmus and phenolphthalein. Litmus turns red in the presence of an acid, while phenolphthalein is colorless.
4. Acids react with some metals to yield hydrogen gas.
5. Acids react with bases to produce a salt and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. Water and an ionic compound called a salt are produced.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base (see figure below), while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.

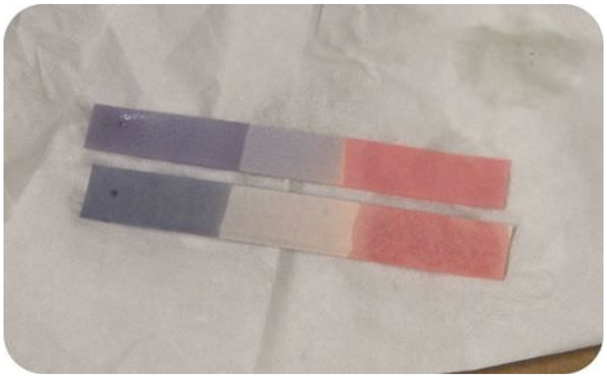


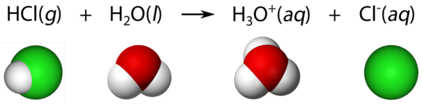
Figure 8.5.2: *Litmus paper has be treated with the plant dye called litmus. It turns red in the presence of an acid and blue in the presence of a base.*

Arrhenius Acids and Bases

The Swedish chemist Svante Arrhenius (1859 - 1927) was the first to propose a theory to explain the observed behavior of acids and bases. Because of their ability to conduct a current, he knew that both acids and bases contained ions in solution. An **Arrhenius acid** *is a compound which ionizes to yield hydrogen ions* $\left(\mathrm{H}^{+}\right)$ *in aqueous solution*. An **Arrhenius base** *is a compound which ionizes to yield hydroxide ions* $\left(\mathrm{OH}^{-}\right)$ *in aqueous solution*.

Arrhenius Acids

Acids are molecular compounds with ionizable hydrogen atoms. Only hydrogen atoms that are part of a highly polar covalent bond are ionizable. The hydrogen atom is attracted to the lone pair of electrons in a water molecule when HCl is dissolved in water. The result is that the $\mathrm{H-Cl}$ bond breaks, with both bonding electrons remaining with the Cl, forming a chloride ion. The H^{+} ion attaches to the water molecule, forming a polyatomic ion called the hydronium ion. The **hydronium ion** $\left(\mathrm{H_3O}^{+}\right)$ can be thought of as *a water molecule with an attached hydrogen ion*.



Equations showing the ionization of an acid in water are frequently simplified by omitting the water molecule.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq) \tag{8.5.1}$$

This is merely a simplification of the previous equation, but it is commonly used. Any hydrogen ions in an aqueous solution will be attached to water molecules as hydronium ions, $\mathrm{H_3O}^{+}$, even if it is written as H^{+}.

Not all hydrogen atoms in molecular compounds are ionizable. In methane $\left(\mathrm{CH_4}\right)$, the hydrogen atoms are covalently bonded to carbon in bonds that are only slightly polar. The hydrogen atoms are not capable of ionizing, and methane has no acidic properties. Acetic acid $\left(\mathrm{CH_3COOH}\right)$ (see figure below) belongs to a class of acids called carboxylic acids. There are four hydrogen atoms in the molecule, but only the one hydrogen that is bonded to an oxygen atom is ionizable.

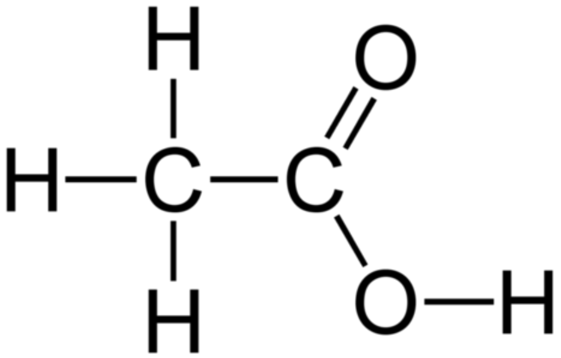


Figure 8.5.3: The $O-H$ bond can be ionized to yield the H^+ ion and the acetate ion. The other hydrogen atoms in this molecule are not acidic.

The table below lists some of the more common acids.

Table 8.5.1: Common Acids

Acid Name	Formula
hydrochloric acid	HCl
nitric acid	HNO_3
sulfuric acid	H_2SO_4
phosphoric acid	H_3PO_4
acetic acid	CH_3COOH
hypochlorous acid	$HClO$

A **monoprotic acid** *is an acid that contains only one ionizable hydrogen*. Hydrochloric acid and acetic acid are monoprotic acids. A **polyprotic acid** *is an acid that contains multiple ionizable hydrogens*. Most common polyprotic acids are either diprotic (such as H_2SO_4), or triprotic (such as H_3PO_4).

Arrhenius Bases

Bases are ionic compounds which yield the hydroxide ion $\left(OH^-\right)$ upon dissociating in water. The table below lists several of the more common bases.

Table 8.5.2: Common Bases

Base Name	Formula
Sodium hydroxide	$NaOH$
Potassium hydroxide	KOH
Magnesium hydroxide	$Mg(OH)_2$
Calcium hydroxide	$Ca(OH)_2$

All of the bases listed in the table are solids at room temperature. Upon dissolving in water, each dissociates into a metal cation and the hydroxide ion.

$$NaOH\,(s) \overset{H_2O}{\rightarrow} Na^+\,(aq) + OH^-\,(aq) \tag{8.5.2}$$

Sodium hydroxide is a very caustic substance also known as lye. Lye is used as a rigorous cleaner and is an ingredient in the manufacture of soaps. Care must be taken with strong bases like sodium hydroxide, as exposure can lead to severe burns (see figure below).

Figure 8.5.4: This foot has sever burns due to prolonged contact with a solution of sodium hydroxide, also known as lye.

Brønsted-Lowry Acids and Bases

The Arrhenius definition of acids and bases is somewhat limited. There are some compounds whose properties suggest that they are either acidic or basic, but which do not qualify according to the Arrhenius definition. An example is ammonia $\left(\mathrm{NH_3}\right)$. An aqueous solution of ammonia turns litmus blue, reacts with acids, and displays various other properties that are common for bases. However, it does not contain the hydroxide ion. In 1923, a broader definition of acids and bases was independently proposed by the Danish chemist Johannes Brønsted (1879 - 1947) and the English chemist Thomas Lowry (1874 - 1936). A **Brønsted-Lowry acid** *is a molecule or ion that donates a hydrogen ion in a reaction*. A **Brønsted-Lowry base** *is a molecule or ion that accepts a hydrogen ion in a reaction*. Because the most common isotope of hydrogen consists of a single proton and a single electron, a hydrogen ion (in which the single electron has been removed) is commonly referred to as a proton. As a result, acids and bases are often called proton donors and proton acceptors, respectively, according to the Brønsted-Lowry definition. All substances that are categorized as acids and bases under the Arrhenius definition are also define as such under the Brønsted-Lowry definition. The new definition, however, includes some substances that are left out according to the Arrhenius definition.

Brønsted-Lowry Acid-Base Reactions

An acid-base reaction according the Brønsted-Lowry definition is a transfer of a proton from one molecule or ion to another. When ammonia is dissolved in water, it undergoes the following reversible reaction.

$$\underset{\text{base}}{\mathrm{NH_3}\,(aq)} + \underset{\text{acid}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{NH_4^+}\,(aq)} + \underset{\text{base}}{\mathrm{OH^-}\,(aq)} \tag{8.5.3}$$

In this reaction, the water molecule is donating a proton to the ammonia molecule. The resulting products are the ammonium ion and the hydroxide ion. The water is acting as a Brønsted-Lowry acid, while the ammonia is acting as a Brønsted-Lowry base. The hydroxide ion that is produced causes the solution to be basic.

We can also consider the reverse reaction in the above equation. In that reaction, the ammonium ion donates a proton to the hydroxide ion. The ammonium ion is a Brønsted-Lowry acid, while the hydroxide ion is a Brønsted-Lowry base. Most Brønsted-Lowry acid-base reactions can be analyzed in this way. There is one acid and one base as reactants, and one acid and one base as products.

In the above reaction, water acted as an acid, which may seem a bit unexpected. Water can also act as a base in a Brønsted-Lowry acid-base reaction, as long as it reacts with a substance that is a better proton donor. Shown below is the reaction of water with the hydrogen sulfate ion.

$$\underset{\text{acid}}{\mathrm{HSO_4^-}\,(aq)} + \underset{\text{base}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{H_3O^+}\,(aq)} + \underset{\text{base}}{\mathrm{SO_4^{2-}}\,(aq)} \tag{8.5.4}$$

Water is capable of being either an acid or a base, a characteristic called amphoterism. An **amphoteric substance** *is one that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.*

Conjugate Acids and Bases

When a substance that is acting as a Brønsted-Lowry acid donates its proton, it becomes a base in the reverse reaction. In the reaction above, the hydrogen sulfate ion $\left(HSO_4^-\right)$ donates a proton to water and becomes a sulfate ion $\left(SO_4^{2-}\right)$. The HSO_4^- and the SO_4^{2-} are linked to one another by the presence or absence of the H^+ ion. A **conjugate acid-base pair** *is a pair of substances related by the loss or gain of a single hydrogen ion*. A **conjugate acid** *is the particle produced when a base accepts a proton*. The hydrogen sulfate ion is the conjugate acid of the sulfate ion. A **conjugate base** *is the particle produced when an acid donates a proton*. The sulfate ion is the conjugate base of the hydrogen sulfate ion.

A typical Brønsted-Lowry acid-base reaction contains two conjugate acid-base pairs as shown below.

$$HNO_2\,(aq)\,PO_4^{3-}\,(aq) \rightleftharpoons NO_2^-\,(aq) + HPO_4^{2-}\,(aq) \tag{8.5.5}$$

One conjugate acid-base pair is HNO_2/NO_2^-, while the other pair is HPO_4^{2-}/PO_4^{3-}.

The buffer systems in blood have conjugate acid-base pairs which help maintain the correct acid and base concentration in blood.

Table 8.5.3: Acid-Base Definitions

Type	Acid	Base
Arrhenius	H^+ ions in solution	OH^- ions in solution
Brønsted-Lowry	H^+ donor	H^+ acceptor

Supplemental Resources

- Acid-Base Theories: http://www.kentchemistry.com/links/A...seTheories.htm
- Theories of Acids and Bases: http://www.chemguide.co.uk/physical/.../theories.html

Contributors and Attributions

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8.6: The pH Concept

Learning Outcomes

- Calculate between pH/$\left[\mathrm{H^+}\right]$ and pOH/$\left[\mathrm{OH^-}\right]$
- Convert among pH, pOH, hydrogen-ion concentration, and hydroxide-ion concentration for a given solution.

Self-Ionization of Water

Water is a molecular compound, so you may not necessarily expect it to break apart into ions. However, sensitive experiments show that water is actually a very weak electrolyte. When two molecules of water collide, there can be a transfer of a hydrogen ion from one molecule to the other. The products are a positively charged hydronium ion and a negatively charged hydroxide ion.

$$\mathrm{H_2O}\left(l\right)+\mathrm{H_2O}\left(l\right)\rightleftharpoons \mathrm{H_3O^+}\left(aq\right)+\mathrm{OH^-}\left(aq\right) \tag{8.6.1}$$

The **self-ionization** *of water is the process in which water ionizes to hydronium ions and hydroxide ions*. As with other aqueous acid-base reactions, the process is often simplified to show the ionization of just one water molecule into a hydrogen ion and a hydroxide ion.

$$\mathrm{H_2O}\left(l\right)\rightleftharpoons \mathrm{H^+}\left(aq\right)+\mathrm{OH^-}\left(aq\right) \tag{8.6.2}$$

Either equation is adequate, though the first is more accurate, since hydrogen ions in aqueous solution will always be attached to water molecules. Further discussion of acids and acid ionizations in this book will primarily show hydrogen ions in aqueous solution as $\mathrm{H^+}$, but keep in mind that this is just a commonly used abbreviation for the more accurate hydronium $\left(\mathrm{H_3O^+}\right)$ structure.

In pure water, the concentrations of hydrogen and hydroxide ions are equal to one another. Any aqueous solution in which $\left[\mathrm{H^+}\right]=\left[\mathrm{OH^-}\right]$ is said to be neutral.

For any neutral solution at $25^\mathrm{o}\mathrm{C}$, each of these ions has a concentration of $1.0\times 10^{-7}\ \mathrm{M}$.

An **acidic solution** *is a solution in which the concentration of hydrogen ions is greater than the concentration of hydroxide ions*. For example, hydrogen chloride ionizes to produce $\mathrm{H^+}$ and $\mathrm{Cl^-}$ ions upon dissolving in water.

$$\mathrm{HCl}\left(g\right)\rightarrow \mathrm{H^+}\left(aq\right)+\mathrm{Cl^-}\left(aq\right) \tag{8.6.3}$$

This increases the concentration of $\mathrm{H^+}$ ions in the solution.

A **basic solution** *is a solution in which the concentration of hydroxide ions is greater than the concentration of hydrogen ions*. Solid potassium hydroxide dissociates in water to yield potassium ions and hydroxide ions.

$$\mathrm{KOH}\left(s\right)\rightarrow \mathrm{K^+}\left(aq\right)+\mathrm{OH^-}\left(aq\right) \tag{8.6.4}$$

The pH Scale

Expressing the acidity of a solution by using the molarity of the hydrogen ion is cumbersome because the quantities are generally very small. Danish scientist Søren Sørensen (1868 - 1939) proposed an easier system for indicating the concentration of $\mathrm{H^+}$ called the pH scale. The letters pH stand for the power of the hydrogen ion. The **pH** *of a solution is the negative logarithm of the hydrogen-ion concentration*.

$$\mathrm{pH}=-\log\left[\mathrm{H^+}\right] \tag{8.6.5}$$

In pure water or a neutral solution $\left[\mathrm{H^+}\right]=1.0\times 10^{-7}\ \mathrm{M}$. Substituting this value into the pH expression:

$$\mathrm{pH}=-\log\left[1.0\times 10^{-7}\right]=-\left(-7.00\right)=7.00 \tag{8.6.6}$$

The pH of pure water or any neutral solution is thus 7.00. Due to the somewhat less intuitive rules for dealing with significant figures in the context of logarithms, only the numbers ot the right of the decimal point in the pH value are the significant figures. Since 1.0×10^{-7} has two significant figures, the pH can be reported as 7.00.

A logarithmic scale condenses the range of acidity to numbers that are easy to use. For example, a solution in which $\left[\mathrm{H^+}\right]=1.0\times 10^{-4}\ \mathrm{M}$ has a hydrogen-ion concentration that is 1000 times higher than in pure water. The pH of such a solution is

4.00, a difference of 3 pH units. Notice that when $\left[\mathrm{H}^{+}\right]$ is written in scientific notation and the coefficient is 1, the pH is simply the exponent with the sign changed. The pH of a solution in which $\left[\mathrm{H}^{+}\right]=1.0\times10^{-2}\ \mathrm{M}$ is 2.0 and the pH of a solution in which $\left[\mathrm{H}^{+}\right]=1.0\times10^{-10}\ \mathrm{M}$ is 10.0. If the coefficient is not equal to 1, a calculator must be used to find the pH. For example, the pH of a solution in which $\left[\mathrm{H}^{+}\right]=2.3\times10^{-5}\ \mathrm{M}$ can be found as shown below.

$$\mathrm{pH}=-\log\left[2.3\times10^{-5}\right]=4.64 \tag{8.6.7}$$

As we saw earlier, a solution in which $\left[\mathrm{H}^{+}\right]$ is higher than $1\times10^{-7}\ \mathrm{M}$ acidic, while a solution in which $\left[\mathrm{H}^{+}\right]$ is lower than $1.0\times10^{-7}\ \mathrm{M}$ is basic. Consequently, solutions with pH values of less than 7 are acidic, while solutions with pH values higher than 7 are basic. The figure below illustrates this relationship, along with some examples of the pH for various solutions.

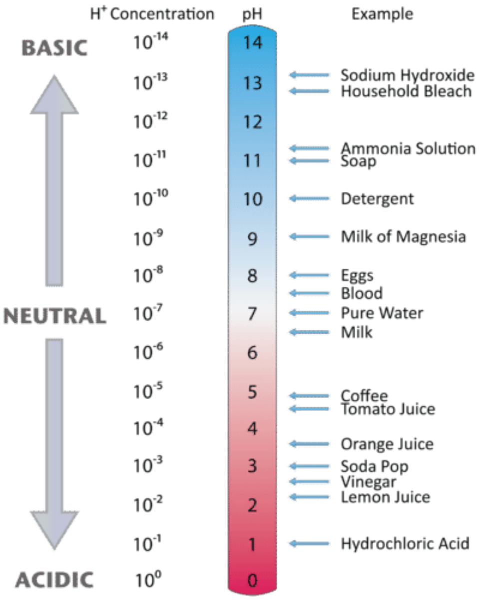

Figure 8.6.1: The pH scale is a logarithmic scale based on the concentration of hydrogen ions. The higher the H^{+} ion concentration is, the lower the pH of the solution.

The pH scale is generally presented as running from 0 to 14, though it is possible to have a pH of less than 0 or greater than 14. For example, a highly concentrated $3.0\ \mathrm{M}$ of HCl has a negative pH.

$$\mathrm{pH}=-\log(3.0)=-0.48 \tag{8.6.8}$$

When the pH of a solution is known, the concentration of the hydrogen ion can be calculated. The inverse of the logarithm (or antilog) is the 10^{x} key on a calculator.

$$\left[\mathrm{H}^{+}\right]=10^{-\mathrm{pH}} \tag{8.6.9}$$

For example, suppose that you have a solution with a pH of 9.14. $\left[\mathrm{H}^{+}\right]$ c an be found as follows:

$$\left[\mathrm{H}^{+}\right]=10^{-\mathrm{pH}}=10^{-9.14}=7.24\times10^{-10}\ \mathrm{M} \tag{8.6.10}$$

The pOH Concept

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** *of a solution is the negative logarithm of the hydroxide-ion concentration*.

$$\mathrm{pOH}=-\log\left[\mathrm{OH}^{-}\right] \tag{8.6.11}$$

The pH of a solution can be related to the pOH. Consider a solution with pH $=4.0$. The pOH of the solution would be 10. This example illustrates the following relationship. This is a result of the concentration of the H^{+} and OH^{-} ions present in pure water which we will not discuss further.

$$\mathrm{pH}+\mathrm{pOH}=14 \tag{8.6.12}$$

The pOH scale is similar to the pH scale, in that a pOH of 7 is indicative of a neutral solution. A basic solution has a pOH of less than 7, while an acidic solution has a pOH of greater than 7. The pOH is convenient to use when finding the hydroxide ion concentration from a solution with a known pH.

Example 8.6.1

Find the hydroxide concentration of a solution with a pH of 4.42.

Solution

Step 1: List the known values and plan the problem.

Known

- $\mathrm{pH} = 4.42$
- $\mathrm{pH} + \mathrm{pOH} = 14$

Unknown

- $\left[\mathrm{OH}^-\right]$

First, the pOH is calculated, followed by the $\left[\mathrm{OH}^-\right]$

Step 2: Solve.

$$\mathrm{pOH} = 14 - \mathrm{pH} = 14 - 4.42 = 9.58 \tag{8.6.13}$$

$$\left[\mathrm{OH}^-\right] = 10^{-\mathrm{pOH}} = 10^{-9.58} = 2.6 \times 10^{-10}\ \mathrm{M} \tag{8.6.14}$$

Step 3: Think about your result.

The pH is that of an acidic solution, and the resulting hydroxide-ion concentration is less than $1 \times 10^{-7}\ \mathrm{M}$. The answer has two significant figures because the given pH has two decimal places.

The diagram below shows all of the interrelationships between $\left[\mathrm{H}^+\right]$, $\left[\mathrm{OH}^-\right]$, pH, and pOH.

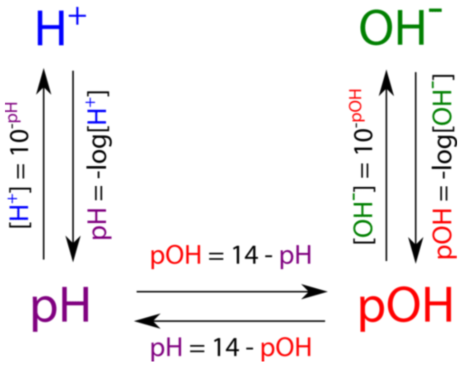

Figure 8.6.2: The flowchart shows the way to convert between hydrogen ion concentration, hydroxide ion concentration, pH, and pOH.

Contributors and Attributions

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8.7: Properties of Solutions (Exercises)

These are homework exercises to accompany Chapter 8 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Answers are below the questions.

Questions

8.1: Concentrations of Solutions

(click here for solutions)

Q8.1.1

How is a concentrated solution different from a dilute solution?

Q8.1.2

What is the molarity of a solution prepared with 0.75 moles NaCl in 250. mL solution?

Q8.1.3

What is the mass percent of an aqueous solution prepared by dissolving 12.0 g of solute into 40.0 g of water?

Q8.1.4

What is the volume percent of a solution prepared by adding enough water to 200. mL of acetone to make a total volume of 1.60 L?

Q8.1.5

What mass of glucose is in 250.0 mL of solution that is 5.00% by mass? Assume the density of the solution is 1.00 g/mL.

Q8.1.6

For a solution that contains 25.0 g of NaCl in 300.0 mL of water, find each of the following. The density of the water is 1.00 g/mL. Assume the NaCl does not contribute to the **volume** of the solution

a. mass percent
b. mass/volume percent

Q8.1.7

For a solution that contains 15.0 mL of methanol 125 mL of ethanol, find each of the following. The density of methanol is 0.792 g/mL and the density of ethanol is 0.789 g/mL.

a. mass percent
b. mass/volume percent
c. volume percent

Q8.1.8

A saline solution has a mass percent concentration of 10.5%. What mass of NaCl is present in 150.0 mL of the solution? Assume the density of the solution is 1.00 g/mL.

Q8.1.9

Calculate the molarity for each solution.

a. 87.2 g of Na_2SO_4 in enough water to make 500. mL of solution
b. 61.8 g of NH_3 in enough water to make 7.00 L of solution
c. 100. mL of ethanol (C_2H_5OH) in 500. mL of solution (The density of ethanol is 0.789 g/mL.)

Q8.1.10

How many moles of KF are contained in 180.0 mL of a 0.250 M solution?

Q8.1.11

Calculate how many grams of each solute would be required in order to make the given solution.

a. 3.40 L of a 0.780 M solution of iron(III) chloride, $FeCl_3$

b. 60.0 mL of a 4.10 M solution of calcium acetate, $Ca(CH_3COO)_2$

Q8.1.12

What volume of a 0.500 M solution of NaI could be prepared with 113 g of solid NaI?

Q8.1.13

Calculate the molarity of the solutions prepared from the following dilutions.

a. 125 mL of 2.00 M HCl is diluted to a volume of 4.00 L.
b. 1.85 mL of 6.30 M $AgNO_3$ is diluted to a volume of 5.00 mL.

Q8.1.14

What volume of 12 M HCl is required to prepare 6.00 L of a 0.300 M solution?

Q8.1.15

What mass of lead is present in 50.0 mL of solution with a lead concentration of 12 ppm?

Q8.1.16

What mass of mercury is present in 175 mL of solution with a mercury concentration of 25 ppb?

Q8.1.17

What is the concentration, in units of ppm, for a solution that contains 34 g of iron in 365 mL of water?

Q8.1.18

How many equivalents are there in 2.0 moles of the ion of each element below?

a. magnesium
b. aluminum
c. sulfur
d. bromine (Br)
e. cesium (Cs)
f. barium (Ba)

Q8.1.19

How many equivalents are present in 2.50 moles of ions for each of the elements in the previous question?

Q8.1.20

How many moles of Ca^{2+} are given to a patient if they receive 250.0 mL of a solution with a concentration of 132 mEq/L?

Q8.1.21

How many grams of K^+ are given to a patient if they receive 500.0 mL of a solution with a concentration of 98 mEq/L?

Q8.1.22

A solution contains 128 mEq/L of Sr^{2+}. What volume of solution is needed to have a total mass of 3.93 g of strontium ions?

8.2: Chemical Equilibrium

(click here for solutions)

Q8.2.1

What is chemical equilibrium?

Q8.2.2

If the reaction $H_2 + I_2 \rightleftharpoons 2HI$ is at equilibrium, do the concentrations of HI, H_2, and I_2 have to be equal?

Q8.2.3

Do the concentrations at equilibrium depend upon how the equilibrium was reached?

Q8.2.4

What does the equilibrium constant tell us?

Q8.2.5

What does it mean if the K_{eq} is > 1?

Q8.2.6

What does it mean if the K_{eq} is < 1?

Q8.2.7

Does the equilibrium state depend on the starting concentrations?

8.3: Le Chatelier's Principle

(click here for solutions)

Q8.3.1

Define Le Chatelier's principle.

Q8.3.2

List the three factors types of changes that can disturb the equilibrium of a system.

Q8.3.3

How will each change affect the reaction?

$PCl_5(g) + \text{heat} \rightleftharpoons PCl_3(g) + Cl_2(g)$

a. Addition of PCl_5
b. Addition of Cl_2
c. Removal of PCl_3
d. Increasing temperature
e. Decreasing temperature
f. Decreasing volume

Q8.3.4

How will each change affect the reaction?

$HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$

a. Removal of HNO_2
b. Addition of HCl (i.e. adding more H^+)
c. Increasing volume
d. Decreasing volume
e. Removal of NO_2^-
f. Addition of OH^- (which will react with and remove H^+)

Q8.3.5

How will each change affect the reaction?

$CO_2(g) + C(s) \rightleftharpoons 2CO(g)\ \Delta H = 172.5\ kJ$

a. Addition of CO_2
b. Removal of CO_2
c. Increasing temperature
d. Decreasing temperature
e. Increasing volume
f. Addition of CO

Q8.3.6

How will each change affect the reaction?

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)\ \Delta H = -9.48\ kJ$

a. Addition of H_2
b. Removal of H_2
c. Increasing temperature
d. Decreasing temperature
e. Increasing volume
f. Decreasing volume

8.4: Osmosis and Diffusion

(click here for solutions)

Q8.4.1

What are some of the features of a semipermeable membrane?

Q8.4.2

Two solutions are separated by a semipermeable membrane. Solution A contains 25.0 g of NaCl in 100.0 mL of water and solution B contains 35.0 g of NaCl in 100.0 mL of water.

a. Which one has a higher concentration?
b. Which way will water molecules flow?
c. Which volume will increase?
d. Which volume will decrease?
e. What will happen to the concentration of solution A?
f. What will happen to the concentration of solution B?

Q8.4.3

What do the prefixes hyper, hypo, and iso mean?

Q8.4.4

Cells are placed in a solution and the cells then undergo hemolysis. What can be said about the relative concentrations of solute in the cell and the solution?

Q8.4.5

Describe the relative concentrations inside and outside a red blood cell when crenation occurs.

Q8.4.6

A saltwater fish is placed in a freshwater tank. What will happen to the fish? Describe the flow of water molecules to explain the outcome.

Q8.4.7

What makes up the "head" region of a phospholipid? Is it hydrophobic or hyrdrophillic?

Q8.4.8

What makes up the "tail" region of a phospholipid? Is it hydrophobic or hyrdrophillic?

8.5: Acid-Base Definitions

(click here for solutions)

Q8.5.1

Which statement below is true? Explain.

a. All Arrhenius bases are also Brønsted-Lowry bases.
b. All Brønsted-Lowry bases are also Arrhenius bases.

Q8.5.2

Classify each of the following as an acid, base, or neither.

a. LiOH
b. $HClO_4$
c. CH_3COOH
d. $Sr(OH)_2$
e. CH_4
f. CH_4OH

Q8.5.3

What does it mean to say that a substance is amphoteric?

Q8.5.4

Identify each reactant in the following reactions as an acid or a base according to the Brønsted-Lowry theory.

a. $HIO_3(aq) + H_2O(l) \rightleftharpoons IO_3^-(aq) + H_3O^+(aq)$
b. $F^-(aq) + HClO(aq) \rightleftharpoons HF(aq) + ClO^-(aq)$
c. $H_2PO_4^-(aq) + OH^-(aq) \rightleftharpoons HPO_4^{2-}(aq) + H_2O(l)$
d. $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

Q8.5.5

Referring to question 4, identify the conjugate acid-base pairs in each reaction.

Q8.5.6

Write the formula of each acid’s conjugate base.

a. HNO_3
b. HSO_3^-
c. H_3AsO_4
d. HCOOH
e. HPO_4^{2-}
f. H_2S
g. HS^-
h. HCO_3^-
i. H_2CO_3
j. H_3PO_4
k. $NaHSO_4$

Q8.5.7

Write the formula of each base’s conjugate acid.

a. BrO_3^-
b. NH_3
c. CH_3COO^-
d. HCO_3^-
e. CN^-
f. HPO_4^{2-}
g. HS^-
h. SO_4^{2-}
i. CO_3^{2-}
j. HCO_3^-
k. PH_3

Q8.5.8

Explain why the hydrogen phosphate ion (HPO_4^{2-}) is amphoteric.

8.6: The pH Concept

(click here for solutions)

Q8.6.1

Describe the process by which water self-ionizes, and explain why pure water is considered to be neutral.

Q8.6.2

Indicate whether solutions with the following pH values are acidic, basic, or neutral.

a. pH = 9.4
b. pH = 7.0
c. pH = 5.0

Q8.6.3

How can the pOH of a solution be determined if its pH is known? (Hint: Write a mathematical expression.)

Q8.6.4

Find pH and pOH of each solution.

a. $[H^+] = 2.3 \times 10^{-4}$ M
b. $[H^+] = 8.7 \times 10^{-10}$ M
c. $[OH^-] = 1.9 \times 10^{-9}$ M
d. $[OH^-] = 0.60$ M

Q8.6.5

Find pH and pOH of each solution.

a. $[H^+] = 1.0 \times 10^{-5}$ M
b. $[H^+] = 2.8 \times 10^{-11}$ M
c. $[OH^-] = 1.0 \times 10^{-2}$ M
d. $[OH^-] = 4.4 \times 10^{-9}$ M

Q8.6.6

Determine $[H^+]$ and $[OH^-]$ in aqueous solutions with the following pH or pOH values.

a. pH = 1.87
b. pH = 11.15
c. pH = 0.95
d. pOH = 6.21
e. pOH = 13.42
f. pOH = 7.03

Q8.6.7

You have prepared 1.00 L of a solution with a pH of 5.00. What is the pH of the solution if 0.100 L of additional water is added to it? (Hint: Calculate the moles of H^+ ions present in the solution.)

Q8.6.8

How much water would need to be added to the original solution in question 8 in order to bring the pH to 6.00?

Answers

8.1: Concentrations of Solutions

Q8.1.1

Concentrated solutions have more solute per unit of solvent or solution.

Q8.1.2

$$M=\frac{mol\ solute}{L\ soln}$$
$$M=\frac{0.75\ mol}{0.250\ L}$$
$$M=3.0\ M$$

Q8.1.3

$$mass\ \%=\frac{g\ solute}{g\ soln}\times 100$$
$$mass\ \%=\frac{12.0\ g}{40.0\ g+12.0\ g}\times 100$$
$$mass\ \%=23.1\ \%$$

Remember, the mass of the solution includes both the solute and solvent.

Q8.1.4

$$volume\ \%=\frac{L\ solute}{L\ soln}\times 100$$
$$volume\ \%=\frac{0.200\ L}{1.60\ L}\times 100$$
$$volume\ \%=12.5\ \%$$

Volumes can also be used in mL (or any other unit) as long as both volumes are in the same unit.

Q8.1.5

Write the concentration in "expanded form" which shows the relationship to then be used in dimensional analysis.

$$5.00\ \%\ m/m=\frac{5.00\ g\ glucose}{100\ g\ solution}$$

$$250.0\ mL\ soln\left(\frac{1.00\ g\ soln}{mL\ soln}\right)\left(\frac{5.00\ g\ glucose}{100\ g\ soln}\right)=12.5\ g\ glucose$$

Q8.1.6

a. $\%\ m/m=\frac{mass\ solute}{mass\ solution}\times 100$
   $\%\ m/m=\frac{25.0\ g}{25.0\ g+300\ g}\times 100$
   $\%\ m/m=7.69\ \%$
b. The NaCl does not contribute to the volume of the solution so only the volume of the water is used for the volume fo the solution. Given the density is 1.00 g/mL, the volume of the solution is 300.0 mL.

   $\%\ m/v=\frac{mass\ solute}{volume\ solution}\times 100$
   $\%\ m/v=\frac{25.0\ g}{25.0\ mL+300\ mL}\times 100$
   $\%\ m/v=8.33\ \%$

Q8.1.7

The parts of this problem require both the volume and mass of solute and solvent. The volume of the solute and solvent are given so first, find the mass of the solute and solvent so all the values are present before we start calculating the concentrations.

$$15.0\ mL\ \text{methanol}\left(\frac{0.792\ g}{mL\ \text{methanol}}\right)=11.9\ g\ \text{methanol}$$

$$125.0\ mL\ \text{ethanol}\left(\frac{0.789\ g}{mL\ \text{ethanol}}\right)=98.6\ g\ \text{ethanol}$$

a. $\%\ m/m=\frac{mass\ solute}{mass\ solution}\times 100$
   $\%\ m/m=\frac{11.9\ g\ \text{methanol}}{11.9\ g+98.6\ g}\times 100$
   $\%\ m/m=10.8\ \%$
b. $\%\ m/v=\frac{mass\ solute}{volume\ solution}\times 100$
   $\%\ m/v=\frac{11.9\ g\ \text{methanol}}{15.0\ mL+125\ mL}\times 100$
   $\%\ m/v=8.50\ \%$
c. $\%\ v/v=\frac{volume\ solute}{volume\ solution}\times 100$
   $\%\ v/v=\frac{15.0\ mL\ \text{methanol}}{15.0\ mL+125\ mL}\times 100$
   $\%\ v/v=10.7$

Q8.1.8

Write the concentration in "expanded form" which shows the relationship to then be used in dimensional analysis.

$$10.5\ \%\ m/m = \frac{10.5\ g\ \mathrm{NaCl}}{100\ g\ soln}$$

$$150.0\ mL\ soln \left(\frac{1.00\ g\ soln}{mL\ soln}\right)\left(\frac{10.5\ g\ \mathrm{NaCl}}{100\ g\ soln}\right) = 15.8\ g\ \mathrm{NaCl}$$

Q8.1.9

a. $M = \frac{0.614\ mol\ \mathrm{Na_2SO_4}}{0.500\ L\ soln} = 1.23\ M$
b. $M = \frac{3.63\ mol\ \mathrm{NH_3}}{7.00\ L\ soln} = 0.519\ M$
c. $M = \frac{1.71\ mol\ \mathrm{EtOH}}{0.500\ L\ soln} = 3.43\ M$

Q8.1.10

$$0.250\ M = \frac{0.250\ mol\ \mathrm{KF}}{1\ L\ soln}$$

$$180.0\ mL \left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.250\ mol\ \mathrm{KF}}{1\ L\ soln}\right) = 0.0450\ mol\ \mathrm{KF}$$

Q8.1.11

a. $3.40\ L \left(\frac{0.780\ mol}{1\ L}\right)\left(\frac{162.2\ g}{mol}\right) = 430.\ g\ \mathrm{FeCl_3}$
b. $60.0\ mL \left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{4.10\ mol}{1\ L}\right)\left(\frac{158.17\ g}{mol}\right) = 38.9\ g\ \mathrm{Ca(CH_3COO)_2}$

Q8.1.12

$$0.500\ M = \frac{0.500\ mol\ \mathrm{NaI}}{1\ L\ soln}$$

$$113\ g\ \mathrm{NaI} \left(\frac{1\ mol}{149.89\ g}\right)\left(\frac{1\ L}{0.500\ mol}\right) = 1.51\ L\ soln$$

Q8.1.13

a. $C_1V_1 = C_2V_2$
   $2.00\ M \cdot 0.125\ L = C_2 \cdot 4.00\ L$
   $C_2 = 0.0625\ M$
b. $C_1V_1 = C_2V_2$
   $6.30\ M \cdot 1.85\ mL = C_2 \cdot 5.00\ mL$
   $C_2 = 2.33\ M$

Q8.1.14

$C_1V_1 = C_2V_2$
$0.300\ M \cdot 6.00\ L = 12\ M\backslash\mathrm{cdotV}_2$
$V_2 = 0.15\ L$

Q8.1.15

$$12\ ppm\ \mathrm{Pb} = \frac{12\ mg\ \mathrm{Pb}}{1\ L\ soln}$$

$$50.0\ mL \left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{12\ mg}{1\ L}\right) = 0.60\ mg\ \mathrm{Pb}$$

Q8.1.16

$$25\ ppb\ \mathrm{Hg} = \frac{25\ \mu g\ \mathrm{Hg}}{1\ L\ soln}$$

$$175\ mL \left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{25\ \mu g}{1\ L}\right) = 4.4\ \mu g\ \mathrm{Hg}$$

Q8.1.17

$$ppm = \frac{mg}{L}$$

$$34\ g\ \mathrm{Fe} \left(\frac{1\ mg}{10^{-3}\ g}\right) = 3.4 \times 10^4\ mg\ \mathrm{Fe}$$

$$365\ mL\left(\frac{10^{-3}\ L}{1\ mL}\right)=0.365\ L$$

$$ppm=\frac{3.4\times 10^{4}\ mg\ \mathrm{Fe}}{0.365\ L}=9.3\times 10^{4}\ ppm\ \mathrm{Fe}$$

Q8.1.18

a. $2.0\ mol\ \mathrm{Mg}^{2+}\left(\frac{2\ Eq}{1\ mol}\right)=4.0\ Eq$
b. $2.0\ mol\ \mathrm{Al}^{3+}\left(\frac{3\ Eq}{1\ mol}\right)=6.0\ Eq$
c. $2.0\ mol\ \mathrm{S}^{2-}\left(\frac{2\ Eq}{1\ mol}\right)=4.0\ Eq$
d. $2.0\ mol\ \mathrm{Br}^{-}\left(\frac{1\ Eq}{1\ mol}\right)=2.0\ Eq$
e. $2.0\ mol\ \mathrm{Cs}^{+}\left(\frac{1\ Eq}{1\ mol}\right)=2.0\ Eq$
f. $2.0\ mol\ \mathrm{Ba}^{2+}\left(\frac{2\ Eq}{1\ mol}\right)=4.0\ Eq$

Q8.1.19

a. $2.50\ mol\ \mathrm{Mg}^{2+}\left(\frac{2\ Eq}{1\ mol}\right)=5.00\ Eq$
b. $2.50\ mol\ \mathrm{Al}^{3+}\left(\frac{3\ Eq}{1\ mol}\right)=7.50\ Eq$
c. $2.50\ mol\ \mathrm{S}^{2-}\left(\frac{2\ Eq}{1\ mol}\right)=5.00\ Eq$
d. $2.50\ mol\ \mathrm{Br}^{-}\left(\frac{1\ Eq}{1\ mol}\right)=2.50\ Eq$
e. $2.50\ mol\ \mathrm{Cs}^{+}\left(\frac{1\ Eq}{1\ mol}\right)=2.50\ Eq$
f. $2.50\ mol\ \mathrm{Ba}^{2+}\left(\frac{2\ Eq}{1\ mol}\right)=5.00\ Eq$

Q8.1.20

$$250.0\ mL\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{132\ mEq}{L}\right)\left(\frac{10^{-3}\ Eq}{1\ mEq}\right)\left(\frac{1\ mol\ \mathrm{Ca}^{2+}}{2\ Eq}\right)=0.0165\ mol\ \mathrm{Ca}^{2+}$$

Q8.1.21

$$500.0\ mL\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{98\ mEq}{L}\right)\left(\frac{10^{-3}\ Eq}{1\ mEq}\right)\left(\frac{1\ mol\ \mathrm{K}^{+}}{1\ Eq}\right)\left(\frac{39.10\ g}{mol}\right)=1.92\ g\ \mathrm{K}^{+}$$

Q8.1.22

$$3.93\ g\ \mathrm{Sr}^{2+}\left(\frac{1\ mol}{87.62\ g}\right)\left(\frac{2\ Eq}{1\ mol}\right)\left(\frac{1\ mEq}{10^{-3}\ Eq}\right)\left(\frac{1\ L}{128\ mEq}\right)=0.701\ L\ soln$$

8.2: Chemical Equilibrium

Q8.2.1

The rate of the forward reaction equals the rate of the reverse reaction.

Q8.2.2

No, the concentrations are constant but the concentrations do not have to be equal.

Q8.2.3

No.

Q8.2.4

The ratio of products and reactants when the system is at equilibrium.

Q8.2.5

More products than reactants are present at equilibrium.

Q8.2.6

More reactants than products present at equilibrium.

Q8.2.7

No. The equilibrium ratio does not depend on the initial concentrations.

8.3: Le Chatelier's Principle

Q8.3.1

Le Chatelier’s principle states that a system at equilibrium is disturbed, it will respond in a way to minimize te disturbance.

Q8.3.2

temperature, change in amount of substance, change in pressure through change in volume

Q8.3.3

a. shift right
b. shift left
c. shift right
d. shift right
e. shift left
f. shift left

Q8.3.4

a. shift left
b. shift left
c. no effect
d. no effect
e. shift right
f. shift right

Q8.3.5

a. shift right
b. shift left
c. shift right
d. shift left
e. shift right
f. shift left

Q8.3.6

a. shift right
b. shift left
c. shift left
d. shift right
e. no effect
f. no effect

8.4: Osmosis and Diffusion

Q8.4.1

A semipermeable membrane allows some substances to pass through but not others.

Q8.4.2

Two solutions are separated by a semipermeable membrane. Solution A contains 25.0 g of NaCl in 100.0 mL of water and solution B contains 35.0 g of NaCl in 100.0 mL of water.

a. Solution B
b. A $\rightarrow$ B
c. B
d. A

e. increase
f. decrease

Q8.4.3

hyper - higher

hypo - lower

iso - same

Q8.4.4

Cells contain fluid with higher concentration than solution outside the cell.

Q8.4.5

Cells contain fluid with a lower concentration than the solution outside the cell.

Q8.4.6

Water molecules will flow from the tank water into the fish because the fish has a higher concentration of salt. If the fish absorbs too much water, it will die.

Q8.4.7

The "head" region is a phosphate group and it is hydrophillic.

Q8.4.8

The "tail" is a hydrocarbon tail and it is hydrophobic.

8.5: Acid-Base Definitions

Q8.5.1

a. TRUE
b. FALSE - Bronsted-Lowry acid-base definitions are broader.

Q8.5.2

Classify each of the following as an acid, base, or neither.

a. base (contains metal and -OH group)
b. acid (formula starts with H and isn't water)
c. acid (contains -COOH which is carboxylic acid functional group)
d. base (contains metal and -OH group)
e. neither
f. neither (-OH group has to be with metal)

Q8.5.3

Amphoteric substances can act as an acid or base.

Q8.5.4

Identify each reactant in the following reactions as an acid or a base according to the Brønsted-Lowry theory.

a. $HIO_3(aq)$ - acid; $H_2O(l)$ - base; $IO_3^-(aq)$ - base ; $H_3O^+(aq)$ - acid
b. $F^-(aq)$ - base; $HClO(aq)$ - acid; $HF(aq)$ - acid; $ClO^-(aq)$ - base
c. $H_2PO_4^-(aq)$ - acid; $OH^-(aq)$ - base; $HPO_4^{2-}(aq)$ - base; $H_2O(l)$ - acid
d. $CO_3^{2-}(aq)$ - base; $H_2O(l)$ - acid; $HCO_3^-(aq)$ - acid; $OH^-(aq)$ - base

Q8.5.5

a. HIO_3/IO_3^- and H_3O^+/H_2O
b. HF/F^- and $HClO/ClO^-$
c. $H_2PO_4^-/HPO_4^{2-}$ and H_2O/OH^-
d. HCO_3^-/CO_3^{2-} and H_2O/OH^-

Q8.5.6

a. NO_3^-
b. SO_3^{2-}
c. $H_2AsO_4^-$
d. $HCOO^-$ (the H that is removed comes from the carboxylic acid functional group)
e. PO_4^{3-}
f. HS^-
g. S^{2-}
h. CO_3^{2-}
i. HCO_3^-
j. $H_2PO_4^-$
k. Na_2SO_4 or $NaSO_4^-$

Q8.5.7

Write the formula of each base's conjugate acid.

a. $HBrO_3$
b. NH_4^+
c. CH_3COOH
d. H_2CO_3
e. HCN
f. $H_2PO_4^-$
g. H_2S
h. HSO_4^-
i. HCO_3^-
j. H_2CO_3
k. PH_4^+

Q8.5.8

HPO_4^{2-} can act as a base and accept a proton to form $H_2PO_4^-$ and it can act as an acid and donate a proton to form PO_4^{3-}.

8.6: The pH Concept

Q8.6.1

$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

It's neutral because there are equal amounts of H^+ and OH^-.

Q8.6.2

Indicate whether solutions with the following pH values are acidic, basic, or neutral.

a. basic
b. neutral
c. acidic

Q8.6.3

pH + pOH = 14

Q8.6.4

a. pH = 3.64; pOH = 10.36
b. pH = 9.06; pOH = 4.94
c. pOH = 8.72; pH = 5.28
d. pOH = 0.22; pH = 13.78

Q8.6.5

a. pH = 5.00; pOH = 9.00

b. pH = 10.55; pOH = 4.94
c. pOH = 8.72; pH = 5.28
d. pOH = 0.22; pH = 13.78

Q8.6.6

a. $[\mathrm{H}^+] = 1.3 \times 10^{-2}\ M, [\mathrm{OH}^-] = 7.4 \times 10^{-13}\ M$
b. $[\mathrm{H}^+] = 7.1 \times 10^{-12}\ M, [\mathrm{OH}^-] = 1.4 \times 10^{-3}\ M$
c. $[\mathrm{H}^+] = 0.11\ M, [\mathrm{OH}^-] = 8.9 \times 10^{-14}\ M$
d. $[\mathrm{OH}^-] = 6.2 \times 10^{-7}\ M, [\mathrm{H}^+] = 1.6 \times 10^{-8}\ M$
e. $[\mathrm{OH}^-] = 3.8 \times 10^{-14}\ M, [\mathrm{H}^+] = 0.26\ M$
f. $[\mathrm{OH}^-] = 9.3 \times 10^{-8}\ M, [\mathrm{H}^+] = 1.1 \times 10^{-7}\ M$

Q8.6.7

Given pH = 5.00, we know $[\mathrm{H}^+] = 1.0 \times 10^{-5}\ M$ which means $M = \frac{1.0\times10^{-5}\ mol\ \mathrm{H}^+}{1.00\ L}$.

If 0.100 L of water is added to 1.00 L, then the volume changes to 1.10 L but the moles of H^+ does not change. The molarity can be calclulated with the same number of moles and the new volume.

$$M = \frac{1.0\times10^{-5}\ mol\ \mathrm{H}^+}{1.00+0.100\ L}$$
$$M = 9.1 \times 10^{-6}\ M$$

$$pH = 5.04$$

Q8.6.8

How much water would need to be added to the original solution in question 8 in order to bring the pH to 6.00?

To get to pH = 6.00, we need $[\mathrm{H}^+] = 1.0 \times 10^{-6}\ M$.

Use the dilution formula to calculate the total volume of solution.

$$C_1V_1 = C_2V_2$$
$$1.0 \times 10^{-5}\ M \cdot 1.00\ L = 1.0 \times 10^{-6}\ M \cdot V_2\ V_2 = 10.\ L$$

The total volume is 10. L so 9 L needs to be added to the original 1 L solution.

8.3: Le Chatelier's Principle

Learning Outcomes

- Define Le Chatelier's principle.
- Predict how the change in amounts of substances, temperature, or pressure will affect amounts of reactants and products present at equilibrium.

Le Chatelier's Principle

Chemical equilibrium was studied by the French chemist Henri Le Chatelier (1850 - 1936) and his description of how a system responds to a stress to equilibrium has become known as **Le Chatelier's principle**: When a chemical system that is at equilibrium is disturbed by a stress, the system will respond in order to relieve the stress. Stresses to a chemical system involve changes in the concentrations of reactants or products, changes in the temperature of the system, or changes in the pressure of the system. We will discuss each of these stresses separately. The change to the equilibrium position in every case is either a favoring of the forward reaction or a favoring of the reverse reaction. When the forward reaction is favored, the concentrations of products increase, while the concentrations of reactants decrease. When the reverse reaction is favored, the concentrations of the products decrease, while the concentrations of reactants increase.

$$\begin{array}{lll} \textbf{Original Equilibrium} & \textbf{Favored Reaction} & \textbf{Result} \\ \text{A} \rightleftharpoons \text{B} & \text{Forward: A} \rightarrow \text{B} & [\text{A}] \text{ decreases; } [\text{B}] \text{ increases} \\ \text{A} \rightleftharpoons \text{B} & \text{Reverse: A} \leftarrow \text{B} & [\text{A}] \text{ increases; } [\text{B}] \text{ decreases} \end{array} \tag{8.3.1}$$

Figure 8.3.1: Henri Le Chatelier.

Effect of Concentration

A change in concentration of one of the substances in an equilibrium system typically involves either the addition or the removal of one of the reactants or products. Consider the Haber-Bosch process for the industrial production of ammonia from nitrogen and hydrogen gases.

$$\text{N}_2 \left(g \right) + 3 \text{H}_2 \left(g \right) \rightleftharpoons 2 \text{NH}_3 \left(g \right) \tag{8.3.2}$$

If the concentration of one substance in a system is increased, the system will respond by favoring the reaction that removes that substance. When more N_2 is added, the forward reaction will be favored because the forward reaction uses up N_2 and converts it to NH_3. The forward reaction speeds up temporarily as a result of the addition of a reactant. The position of equilibrium shifts as more NH_3 is produced. The concentration of NH_3 increases, while the concentrations of N_2 and H_2 decrease. After some time passes, equilibrium is reestablished with new concentrations of all three substance. As can be seen in the figure below, if more N_2 is added, a new equilibrium is achieved by the system. The new concentration of NH_3 is higher because of the favoring of the forward reaction. The new concentration of the H_2 is lower .The concentration of N_2 is higher than in the original equilibrium, but went down slightly following the addition of the N_2 that disturbed the original equilibrium. By responding in this way, the value of the equilibrium constant for the reaction, K_eq, does not change as a result of the stress to the system.

In other words, the amount of each substance is different but the ratio of the amount of each remains the same.

If more NH_3 were added, the reverse reaction would be favored. This "favoring" of a reaction means temporarily speeding up the reaction in that direction until equilibrium is reestablished. Recall that once equilibrium is reestablished, the rates of the forward and reverse reactions are again equal. The addition of NH_3 would result in increased formation of the reactants, N_2 and H_2.

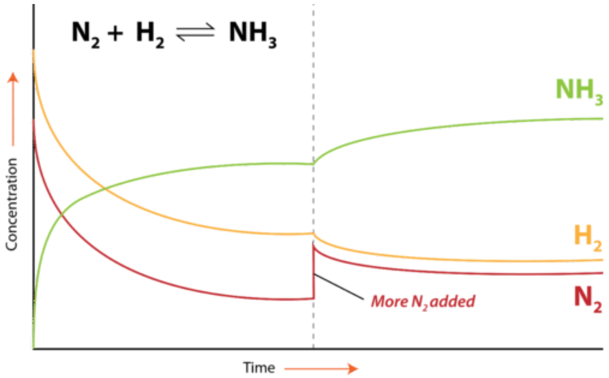

Figure 8.3.2: The Haber-Bosch process is an equilibrium between reactants N_2 and H_2 and product NH_3.

An equilibrium can also be disrupted by the removal of one of the substances. If the concentration of a substance is decreased, the system will respond by favoring the reaction that replaces that substance. In the industrial Haber-Bosch process, NH_3 is removed from the equilibrium system as the reaction proceeds. As a result, the forward reaction is favored so that more NH_3 is produced. The concentrations of N_2 and H_2 decrease. Continued removal of NH_3 will eventually force the reaction to go to completion until all of the reactants are used up. If either N_2 or H_2 were removed from the equilibrium system, the reverse reaction would be favored and the concentration of NH_3 would decrease.

The effect of changes in concentration on an equilibrium system according to Le Chatelier's principle is summarized in the table below.

Table 8.3.1

Stress	Response
addition of reactant	forward reaction favored
addition of product	reverse reaction favored
removal of reactant	reverse reaction favored
removal of product	forward reaction favored

Effect of Temperature

Increasing or decreasing the temperature of a system at equilibrium is also a stress to the system. The equation for the Haber-Bosch process is written again below, as a thermochemical equation (i.e. it contains information about the energy gained or lost when the reaction occurs).

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightleftharpoons 2\mathrm{NH_3}\,(g) + 91\ \mathrm{kJ} \tag{8.3.3}$$

The forward reaction is the exothermic direction: the formation of NH_3 releases heat which is why that is shown as a product. The reverse reaction is the endothermic direction: as NH_3 decomposes to N_2 and H_2, heat is absorbed. An increase in the temperature for this is like adding a product because heat is being released by the reaction. If we add a product then the reaction proceeds towards the formation of more reactants. Reducing the temperature for this system would be similar to removing a product which would favor the formation of more products. The amount of NH_3 will increase and the amount of N_2 and H_2 will decrease.

For changes in concentration, the system responds in such a way that the value of the equilibrium constant, K_{eq} is unchanged. However, a change in temperature shifts the equilibrium and the K_{eq} value either increases or decreases. As discussed in the previous section, values of K_{eq} are dependent on the temperature. When the temperature of the system for the Haber-Bosch process is increased, the resultant shift in equilibrium towards the reactants means that the K_{eq} value decreases. When the temperature is decreased, the shift in equilibrium towards the products means that the K_{eq} value increases.

Le Chatelier's principle as related to temperature changes can be illustrated easily be the reaction in which dinitrogen tetroxide is in equilibrium with nitrogen dioxide.

$$\mathrm{N_2O_4}\,(g) + \mathrm{heat} \rightleftharpoons 2\mathrm{NO_2}\,(g) \tag{8.3.4}$$

Dinitrogen tetroxide $\left(N_2O_4\right)$ is colorless, while nitrogen dioxide $\left(NO_2\right)$ is dark brown in color. When N_2O_4 breaks down into NO_2, heat is absorbed (endothermic) according to the forward reaction above. Therefore, an increase in temperature (adding heat) of the system will favor the forward reaction. Conversely, a decrease in temperature (removing heat) will favor the reverse reaction.

Effect of Pressure

Changing the pressure of an equilibrium system in which gases are involved is also a stress to the system. A change in the pressure on a liquid or a solid has a negligible effect. We will return again the equilibrium for the Haber-Bosch process. Imagine the gases are contained in a closed system in which the volume of the system is controlled by an adjustable piston as shown in the figure below.

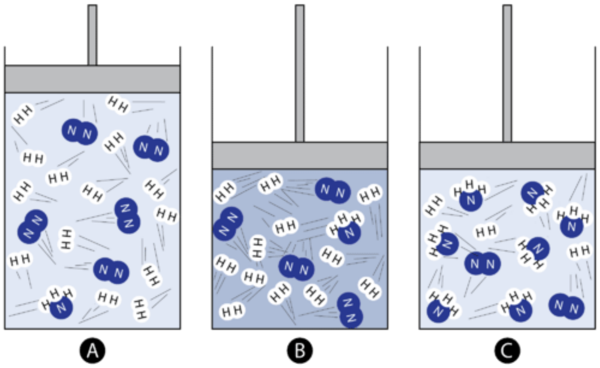

Figure 8.3.3: Effect of pressure on equilibrium. (A) is at equilibrium, (B) increase of pressure through decreased volume, and (C) equilibrium is reestablished.

On the far left, the reaction system contains primarily N_2 and H_2, with only one molecule of NH_3 present. As the piston is pushed inwards, the pressure of the system increases according to Boyle's law. This is a stress to the equilibrium. In the middle image, the same number of molecules is now confined in a smaller space and so the pressure has increased. According to Le Chatelier's principle, the system responds in order to relieve the stress. In the image on the right, the forward reaction has been favored and more NH_3 is produced. The overall result is a decrease in the number of gas molecules in the entire system. This in turn decreases the pressure and provides a relief to the original stress of a pressure increase. An increase in pressure on an equilibrium system favors the reaction which products fewer total moles of gas. In this case, it is the forward reaction that is favored.

A decrease in pressure on the above system could be achieved by pulling the piston outward, increasing the container volume. The equilibrium would respond by favoring the reverse reaction in which NH_3 decomposes to N_2 and H_2. This is because the overall number of gas molecules would increase and so would the pressure. A decrease in pressure on an equilibrium system favors the reaction which produces more total moles of gas. This is summarized in the table below.

Table 8.3.2

Stress	Response
pressure increase	reaction produces fewer gas molecules
pressure decrease	reaction produces more gas molecules

Like changes in concentration, the K_{eq} value for a given reaction is unchanged by a change in pressure. The amounts of each substance will change but the ratio will not. It is important to remember when analyzing the effect of a pressure change on equilibrium that only gases are affected. If a certain reaction involves liquids or solids, they should be ignored. For example, calcium carbonate decomposes according to the equilibrium reaction:

$$\mathrm{CaCO_3}\,(s) \rightleftharpoons \mathrm{CaO}\,(s) + \mathrm{O_2}\,(g) \tag{8.3.5}$$

Oxygen is the only gas in the system. An increase in the pressure of the system slows the rate of decomposition of $\mathrm{CaCO_3}$ because the reverse reaction is favored. When a system contains equal moles of gas on both sides of the equation, pressure has no effect on the equilibrium position, as in the formation of HCl from H_2 and Cl_2.

$$\mathrm{H_2}\,(g) + \mathrm{Cl_2}\,(g) \rightleftharpoons 2\mathrm{HCl}\,(g) \tag{8.3.6}$$

Contributors and Attributions

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CHAPTER OVERVIEW

9: Equilibrium Applications

9.1: Acid and Base Strength

Learning Outcomes

- Define weak acids and bases.
- Write an equation representing the behavior of a weak acid.
- Explain differences between strong and weak acids and strong and weak bases.
- List the 6 strong acids.
- Calculate $\mathrm{p}K_\mathrm{a}$ and $\mathrm{p}K_\mathrm{b}$ values.
- Rank acids in order of strength based on their K_a and $\mathrm{p}K_\mathrm{a}$ values.
- Rank bases in order of strength based on their K_b and $\mathrm{p}K_\mathrm{b}$ values.

So far, we have primarily been defining acids by their ability to donate an $\mathrm{H^+}$ ion and bases by their ability to accept an $\mathrm{H^+}$ ion. However, acids and bases vary in their relative ability to undergo these processes. Which was mentioned when we talked about weak electrolytes.

In general, acids can be classified as strong or weak based on the extent to which they produce $\mathrm{H_3O^+}$ when dissolved in water. For a generic acid, we can write the following equilibrium reaction:

$$\mathrm{HA}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{A^-}\,(aq) \tag{9.1.1}$$

Using the usual shorthand notation, this equation can also be written as follows:

$$\mathrm{HA}\,(aq) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{A^-}\,(aq) \tag{9.1.2}$$

This type of equilibrium, in which a proton is being transferred to water, is often indicated by writing the equilibrium constant as K_a. The relative position of this equilibrium for a given acid determines whether it will be considered strong or weak. When dissolved in water, a **strong acid** will completely transfer its proton to the solvent. In terms of the equilibrium above, the products will be heavily favored $(K_\mathrm{a} \gg 1)$. In fact, the products are so heavily favored that the reverse reaction is often not even considered, and the proton transfer is written as unidirectional. For example, the strong acid HCl can dissociated in water according to the following reaction:

$$\mathrm{HCl}\,(aq) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{H_3O^+}\,(aq) + \mathrm{Cl^-}\,(aq) \tag{9.1.3}$$

which is sometimes written as

$$\mathrm{HCl}\,(aq) \rightarrow \mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq) \tag{9.1.4}$$

to simplify the equation by eliminating the water in the equation because the "*aq*" indicates that water is present. At equilibrium, essentially no intact HCl molecules are still present in solution.

In contrast, the equilibrium for a **weak acid** favors the reactants. A particularly common type of weak acid is an organic molecule that contains a carboxyl group COOH. For example, acetic acid (the acidic component of vinegar) has the formula $\mathrm{CH_3COOH}$. Its dissociation equation can be written as follows:

$$\mathrm{CH_3COOH}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{CH_3COO^-}\,(aq) \tag{9.1.5}$$

sometimes written as $\mathrm{CH_3COOH}\,(aq) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{CH_3COO^-}\,(aq)$. Because we are dealing with a weak acid, K_a for this equilibrium is much less than 1. At equilibrium, most of the acetic acid molecules are still intact, and only a small percentage have transferred their protons to the solvent. The K_a values for some weak acids are listed in the table below. All weak acids are not equally

Table 9.1.1

Acid Name	Structure	K_a
hydrofluoric acid	$\mathrm{H{-}F}$	7.1×10^{-4}
nitrous acid	$\mathrm{O{=}N{-}O{-}H}$	4.5×10^{-4}
formic acid	HCOOH	1.7×10^{-4}

Acid Name	Structure	K_a
acetic acid	CH_3COOH	1.8×10^{-5}
hydrocyanic acid	$H{-}CN$	4.9×10^{-10}

The nature of pK_a is also used to indicate the strength of an acid. pK_a is determined much like pH by taking the negative logarithm of K_a. As with pH, it is used to make values easier to manage. As an acid's strength increases, its K_a value increases and its pK_a value decreases as shown in the table above. Most acids that you will encounter in general chemistry courses are weak acids. There are six common strong acids (see table below). If you recognize these six then you can assume any other acids are weak.

Acids

- Hydrochloric acid, HCl
- Hydrobromic acid, HBr
- Hydroiodic acid, HI
- Perchloric acid, $HClO_4$
- Nitric acid, HNO_3
- Sulfuric acid, H_2SO_4

Strong vs. Weak Bases

Analogous to the acid dissociation reaction from the previous section, we can write the reaction between a generic base and water as follows:

$$B\,(aq) + H_2O\,(l) \rightleftharpoons BH^+\,(aq) + OH^-\,(aq) \tag{9.1.6}$$

The equilibrium constant for a reaction in which a base is deprotonating water (taking water's hydrogen atom) is often given the symbol K_b. Strong bases and weak bases can then be defined based on the position of this equilibrium. A **weak base** would have a very small K_b value (much less than 1), indicating that most molecules of the base do *not* remove a proton from water. Conversely, a **strong base** would have a K_b value greater than or equal to 1.

Nitrogen-containing compounds are a common type of weak base. The lone pair on the nitrogen atom can accept a proton from water as follows:

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq) \tag{9.1.7}$$

The equilibrium constant for this reaction is quite low, so most of the NH_3 molecules will not remove a proton from water. K_b and pK_b values for a few weak bases are listed in the table below.

Table 9.1.3

Base	K_b
ethylamine $(CH_3CH_2NH_2)$	5.6×10^{-4}
methylamine (CH_3NH_2)	4.4×10^{-4}
ammonia (NH_3)	1.8×10^{-5}

The only strong bases that are commonly used in general chemistry courses are ionic compounds composed of metal cations and hydroxide anions, such as $NaOH$, KOH, or $Ba(OH)_2$.

Contributors and Attributions

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9.2: Buffers

Learning Outcomes

- Define buffer.
- Define buffer capacity.
- Describe how a buffer controls pH.
- Identify the components of a buffer solution.

Diabetes mellitus is a disorder of glucose metabolism in which insulin production by the pancreas is impaired. Since insulin helps glucose enter the cells, a decrease of this hormone means that glucose cannot be used in its normal fashion. When this happens, the body begins to break down fats, producing a decrease in blood pH. Chemical systems in the body can balance this pH shift for a while, but excessive acid production can create serious problems if not corrected by administering insulin to restore normal glucose use.

Buffers

If only 1.0 mL of 0.10 M hydrochloric acid is added to 1.0 L of pure water the pH drops drastically from 7.0 to 4.0. This is a 1000-fold increase in the acidity of the solution. For many purposes, it is desirable to have a solution which is capable of resisting such large changes in pH when relatively small amounts of acid or base are added to them. Such a solution is called a buffer. A **buffer** is a solution of a weak acid or a base and its salt. Both components must be present for the system to act as a buffer to resist changes in pH. The salt is the conjugate of the weak acid or of the weak base. It can be shown as the ion or with the counter ion. $\left(\text{COO}^- \text{ or } \text{COONa} \right)$

Some common buffer systems are listed in the table below. Note that the two components of the buffer system differ by only one hydrogen ion $\left(\text{H}^+ \right)$.

Table 9.2.1: Some Common Buffers

Buffer system	Buffer components	pH of buffer (equal molarities of both components)
Acetic acid/acetate ion	$\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$	4.74
Carbonic acid/hydrogen carbonate ion	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.38
Dihydrogen phosphate ion/hydrogen phosphate ion	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
Ammonia/ammonium ion	$\text{NH}_3/\text{NH}_4^+$	9.25

One example of a buffer is a solution made of acetic acid (the weak acid) and sodium acetate (a conjugate of the acid). The pH of a buffer consisting of $0.50 \text{ M CH}_3\text{COOH}$ and $0.50 \text{ M CH}_3\text{COONa}$ is 4.74. If 10.0 mL of 1.0 M HCl is added to 1.0 L of the buffer, the pH only decreases to 4.73. This ability to "soak up" the additional hydrogen ions from the HCl that was added is due to the reaction below.

$$\text{CH}_3\text{COO}^- \left(aq \right) + \text{H}^+ \left(aq \right) \rightleftharpoons \text{CH}_3\text{COOH} \left(aq \right) \tag{9.2.1}$$

Since both the acetate ion and the acetic acid were already present in the buffer, the only thing that changes is the ratio of one to the other. Small changes in that ratio have only very minor effects on the pH.

If 10.0 mL of 1.0 M NaOH were added to another 1.0 L of the same buffer, the pH would only increase to 4.76. In this case, the buffer takes up the additional hydroxide ions.

$$\text{CH}_3\text{COOH} \left(aq \right) + \text{OH}^- \left(aq \right) \rightleftharpoons \text{CH}_3\text{COO}^- \left(aq \right) + \text{H}_2\text{O} \left(l \right) \tag{9.2.2}$$

Again the ratio of acetate ion to acetic acid changes only slightly, this time causing a very small increase in the pH. All buffers follow Le Chatelier's principle and respond to a stress on the system by responding to minimize the stress.

It is possible to add so much acid or base to a buffer that its ability to resist a significant change in pH is overwhelmed. The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs. The buffer capacity is exceeded when the number of moles of $\mathrm{H^+}$ or $\mathrm{OH^-}$ that are added to the buffer exceeds the number of moles of the buffer components.

Human blood contains a buffer of carbonic acid $\left(\mathrm{H_2CO_3}\right)$ and bicarbonate anion $\left(\mathrm{HCO_3^-}\right)$ in order to maintain blood pH between 7.35 and 7.45, as a value higher than 7.8 (alkalosis) or lower than 6.8 (acidosis) can lead to death. In this buffer, hydronium and bicarbonate anion are in equilibrium with carbonic acid. The bicarbonate neutralizes excess acids in the blood while the carbonic acid neutralizes excess bases.

Furthermore, carbonic acid can decompose into $\mathrm{CO_2}$ gas and water, resulting in a second equilibrium system between carbonic acid and water. Because $\mathrm{CO_2}$ is an important component of the blood buffer, its regulation in the body, as well as that of $\mathrm{O_2}$, is extremely important. The effect of this can be important when the human body is subjected to strenuous conditions.

$$\mathrm{CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons HCO_3^- + H^+} \tag{9.2.3}$$

Contributors and Attributions

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9.3: Equilibrium Applications (Exercises)

These are homework exercises to accompany Chapter 9 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Answers are below the questions.

Questions

9.1: Acid and Base Strength

(click here for solutions)

Q9.1.1

Describe the difference between a strong acid and a weak acid.

Q9.1.2

Describe the difference between a strong base and a weak base.

Q9.1.3

Identify each of the following as a strong acid, weak acid, strong base, or weak base.

a. HCl
b. $NaOH$
c. KOH
d. HNO_2
e. HNO_3
f. HF
g. NH_3
h. $Ba(OH)_2$
i. CH_3CH_2COOH

Q9.1.4

Write an equation representing the behavior of each substance in question 3. Pay attention to the type of arrow used in the equation.

Q9.1.5

A solution is prepared by dissolving 15.0 grams of NaOH in enough water to make 500.0 mL of solution. Calculate the pH of the solution.

Q9.1.6

A solution is prepared by dissolving 22.0 grams of HCl in enough water to make 300.0 mL of solution. Calculate the pH of the solution.

9.2: Buffers

(click here for solutions)

Q9.2.1

What is a buffer?

Q9.2.2

What is the purpose of a buffer?

Q9.2.3

Determine whether or not each of these pairs can act as a buffer.

a. HCl/Cl^-
b. HF/F^-
c. H_2SO_4/HSO_4^-
d. HSO_4^-/SO_4^{2-}

e. H_2O/NaOH
f. HNO_2/KNO_2
g. HCl/NaOH

Q9.2.4

Write the formula of the conjugate base needed to form a buffer with each of the following weak acids.

a. $HClO_3$
b. $H_2PO_4^-$
c. CH_3COOH

Q9.2.5

Write the formula of the conjugate acid needed to form a buffer with each of the following weak bases.

a. NH_3
b. CH_3NH_2

Q9.2.6

Describe buffer capacity.

Answers

9.1: Acid and Base Strength

Q9.1.1

A strong acid completely dissociates into ions and a weak acid doesn't completely dissociate into ions.

Q9.1.2

A strong base is a base (metal with an -OH) group that dissociates completely into ions. A weak base is a proton acceptor but not ll of the molecules will accept a proton.

Q9.1.3

Identify each of the following as a strong acid, weak acid, strong base, or weak base.

a. HCl is a strong acid. See list of 6 strong acids.
b. NaOH is a strong base. It has a metal with an -OH group.
c. KOH is a strong base. It has a metal and an -OH group.
d. HNO_2 is a weak acid. The formula starts with H but it's not water so it's an acid. Recognize it is weak because it is not on the list of 6 strong acids.
e. HNO_3 is a strong acid. The formula starts with H but it's not water so it's an acid. See the list of 6 strong acids.
f. HF is a weak acid. the formula starts with H but it's not water so it's an acid. Recognize it is weak becauce it is not on the list of 6 strong acids.
g. NH_3 is a weak base. Amines are weak bases.
h. $Ba(OH)_2$ is a strong base. It has a metal with an -OH group.
i. CH_3CH_2COOH is a weak base. It has a carboxylic acid functional group so it's an acid. Carboxylic acids are all weak (also, not on the list of 6 strong acids).

Q9.1.4

Write an equation representing the behavior of each substance in question 3. Pay attention to the type of arrow used in the equation.

a. $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
b. $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
c. $KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$
d. $HNO_2(aq) \rightleftharpoons H^+(aq) + NO_2^-(aq)$
e. $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
f. $HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$
g. $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

h. $Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$
i. $CH_3CH_2COOH(aq) \rightleftharpoons H^+(aq) + CH_3CH_2COO^-(aq)$

Q9.1.5

NaOH is a strong base and completely dissociates (see reaction). Since it completely dissociates, the concentration of NaOH equals the concentration of OH^-. We need to calculate the concentration of NaOH.

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

$$15.0\ g\ \mathrm{NaOH}\left(\frac{1\ mol}{40.00\ g}\right) = 0.375\ mol\ \mathrm{NaOH}$$

$$M = \frac{mol\ solute}{L\ soln} = \frac{0.375\ mol}{0.500\ L} = 0.750\ M\ NaOH$$

$[NaOH] = [OH^-] = 0.750\ M$

Use $[OH^-]$ to find pOH.

$$p\mathrm{OH} = -log[\mathrm{OH}^-] = -log[0.750] = 0.125$$

Now, convert from pOH to pH. Note that 14 is an exact number in this context so it does not affect significant figures.

$$p\mathrm{H} + p\mathrm{OH} = 14$$
$$p\mathrm{H} = 14 - p\mathrm{OH}$$
$$p\mathrm{H} = 14 - 0.125$$
$$p\mathrm{H} = 13.875$$

Q9.1.6

HCl is a strong acid and completely dissociates (see reaction). Since it completely dissociates, the concentration of HCl equals the concentration of H^+. We need to calculate the concentration of HCl.

$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

$$22.6\ g\ \mathrm{HCl}\left(\frac{1\ mol}{36.46\ g}\right) = 0.603\ mol\ \mathrm{HCl}$$

$$M = \frac{mol\ solute}{L\ soln} = \frac{0.603\ mol}{0.300\ L} = 2.01\ M\ HCl$$

$[HCl] = [H^+] = 2.01\ M$

$p\mathrm{H} = -log[\mathrm{H}^+] = -log[2.01] = -0.303$ (pH can be less than zero if it is a strong acid with a concentration greater than 1 M)

9.2: Buffers

Q9.2.1

A buffer is a weak acid and its conjugate base (or a weak base and its conjugate acid) that helps maintain the pH of a solution.

Q9.2.2

The purpose of a buffer is to resist change of pH in a solution.

Q9.2.3

a. HCl/Cl^- cannot because HCl is a strong acid.
b. HF/F^- can because HF is a weak acid and F^- is its conjugate base. It will be added to mixture as a salt (ionic compound) of F^- (i.e. NaF, KF, etc)
c. H_2SO_4/HSO_4^- cannot because H_2SO_4 is a strong acid.
d. HSO_4^-/SO_4^{2-} can because HSO_4^- is a weak acid and SO_4^{2-} is its conjugate base. Both compounds will be added to the mixture as salts (i.e. $NaHSO_4$ and Na_2SO_4).
e. $H_2O/NaOH$ cannot becase NaOH is a strong base and NaOH and H_2O are not a conjugate acid-base pair.
f. HNO_2/KNO_2 can because HNO_2 is a weak acid and KNO_2 contains its conjugate base. Note that KNO_2 is a strong electrolyte so it dissociates into K^+ and NO_2^-. HNO_2 and NO_2^- form a conjugate acid-base pair.
g. HCl/NaOH cannot because HCl is a strong acid, NaOH is a strong base, and they do not form a conjugate acid-base pair.

Q9.2.4

Write the formula of the conjugate base needed to form a buffer with each of the following weak acids.

a. ClO_3^- (or a salt such as $NaClO_3$)
b. HPO_4^{2-} (or a salt such as Na_2HPO_4)
c. CH_3COO^- (or a salt such as CH_3COONa)

Q9.2.5

Write the formula of the conjugate acid needed to form a buffer with each of the following weak acids.

a. NH_4^+ (or a salt such as NH_4Cl)
b. $CH_3NH_3^+$ (or a salt such as CH_3NH_3Cl)

Q9.2.6

Buffer capacity is the amount of acid or base that can be added to a buffer solution before it can no longer resist significant changes in the pH of the solution. Adding small amounts of acid or base will change the pH of a buffer by a small amount and the buffer continues to be effective. If larger amounts of acid or base are added, the buffer capacity is reached and the solution can no longer resist changes in pH.

CHAPTER OVERVIEW

10: Nuclear and Chemical Reactions

10.1: Nuclear Radiation

Learning Outcomes

- Define radioactivity.
- Describe a radioisotope.
- Explain how radioactivity involves a change in the nucleus of a radioisotope.
- Explain the characteristics of the forms of radiation.
- Describe and write equations for the primary types of radioactive decay.

Marie Curie (1867 - 1934) was a Polish scientist who pioneered research into nuclear radiation (Figure 10.1.1). She was awarded the Nobel Prize in physics in 1903 along with her husband Pierre and Antoine Henri Becquerel for their work on radioactivity. She was awarded a second Nobel Prize in 1911, this time in chemistry, for her continued research on radioactive elements. In this lesson, you will learn about radioactivity, the reasons why certain elements and isotopes are radioactive, and the most common types of radioactive decay processes.

Figure 10.1.1: Marie Curie was one of the leading scientists in the field of radioactivity. She discovered two radioactive elements and was awarded two Nobel Prizes for her work.

Radioactivity

Radioactivity was discovered quite by accident. In 1896, Henri Becquerel was studying the effect of certain uranium salts on photographic film plates. He believed that the salts had an effect on the film only when they had been exposed to sunlight. He accidentally found that uranium salts that had not been exposed to sunlight still had an effect on the photographic plates. The Curies, associates of Becquerel at the time, showed that the uranium was emitting a type of ray that interacted with the film. Marie Curie called this radioactivity. **Radioactivity** *is the spontaneous breakdown of an atom's nucleus by the emission of particles and/or radiation*. **Radiation** *is the emission of energy through space in the form of particles and/or waves*.

Nuclear reactions are very different from chemical reactions. In chemical reactions, atoms become more stable by participating in a transfer of electrons or by sharing electrons with other atoms. In nuclear reactions, it is the nucleus of the atom that gains stability by undergoing a change of some kind. Some elements have no stable isotopes, which means that any atom of that element is radioactive. For some other elements, only certain isotopes are radioactive. A **radioisotope** *is an isotope of an element that is unstable and undergoes radioactive decay*. The energies that are released in nuclear reactions are many orders of magnitude greater than the energies involved in chemical reactions. Unlike chemical reactions, nuclear reactions are not noticeably affected by changes in environmental conditions, such as temperature or pressure.

The discovery of radioactivity and its effects on the nuclei of elements disproved Dalton's assumption that atoms are indivisible. A **nuclide** *is a term for an atom with a specific number of protons and neutrons in its nucleus*. As we will see, when nuclides of one type emit radiation, they are changed into different nuclides. Radioactive decay is spontaneous and does not required an input of energy to occur. The stability of a particular nuclide depends on the composition of its nucleus, including the number of protons, the number of neutrons, and the proton-to-neutron ratio.

The Band of Stability

Carbon-12, with six protons and six neutrons, is a stable nucleus, meaning that it does not spontaneously emit radioactivity. Carbon-14, with six protons and eight neutrons, is unstable and naturally radioactive. Among atoms with lower atomic numbers, the ideal ratio of neutrons to protons is approximately 1:1. As the atomic number increases, the stable neutron-proton ratio gradually increases to about 1.5:1 for the heaviest known elements. For example, lead-206 is a stable nucleus that contains 124 neutrons and 82 protons, a ratio of 1.51 to 1.

This observation is shown in the figure below. The **band of stability** *is the range of stable nuclei on a graph that plots the number of neutrons in a nuclide against the number of protons.* Known stable nuclides are shown with individual blue dots, while the 1:1 and 1.5:1 ratios are shown with a solid red line and a green line, respectively.

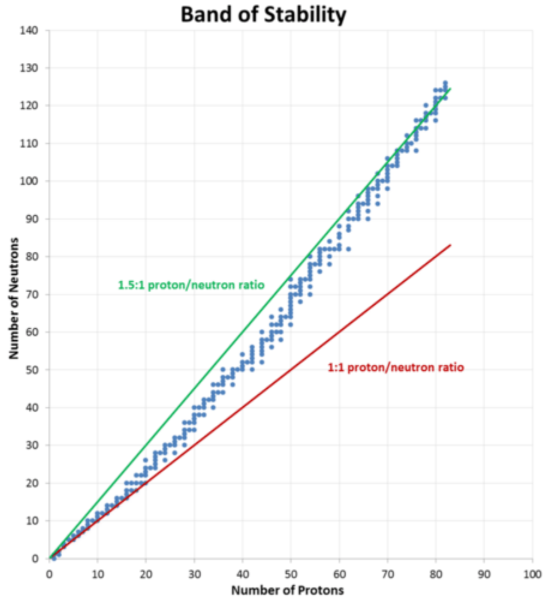


Figure 10.1.2: A graph of the number of neutrons in a nucleus as a function of the number of protons. Each known stable nucleus is represented by a blue dot. The ideal neutron to proton ratio changes from 1:1 for light nuclei to 1.5:1 for the heaviest nuclei.

It should be noted that just because a nucleus is "unstable" (able to undergo spontaneous radioactive decay) does not mean that it will rapidly decompose. For example, uranium-238 is unstable because it spontaneously decays over time, but if a sample of uranium-238 is allowed to sit for 1000 years, only $0.0000155\%$of the sample will have decayed. However, other unstable nuclei, such as berkelium-243, will be almost completely gone (>99.9999%decayed) in less than a day.

Radioactive Decay

Unstable nuclei spontaneously emit radiation in the form of particles and energy. This generally changes the number of protons and/or neutrons in the nucleus, resulting in a more stable nuclide. A **nuclear reaction** *is a reaction that affects the nucleus of an atom*. One type of a nuclear reaction is **radioactive decay**, *a reaction in which a nucleus spontaneously disintegrates into a slightly lighter nucleus, accompanied by the emission of particles, energy, or both*. An example is shown below, in which the nucleus of a polonium atom radioactively decays into a lead nucleus.

$$^{210}_{84}\mathrm{Po} \rightarrow {}^{206}_{82}\mathrm{Pb} + {}^{4}_{2}\mathrm{He} \tag{10.1.1}$$

Note that in a balanced nuclear equation, the sum of the atomic numbers (subscripts) and the sum of the mass numbers (superscripts) must be equal on both sides of the equation. Recall the notation system for isotopes, which shows both the atomic number and mass number along with the chemical symbol.

Because the number of protons changes as a result of this nuclear reaction, the identity of the element changes. **Transmutation** *is a change in the identity of a nucleus as a result of a change in the number of protons*. There are several different types of naturally occurring radioactive decay, and we will examine each separately.

Alpha Decay

An **alpha particle** (α) *is a helium nucleus with two protons and two neutrons.* Alpha particles are emitted during some types of radioactive decay. The net charge of an alpha particle is $2+$, and its mass is approximately $4\ \mathrm{amu}$. The symbol for an alpha particle in a nuclear equation is usually $^{4}_{2}\mathrm{He}$, though sometimes α is used. Alpha decay typically occurs for very heavy nuclei in which the nuclei are unstable due to large numbers of nucleons. For nuclei that undergo alpha decay, their stability is increased by the subtraction of two protons and two neutrons. For example, uranium-238 decays into thorium-234 by the emission of an alpha particle (see figure below).

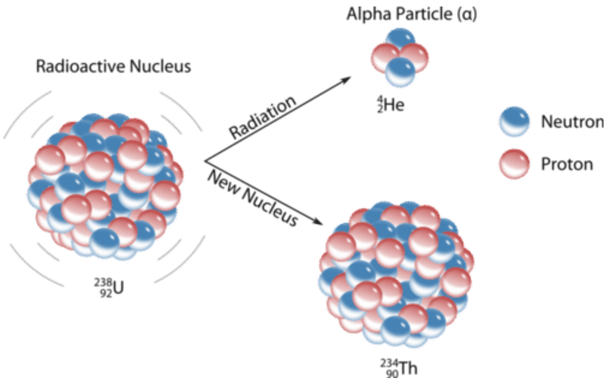


Figure 10.1.3: The unstable uranium-238 nucleus spontaneously decays into a thorium-234 nucleus by emitting an alpha particle.

Beta Decay

Nuclei above the band of stability are unstable because their neutron to proton ratio is too high. To decrease that ratio, a neutron in the nucleus is capable of turning into a proton and an electron. The electron is immediately ejected at a high speed from the nucleus. A **beta particle** (β) *is a high-speed electron emitted from the nucleus of an atom during some kinds of radioactive decay* (see figure below). The symbol for a beta particle in an equation is either β or $^{0}_{-1}$e. Carbon-14 undergoes beta decay, transmutating into a nitrogen-14 nucleus.

$$^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\mathrm{e} \tag{10.1.2}$$

Note that beta decay increases the atomic number by one, but the mass number remains the same.

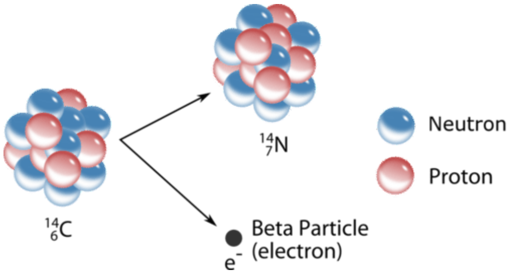


Figure 10.1.4: The beta decay of a carbon-14 nuclide involves the conversion of a neutron to a proton and an electron, with the electron being emitted from the nucleus.

Positron Emission

Nuclei below the band of stability are unstable because their neutron to proton ratio is too low. One way to increase that ratio is for a proton in the nucleus to turn into a neutron and another particle called a positron. A **positron** *is a particle with the same mass as an electron, but with a positive charge*. Like the beta particle, a positron is immediately ejected from the nucleus upon its formation. The symbol for a positron in an equation is $^{0}_{+1}$e. For example, potassium-38 emits a positron, becoming argon-38.

$$^{38}_{19}\mathrm{K} \rightarrow {}^{38}_{18}\mathrm{Ar} + {}^{0}_{1}\mathrm{e} \tag{10.1.3}$$

Positron emission decreases the atomic number by one, but the mass number remains the same.

Electron Capture

An alternate way for a nuclide to increase its neutron to proton ratio is by a phenomenon called electron capture. In electron capture, an electron from an inner orbital is captured by the nucleus of the atom and combined with a proton to form a neutron. For example, silver-106 undergoes electron capture to become palladium-106.

$$^{106}_{47}\mathrm{Ag} + {}^{0}_{-1}\mathrm{e} \rightarrow {}^{106}_{46}\mathrm{Pd} \tag{10.1.4}$$

Note that the overall result of electron capture is identical to positron emission. The atomic number decreases by one while the mass number remains the same.

Gamma Ray Emission

Gamma rays (γ) *are very high energy electromagnetic waves emitted from a nucleus.* Gamma rays are emitted by a nucleus when nuclear particles undergo transitions between nuclear energy levels. This is analogous to the electromagnetic radiation emitted when excited electrons drop from higher to lower energy levels; the only difference is that nuclear transitions release much more energetic radiation. Gamma ray emission often accompanies the decay of a nuclide by other means.

$$^{230}_{90}\text{Th} \rightarrow {}^{226}_{88}\text{Ra} + {}^{4}_{2}\text{He} + \gamma \tag{10.1.5}$$

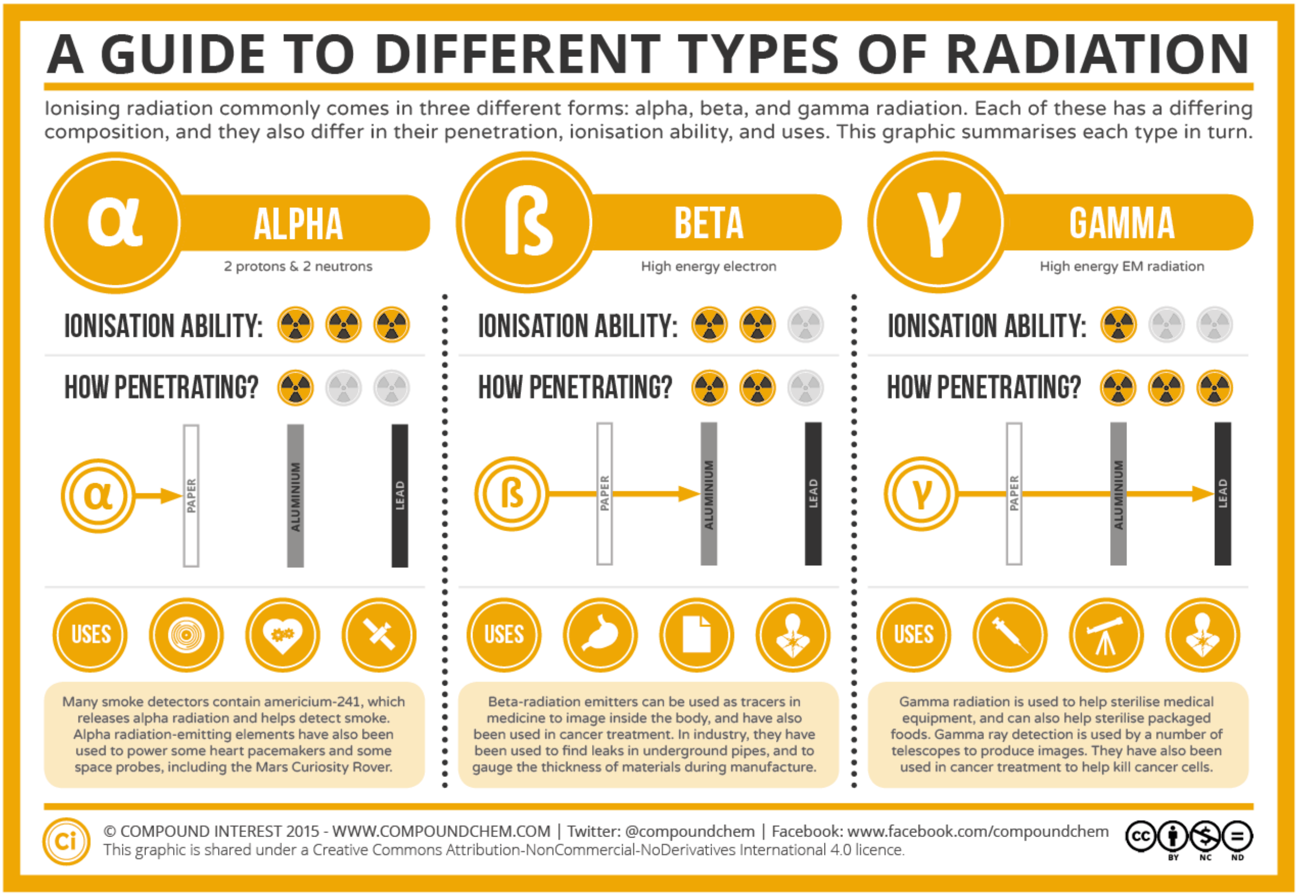
The emission of gamma radiation has no effect on the atomic number or mass number of the products, but it reduces their energy.

Summary of Nuclear Radiation

The table below summarizes the main types of nuclear radiation, including charge, mass, symbol, and penetrating power. Penetrating power refers to the relative ability of the radiation to pass through common materials. Radiation with high penetrating power is potentially more dangerous because it can pass through skin and do cellular damage.

Table 10.1.1 *Summary of types of nuclear radiation.*

Type	Symbol	Mass number	Charge	Penetration Power	Shielding
Alpha particle	${}^{4}_{2}\text{He}$ or α	4	2+	Low	Paper, skin
Beta particle	${}^{0}_{-1}\text{e}$ or β	0	1−	Moderate	Heavy cloth, plastic
Positron	${}^{0}_{1}\text{e}$ or β^+	0	1+	Moderate	Heavy cloth, plastic
Gamma ray	γ or ${}^{0}_{0}\gamma$	0	0	High	Lead, concrete
Neutron	${}^{1}_{0}\text{n}$	1	0	High	Water, lead


A GUIDE TO DIFFERENT TYPES OF RADIATION

Ionising radiation commonly comes in three different forms: alpha, beta, and gamma radiation. Each of these has a differing composition, and they also differ in their penetration, ionisation ability, and uses. This graphic summarises each type in turn.

Many smoke detectors contain americium-241, which releases alpha radiation and helps detect smoke. Alpha radiation-emitting elements have also been used to power some heart pacemakers and some space probes, including the Mars Curiosity Rover.

Beta-radiation emitters can be used as tracers in medicine to image inside the body, and have also been used in cancer treatment. In industry, they have been used to find leaks in underground pipes, and to gauge the thickness of materials during manufacture.

Gamma radiation is used to help sterilise medical equipment, and can also help sterilise packaged foods. Gamma ray detection is used by a number of telescopes to produce images. They have also been used in cancer treatment to help kill cancer cells.

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Figure 10.1.5: Infographic showing the different properties of radiation. "A Guide to Different Types of Radiation" is by Compound Interest is licensed under CC BY NC ND 4.0.

Supplemental Resources

- Atomic structure review: www.sciencegeek.net/Chemistry...t1Numbers2.htm
- Balancing Nuclear equations: www.sciencegeek.net/Chemistry...rEquations.htm
- Nuclear decay: www.sciencegeek.net/Chemistry...cleardecay.htm

Contributors and Attributions

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

10.2: Fission and Fusion

Learning Outcomes

- Define fission.
- Describe a nuclear chain reaction and how it is applied in both a fission bomb and in a nuclear power plant.
- Define fusion.

Nuclear Fission

The most stable nuclei are of intermediate mass. To become more stable, the heaviest nuclei are capable of splitting into smaller fragments. **Nuclear fission** *is a process in which a very heavy nucleus (mass > 200) splits into smaller nuclei of intermediate mass*. Because the smaller nuclei are more stable, the fission process releases tremendous amounts of energy. Nuclear fission may occur spontaneously or may occur as a result of bombardment. When uranium-235 is hit with a slow-moving neutron, it absorbs it and temporarily becomes the very unstable uranium-236. This nucleus splits into two medium-mass nuclei while also emitting more neutrons. The mass of the products is less than the mass of the reactants, with the lost mass being converted to energy.

Nuclear Chain Reactions

Because the fission process produces more neutrons, a chain reaction can result. A **chain reaction** *is a reaction in which the material that starts the reaction is also one of the products and can start another reaction*. Illustrated below is a nuclear chain reaction for the fission of uranium-235.

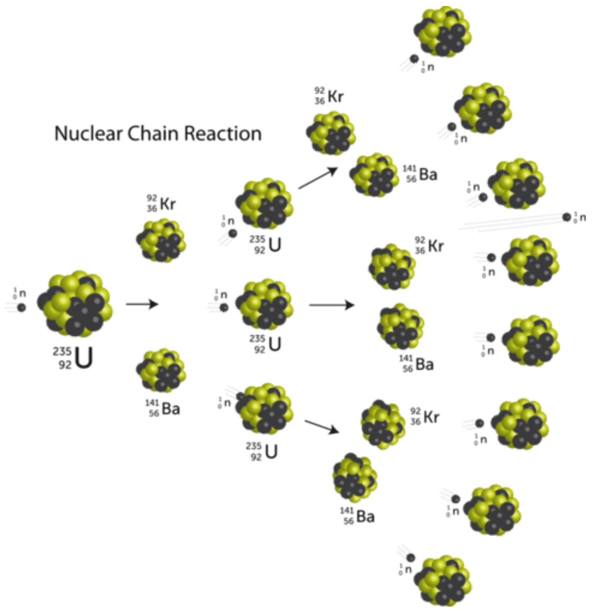

Figure 10.2.1: The nuclear chain reaction is a series of fission processes that sustains itself due to the continuous production of neutrons in each reaction.

The original uranium-235 nucleus absorbs a neutron, splits into a krypton-92 nucleus and a barium-141 nucleus, and releases three more neutrons upon splitting.

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{92}_{36}\mathrm{Kr} + {}^{141}_{56}\mathrm{Ba} + 3\,{}^{1}_{0}\mathrm{n} \tag{10.2.1}$$

Those three neutrons are then able to cause the fission of three more uranium-235 nuclei, each of which release more neutrons, and so on. The chain reaction continues until all of the uranium-235 nuclei have been split, or until the released neutrons escape the sample without striking any more nuclei. If the size of the original sample of uranium-235 is sufficiently small, too many neutrons escape without striking other nuclei, and the chain reaction quickly ceases. The **critical mass** *is the minimum amount of fissionable material needed to sustain a chain reaction*. Atomic bombs and nuclear reactors are two ways to harness the large energy released during nuclear fission.

Atomic Bombs - Uncontrolled Nuclear Reactions

In an atomic bomb, or fission bomb, the nuclear chain reaction is designed to be uncontrolled, releasing huge amounts of energy in a short amount of time. A critical mass of fissionable plutonium is contained within the bomb, but not at a sufficient density. Conventional explosives are used to compress the plutonium, causing it to go critical and trigger a nuclear explosion.

Nuclear Power Plants- Controlled Nuclear Reactions

A nuclear power plant (see figure below) uses a controlled fission reaction to produce large amounts of heat. The heat is then used to generate electrical energy.

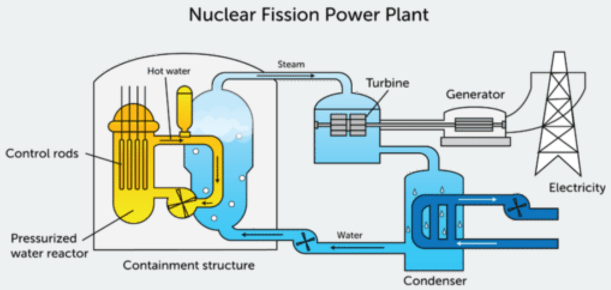

Figure 10.2.2: A nuclear reactor harnesses the energy of nuclear fission to generate electricity.

Uranium-235, the usual fissionable material in a nuclear reactor, is first packaged into fuel rods. In order to keep the chain reaction from processing unchecked, moveable control rods are placed in between the fuel rods. **Control rods** *limit the amount of available neutrons by absorbing some of them and preventing the reaction from proceeding too rapidly*. Common control rod materials include alloys with various amounts of silver, indium, cadmium, or boron. A **moderator** *is a material that slows down high-speed neutrons*. This is beneficial because slow-moving neutrons are more efficient at splitting nuclei. Water is often used as a moderator. The heat released by the fission reaction is absorbed by constantly circulating coolant water. The coolant water releases its heat to a steam generator, which turns a turbine and generates electricity. The core of the reactor is surrounded by a containment structure that absorbs radiation.

Nuclear Fusion

The lightest nuclei are also not as stable as nuclei of intermediate mass. **Nuclear fusion** *is a process in which light-mass nuclei combine to form a heavier and more stable nucleus*. Fusion produces even more energy than fission. In the sun and other stars, four hydrogen nuclei combine at extremely high temperatures and pressures to produce a helium nucleus. The concurrent loss of mass is converted into extraordinary amounts of energy (see figure below).

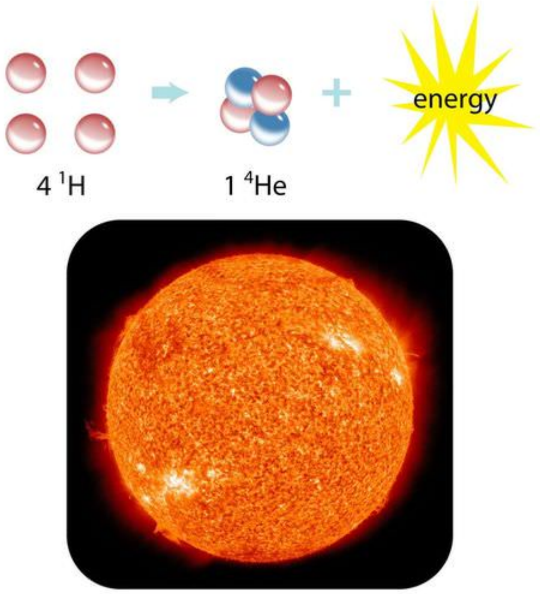

Figure 10.2.3: Nuclear fusion takes place when small nuclei combine to make larger ones. The enormous amounts of energy produced by fusion powers our sun and other stars.

Fusion is even more appealing than fission as an energy source because no radioactive waste is produced and the only reactant needed is hydrogen. However, fusion reactions only occur at very high temperatures - in excess of $40,000,000^{\text{o}}\text{C}$. No known materials can withstand such temperatures, so there is currently no feasible way to harness nuclear fusion for energy production, although research is ongoing.

Uses of Radiation

As we saw earlier, different types of radiation vary in their abilities to penetrate through matter. Alpha particles have very low penetrating ability and are stopped by skin and clothing. Beta particles have a penetrating ability that is about 100 times that of alpha particles. Gamma rays have very high penetrating ability, and great care must be taken to avoid overexposure to gamma rays.

Exposure and Detection

Radiation emitted by radioisotopes is called ionizing radiation. **Ionizing radiation** *is radiation that has enough energy to knock electrons off the atoms of a bombarded substance and produce ions*. The **roentgen** *is a unit that measures nuclear radiation and is equal to the amount of radiation that produces* 2×10^{9} *ion pairs when it passes through* 1 cm^{3} *of air*. The primary concern is that ionizing radiation can do damage to living tissues. Radiation damage is measured in rems, which stands for roentgen equivalent man . A **rem** *is the amount of ionizing radiation that does as much damage to human tissue as is done by 1 roentgen of high-voltage x-rays*. Tissue damage from ionizing radiation can cause genetic mutations due to interactions between the radiation and DNA, which can lead to cancer.

You are constantly being bombarded with background radiation from space and from geologic sources that vary depending on where you live. Average exposure is estimated to be about 0.1 rem per year. The maximum permissible does of radiation exposure for people in the general population is 0.5 rem per year. Some people are naturally at higher risk because of their occupations, so reliable instruments to detect radiation exposure have been developed. A **Geiger counter** *is a device that uses a gas-filled metal tube to detect radiation* (see figure below). When the gas is exposed to ionizing radiation, it conducts a current, and the Geiger counter registers this as audible clicks. The frequency of the clicks corresponds to the intensity of the radiation.

Figure 10.2.4: A Geiger counter is used to detect radiation.

A **scintillation counter** *is a device that uses a phosphor-coated surface to detect radiation by the emission of bright bursts of light*. Workers who are at risk of exposure to radiation wear small portable film badges. A **film badge** *consists of several layers of photographic film that can measure the amount of radiation to which the wearer has been exposed*. Film badges are removed and analyzed at periodic intervals to ensure that the person does not become overexposed to radiation on a cumulative bases.

Medicine and Agriculture

Radioactive nuclides, such as cobalt-60, are frequently used in medicine to treat certain types of cancers. The faster growing cancer cells are exposed to the radiation and are more susceptible to damage than healthy cells. Thus, the cells in the cancerous area are killed by the exposure to high-energy radiation. Radiation treatment is risky because some healthy cells are also killed, and cells at the center of a cancerous tumor can become resistant to the radiation.

Radioactive tracers *are radioactive atoms that are incorporated into substances so that the movement of these substances can be tracked by a radiation detector*. Tracers are used in the diagnosis of cancer and other diseases. For example, iodine-131 is used to detect problems with a person's thyroid. A patient first ingests a small amount of iodine-131. About two hours later, the iodine uptake by the thyroid is determined by a radiation scan of the patient's throat. In a similar way, technetium-99 is used to detect brain tumors and liver disorders, and phosphorus-32 is used to detect skin cancer.

Radioactive tracers can be used in agriculture to test the effectiveness of various fertilizers. The fertilizer is enriched with a radioisotope, and the uptake of the fertilizer by the plant can be monitored by measuring the emitted radiation levels. Nuclear radiation is also used to prolong the shelf life of produce by killing bacteria and insects that would otherwise cause the food to spoil faster.

Contributors and Attributions

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

10.3: Half-Life

Learning Outcomes

- Define half-life as it relates to radioactive nuclides and solve half-life problems.
- Describe the general process by which radioactive dating is used to determine the age of various objects.
- Calculate the time for a sample to decay.
- Complete dosage calculations based on nuclide activity.

The rate of radioactive decay is often characterized by the half-life of a radioisotope. **Half-life** $\left(t_{1/2}\right)$ *is the time required for one half of the nuclei in a sample of radioactive material to decay.* After each half-life has passed, one half of the radioactive nuclei will have transformed into a new nuclide (see table below). The rate of decay and the half-life do not depend on the original size of the sample. They also do not depend upon environmental factors such as temperature and pressure.

Table 10.3.1

Number of Half-Lives Passed	Fraction Remaining	Percentage Remaining	Mass remaining starting with 80 g
1	1/2	50	40 g
2	1/4	25	20 g
3	1/8	12.5	10 g
4	1/16	6.25	5.0 g
5	1/32	3.125	2.5 g

As an example, iodine-131 is a radioisotope with a half-life of 8 days. It decays by beta particle emission into xenon-131.

$$^{131}_{53}\text{I} \rightarrow {}^{131}_{54}\text{Xe} + {}^{0}_{-1}\text{e} \tag{10.3.1}$$

After eight days have passed, half of the atoms of any sample of iodine-131 will have decayed, and the sample will now be 50% iodine-131 and 50% xenon-131. After another eight days pass (a total of 16 days or 2 half-lives), the sample will be 25% iodine-131 and 75% xenon-131. This continues until the entire sample of iodine-131. has completely decayed (see figure below).

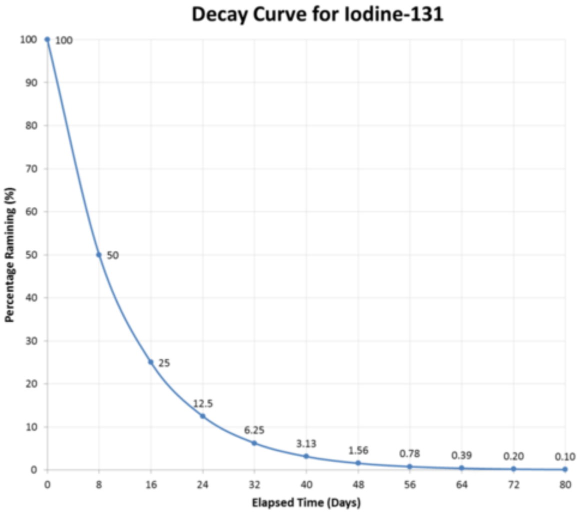

Figure 10.3.1: The half-life of iodine-131 is eight days. Half of a given sample of iodine-131 decays after each eight-day time period elapses.

Half-lives have a very wide range, from billions of years to fractions of a second. Listed below (see table below) are the half-lives of some common and important radioisotopes. Those with half-lives on the scale of hours or days are the ones most suitable for use in medical treatment.

Table 10.3.2

Nuclide	Half-Life $(t_{1/2})$	Decay Mode
Carbon-14	5730 years	β^-
Cobalt-60	5.27 years	β^-
Francium-220	27.5 seconds	α
Hydrogen-3	12.26 years	β^-
Iodine-131	8.07 days	β^-
Nitrogen-16	7.2 seconds	β^-
Phosphorus-32	14.3 days	β^-
Plutonium-239	24,100 years	α
Potassium-40	1.28×10^9 years	β^- and e^- capture
Radium-226	1600 years	α
Radon-222	3.82 days	α
Strontium-90	28.1 days	β^-
Technetium-99	2.13×10^5 years	β^-
Thorium-234	24.1 days	β^-
Uranium-235	7.04×10^8 years	α
Uranium-238	4.47×10^9 years	α

The following example illustrates how to use the half-life of a sample to determine the amount of radioisotope that remains after a certain period of time has passed.

Example 10.3.1: Strontium-90

Strontium-90 has a half-life of 28.1 days. If you start with a 5.00 mg sample of the isotope, how much remains after 140.5 days have passed?

Solution

Step 1: List the known values and plan the problem.

Known

- Original mass $= 5.00 \text{ mg}$
- $t_{1/2} = 28.1$ days
- Time elapsed $= 140.5$ days

Unknown

- Final mass of $\text{Sr-90} = ? \text{ mg}$

First, find the number of half-lives that have passed by dividing the time elapsed by the half-life. Then, reduce the amount of Sr-90 by half, once for each half-life.

Step 2: Solve.

Number of half-lives, days, mass

0 half-lives, 0 days, 5.00 mg

1 half-life, 28.1 days, 2.50 mg

2 half-lives, 56.2 days, 1.25 mg

3 half-lives, 84.3 days, 0.613 mg

4 half-lives, 112.4 days, 0.313 mg

5 half-lives, 140.5 days, 0.156 mg

Step 3: Think about your result.

According to the data above, the passage of 5 half-lives means 0.156 mg of the original Sr-90 remains. The remaining 4.844 mghas decayed by beta particle emission to yttrium-90.

Radioactive Dating

Radioactive dating *is a process by which the approximate age of an object is determined through the use of certain radioactive nuclides*. For example, carbon-14 has a half-life of 5,730 years and is used to measure the age of organic material. The ratio of carbon-14 to carbon-12 in living things remains constant while the organism is alive because fresh carbon-14 is entering the organism whenever it consumes nutrients. When the organism dies, this consumption stops, and no new carbon-14 is added to the organism. As time goes by, the ratio of carbon-14 to carbon-12 in the organism gradually declines, because carbon-14 radioactively decays while carbon-12 is stable. Analysis of this ratio allows archaeologists to estimate the age of organisms that were alive many thousands of years ago. Carbon dating is effective until about 50,000 years. The ages of many rocks and minerals are far greater than the ages of fossils. Uranium-containing minerals that have been analyzed in a similar way have allowed scientists to determine that the Earth is over 4 billion years old.

Decay Series

In many instances, the decay of an unstable radioactive nuclide simply produces another radioactive nuclide. It may take several successive steps to reach a nuclide that is stable. A **decay series** *is a sequence of successive radioactive decays that proceeds until a stable nuclide is reached*. The terms reactant and product are generally not used for nuclear reactions. Instead, the terms parent and daugher nuclide are used to to refer to the starting and ending isotopes in a decay process. The figure below shows the decay series for uranium-238.

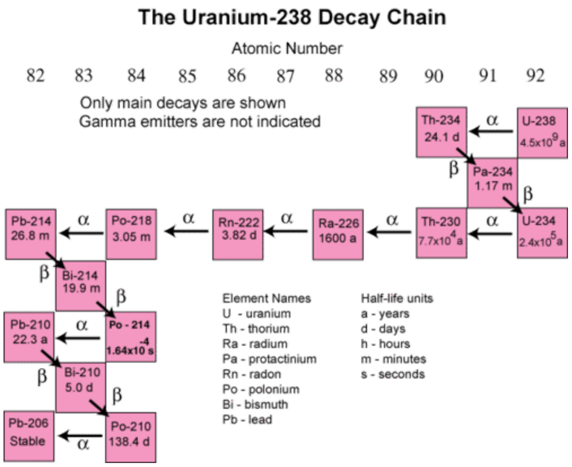

Figure 10.3.2: The decay of uranium-238 proceeds along many steps until a stable nuclide, lead-206, is reached. Each decay has its own characteristic half-life.

In the first step, uranium-238 decays by alpha emission to thorium-234 with a half-life of 4.5×10^9 years. This decreases its atomic number by two. The thorium-234 rapidly decays by beta emission to protactinium-234 ($t_{1/2} =$ 24.1 days). The atomic number increases by one. This continues for many more steps until eventually the series ends with the formation of the stable isotope lead-206.

Artificial Transmutation

As we have seen, transmutation occurs when atoms of one element spontaneously decay and are converted to atoms of another element. **Artificial transmutation** *is the bombardment of stable nuclei with charged or uncharged particles in order to cause a nuclear reaction*. The bombarding particles can be protons, neutrons, alpha particles, or larger atoms. Ernest Rutherford performed some of the earliest bombardments, including the bombardment of nitrogen gas with alpha particles to produce the unstable fluorine-18 isotope.

$$^{14}_{7}\mathrm{N} + ^{4}_{2}\mathrm{He} \rightarrow ^{18}_{9}\mathrm{F} \tag{10.3.2}$$

Fluorine-18 quickly decays to the stable nuclide oxygen-17 by releasing a proton.

$$^{18}_{9}\mathrm{F} \rightarrow ^{17}_{8}\mathrm{O} + ^{1}_{1}\mathrm{H} \tag{10.3.3}$$

When beryllium-9 is bombarded with alpha particles, carbon-12 is produced with the release of a neutron.

$$^{9}_{4}\mathrm{Be} + ^{4}_{2}\mathrm{He} \rightarrow ^{12}_{6}\mathrm{C} + ^{1}_{0}\mathrm{n} \tag{10.3.4}$$

Transuranium Elements

Many, many radioisotopes that do not occur naturally have been generated by artificial transmutation. The elements technetium and promethium have been produced, since these elements no longer occur in nature. All of their isotopes are radioactive and have half-lives short enough that any amount of the elements that once existed have long since disappeared through natural decay. The **transuranium elements** *are elements with atomic numbers greater than 92*. All isotopes of these elements are radioactive and none occur naturally.

Dosing

Half-life calculations can be based on mass, percent remaining, or dose. Regardless of which one, the concept is still the same. Understanding the radioactivity and half-life of a sample is important for calculating the correct dose for a patient and determining the levels and duration of radioactive emission from a patient after treatment is received.

Frequently, dosages for radioactive isotopes are given the activity in volume. For example, the concentration of I-137 is given as $50\ \mu\mathrm{Ci/mL}$ (microCurie per milliliter). This relationship can be used to calculate the volume needed for a particular dose. For example, a patient needs $125\ \mu\mathrm{Ci}$ of $\mathrm{I}1$-51. What volume of a $50\mu\mathrm{Ci}$ per $10\ \mathrm{mL}$ solution should be given?

$$125\ \mu\mathrm{Ci} \left(\frac{10\ \mathrm{mL}}{50\ \mu\mathrm{Ci}}\right) = 25\ \mathrm{mL} \tag{10.3.5}$$

Example 10.3.2

A patient is given I-131 to treat thyroid cancer. The patient receives $5.50\ \mathrm{mL}$ of a solution containing $50\ \mathrm{mCi}$ (milliCurie) in $2\ \mathrm{mL}$ (assume concentration is an exact number). What does (in mCi) is given to the patient? What will be the activity in mCi after 24.21 days given that the half-life of I-131 is 8.07 days?

Solution

The first part of the problem is to find the dose given to the patient. We are given the volume and the concentration (in units of radioactivity over volume).

$$5.50\ \mathrm{mL\ solution} \left(\frac{50\ \mathrm{mCi}}{2\ \mathrm{mL}}\right) = 138\ \mathrm{mCi} \tag{10.3.6}$$

The patient is given a does of $138\ \mathrm{mCi}$

Now, we need to find the activity after 24.21 days have passed. After one half-life (8.07 days), the sample will have half as much activity $(138/2 = 69.0\ \mathrm{mCi})$ After two half-lives (total of 16.14 days), the sample will have half as much activity as after the first half-life $(69.0/2 = 34.5\ \mathrm{mCi})$ After three half-lives (total of 24.21 days), the sample will have half as much activity as after the second half-life $(34.5/2 = 17.3\ \mathrm{mCi})$ Therefore, after 24.21 days (3 half-lives), the radioactivity will be $17.3\ \mathrm{mCi}$

Contributors and Attributions

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

10.4: Physical and Chemical Changes

Learning Outcomes

- Distinguish between physical and chemical changes.
- Give examples of physical and chemical changes.

Physical Changes

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. **Melting** is an example of a **physical change**. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.

Figure 10.4.1: Melting ice in the Beaufort Sea.

When a piece of wood is ground into sawdust, that change is irreversible since the sawdust could not be reconstituted into the same piece of wood that it was before. Cutting the grass or pulverizing a rock would be other irreversible physical changes. Firewood also represents an irreversible physical change since the pieces cannot be put back together to form the tree.

Figure 10.4.2: Firewood being cut is a physical change because the composition doesn't change when being cut.

Chemical Changes

When exposed to air, an object made of iron will eventually begin to rust (see figure below).

Figure 10.4.3: Rust (iron oxide) forms on an unprotected iron surface.

As the rust forms on the surface of the iron, it flakes off to expose more iron, which will continue to rust. Rust is clearly a substance that is different from iron. Rusting is an example of a chemical change.

A **chemical property** *describes the ability of a substance to undergo a specific chemical change*. A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust. A more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are *burn*, *rot*, *explode*, and *ferment*. Chemical properties are very useful as a way of identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

A chemical change is also called a chemical reaction. A **chemical reaction** *is a process that occurs when one or more substances are changed into one or more new substances*. Zinc (Zn) is a silver-gray element that can be ground into a powder. If zinc is mixed at room temperature with powdered sulfur (S), a bright yellow element, the result will simply be a mixture of zinc and sulfur. No chemical reaction occurs. However, if energy is provided to the mixture in the form of heat, the zinc will chemically react with the sulfur to form the compound zinc sulfide (ZnS). Pictured below are the substances involved in this reaction.

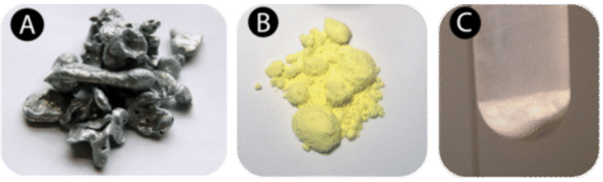

Figure 10.4.4: Zinc (A) and sulfur (B) are two elements that undergo a chemical reaction when heated to form the the compound zinc sulfide (C).

The reaction between zinc and sulfur can be depicted in something called a chemical equation. In words, we could write the reaction as:

$$\text{zinc} + \text{sulfur} \rightarrow \text{zinc sulfide} \tag{10.4.1}$$

A more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved:

$$\mathrm{Zn} + \mathrm{S} \rightarrow \mathrm{ZnS} \tag{10.4.2}$$

The substance(s) to the left of the arrow in a chemical equation are called reactants. A **reactant** *is a substance that is present at the start of a chemical reaction.* The substance(s) to the right of the arrow are called products. A **product** *is a substance that is present at the end of a chemical reaction*. In the equation above, zinc and sulfur are the reactants that chemically combine to form zinc sulfide as a product.

Recognizing Chemical Reactions

How can you tell if a chemical reaction is taking place? Certain visual clues indicate that a chemical reaction is likely (but not necessarily) occurring, including the following examples:

1. A change of color occurs during the reaction.
2. A gas is produced during the reaction.
3. A solid product, called a precipitate, is produced in the reaction.
4. A visible transfer of energy occurs in the form of light as a result of the reaction.

When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced (see figure below). The production of a gas is also an indication that a chemical reaction may be occurring.

Figure 10.4.5: Zinc reacts with hydrochloric acid to produce bubbles of hydrogen gas.

When a colorless solution of lead (II) nitrate is added to a colorless solution of potassium iodide, a yellow solid called a precipitate is instantly produced (see figure below). A **precipitate** *is a solid product that forms from a reaction and settles out of a liquid mixture*. The formation of a precipitate may also indicate the occurrence of a chemical reaction.

$$\mathrm{Pb(NO_3)_2}\,(aq) + 2\mathrm{KI}\,(aq) \rightarrow \mathrm{PbI_2}\,(s) + 2\mathrm{KNO_3}\,(aq) \tag{10.4.3}$$

Figure 10.4.6: A yellow precipitate of solid lead (II) iodide forms immediately when solutions of lead (II) nitrate and potassium iodide are mixed.

Contributors and Attributions

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

10.5: Chemical Equations

Learning Outcomes

- Describe chemical reactions using word equations.
- Write equations for chemical reactions.
- Use coefficients to balance chemical equations.

Chemical reactions are occurring all around you. Plants use sunlight to drive their photosynthetic process and produce energy. Cars and other vehicles burn gasoline in order to power their engines. Batteries use electrochemical reactions to produce energy and power many everyday devices. Many chemical reactions are going on inside you as well, especially during the digestion of food.

In math class, you have written and solved many mathematical equations. Chemists keep track of chemical reactions by writing equations as well. In any chemical reaction one or more substances, called reactants, are converted into one or more new substance, called products. The general form of the equation for such a process looks like this.

$$\text{Reactants} \rightarrow \text{Products} \tag{10.5.1}$$

Unlike in a math equation, a chemical equation does not use an equal sign. Instead the arrow is called a yield sign and so the equation is described as "reactants yield products".

Word Equations

You can describe a chemical reaction by writing a word equation. When silver metal is exposed to sulfur it reacts to form silver sulfide. Silver sulfide is commonly known as tarnish and turns the surface of silver objects dark and streaky black (see figure below). The sulfur that contributes to tarnish can come from traces of sulfur in the air or from food such as eggs. The word equation for the process is:

$$\text{Silver} + \text{sulfur} \rightarrow \text{Silver sulfide} \tag{10.5.2}$$

The silver and the sulfur are the reactants in the equation, while the silver sulfide is the product.

Figure 10.5.1: The coffee percolator on the left has been tarnished from exposure to sulfur. Tarnish is the chemical compound silver sulfide. The same percolator on the right has been polished with a tarnish removal product in order to restore its silver finish.

Another common chemical reaction is the burning of methane gas. Methane is the major component of natural gas and is commonly burned on a gas stove or in a Bunsen burner (see figure below). Burning is a chemical reaction in which some type of fuel is reacted with oxygen gas. The products of the reaction in the burning of methane as well as other fuels are carbon dioxide and water. The word equation for this reaction is:

$$\text{Methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water} \tag{10.5.3}$$

Figure 10.5.2 *A Bunsen burner is commonly used to heat substances in a chemistry lab. Methane is reacted with oxygen to form carbon dioxide and water.*

Chemical Equations

Word equations are time-consuming to write and will not prove to be convenient for many of the things that chemists need to do with equations. A **chemical equation** *is a representation of a chemical reaction that displays the reactants and products with chemical formulas*. The chemical equation for the reaction of methane with oxygen is shown:

$$\mathrm{CH_4 + O_2 \rightarrow CO_2 + H_2O} \tag{10.5.4}$$

The equation above, called a **skeleton equation**, *is an equation that shows only the formulas of the reactants and products with nothing to indicate the relative amounts*. The first step in writing an accurate chemical equation is to write the skeleton equation, making sure that the formulas of all substances involved are written correctly. All reactants are written to the left of the yield arrow, separated from one another by a plus sign. Likewise, products are written to the right of the yield arrow, also separated with a plus sign.

It is often important to know the physical states of the reactants and products taking part in a reaction. To do this, put the appropriate symbol in parentheses after each formula: (s) for solid, (l), for liquid, (g) for gas, and (aq) for an aqueous (water-based) solution. At room temperature, the components of the previous reaction are in the following states:

$$\mathrm{CH_4}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l) \tag{10.5.5}$$

The table below shows a listing of symbols used in chemical equations.

Table 10.5.1: Symbols Used in Chemical Equations

Symbol	Description
$+$	used to separate multiple reactants or products
$\rightarrow$	yield sign; separates reactants from products
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium
(s)	reactant or product in the solid state
(l)	reactant or product in the liquid state
(g)	reactant or product in the gas state
(aq)	reactant or product in an aqueous solution (dissolved in water)
$\xrightarrow{\mathrm{Pt}}$	formula written above the arrow is used as a catalyst in the reaction
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated

Balancing Chemical Equations

Suppose you were to write a word equation for building the ideal ham sandwich (see figure below). Perhaps you might come up with this:

$$\text{Ham} + \text{cheese} + \text{tomato} + \text{pickles} + \text{bread} \rightarrow \text{ham sandwich} \tag{10.5.6}$$

The reactants are the "parts" or ingredients of the ham sandwich while the sandwich itself is the product. There is something missing from your equation, however. There is no indication how many of each "reactant" is required to make the "product". For one thing, you would certainly need two slices of bread to make a conventional sandwich.

Figure 10.5.3: A ham sandwich can be thought of as the product of a reaction while all of the individual ingredients are the reactants.

Let's say that the perfect ham sandwich (HS) is composed of 2 slices of ham (H), a slice of cheese (C), 1 slice of tomato (T), 5 pickles (P), and 2 slices of bread (B). Accounting for the numbers of each reactant, as well as substituting symbols for words, your equation would become:

$$2\mathrm{H}+\mathrm{C}+\mathrm{T}+5\mathrm{P}+2\mathrm{B}\rightarrow\mathrm{HS} \tag{10.5.7}$$

This now shows the correct quantities of the reactants. As one final improvement, we will change the "formula" of the product. Since the final sandwich contains all the reactants that went into it, its formula should reflect that.

$$2\mathrm{H}+\mathrm{C}+\mathrm{T}+5\mathrm{P}+2\mathrm{B}\rightarrow\mathrm{H_2CTP_5B_2} \tag{10.5.8}$$

The subscript after each symbol in the product stands for the number of that particular reactant found on the reactant side of the equation: 2 for H, 1 for C, etc.

Since the equation now shows equal numbers of each sandwich part on both sides of the equation, we say that the equation is balanced. Chemical equations must also be balanced in a similar way. A **balanced equation** *is a chemical equation in which mass is conserved and there are equal numbers of atoms of each element on both sides of the equation.*

We can write a chemical equation for the reaction of carbon with hydrogen gas to form methane $(\mathrm{CH_4})$.

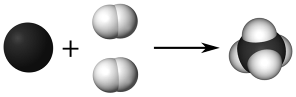
$$\begin{array}{ccccc}\mathrm{C}\,(s) & + & \mathrm{H_2}\,(g) & \rightarrow & \mathrm{CH_4}\,(g) \\ 1\,\mathrm{C\ atom} & & 2\,\mathrm{H\ atoms} & & 1\,\mathrm{C\ atom},\ 4\,\mathrm{H\ atoms}\end{array} \tag{10.5.9}$$

In order to write a correct equation, you must first write the correct skeleton equation with the correct chemical formulas. Recall that hydrogen is a diatomic molecule and so is written as $\mathrm{H_2}$. When we count the number of atoms of both elements, shown under the equation, we see that the equation is not balanced. There are only 2 atoms of hydrogen on the reactant side of the equation, while there are 4 atoms of hydrogen on the product side. This violates the law of conservation of mass, which states that mass must be conserved in any chemical reaction or physical process. Another common way to express the law of conservation of mass is that matter cannot be created or destroyed.

John Dalton's atomic theory stated that chemical reactions are separations, combinations, or rearrangements of atoms. Atoms themselves cannot be created or destroyed. Dalton's theory explains the law of conservation of mass and the process of balancing an equation ensures that the law is followed. We can balance the above equation by adding a coefficient of 2 in front of the formula for hydrogen.

$$\mathrm{C}\,(s)+2\mathrm{H_2}\,(g)\rightarrow\mathrm{CH_4}\,(g) \tag{10.5.10}$$

A **coefficient** *is a small whole number placed in front of a formula in an equation in order to balance it*. The 2 in front of the $\mathrm{H_2}$ means that there are a total of $2\times2=4$ atoms of hydrogen as reactants. Visually, the reaction looks like:



In the balanced equation, there is one atom of carbon and four atoms of hydrogen on both sides of the arrow. Below are guidelines for writing and balancing chemical equations.

1. Determine the correct chemical formulas for each reactant and product.
2. Write the skeleton equation by placing the reactant(s) on the left side of the yield sign $(\rightarrow)$ and the product(s) on the right side. If there is more than one reactant or product, separate with plus signs.
3. Count the number of atoms of each element that appears as a reactant and as a product. If a polyatomic ion is unchanged on both sides of the equation, count it as a unit.
4. Balance each element one at a time by placing coefficients in front of the formula. No coefficient is written for a 1. It is best to begin by balancing elements that only appear in one formula on each side of the equation. You can only balance equations by using coefficients, NEVER change the subscripts in a chemical formula.
5. Check each atom or polyatomic ion to be sure that they are equal on both sides of the equation.
6. Make sure that all coefficients are in the lowest possible ratio. If necessary, reduce to the lowest ratio.

Example 10.5.1

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the chemical equation for this reaction.

Solution

Step 1: Plan the problem.

Follow the steps for writing and balancing a chemical equation.

Step 2: Solve.

Write the skeleton equation with the correct formulas.

$$\mathrm{Pb(NO_3)_2}\,(aq) + \mathrm{NaCl}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{PbCl_2}\,(s) \tag{10.5.11}$$

Count the number of each atom or polyatomic ion on both sides of the equation (see table below).

The nitrate ions and the chlorine atoms are unbalanced. Start by placing a 2 in front of the NaCl. This increases the reactant counts to 2 Na atoms and 2 Cl atoms. Then place a 2 in front of the $\mathrm{NaNO_3}$. The result is:

$$\mathrm{Pb(NO_3)_2}\,(aq) + 2\mathrm{NaCl}\,(aq) \rightarrow 2\mathrm{NaNO_3}\,(aq) + \mathrm{PbCl_2}\,(s) \tag{10.5.12}$$

The new count for each atom and polyatomic ion becomes (see table below).

Step 3: Think about your result.

The equation is now balanced since there are equal numbers of atoms of each element on both sides of the equation.

Count the number of each atom or polyatomic ion on both sides of the equation.

Table 10.5.2: Unbalanced Reactants/Products

Reactants	Products
1 Pb atom	1 Pb atom
2 $\mathrm{NO_3^-}$	1 $\mathrm{NO_3^-}$
2 Na atoms	2 Na atoms
2 Cl atoms	1 Cl atom

The new count for each atom and polyatomic ion becomes:

Table 10.5.3: Balanced Reactants/Products

Reactants	Products
1 Pb atom	1 Pb atom

Reactants	Products
2 NO_3^-	2 NO_3^-
2 Na atoms	2 Na atoms
2 Cl atoms	2 Cl atoms

Some equations provide a challenge to balancing when one or more of the elements can't be balanced simply by using one coefficient. Aluminum reacts with oxygen gas to form aluminum oxide according to the equation:

$$\mathrm{Al}\,(s) + \mathrm{O_2}\,(g) \rightarrow \mathrm{Al_2O_3}\,(s) \tag{10.5.13}$$

Since there are two oxygen atoms on the reactant side and 3 oxygen atoms on the product side, no single whole-number coefficient will balance the oxygen atoms. Find the lowest common multiple of 2 and 3, which is 6. Placing a 3 in front of the O_2 and a 2 in front of the Al_2O_3 will result in 6 oxygen atoms on both sides. Finish by balancing the aluminum with a 4.

$$4\mathrm{Al}\,(s) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{Al_2O_3}\,(s) \tag{10.5.14}$$

The equation is balanced with 4 Al atoms and 6 O atoms on each side.

Finally, we will return to the equation from earlier where methane was reacted with oxygen to form carbon dioxide and water.

$$\mathrm{CH_4}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l) \tag{10.5.15}$$

The element oxygen appears in two different places on the product side of the equation, so you should not start by trying to balance the oxygen. Instead, balance the carbon and the hydrogen first. The carbon is already balanced, but the hydrogen is balanced by placing a 2 in front of the water.

$$\mathrm{CH_4}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) \tag{10.5.16}$$

Now count the total number of oxygen atoms on the product side: two from the CO_2 and two from the 2 H_2O to give a total of four. Place a 2 in front of the O_2.

$$\mathrm{CH_4}\,(g) + 2\mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + 2\mathrm{H_2O}\,(l) \tag{10.5.17}$$

Balancing difficult equations can be a trial-and-error process and is a skill that requires practice. If you find that one particular strategy with a tough equation isn't working, start over and balance a different element first. Persistence will lead you to the correct balanced equation.

Supplemental Resources

- http://phet.colorado.edu/en/simulati...ical-equations
- "Balancing Act": http://education.jlab.org/elementbalancing/index.html
- Balancing Chemical Equations: www.sciencegeek.net/Chemistry...nBalancing.htm

Contributors and Attributions

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10.6: Nuclear and Chemical Reactions (Exercises)

These are homework exercises to accompany Chapter 10 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Answers are below the questions.

Questions

10.1: Nuclear Radiation

(click here for solutions)

Q10.1.1

Write the symbol for the isotope described.

a. 12 protons, 12 electrons, 13 neutrons
b. 17 protons, 17 electrons, 20 neutrons
c. 53 protons, 53 electrons, 78 neutrons
d. 92 protons, 92 electrons, 146 neutrons

Q10.1.2

Determine the number of protons, neutrons, and electrons in each isotope.

a. ${}^{195}_{77}\text{Ir}$
b. ${}^{209}_{82}\text{Pb}$
c. ${}^{211}_{84}\text{Po}$
d. ${}^{237}_{93}\text{Np}$

Q10.1.3

Fill in the missing numbers in each equation.

a. ${}^{196}_{82}\text{Pb} + {}^{0}_{-1}\text{e} \rightarrow {}^{\_\_}_{\_\_}\text{Tl}$
b. ${}^{28}_{15}\text{P} \rightarrow {}^{\_\_}_{\_\_}\text{Si} + {}^{0}_{1}\text{e}$
c. ${}^{226}_{88}\text{Ra} \rightarrow {}^{\_\_}_{\_\_}\text{Rn} + {}^{4}_{2}\alpha$
d. ${}^{73}_{30}\text{Zn} \rightarrow {}^{\_\_}_{\_\_}\text{Ga} + {}^{0}_{-1}\text{e}$

Q10.1.4

Fill in the blanks for each of the nuclear reactions below. State the type of decay in each case.

a. ${}^{198}_{79}\text{Au} \rightarrow \_\_\_\_\_\_\_\_ + {}^{0}_{-1}\text{e}$
b. $ce^{57}_{27}Co + {}^{0}_{-1}\text{e} \rightarrow \_\_\_\_\_\_\_\_$
c. ${}^{230}_{92}\text{U} \rightarrow \_\_\_\_\_\_\_\_ + {}^{4}_{2}\text{He}$
d. ${}^{128}_{56}\text{Ba} \rightarrow \_\_\_\_\_\_\_\_ + {}^{0}_{1}\text{e}$
e. ${}^{131}_{53}\text{I} \rightarrow {}^{131}_{54}\text{Xe} + \_\_\_\_\_\_\_\_$
f. ${}^{239}_{94}\text{Pu} \rightarrow {}^{235}_{92}\text{U} + \_\_\_\_\_\_\_\_$

Q10.1.5

Write balanced nuclear reactions for each of the following.

a. Francium-220 undergoes alpha decay.
b. Arsenic-76 undergoes beta decay.
c. Uranium-231 captures an electron.
d. Promethium-143 emits a positron.

10.2: Fission and Fusion

(click here for solutions)

Q10.2.1

Describe the main difference between fission and fusion.

Q10.2.2

What is the difference between the fission reactions used in nuclear power plants and nuclear weapons?

Q10.2.3

How do the doses of radioisotopes used in diagnostic procedures and therapeutic treatment compare to one another?

10.3: Half-Life

(click here for solutions)

Q10.3.1

What percent of a sample remains after one half-life? Three half-lives?

Q10.3.2

The half-life of polonium-218 is 3.0 min. How much of a 0.540 mg sample would remain after 9.0 minutes have passed?

Q10.3.3

The half-life of hydrogen-3, commonly known as tritium, is 12.26 years. If 4.48 mg of tritium has decayed to 0.280 mg, how much time has passed?

Q10.3.4

The half-life of protactinium-234 is 6.69 hours. If a 0.812 mg sample of Pa-239 decays for 40.14 hours, what mass of the isotope remains?

Q10.3.5

2.86 g of a certain radioisotope decays to 0.358 g over a period of 22.8 minutes. What is the half-life of the radioisotope?

Q10.3.6

Use Table 10.3.2 above to determine the time it takes for 100. mg of carbon-14 to decay to 6.25 mg.

Q10.3.7

A radioisotope decays from 55.9 g to 6.99 g over a period of 72.5 hours. What is the half-life of the isotope?

Q10.3.8

A sample of a radioisotope with a half-life of 9.0 hours has an activity of 25.4 mCi after 36 hours. What was the original activity of the sample?

Q10.3.9

What volume of a radioisotope should be given if a patient needs 125 mCi of a solution which contains 45 mCi in 5.0 mL?

Q10.3.10

Sodium-24 is used to treat leukemia. A 36-kg patient is prescribed 145 μCi/kg and it is supplied to the hospital in a vial containing 250 μCi/mL. What volume should be given to the patient?

Q10.3.11

Using information from the previous question and knowing the half-life of Na-24 is 15 hours, calculate the total dose in μCi given to the patient. How long will it take for the radioactivity to be approximately 80 μCi?

Q10.3.12

Lead-212 is one of the radioisotopes used in the treatment of breast cancer. A patient needs a 15 μCi dose and it is supplied as a solution with a concentration of 2.5 μCi/mL. What volume does the patient need? Given the half-life of lead is 10.6 hours, what will be the radioactivity of the sample after approximately four days?

10.4: Physical and Chemical Changes

(click here for solutions)

Q10.4.1

Identify each of the following as a physical or chemical change.

a. melting ice
b. boiling water
c. cooking eggs
d. dissolving salt in water
e. burning match
f. metal reacting with HCl
g. mixing NaCl and KCl
h. decomposition of hydrogen peroxide

Q10.4.2

Give two signs that indicate a chemical change is occurring.

Q10.4.3

What doesn't change when a substance undergoes a physical change?

10.5: Chemical Equations

(click here for solutions)

Q10.5.1

Identify the reactants and products in each chemical reaction.

a. In photosynthesis, carbon dioxide and water react to form glucose and oxygen.
b. Magnesium oxide forms when magnesium is exposed to oxygen gas.

Q10.5.2

Write grammatically correct sentences that completely describe the chemical reactions shown in each equation. You may need to look up the names of elements or compounds.

a. $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$
b. $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
c. $2Cs(s) + 2H_2O(l) \rightarrow 2CsOH(aq) + H_2(g)$

Q10.5.3

How many atoms of each element are represented by the following combinations of coefficients and chemical formulas?

a. $5Br_2$
b. $2NH_3$
c. $4(NH_4)_2SO_4$
d. $2CH_3COOH$
e. $3Fe(NO_3)_3$
f. $2K_3PO_4$

Q10.5.4

Balance the following equations.

a. $Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
b. $Li(s) + N_2(g) \rightarrow Li_3N(s)$
c. $Ca(OH)_2 + HBr \rightarrow CaBr_2 + H_2O$
d. $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$
e. $NH_3 + CuO \rightarrow Cu + N_2 + H_2O$

Q10.5.5

Balance the following equations.

a. $Fe(s) + Cl_2(g) \rightarrow FeCl_3(g)$
b. $C_4H_{10}O + O_2 \rightarrow CO_2 + H_2O$
c. $As + NaOH \rightarrow Na_3AsO_3 + H_2$
d. $SiO_2 + HF \rightarrow SiF_4 + H_2O$
e. $N_2 + O_2 + H_2O \rightarrow HNO_3$

Answers

10.1: Nuclear Radiation

Q10.1.1

Write the symbol for the isotope described.

a. ${}^{25}_{12}\text{Mg}$
b. ${}^{37}_{17}\text{Cl}$
c. ${}^{131}_{53}\text{I}$
d. ${}^{238}_{92}\text{U}$

Q10.1.2

a. 77 protons, 77 electrons, 118 neutrons
b. 82 protons, 82 electrons, 127 neutrons
c. 84 protons, 84 electrons, 127 neutrons
d. 93 protons, 93 electrons, 144 neutrons

Q10.1.3

a. ${}^{196}_{82}\text{Pb} + {}^{0}_{-1}\text{e} \rightarrow {}^{196}_{81}\text{Tl}$
b. ${}^{28}_{15}\text{P} \rightarrow {}^{28}_{14}\text{Si} + {}^{0}_{1}\text{e}$
c. ${}^{226}_{88}\text{Ra} \rightarrow {}^{222}_{86}\text{Rn} + {}^{4}_{2}\alpha$
d. ${}^{73}_{30}\text{Zn} \rightarrow {}^{73}_{31}\text{Ga} + {}^{0}_{-1}\text{e}$

Q10.1.4

a. ${}^{198}_{79}\text{Au} \rightarrow {}^{198}_{80}\text{Hg} + {}^{0}_{-1}\text{e}$, beta
b. ${}^{57}_{27}\text{Co} + {}^{0}_{-1}\text{e} \rightarrow {}^{57}_{26}\text{Fe}$, electron capture
c. ${}^{230}_{92}\text{U} \rightarrow {}^{226}_{90}\text{Th} + {}^{4}_{2}\text{He}$, alpha
d. ${}^{128}_{56}\text{Ba} \rightarrow {}^{128}_{55}\text{Cs} + {}^{0}_{1}\text{e}$, positron
e. ${}^{131}_{53}\text{I} \rightarrow {}^{131}_{54}\text{Xe} + {}^{0}_{-1}e$, beta
f. ${}^{239}_{94}\text{Pu} \rightarrow {}^{235}_{92}\text{U} + {}^{4}_{2}\alpha$ (or can show as ${}^{4}_{2}\text{He}$), alpha

Q10.1.5

a. ${}^{220}_{87}\text{Fr} \rightarrow {}^{4}_{2}\text{He} + {}^{216}_{85}\text{At}$
b. ${}^{76}_{33}\text{As} \rightarrow {}^{0}_{-1}e + {}^{76}_{34}\text{Se}$
c. ${}^{231}_{92}\text{U} + {}^{0}_{-1}e \rightarrow {}^{231}_{91}\text{Pa}$
d. ${}^{143}_{61}\text{Pm} \rightarrow {}^{0}_{1}e + {}^{143}_{60}\text{Nd}$

10.2: Fission and Fusion

Q10.2.1

During fission, big nuclei split into smaller nuclei. During fusion, nuclei combine to form large nuclei.

Q10.2.2

Fission in nuclear power plants is controlled through limiting the availability of neutrons. Nuclear weapons are uncontrolled once the process initiates.

Q10.2.3

Diagnostic amounts are much smaller than therapeutic amounts.

10.3: Half-Life

Q10.3.1

1 half-life: 50%

3 half-lives: 12.5%

Q10.3.2

Time	Half-lives	Amount			
0 minutes		0.540 mg			
3 minutes	1	0.270 mg			
6 minutes	2	0.135 mg			
9 minutes	3	0.0675 mg			

Q10.3.3

Amount	Half-lives	Time			
4.48 mg		0 years			
2.24 mg	1	12.26 years			
1.12 mg	2	24.52 years			
0.560 mg	3	36.78 years			
0.280 mg	4	49.04 years			

Q10.3.4

Time	Half-lives	Amount			
0 hours		0.812 mg			
6.69 hours	1	0.406 mg			
13.38 hours	2	0.203 mg			
20.07 hours	3	0.102 mg			
26.76 hours	4	0.0508 mg			
33.45 hours	5	0.0254 mg			
40.14 hours	6	0.0127 mg			

Q10.3.5

Amount	Half-lives				
2.86 g					
1.43 g	1				
0.715 g	2				
0.358 g	3				

It takes three half-lives to go from 2.86 g to 0.358 g in a total time of 22.8 minutes.

$22.8\ min\ \div\ 3\ = 7.60\ min$

One half-life is 7.60 minutes.

Q10.3.6

Amount	Half-lives	Time			
100. mg		0 years			
50.0 mg	1	5730 years			
25.0 mg	2	11460 years			
12.5 mg	3	17190 years			
6.25 mg	4	22920 years			

Q10.3.7

Amount	Half-lives				
55.9 g					
28.0 g	1				
14.0 g	2				
6.99 g	3				

It takes three half-lives to go from 55.9 g to 6.99 g in a total time of 72.5 hours.

$72.5\ hr\ \div\ 3\ = 24.2\ hr$

One half-life is 24.2 hours.

Q10.3.8

Fill in the time and half-lives from top to bottom. Start at the bottom of the amount column to fill it in because we know where we end up but not where we started.

Time	Half-lives	Activity			
0 hours		406 mCi			
9.0 hours	1	203 mCi			
18 hours	2	102 mCi			
27 hours	3	50.8 mCi			
36 hours	4	25.4 mCi	← **START HERE**		

Q10.3.9

$$125\ mCi\left(\frac{5.0\ mL}{45\ mCi}\right) = 14\ mL$$

Q10.3.10

Sodium-24 is used to treat leukemia. A 36-kg patient is prescribed 145 μCi/kg and it is supplied to the hospital in a vial containing 250 μCi/mL. What volume should be given to the patient?

$$36\ kg\left(\frac{145\ \mu Ci}{kg}\right)\left(\frac{1\ mL}{250\ \mu Ci}\right) = 21\ mL$$

Q10.3.11

$21\ mL\left(\frac{250\ \mu Ci}{mL}\right) = 5250\ \mu Ci$is the total dose received

Amount	Half-lives	Time			
5250 μCi		0 hours			
2625 μCi	1	15 hours			
1313 μCi	2	30 hours			
656 μCi	3	45 hours			
328 μCi	4	60 hours			
164 μCi	5	75 hours			
82 μCi	6	90 hours			

Q10.3.12

Lead-212 is one of the radioisotopes used in the treatment of breast cancer. A patient needs a 15 μCi dose and it is supplied as a solution with a concentration of 2.5 μCi/mL. What volume does the patient need? Given the half-life of lead is 10.6 hours, what will be the radioactivity of the sample after approximately four days?

Volume given: $15\ \mu Ci\left(\frac{1\ mL}{2.5\ \mu Ci}\right) = 6.0\ mL$

Elapsed time in hours: $4\ days\left(\frac{24\ hr}{day}\right) = 96\ hr$

Time	Half-lives	Activity			
0 hours		15 μCi			
10.6 hours	1	7.5 μCi			
21.2 hours	2	3.8 μCi			
31.8 hours	3	1.9 μCi			
42.4 hours	4	0.94 μCi			
53.0 hours	5	0.47 μCi			
63.6. hours	6	0.23 μCi			
74.2 hours	7	0.12 μCi			
84.8 hours	8	0.059 μCi			
95.6 hours	9	0.029 μCi			

10.4: Physical and Chemical Changes

Q10.4.1

a. physical
b. physical
c. chemical
d. physical
e. chemical
f. chemical
g. physical
h. chemical

Q10.4.2

Any two from change in color, formation of gas (i.e. bubbles), formation of precipitate, odor, change in temperature.

Q10.4.3

chemical composition (i.e. chemical formula is the same)

10.5: Chemical Equations

Q10.5.1

a. reactants: carbon dioxide and water; products: glucose and oxygen
b. reactants: magnesium and oxygen; product: magnesium oxide

Q10.5.2

Descriptions may vary.

a. Two moles of liquid hydrogen peroxide decomposes to form two moles of liquid water and one mole of gaseous hydrogen.
b. One mole of solid copper(II) carbonate decomposes to form one mole each of solid copper(II) oxide and gaseous carbon dioxide.
c. Two moles of solid cesium react with 2 moles of liquid water to form 2 moles of aqueous cesium hydroxide and 1 mole of gaseous hydrogen.

Q10.5.3

a. 10 Br
b. 2 N, 6 H
c. 8 N, 32 H, 4 S, 16 O
d. 4 C, 8 H, 4 O
e. 3 Fe, 9 N, 27 O
f. 6 K, 2 P, 8 O

Q10.5.4

a. $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
b. $6\ Li(s) + N_2(g) \rightarrow 2\ Li_3N(s)$
c. $Ca(OH)_2 + 2\ HBr \rightarrow CaBr_2 + 2\ H_2O$
d. $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$
e. $2\ NH_3 + 3\ CuO \rightarrow 3\ Cu + N_2 + 3\ H_2O$

Q10.5.5

a. $2\ Fe(s) + 3\ Cl_2(g) \rightarrow 2\ FeCl_3(g)$
b. $C_4H_{10}O + 6\ O_2 \rightarrow 4\ CO_2 + 5\ H_2O$
c. $2\ As + 6\ NaOH \rightarrow 2\ Na_3AsO_3 + 3\ H_2$
d. $SiO_2 + 4\ HF \rightarrow SiF_4 + 2\ H_2O$
e. $2\ N_2 + 5\ O_2 + 2\ H_2O \rightarrow 4\ HNO_3$

CHAPTER OVERVIEW

11: Properties of Reactions

11.1: Oxidation Numbers

Learning Outcomes

- Assign oxidation numbers to free elements or elements in a compound or ion.

Oxidation Numbers

An **oxidation number** *is a positive or negative number that is assigned to an atom to indicate its degree of oxidation or reduction*. The term oxidation state is often used interchangeably with oxidation number. A partial electron transfer is a shift in the electron density near an atom as a result of a change in the other atoms to which it is covalently bonded. That charge shift is based on the relative electronegativities of the atoms involved in the bond.

Overall, the oxidation number of an atom in a molecule is the charge that the atom would have if all polar covalent and ionic bonds resulted in a complete transfer of electrons from the less electronegative atom to the more electronegative one. Oxidation numbers can be assigned using the set of rules outlined below.

1. The oxidation number of an atom in a neutral free element is zero. A free element is considered to be any element in an uncombined state, whether monatomic or polyatomic. For example, the oxidation number of each atom in Fe, Li, N_2, Ar, and P_4 is zero.
2. The oxidation number of a monatomic (composed of one atom) ion is the same as the charge of the ion. For example, the oxidation numbers of K^+, Se^{2-}, and Au^{3+} are $+1$, -2, and $+3$, respectively.
3. The oxidation number of oxygen in most compounds is -2.
4. The oxidation number of hydrogen in most compounds is $+1$.
5. The oxidation number of fluorine in all compounds is -1. Other halogens usually have an oxidation number of -1 in binary compounds, but can have variable oxidation numbers depending on the bonding environment.
6. In a neutral molecule, the sum of the oxidation numbers of all atoms is zero. For example, in H_2O, the oxidation numbers of H and O are $+1$ and -2, respectively. Because there are two hydrogen atoms in the formula, the sum of all the oxidation numbers in H_2O is $2\,(+1) + 1\,(-2) = 0$.
7. In a polyatomic ion, the sum of the oxidation numbers of all atoms is equal to the overall charge on the ion. For example, in SO_4^{2-}, the oxidation numbers of S and O are $+6$ and -2, respectively. The sum of all oxidation numbers in the sulfate ion would be $1\,(+6) + 4\,(-2) = -2$, which is the charge of the ion.

An examination of the rules for assigning oxidation numbers reveals that there are many elements for which there are no specific rules, such as nitrogen, sulfur, and chlorine. These elements, as well as some others, can have variable oxidation numbers depending on the other atoms to which they are covalently bonded in a molecular compound. It is useful to analyze a few molecules in order to see the strategy to follow in assigning oxidation numbers to other atoms.

Oxidation numbers for the atoms in a binary ionic compound are easy to assign because they are equal to the charge of the ion (rule 2). In FeCl_3, the oxidation number of iron is $+3$, while the oxidation number of chlorine is -1. In Ca_3P_2, the calcium is $+2$, while the phosphorus is -3. This is because an ionic compound is in the form of a crystal lattice that is actually composed of these ions.

Assigning oxidation numbers for molecular compounds is trickier. The key is to remember rule 6: that the sum of all the oxidation numbers for any neutral species must be zero. Make sure to account for any subscripts which appear in the formula. As an example, consider the compound nitric acid, HNO_3. According to rule 4, the oxidation number of hydrogen is $+1$. According to rule 3, the oxidation number of oxygen is -2. There is no rule regarding nitrogen, but its oxidation number can be calculated as follows.

$$1\,(+1) + x + 3\,(-2) = 0,\ \text{where } x \text{ is the oxidation number of nitrogen} \tag{11.1.1}$$

$$\text{Solving: } x = 0 - 1 - (-6) = +5 \tag{11.1.2}$$

The oxidation number of the nitrogen atom in HNO_3 is $+5$. Often when assigning oxidation numbers, it is convenient to write it above the symbol within the formula.

$$\overset{+1}{\text{H}}\overset{+5}{\text{N}}\overset{-2}{\text{O}}_3 \tag{11.1.3}$$

You may wonder if there are any limits on the value of oxidation numbers. The key point to consider is the octet rule. Since nitrogen has 5 valence electrons, the most that it can "lose" while forming bonds in a molecule is 5, so its highest possible oxidation number is $+5$. Alternatively, it could gain up to 3 electrons, and so its lowest (most negative) possible oxidation number is -3. Similarly, chlorine can have oxidation numbers ranging from -1 to $+7$.

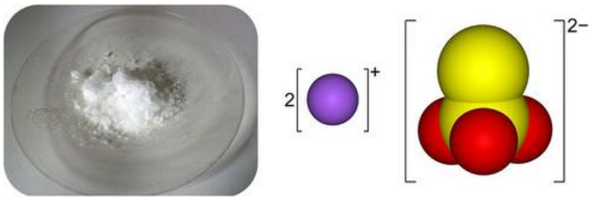

Figure 11.1.1: Sodium thiosulfate is a white crystalline compound (left) composed of two sodium ions $\left(\mathrm{Na^+}\right)$ for every one thiosulfate ion $\left(\mathrm{S_2O_3^{2-}}\right)$ (right).

Now consider the ionic compound sodium thiosulfate, $\mathrm{Na_2S_2O_3}$ (Figure 11.1.1). It contains the thiosulfate polyatomic ion, $\mathrm{S_2O_3^{2-}}$. The sodium is not part of the covalently bonded polyatomic ion, and so its oxidation number is the same as it would be in a binary ionic compound, $+1$. The sulfur is the atom whose oxidation number is not covered by one of the rules. The oxidation number of sulfur is assigned the variable x in the following calculation. Remember the sum of the oxidation numbers of all the elements must equal zero because $\mathrm{Na_2S_2O_3}$ is a neutral compound.

$$2\left(+1\right)+2\left(x\right)+3\left(-2\right) = 0 \tag{11.1.4}$$

$$2+2x-6 = 0 \tag{11.1.5}$$

$$-4+2x = 0 \tag{11.1.6}$$

$$2x = +4 \tag{11.1.7}$$

$$x = +2 \tag{11.1.8}$$

Sulfur has an oxidation number of $+2$ in $\mathrm{Na_2S_2O_3}$. Notice how the subscript of 2 for the S atom had to be accounted for by dividing the result of the subtraction by 2. When assigning oxidation numbers, you do so for each individual atom. In the above example, the oxidation number of sulfur could also have been determined by looking at just the thiosulfate ion, $\mathrm{S_2O_3^{2-}}$.

$$2\left(x\right)+3\left(-2\right) = -2 \tag{11.1.9}$$

$$2x-6 = -2 \tag{11.1.10}$$

$$2x = +4 \tag{11.1.11}$$

$$x = +2 \tag{11.1.12}$$

Contributors and Attributions

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11.2: The Nature of Oxidation and Reduction

Learning Outcomes

- Define oxidation and reduction in terms of a gain or loss of oxygen or hydrogen atoms.
- Define oxidation and reduction in terms of a gain or loss of electrons.
- Identify the substances involved in oxidation and reduction in a reaction.
- Identify the oxidizing and reducing agents in a redox reaction.
- Write the oxidation or reduction sequence for a molecule (alcohol, aldehyde/ketone, carboxylic acid)

Oxygen is an element that has been known for centuries. In its pure elemental form, oxygen is highly reactive, and it readily makes compounds with most other elements. It is also the most abundant element by mass in the Earth's crust. The class of reactions called oxidation and reduction were originally defined with respect to the element oxygen.

Oxygen's Role in Reactions

Many elements simply combine with oxygen to form the oxide of that element. Heating magnesium in air allows it to combine with oxygen to form magnesium oxide.

$$2\mathrm{Mg}\,(s)+\mathrm{O_2}\,(g)\rightarrow 2\mathrm{MgO}\,(s) \tag{11.2.1}$$

Many compounds react with oxygen as well, often in very exothermic processes that are generally referred to as combustion reactions. For example, when methane burns, carbon dioxide and water are produced.

$$\mathrm{CH_4}\,(g)+2\mathrm{O_2}\,(g)\rightarrow \mathrm{CO_2}\,(g)+2\mathrm{H_2O}\,(g) \tag{11.2.2}$$

Carbon dioxide is an oxide of carbon, while water is an oxide of hydrogen. Early scientists viewed oxidation as a process in which a substance was reacted with oxygen to produce one or more oxides. In the previous examples, magnesium and methane are being oxidized.

Oxidation does not necessarily require heating. Iron that is exposed to air and water slowly oxidizes in a process commonly known as rusting. Bleaches contain various compounds such as sodium hypochlorite (NaClO) that can oxidize stains by the transfer of oxygen atoms, making the molecules in the stains more water-soluble and therefore easier to rinse off. Hydrogen peroxide $(\mathrm{H_2O_2})$ releases oxygen as it spontaneously decomposes. It also acts as a bleach and is ued as an antiseptic that kills bacteria by oxidizing them.

The opposite of oxidation is called reduction. Since oxidation was originally defined as the addition of oxygen, reduction was therefore the removal of oxygen from a substance. Many naturally occurring metal ores are present as oxides. The pure metals can be extracted by reduction. For example, pure iron is obtained from iron (III) oxide by reacting it with carbon at high temperatures.

$$2\mathrm{Fe_2O_3}\,(s)+3\mathrm{C}\,(s)\rightarrow 4\mathrm{Fe}\,(s)+3\mathrm{CO_2}\,(g) \tag{11.2.3}$$

The removal of oxygen from the $\mathrm{Fe_2O_3}$ molecule means that it is being reduced to Fe. Note that an oxidation process is also occurring simultaneously in this reaction; the carbon reactant is being oxidized to $\mathrm{CO_2}$. This is an important concept. In looking at organic compounds, we can describe oxidation and reduction in terms of the loss or gain of oxygen or the loss or gain of hydrogen.

If we look at the oxidation of ethanol ($\mathrm{CH_3CH_2OH}$, an alcohol) to form ethanal ($\mathrm{CH_3CH_2CHO}$, an aldehyde), we see that the number of bonds to oxygen has increased and the number of hydrogen atoms has decreased from six to four. Either or both of these indicate that an oxidation has occurred.

For a reduction, we look for the removal of oxygen or bonds to oxygen or the addition or hydrogen atoms. In the following reaction, we see ethanoic acid ($\mathrm{CH_3COOH}$, a carboxylic acid) being reduced to ethanal ($\mathrm{CH_3CH_2CHO}$, an aldehyde). Note that the number of oxygen atoms is reduced from two to one. There is no change in the number of hydrogen atoms, but both changes do not need to happen for evidence of oxidation or reduction.

If we see evidence of an oxidation or reduction, we know the other must have happened as well. Oxidation and reduction must happen together. Neither can happen alone in a reaction.

The Electron in Redox Reactions

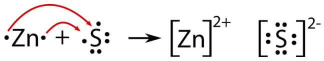
The definitions of oxidation and reduction were eventually broadened to include similar types of reactions that do not necessarily involve oxygen. Oxygen is more electronegative than any element except for fluorine. Therefore, when oxygen is bonded to any element other than fluorine, electrons from the other atom are shifted away from that atom and toward the oxygen atom. An **oxidation-reduction reaction** (sometimes abbreviated as a **redox reaction**) *is a reaction that involves the full or partial transfer of electrons from one reactant to another*. **Oxidation** *involves a full or partial loss of electrons*, while **reduction** *involves a full or partial gain of electrons*.

An easy saying to remember the definitions of oxidation and reduction is "LEO the lion says GER" (see figure below)! LEO stands for Losing Electrons is Oxidation, while GER stands for Gaining Electrons is Reduction.

Figure 11.2.1: This guy knows his redox reactions.

Redox Reactions and Ionic Compounds

In the course of a chemical reaction between a metal and a nonmetal, electrons are transferred from the metal atoms to the nonmetal atoms. For example, when zinc metal is heated in the presence of sulfur, the compound zinc sulfide is produced (see figure below). Two valence electrons from each zinc atom are transferred to each sulfur atom.



Since zinc is losing electrons during this reaction, it is being oxidized. The sulfur is gaining electrons, so it is being reduced. Each of these processes can be shown in a separate equation called a half-reaction. A **half-reaction** *is an equation that shows either the oxidation or the reduction reaction that occurs during a redox reaction.*

$$\text{Oxidation: } \mathrm{Zn} \rightarrow \mathrm{Zn}^{2+} + 2\mathrm{e}^- \tag{11.2.4}$$

$$\text{Reduction: } \mathrm{S} + 2\mathrm{e}^- \rightarrow \mathrm{S}^{2-} \tag{11.2.5}$$

It is important to remember that the two half-reactions occur simultaneously. The resulting ions that are formed are then attracted to one another, forming an ionic bond.

Redox reactions are also commonly described in terms of oxidizing and reducing agents. In the reaction above, the zinc is being oxidized by losing electrons. However, an isolated ionization would be a very high-energy process, so there must be another substance present to gain those lost electrons. In this case, the electrons are gained by sulfur. In other words, the sulfur is causing the zinc to be oxidized. Consequently, sulfur is referred to as the oxidizing agent. Conversely, the zinc causes the sulfur to gain electrons and become reduced, so zinc is the reducing agent. An **oxidizing agent** *is a substance that causes oxidation by accepting electrons*, and a **reducing agent** *is a substance that causes reduction by losing electrons*. Said another way, the oxidizing agent is the substance that is reduced, while the reducing agent is the substance that is oxidized.

Example 11.2.1

When chlorine gas is bubbled into a solution of sodium bromide, a reaction occurs that produces aqueous sodium chloride and elemental bromine.

a. Determine what is being oxidized and what is being reduced.
b. Identify the oxidizing and reducing agents.

Solution

First, look at each element in the reaction and assign oxidation numbers to each element.

$$\mathrm{Cl_2}\,(g) + 2\mathrm{NaBr}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{Br_2}\,(l) \tag{11.2.6}$$

Chlorine in $\mathrm{Cl_2}$ and bromine in $\mathrm{Br_2}$ both have oxidation numbers of 0. NaBr and NaCl are ionic compounds so their ionic charges will be their oxidation numbers.

$$\overset{0}{\mathrm{Cl}}_2\,(g) + 2\overset{+1}{\mathrm{Na}}\overset{-1}{\mathrm{Br}}\,(aq) \rightarrow 2\overset{+1}{\mathrm{Na}}\overset{-1}{\mathrm{Cl}}\,(aq) + \overset{0}{\mathrm{Br}}_2\,(l) \tag{11.2.7}$$

a. The oxidation state of chlorine goes from 0 to -1 so it is gaining an electron and being reduced. The oxidation state of bromine goes from -1 to 0 so it is losing an electron and being oxidized. The oxidation state of sodium does not change so it is not involved in either the oxidation or the reduction.

b. The $\mathrm{Cl_2}$ is being reduced and is the oxidizing agent. The Br^- is being oxidized and is the reducing agent.

Figure 11.2.2: Rusting is a form of corrosion, a process in which metals are oxidized.

Redox Reactions and Molecular Compounds

The loss or gain of electrons is easy to see in a reaction in which ions are formed. However, in many reactions, electrons are not transferred completely. Recall that in a molecular compound, electrons are shared between atoms in a type of bond called a covalent bond. Even though electrons are not completely transferred between atoms, it is still common for reactions involving molecular compounds to be classified as redox reactions.

For example, when hydrogen gas is reacted with oxygen gas, water is formed as the product.

$$2\mathrm{H_2}\,(g) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{H_2O}\,(l) \tag{11.2.8}$$

We assign oxidation numbers to determine which species are gaining and losing electrons. Remember oxidation numbers tell us what the charge would be if the electrons were transferred completely even though they are shared, not transferred, in covalent bonds. For the synthesis of water, the oxidation numbers of the free elements (hydrogen and oxygen) are zero. For water, oxygen has a -2 oxidation number while hydrogen has a $+1$.

$$2\overset{0}{\mathrm{H}}_2\,(g) + \overset{0}{\mathrm{O}}_2\,(g) \rightarrow 2\overset{+1}{\mathrm{H}}_2\overset{-2}{\mathrm{O}}\,(l) \tag{11.2.9}$$

The oxidation state from hydrogen goes from 0 to $+1$, indicating oxidation. The oxidation state of oxygen goes from 0 to -2, indicating reduction. Therefore hydrogen is oxidized and is the reducing agent. Oxygen is *reduced* and is the *oxidizing agent*.

Gain or Loss of Hydrogen and Oxygen

Oxidation can also be defined as the addition of oxygen to a molecule or the removal of hydrogen. Reduction is therefore the addition of hydrogen or the removal of oxygen. Several classes of organic compounds are related to one another by oxidation and reduction reactions. Alkanes, alkenes, and alkynes represent different levels of oxidation of a hydrocarbon. When an alkane is heated in the presence of an appropriate catalyst, it can be oxidized to the corresponding alkene in a reaction called a dehydrogenation reaction. Two hydrogen atoms are removed in the process. The alkene can be further oxidized to an alkyne by the removal of two more hydrogen atoms. This is also considered an oxidation according to the modern definition, because the oxidation number of each carbon atom goes from -3 to -2 to -1.

$$\text{oxidation: } \mathrm{CH_3CH_3} \overset{-\mathrm{H_2}}{\rightarrow} \mathrm{CH_2{=}CH_2} \overset{-\mathrm{H_2}}{\rightarrow} \mathrm{CH\equiv CH} \tag{11.2.10}$$

The reactions are reversible, so an alkyne can be reduced first to an alkene and then to an alkane. These are addition reactions, as seen in the previous section.

$$\text{reduction: } \mathrm{CH\equiv CH} \overset{+\mathrm{H_2}}{\rightarrow} \mathrm{CH_2{=}CH_2} \overset{+\mathrm{H_2}}{\rightarrow} \mathrm{CH_3CH_3} \tag{11.2.11}$$

The alkane is the most reduced form of a hydrocarbon, while the alkyne is the most oxidized form.

The changes in electrons, oxidation number, number of oxygen atoms, and number of hydrogen atoms is summarized below for oxidation and reduction reactions.

Oxidation reactions in organic chemistry often involve the addition of oxygen to a compound (or an increase in the number of bonds to oxygen), which changes the functional group that is present. The following sequence shows how methane can be oxidized first to methanol (alcohol), then to methanal (aldehyde), then to methanoic acid (carboxylic acid), and finally to carbon dioxide. Each step in the process is either a gain of oxygen or a loss of hydrogen. Each step also represents energy, which explains why the complete combustion of alkanes to carbon dioxide is an extremely exothermic reaction.

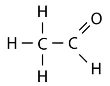
$$\underset{\text{alkane}}{\mathrm{CH_4}} \overset{\text{gain of oxygen}}{\rightarrow} \underset{\text{alcohol}}{\mathrm{CH_3OH}} \overset{\text{loss of hydrogen}}{\rightarrow} \underset{\text{aldehyde}}{\mathrm{CH_2O}} \overset{\text{gain of oxygen}}{\rightarrow} \underset{\text{carboxylic acid}}{\mathrm{HCOOH}} \overset{\text{loss of hydrogen}}{\rightarrow} \underset{\text{carbon dioxide}}{\mathrm{CO_2}} \tag{11.2.12}$$

The opposite process is the reduction of a carboxylic acid to an alkane which involves the loss of oxygen and the gain of hydrogen.

$$\underset{\text{alkane}}{\mathrm{CH_4}} \overset{\text{loss of oxygen}}{\leftarrow} \underset{\text{alcohol}}{\mathrm{CH_3OH}} \overset{\text{gain of hydrogen}}{\leftarrow} \underset{\text{aldehyde}}{\mathrm{CH_2O}} \overset{\text{loss of oxygen}}{\leftarrow} \underset{\text{carboxylic acid}}{\mathrm{HCOOH}} \tag{11.2.13}$$

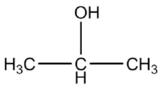
The oxidation of an alcohol can produce either an aldehyde or a ketone. A primary alcohol oxidized to form an aldehyde.

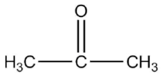
$$\mathrm{CH_3CH_2OH} \underset{\mathrm{H^+}}{\overset{\mathrm{Cr_2O_7^{2-}}}{\rightarrow}} \mathrm{CH_3CHO} \tag{11.2.14}$$



A secondary alcohol is oxidized to a ketone rather than an aldehyde. The oxidation of the simplest secondary alcohol, 2-propanol, yields propanone.

$$\mathrm{CH_3CHOHCH_3} \underset{\mathrm{H^+}}{\overset{\mathrm{Cr_2O_7^{2-}}}{\rightarrow}} \mathrm{CH_3COCH_3} \tag{11.2.15}$$





Tertiary alcohols cannot be oxidized in this way because the carbon to which the hydroxyl group is attached does not have another hydrogen atom attached to it and cannot undergo oxidation.

When a primary alcohol is oxidized to an aldehyde in the presence of water, the reaction can be difficult to stop because the aldehyde can be further oxidized to the corresponding carboxylic acid. For example, the oxidation of ethanal produces ethanoic (acetic) acid. Ethanol-containing beverages such as wine are susceptible to such oxidation if kept for long periods of time after having been opened and exposed to the air. Wine that has become oxidized will have an unpleasant vinegary taste due to the production of acetic acid. Unlike aldehydes, ketones are resistant to further oxidation because the carbonyl group is in the middle of the carbon chain, so the ketone cannot be converted to a carboxylic acid.

$$\mathrm{CH_3CHO} \underset{\mathrm{H^+}}{\overset{\mathrm{Cr_2O_7^{2-}}}{\rightarrow}} \mathrm{CH_3COOH} \tag{11.2.16}$$

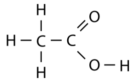


Table 11.2.1

Process	Changes in electrons (always)	Change in oxidation number (always)	Change in oxygen (some reactions)	Change in hydrogen (some reactions)
Oxidation	lose	increase	gain	lose
Reduction	gain	decrease	lose	gain

Figure 11.2.3: The sequence of oxidation and reduction in oxygen- and nitrogen-containing compounds.

Contributors and Attributions

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11.3: Types of Inorganic Reactions

Learning Outcomes

- Classify a reaction as combination, decomposition, single-replacement, double-replacement, or combustion.
- Determine whether a reaction is an oxidation-reduction reaction.
- Predict the products and balance a combustion reaction.

Many chemical reactions can be classified as one of five basic types. Having a thorough understanding of these types of reactions will be useful for predicting the products of an unknown reaction. The five basic types of chemical reactions are combination, decomposition, single-replacement, double-replacement, and combustion. Analyzing the reactants and products of a given reaction will allow you to place it into one of these categories. Some reactions will fit into more than one category. Identifying a reaction as one of these types does not preclude it from also being an oxidation-reduction (redox) reaction. For example, many combination reactions can also be classified as a redox reaction.

Combination Reactions

A **combination reaction**, also known as a **synthesis reaction**, *is a reaction in which two or more substances combine to form a single new substance.* Combination reactions can also be called synthesis reactions .The general form of a combination reaction is:

$$\mathrm{A+B\rightarrow AB} \tag{11.3.1}$$

One example of a combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to product solid sodium chloride.

$$\mathrm{2Na}\,(s)+\mathrm{Cl_2}\,(g)\rightarrow \mathrm{2NaCl}\,(s) \tag{11.3.2}$$

This reaction can also be classified as a redox reaction due to the changes in oxidation states. Sodium goes from a 0 to +1 oxidation state while chlorine goes from a 0 to a -1 oxidation state.

Notice that in order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2).

One sort of combination reaction that occurs frequently is the reaction of an element with oxygen to form an oxide. Metals and nonmetals both react readily with oxygen under most conditions. Magnesium reacts rapidly and dramatically when ignited, combining with oxygen from the air to produce a fine powder of magnesium oxide.

$$\mathrm{2Mg}\,(s)+\mathrm{O_2}\,(g)\rightarrow \mathrm{2MgO}\,(s) \tag{11.3.3}$$

Decomposition Reactions

A **decomposition reaction** *is a reaction in which a compound breaks down into two or more simpler substances.* The general form of a decomposition reaction is:

$$\mathrm{AB\rightarrow A+B} \tag{11.3.4}$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas.

$$\mathrm{2HgO}\,(s)\rightarrow \mathrm{2Hg}\,(l)+\mathrm{O_2}\,(g) \tag{11.3.5}$$

A reaction is also considered to be a decomposition reaction even when one or more of the products is still a compound. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide.

$$\mathrm{CaCO_3}\,(s)\rightarrow \mathrm{CaO}\,(s)+\mathrm{CO_2}\,(g) \tag{11.3.6}$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water.

$$2\text{NaOH}\,(s) \rightarrow \text{Na}_2\text{O}\,(s) + \text{H}_2\text{O}\,(g) \tag{11.3.7}$$

Single-Replacement Reactions

A **single-replacement reaction**, also known as single-displacement or substitution reaction, *is a reaction in which one element replaces a similar element in a compound*. The general form of a single-replacement reaction is:

$$\text{A} + \text{BC} \rightarrow \text{AC} + \text{B} \tag{11.3.8}$$

In this general reaction, element A is a metal and replaces element B, also a metal, in the compound. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$\text{Y} + \text{XZ} \rightarrow \text{XY} + \text{Z} \tag{11.3.9}$$

Y is a nonmetal and replaces the nonmetal Z in the compound with X.

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$\text{Mg}\,(s) + \text{Cu(NO}_3)_2\,(aq) \rightarrow \text{Mg(NO}_3)_2\,(aq) + \text{Cu}\,(s) \tag{11.3.10}$$

Many metals react easily with acids, and, when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (see figure below).

$$\text{Zn}\,(s) + 2\text{HCl}\,(aq) \rightarrow \text{ZnCl}_2\,(aq) + \text{H}_2\,(g) \tag{11.3.11}$$

Figure 11.3.1: Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-replacement reaction.

Figure 11.3.2: Pictured here is about 3 pounds of sodium metal reacting with water. Sodium metal reacts vigorously when dropped into a container of water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite.

Double-Replacement Reactions

A **double-replacement reaction**, also known as double-displacement, *is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds*. The general form of a double-replacement reaction is:

$$\mathrm{AB} + \mathrm{CD} \rightarrow \mathrm{AD} + \mathrm{CB} \tag{11.3.12}$$

In this reaction, A and C are positively-charged cations, while B and D are negatively-charged anions. Double-replacement reactions generally occur between substances in aqueous solution. In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water.

A precipitate forms in a double-replacement reaction when the cations from one of the reactants combine with the anions from the other reactant to form an insoluble ionic compound. When aqueous solutions of potassium iodide and lead (II) nitrate are mixed, the following reaction occurs.

$$2\mathrm{KI}\,(aq) + \mathrm{Pb(NO_3)_2}\,(aq) \rightarrow 2\mathrm{KNO_3}\,(aq) + \mathrm{PbI_2}\,(s) \tag{11.3.13}$$

Figure 11.3.3: When a few drops of lead (II) nitrate are added to a solution of potassium iodide, a yellow precipitate of lead (II) iodide immediately forms in a double-replacement reaction.

Combustion Reactions

A **combustion reaction** *is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat.* Combustion reactions must involve $\mathrm{O_2}$ as one reactant. The combustion of hydrogen gas produces water vapor (see figure below).

$$2\mathrm{H_2}\,(g) + \mathrm{O_2}\,(g) + 2\mathrm{H_2O}\,(g) \tag{11.3.14}$$

Notice that this reaction also qualifies as a combination reaction.

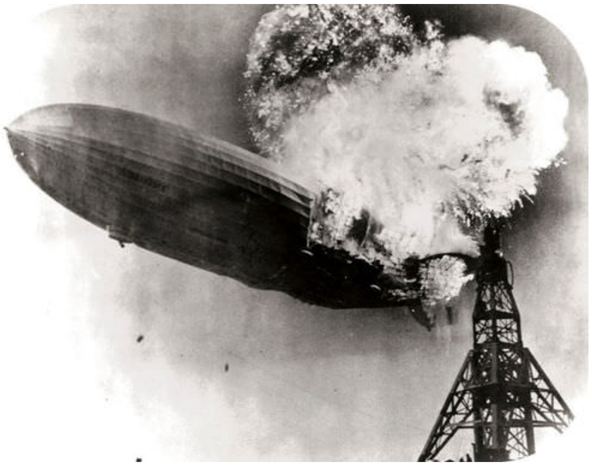

Figure 11.3.4: The Hindenburg *was a hydrogen-filled airship that suffered an accident upon its attempted landing in New Jersey in 1937. The hydrogen immediately combusted in a huge fireball, destroying the airship and killing 36 people. The chemical reaction was a simple one: hydrogen combining with oxygen to produce water.*

Many combustion reactions occur with a hydrocarbon, a compound made up solely of carbon and hydrogen. The products of the combustion of hydrocarbons are always carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amount of heat energy. Propane $\left(\mathrm{C_3H_8}\right)$ is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$\mathrm{C_3H_8}\,(g) + 5\mathrm{O_2}\,(g) \rightarrow 3\mathrm{CO_2}\,(g) + 4\mathrm{H_2O}\,(g) \tag{11.3.15}$$

Example 11.3.1

Ethanol can be used as a fuel source in an alcohol lamp. The formula for ethanol is $\mathrm{C_2H_5OH}$. Write the balanced equation for the combustion of ethanol.

Solution

Step 1: Plan the problem.

Ethanol and oxygen are the reactants. As with a hydrocarbon, the products of the combustion of an alcohol are carbon dioxide and water.

Step 2: Solve.

Write the skeleton equations: $\mathrm{C_2H_5OH}\,(l) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(g)$

Balance the equation.

$$\mathrm{C_2H_5OH}\,(l) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{CO_2}\,(g) + 3\mathrm{H_2O}\,(g) \tag{11.3.16}$$

Step 3: Think about your result.

Combustion reactions must have oxygen as a reactant. Note that the water that is produced is in the gas state rather that the liquid state because of the high temperatures that accompany a combustion reaction.

Supplemental Resources

- Simulation of the synthesis of water: www.dit.ncssm/edu/core/Chapte...Synthesis.html
- View the synthesis of calcium oxide at http://www.youtube.com/watch?v-dszSKIM5rqk
- View the reaction between copper and chlorine gas at http://www.youtube.com/watch?v-edLpxdERQZc
- Watch the decomposition of hydrogen peroxide at http://www.youtube.com/watch?v=oX5FyaqNx54
- Watch the decomposition of potassium chlorate at http://www.youtube.com/watch?v=svRIg_kzE68
- A video experiment of hydrogen replacement by calcium can be seen at http://www.youtube.com/watch?v=hjB96do_fRw
- A video experiment of magnesium metal reacting with hydrochloric acid can be viewed at http://www.youtube.com/watch?v=OBdgeJFzSec

- Watch an animation of a double-replacement reaction at http://www.dlt.ncssm.edu/core/Chapte...cidToBase.html
- A video experiment in which a precipitate is formed by reacting sodium chloride with silver nitrate can be seen at http://www.youtube.com/watch?v=eFF3El4mwok
- A video experiment of the double-replacement reaction between copper (II) sulfate and sodium sulfide can be viewed at http://www.youtube.com/watch?v=KkKBDcFfZWo
- A video of the double-replacement reaction between sodium sulfate and barium chloride can be viewed at http://www.youtube.com/watch?v=XaMyfjYLhxU
- View an in-depth explanation of the combustion reaction that occurs when you strike a match at http://www.pbs.org/wgbh/nova/cigarette/onfire.html
- View an exciting video demonstration that confirms the importance of oxygen in a combustion reaction at education.jlab.org/frost/life_candle.html
- View another video demonstration that confirms the importance of oxygen in a combustion reaction at http://education.jlab.org/frost/combustion.html
- Reaction Identification Practice: http://www.sciencegeek.net/Chemistry...tification.htm

Contributors and Attributions

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11.4: Entropy and Enthalpy

Learning Outcomes

- Recall the meaning of exothermic and endothermic.
- Define entropy.
- Predict whether entropy change for a reaction is increasing or decreasing.

Previously, you learned that chemical reactions either absorb or release energy as they occur. The change in energy is one factor that allows chemists to predict whether a certain reaction will occur. In this lesson, you will learn about a second driving force for chemical reactions called entropy.

Enthalpy as a Driving Forces

The vast majority of naturally occurring reactions are exothermic. In an exothermic reaction, the reactants have a relatively high quantity of energy compared to the products. As the reaction proceeds, energy is released into the surroundings. Low energy can be thought of as providing a greater degree of stability to a chemical system. Since the energy of the system decreases during an exothermic reaction, the products of the system are more stable than the reactants. We can say that an exothermic reaction is an energetically favorable reaction.

If the drive toward lower energy were the only consideration for whether a reaction is able to occur, we would expect that endothermic reactions could never occur spontaneously. In an endothermic reaction, energy is absorbed during the reaction, and the products thus have a larger quantity of energy than the reactants. This means that the products are less stable than the reactants. Therefore, the reaction would not occur without some outside influence such as persistent heating. However, endothermic reactions do occur spontaneously, or naturally. There must be another driving force besides enthalpy change which helps promote spontaneous chemical reaction.

Entropy as a Driving Force

A very simple endothermic process is that of a melting ice cube. Energy is transferred from the room to the ice cube, causing it to change from the solid to the liquid state.

$$\mathrm{H_2O}\,(s) + 6.01\ \mathrm{kJ} \rightarrow \mathrm{H_2O}\,(l) \tag{11.4.1}$$

The solid state of water, ice, is highly ordered because its molecules are fixed in place. The melting process frees the water molecules from their hydrogen-bonded network and allows them a greater degree of movement. Water is more disordered than ice. The change from the solid to the liquid state of any substance corresponds to an increase in the disorder of the system.

There is a tendency in nature for systems to proceed toward a state of greater disorder or randomness. **Entropy** *is a measure of the degree of randomness or disorder of a system*. Entropy is an easy concept to understand when thinking about everyday situations. When the pieces of a jigsaw puzzle are dumped from the box, the pieces naturally hit the table in a very random state. In order to put the puzzle together, a great deal of work must be dome in order to overcome the natural entropy of the pieces. The entropy of a room that has been recently cleaned and organized is low. As time goes by, it likely will become more disordered, and thus its entropy will increase (see figure below). The natural tendency of a system is for its entropy to increase.

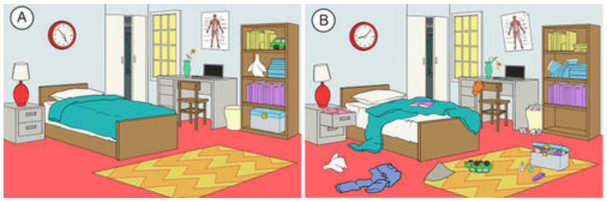

Figure 11.4.1: The messy room on the right has more entropy than the highly ordered room on the left. The drive toward an increase in entropy is the natural direction for all processes.

Chemical reactions also tend to proceed in such a way as to increase the total entropy of the system. How can you tell if a certain reaction shows an increase or a decrease in entropy? The states of the reactants and produces provide certain clues. The general cases below illustrate entropy at the molecular level.

1. For a given substance, the entropy of the liquid state is greater than the entropy of the solid state. Likewise, the entropy of the gas is greater than the entropy of the liquid. Therefore, entropy increases in processes in which solid or liquid reactants form gaseous products. Entropy also increases when solid reactants form liquid products.
2. Entropy increases when a substance is broken up into multiple parts. The process of dissolving increases entropy because the solute particles become separated from one another when a solution is formed.
3. Entropy increases as temperature increases. An increase in temperature means that the particles of the substance have greater kinetic energy. The faster moving particles have more disorder than particles that are moving more slowly at a lower temperature.
4. Entropy generally increases in reactions in which the total number of product molecules is greater than the total number of reactant molecules. An exception to this rule is when nongaseous products are formed from gaseous reactants.

The examples below will serve to illustrate how the entropy change in a reaction can be predicted.

$$\mathrm{Cl_2}\,(g) \rightarrow \mathrm{Cl_2}\,(l) \tag{11.4.2}$$

The entropy is decreasing because a gas is becoming a liquid.

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \tag{11.4.3}$$

The entropy is increasing because a gas is being produced, and the number of molecules is increasing.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightarrow 2\mathrm{NH_3}\,(g) \tag{11.4.4}$$

The entropy is decreasing because four total reactant molecules are forming two total product molecules. All are gases.

$$\mathrm{AgNO_3}\,(aq) + \mathrm{NaCl}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{AgCl}\,(s) \tag{11.4.5}$$

The entropy is decreasing because a solid is formed from aqueous reactants.

$$\mathrm{H_2}\,(g) + \mathrm{Cl_2}\,(g) \rightarrow 2\mathrm{HCl}\,(g) \tag{11.4.6}$$

The entropy change is unknown (but likely not zero) because there are equal numbers of molecules on both sides of the equation, and all are gases.

Contributors and Attributions

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

11.5: Spontaneous Reactions and Free Energy

Learning Outcomes

- Describe the meaning of a spontaneous reaction in terms of enthalpy and entropy changes.
- Define free energy.
- Determine the spontaneity of a reaction based on the value of its change in free energy at high and low temperatures.

The change in enthalpy and change in entropy of a reaction are the driving forces behind all chemical reactions. In this lesson, we will examine a new function called free energy, which combines enthalpy and entropy and can be used to determine whether or not a given reaction will occur spontaneously.

Spontaneous Reactions

A **spontaneous reaction** *is a reaction that favors the formation of products at the conditions under which the reaction is occurring*. A roaring bonfire (see figure below) is an example of a spontaneous reaction. A fire is exothermic, which means a decrease in the energy of the system as energy is released to the surroundings as heat. The products of a fire are composed mostly of gases such as carbon dioxide and water vapor, so the entropy of the system increases during most combustion reactions. This combination of a decrease in energy and an increase in entropy means that combustion reactions occur spontaneously.

Figure 11.5.1: Combustion reactions, such as this fire, are spontaneous reactions. Once the reaction begins, it continues on its own until one of the reactants (fuel or oxygen) is gone.

A **nonspontaneous reaction** *is a reaction that does not favor the formation of products at the given set of conditions*. In order for a reaction to be nonspontaneous, one or both of the driving forces must favor the reactants over the products. In other words, the reaction is endothermic, is accompanied by a decrease in entropy, or both. Out atmosphere is composed primarily of a mixture of nitrogen and oxygen gases. One could write an equation showing these gases undergoing a chemical reaction to form nitrogen monoxide.

$$\mathrm{N_2}\left(g\right)+\mathrm{O_2}\left(g\right)\rightarrow 2\mathrm{NO}\left(g\right) \tag{11.5.1}$$

Fortunately, this reaction is nonspontaneous at normal temperatures and pressures. It is a highly endothermic reaction with a slightly positive entropy change (ΔS). However, nitrogen monoxide is capable of being produced at very high temperatures, and this reaction has been observed to occur as a result of lightning strikes.

One must be careful not to confuse the term spontaneous with the notion that a reaction occurs rapidly. A spontaneous reaction is one in which product formation is favored, even if the reaction is extremely slow. You do not have to worry about a piece of paper on your desk suddenly bursting into flames, although its combustion is a spontaneous reaction. What is missing is the required activation energy to get the reaction started. If the paper were to be heated to a high enough temperature, it would begin to burn, at which point the reaction would proceed spontaneously until completion.

In a reversible reaction, one reaction direction may be favored over the other. Carbonic acid is present in carbonated beverages. It decomposes spontaneously to carbon dioxide and water according to the following reaction.

$$\mathrm{H_2CO_3}\left(aq\right)\rightleftharpoons \mathrm{CO_2}\left(g\right)+\mathrm{H_2O}\left(l\right) \tag{11.5.2}$$

If you were to start with pure carbonic acid in water and allow the system to come to equilibrium, more than 99% of the carbonic acid would be converted into carbon dioxide and water. The forward reaction is spontaneous because the products of the forward reaction are favored at equilibrium. In the reverse reaction, carbon dioxide and water are the reactants, and carbonic acid is the product. When carbon dioxide is bubbled into water (see figure below), less than 1% is converted to carbonic acid when the reaction reaches equilibrium. The reverse of the above reaction is not spontaneous. This illustrates another important point about spontaneity. Just because a reaction is not spontaneous does not mean that it does not occur at all. Rather, it means that the reactants will be favored over the products at equilibrium, even though some products may indeed form.

Figure 11.5.2: A home soda making machine is shown with a bottle of water and a $\mathrm{CO_2}$ cartridge. When the water is carbonated, only a small amount of carbonic acid is formed because the reaction is nonspontaneous. (Public Domain; Baruchlanda)

Gibbs Free Energy

Many chemical reactions and physical processes release energy that can be used to do other things. When the fuel in a car is burned, some of the released energy is used to power the vehicle. **Free energy** *is energy that is available to do work*. Spontaneous reactions release free energy as they proceed. Recall that the determining factors for spontaneity of a reaction are the enthalpy and entropy changes that occur for the system. The free energy change of a reaction is a mathematical combination of the enthalpy change and the entropy change.

$$\Delta G^\text{o} = \Delta H^\text{o} - T \Delta S^\text{o} \tag{11.5.3}$$

The symbol for free energy is G, in honor of American scientist Josiah Gibbs (1839 - 1903), who made many contributions to thermodynamics. The change in Gibbs free energy is equal to the change in enthalpy minus the mathematical product of the change in entropy multiplied by the Kelvin temperature. Each thermodynamic quantity in the equation is for substances in their standard states, as indicated by the $^\text{o}$ superscripts.

A spontaneous reaction is one that releases free energy, and so the sign of ΔG must be negative. Since both ΔH and ΔS can be either positive or negative, depending on the characteristics of the particular reaction, there are four different possible combinations. The outcomes for ΔG based on the signs of ΔH and ΔS are outlined in the table below. Recall that $-\Delta \text{H}$ indicates that the reaction is exothermic and a $+\Delta \text{H}$ means the reaction is endothermic. For entropy, $+\Delta \text{S}$ means the entropy is increasing and the system is becoming more disordered. A $-\Delta \text{S}$ means that entropy is decreasing and the system is becoming less disordered (more ordered).

Table 11.5.1: Enthalpy, Entropy, and Free Energy Changes.

ΔH	ΔS	ΔG
negative	positive	always negative
positive	positive	negative at higher temperatures, positive at lower temperatures
negative	negative	negative at lower temperatures, positive at higher temperatures
positive	negative	always positive

Keep in mind that the temperature in the Gibbs free energy equation is the Kelvin temperature, so it can only have a positive value. When ΔH is negative and ΔS is positive, the sign of ΔG will always be negative, and the reaction will be spontaneous at all temperatures. This corresponds to both driving forces being in favor of product formation. When ΔH is positive and ΔS is negative, the sign of ΔG will always be positive, and the reaction can never be spontaneous. This corresponds to both driving forces working against product formation.

When one driving force favors the reaction, but the other does not, it is the temperature that determines the sign of ΔG. Consider first an endothermic reaction (positive ΔH) that also displays an increase in entropy (positive ΔS). It is the entropy term that favors the reaction. Therefore, as the temperature increases, the $T\Delta S$ term in the Gibbs free energy equation will begin to predominate and ΔG will become negative. A common example of a process which falls into this category is the melting of ice (see figure below). At a relatively low temperature (below 273 K), the melting is not spontaneous because the positive ΔH term "outweighs" the $T\Delta S$ term. When the temperature rises above 273 K, the process becomes spontaneous because the larger T value has tipped the sign of ΔG over to being negative.

Figure 11.5.3: Ice melts spontaneously only when the temperature is above 0^oC. The increase in entropy is then able to drive the unfavorable endothermic process.

When the reaction is exothermic (negative ΔH) but undergoes a decrease in entropy (negative ΔS), it is the enthalpy term which favors the reaction. In this case, a spontaneous reaction is dependent upon the $T\Delta S$ term being small relative to the ΔH term, so that ΔG is negative. The freezing of water is an example of this type of process. It is spontaneous only at a relatively low temperature. Above $273.\text{ K}$, the larger $T\Delta S$ value causes the sign of ΔG to be positive, and freezing does not occur.

Contributors and Attributions

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11.6: Rates of Reactions

Learning Outcomes

- Define reaction rate.
- Define catalyst and explain its behavior in a chemical reaction.
- Describe how temperatures, concentration or pressure of reactant, and a catalyst affect the reaction rate.
- Explain the concept of activation energy.
- Label a diagram with reactants, products, enthalpy of forward and reverse reactions, activation energy of forward and reverse reactions, and activated complex.
- Calculate the enthalpy and activation energy with data on a reaction diagram.

Chemical kinetics is the study of the rates of chemical reactions. In this lesson, you will learn how to express the rate of a chemical reaction and about various factors that influence reaction rates.

Expressing Reaction Rate

Chemical reactions vary widely in the speeds with which they occur. Some reactions occur very quickly. It a lighted match is brought into contact with lighter fluid or another flammable liquid, it erupts into flame instantly and burns fast. Other reactions occur very slowly. A container of milk in the refrigerator will be good to drink for weeks before it begins to turn sour. Millions of years were required for dead plants under Earth's surface to accumulate and eventually turn into fossil fuels such as coal and oil.

Chemists need to be concerned with the rates at which chemical reactions occur. Rate is another word for speed. If a sprinter takes 11.0 seconds (s) to run a 100 meter (\left(m \right)\) dash, his rate or speed is given by the distance traveled divided by the time (see figure below).

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100 \text{ m}}{11.0 \text{ s}} = 9.09 \text{ m/s} \tag{11.6.1}$$

The sprinter's average running rate for the race is 9.09 m/s We say that it is his average rate because he did not run at that speed for the entire race. At the very beginning of the race, while coming from a standstill, his rate must be slower until he is able to get up to his top speed. His top speed must the be greater than 9.09 m/s so that taken over the entire race, the average ends up at 9.09 m/s

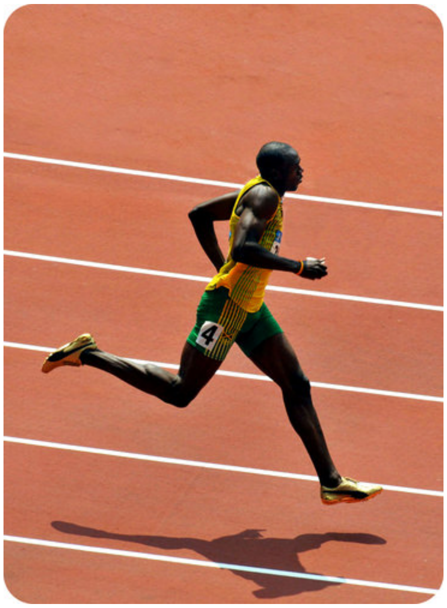

Figure 11.6.1: Usain Bolt set the world record for the 100 meter dash in 2009 with a time of 9.58 seconds. His average running rate over the course of this race was 10.4 m/s, or 23.4 mph.

Chemical reactions can't be measured in units of meters per second, as that would not make any sense. A **reaction rate** *is the change in concentration of a reactant or product with time*. Suppose that a simple reaction were to take place in which a 1.00 molar

(M) aqueous solution of substance A was converted to substance B.

$$\mathrm{A}\,(aq) \rightarrow \mathrm{B}\,(aq) \tag{11.6.2}$$

Suppose that after 20.0 seconds, the concentration of A had dropped from $1.00\ \mathrm{M}$ to $0.72\ \mathrm{M}$ as it was being converted to substance B. We can express the rate of this reaction as the change in concentration of A divided by the time.

$$\text{rate} = -\frac{\Delta\,[\mathrm{A}]}{\Delta t} = -\frac{[\mathrm{A}]_{\text{final}} - [\mathrm{A}]_{\text{initial}}}{\Delta t} \tag{11.6.3}$$

A bracket around a symbol or formula means the concentration in molarity of that substance. The change in concentration of A is its final concentration minus its initial concentration. Because the concentration of A is decreasing over time, the negative sign is used. Thus, the rate for the reaction is positive, and the units are molarity per second or $\mathrm{M/s}$.

$$\text{rate} = -\frac{0.72\ \mathrm{M} - 1.00\ \mathrm{M}}{20.0\ \mathrm{s}} = -\frac{-0.28\ \mathrm{M}}{20.0\ \mathrm{s}} = 0.041\ \mathrm{M/s} \tag{11.6.4}$$

Over the first 20.0 seconds of this reaction, the molarity of A decreases by an average rate of $0.041\ \mathrm{M}$ every second. In summary, the rate of a chemical reaction is measured by the change in concentration over time for a reactant or product. The unit of measurement for a reaction rate is molarity per second $(\mathrm{M/s})$.

Collision Theory

The behavior of the reactant atoms, molecules, or ions is responsible for the rates of a given chemical reaction. **Collision theory** *is a set of principles based around the idea that reactant particles form products when they collide with one another, but only when those collisions have enough kinetic energy and the correct orientation to cause a reaction.* Particles that lack the necessary kinetic energy may collide, but the particles will simply bounce off one another unchanged. The figure below illustrates the difference. In the first collision, the particles bounce off one another, and no rearrangement of atoms has occurred. The second collision occurs with greater kinetic energy, and so the bond between the two red atoms breaks. One red atom bonds with the other molecule as one product, while the single red atom is the other produce. The first collision is called an ineffective collision, while the second collision is called an effective collision.

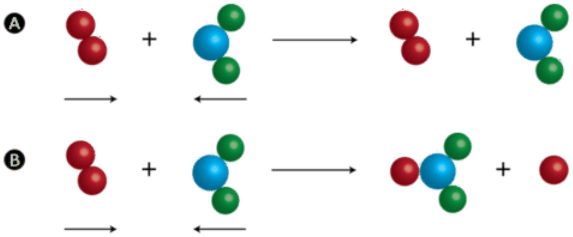

Figure 11.6.2: (A) An ineffective collision is one that does not result in product formation. (B) An effective collision is one in which chemical bonds are broken, and a product is formed.

Supplying reactant particles with energy causes the bonds between the atoms to vibrate with a greater frequency. This increase in vibrational energy makes a chemical bond more likely to break and a chemical reaction more likely to occur when those particles collide with other particles. Additionally, more energetic particles have more forceful collisions, which also increases the likelihood that a rearrangement of atoms will take place. The **activation energy** *for a reaction is the minimum energy that colliding particles must have in order to undergo a reaction*. Some reactions occur readily at room temperature because most of the reacting particles already have the requisite activation energy at that temperature. Other reactions only occur when heated because the particles do not have enough energy to react unless more is provided by an external source of heat.

Potential Energy Diagrams

Then energy changes that occur during a chemical reaction can be shown in a diagram called a potential energy diagram, sometimes called a reaction progress curve. A **potential energy diagram** *shows the change in the potential energy of a system as reactants are converted into products*. The figure below shows basic potential energy diagrams for an endothermic (left) and an exothermic (right) reaction. Recall that the enthalpy change (ΔH) is positive for an endothermic reaction and negative for an exothermic reaction. This can be seen in the potential energy diagrams. The total potential energy of the system increases for the endothermic reaction as the system absorbs energy from the surroundings. The total potential energy of the system decreases for the exothermic reaction as the system releases energy to the surroundings.

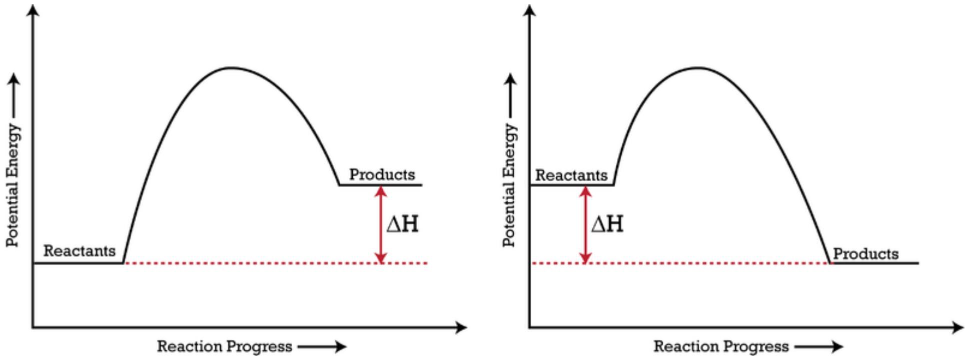

Figure 11.6.3: A potential energy diagram shows the total potential energy of a reacting system as the reaction proceeds. (Left) In an endothermic reaction, the energy of the products is greater than the energy of the reactants, and ΔH is positive. (Right) In an exothermic reaction, the energy of the products is lower than the energy of the reactants, and ΔH is negative.

The activation energy for a reaction is illustrated in the potential energy diagram by the height of the hill between the reactants and the products. For this reason, the activation energy of a reaction is sometimes referred to as the activation energy barrier. Reacting particles must have enough energy so that when they collide, they can overcome this barrier (see figure below).

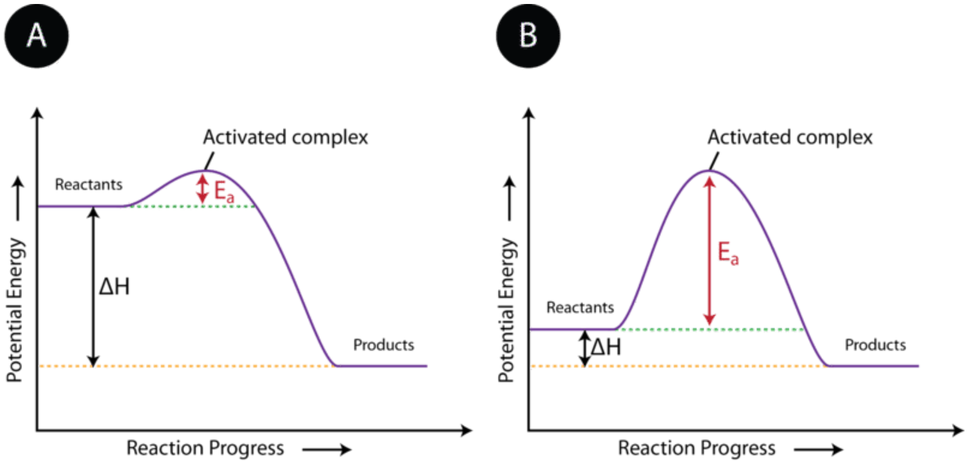

Figure 11.6.4: The activation energy (E_a) of a reaction is the barrier that must be overcome in order for the reactants to become products. (A) The activation energy is low, meaning that the reaction is likely to be fast. (B) The activation energy is high, meaning that the reaction is likely to be slow.

As discussed earlier, reactant particles sometimes collide with one another and yet remain unchanged by the collision. Other times, the collision leads to the formation of products. The state of the particles that is in between the reactants and products is called the activated complex. An **activated complex** *is an unstable arrangement of atoms that exists momentarily at the peak of the activation energy barrier*. Because of its high energy, the activated complex exists only for an extremely short period of time (about 10^{-13} s). The activated complex is equally likely to either reform the original reactants or go on to form the products. The figure below shows the formation of a possible activated complex between colliding hydrogen and oxygen molecules. Because of their unstable nature and brief existence, very little is known about the exact structures of most activated complexes.

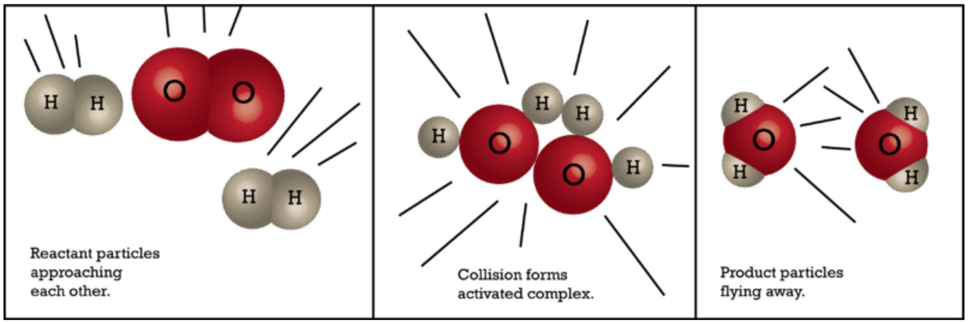

Figure 11.6.5: An activated complex is a short-lived state in which the colliding particles are at the peak of the potential energy curve.

Factors Affecting Reaction Rates

By their nature, some reactions occur very quickly, while others are very slow. However, certain changes in the reaction conditions can have an effect on the rate of a given chemical reaction. Collision theory can be utilized to explain these rate effects.

Concentration

Increasing the concentration of one or more of the reacting substances generally increases the reaction rate. When more particles are present in a given amount of space, a greater number of collisions will naturally occur between those particles. Since the rate of a reaction is dependent on the frequency of collisions between the reactants, the rate increases as the concentration increases.

Pressure

When the pressure of a gas is increased, its particles are forced closer together, decreasing the amount of empty space between them. Therefore, an increase in the pressure of a gas is also an increase in the concentration of the gas. For gaseous reactions, an increase in pressure increases the rate of reaction for the same reasons as described above for an increase in concentration. Higher gas pressure leads to a greater frequency of collisions between reacting particles.

Surface Area

A large log placed in a fire will burn relatively slowly. If the same mass of wood were added to the fire in the form of small twigs, they would burn much more quickly. This is because the twigs provide a greater surface are than the log does. An increase in the surface area of a reactant increases the rate of a reaction. Surface area is larger when a given amount of a solid is present as smaller particles. A powdered reactant has a greater surface area than the same reactant as a solid chunk. In order to increase the surface area of a substance, it may be ground into smaller particles or dissolved into a liquid. In solution, the dissolved particles are separated from each other and will react more quickly with other reactants. The figure below shows the unfortunate result of high surface area in an unwanted combustion reaction. Small particles of grain dust are very susceptible to rapid reactions with oxygen, which can result in violent explosions and quick-burning fires.

Figure 11.6.6: This grain elevator in Kansas exploded in 1998. The tiny size of the reacting particles (grain dust) caused the reaction with oxygen in the air to be violently explosive.

Temperature

Raising the temperature of a chemical reaction results in a higher reaction rate. When the reactant particles are heated, they move faster and faster, resulting in a greater frequency of collisions. An even more important effect of the temperature increase is that the collisions occur with a greater force, which means the reactants are more likely to surmount the activation energy barrier and go on to form products. Increasing the temperature of a reaction increases not only the frequency of collisions, but also the percentage of those collisions that are effective, resulting in an increased reaction rate.

Paper is certainly a highly combustible material, but paper does not burn at room temperature because the activation energy for the reaction is too high. The vast majority of collisions between oxygen molecules and the paper are ineffective. However, when the paper is heated by the flame from a match, it reaches a point where the molecules now have enough energy to react. The reaction is very exothermic, so the heat released by the initial reaction will provide enough energy to allow the reaction to continue, even if the match is removed. The paper continues to burn rapidly until it is gone.

Catalysts

The rates of some chemical reactions can be increased dramatically by introducing certain other substances into the reaction mixture. Hydrogen peroxide is used as a disinfectant for scrapes and cuts, and it can be found in many medicine cabinets as a 3% aqueous solution. Hydrogen peroxide naturally decomposes to produce water and oxygen gas, but the reaction is very slow. A bottle of hydrogen peroxide will last for several years before it needs to be replaced. However, the addition of just a small amount of manganese (IV) oxide to hydrogen peroxide will cause it to decompose completely in just a matter of minutes. A **catalyst** *is a* *substance that increases the rate of a chemical reaction without being used up in the reaction*. It accomplishes this task by providing an alternate reaction pathway that has a lower activation energy barrier. After the reaction occurs, a catalyst returns to its original state, so catalysts can be used over and over again. Because it is neither a reactant nor a product, a catalyst is shown in a chemical equation by being written above the yield arrow.

$$2\mathrm{H_2O_2}\,(aq) \overset{\mathrm{MnO_2}}{\rightarrow} 2\mathrm{H_2O}\,(l) + \mathrm{O_2}\,(g) \tag{11.6.5}$$

A catalyst works by changing the mechanism of the reaction, which can be though of as the specific set of smaller steps by which the reactants become products. The important point is that the use of a catalyst lowers the overall activation energy of the reaction (see figure below). With a lower activation energy barrier, a greater percentage of reactant molecules are able to have effective collisions, and the reaction rate increases.

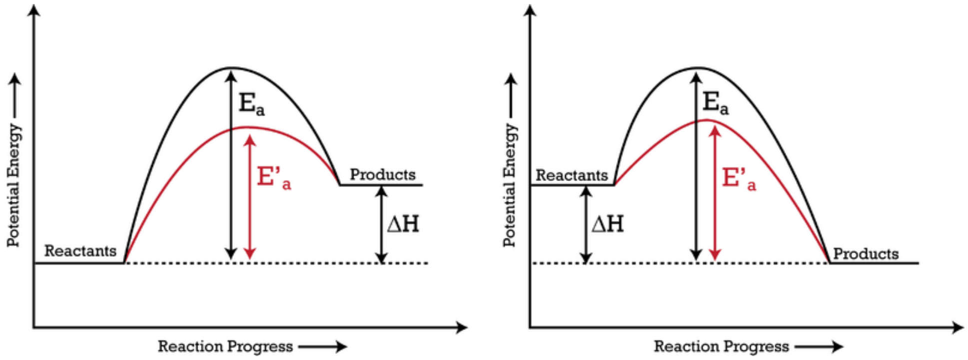

Figure 11.6.7: The addition of a catalyst to a reaction lowers the activation energy, increasing the rate of the reaction. The activation energy of the uncatalyzed reaction is shown by E_a, while the catalyzed reaction is shown by E_a'. The heat of reaction (ΔH) is unchanged by the presence of the catalyst.

Catalysts are extremely important parts of many chemical reactions. Enzymes in your body act as nature's catalysts, allowing important biochemical reactions to occur at reasonable rates. Chemical companies constantly search for new and better catalysts to make reactions go faster and thus make the company more profitable.

Contributors and Attributions

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11.7: Properties of Reactions (Exercises)

These are homework exercises to accompany Chapter 11 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health.

Questions

11.1: Oxidation Numbers

(click here for solutions)

Q11.1.1

What is a free element, and what is the oxidation number for atoms that exist as a free element?

Q11.1.2

What is the highest oxidation number that sulfur can have? The lowest? Consider the number of valence electrons it has.

Q11.1.3

Determine the oxidation numbers of each of the atoms in the following.

a. $KMnO_4$
b. OCl_2
c. $H_2C_2O_4$
d. Li_3PO_4
e. $NaClO$
f. Br_2
g. ClF_3
h. $CaCl_2$
i. K_2O

Q11.1.4

Determine the oxidation number of each of the atoms in the following ions.

a. NO_2^-
b. NO_3^-
c. $Cr_2O_7^{2-}$
d. BrO_3^-
e. ClO_3^-
f. BO_3^{3-}
g. CO_3^{2-}
h. NH_4^+
i. CrO_4^{2-}
j. $S_2O_3^{2-}$

11.2: The Nature of Oxidation and Reduction

(click here for solutions)

Q11.2.1

Explain why oxidation and reduction always occur together in a reaction.

Q11.2.2

What happens to the oxidizing agent in a redox reaction?

Q11.2.3

What happens to the reducing agent in a redox reaction?

Q11.2.4

Identify each process as an oxidation or a reduction.

a. $Rb \rightarrow Rb^+$
b. $Te \rightarrow Te^{2-}$
c. $2H^+ \rightarrow H_2$
d. $P^{3-} \rightarrow P$
e. $2Cl^- \rightarrow Cl_2$
f. $Sn^{4+} \rightarrow Sn^{2+}$
g. $Br_2 \rightarrow Br^-$
h. $Fe^{2+} \rightarrow Fe^{3+}$

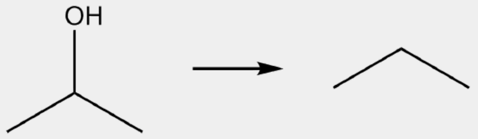
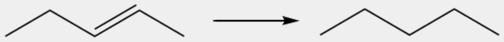
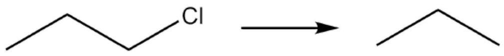
Q11.2.5

For each equation, 1) identify the oxidation numbers of each element, 2) determine if it is a redox reaction or not, and for redox reactions, 3) identify the species being oxidized and the species being reduced, and 4) identify the oxidizing and reducing agents.

a. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
b. $H_2(g) + CuO(s) \rightarrow Cu(s) + H_2O(l)$
c. $2Al(s) + 3Mg(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3Mg(s)$
d. $2HNO_3(aq) + 6HI(aq) \rightarrow 2NO(g) + 3I_2(s) + 4H_2O(l)$
e. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
f. $2FeCl_3(aq) + H_2S(g) \rightarrow 2FeCl_2(aq) + 2HCl(aq) + S(s)$

Q11.2.6

Identify the change occurring with respect to the gain or loss of hydrogen or oxygen.

a.	
b.	
c.	
d.	
e.	

11.3: Types of Chemical Reactions

(click here for solutions)

Q11.3.1

Classify the following reactions as combination, decomposition, single replacement, double replacement, or combustion. For each, indicate if it is a redox reaction or not.

a. $Cd(s) + H_2SO_4(aq) \rightarrow CdSO_4(aq) + H_2(g)$
b. $2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$
c. $C_7H_8(l) + 9O_2(g) \rightarrow 7CO_2(g) + 4H_2O(g)$
d. $2NH_4NO_3(s) \rightarrow 2N_2(g) + O_2(g) + 4H_2O(g)$
e. $2CoCl_3(aq) + 3Pb(NO_3)_2(aq) \rightarrow 2Co(NO_3)_3(aq) + 3PbCl_2(s)$

Q11.3.2

Balance the following equations and classify the following reactions as combination, decomposition, single replacement, double replacement, or combustion. For each, indicate if it is a redox reaction or not.

a. $Na + Cl_2 \rightarrow NaCl$
b. $Na_3PO_4 + KOH \rightarrow NaOH + K_3PO_4$
c. $P_4 + O_2 \rightarrow P_2O_3$
d. $N_2 + H_2 \rightarrow NH_3$
e. $Al + HCl \rightarrow H_2 + AlCl_3$
f. $H_2O_2 \rightarrow H_2O + O_2$
g. $NH_3 + CuO \rightarrow Cu + N_2 + H_2O$
h. $NH_4NO_3 \rightarrow N_2O + H_2O$

Q11.3.3

What do the products of combustion reactions have in common?

Q11.3.4

Write and balance combustion reactions for the following compounds.

a. methane (CH_4)
b. propane (C_3H_8)
c. octane (C_8H_{18})
d. ethanol (CH_3CH_2OH)
e. sucrose ($C_{12}H_{22}O_{11}$)

11.4: Entropy

(click here for solutions)

Q11.4.1

In a certain reaction, the energy of the reactants is less than the energy of the products (reaction consumes energy). Is the reaction endothermic or exothermic?

Q11.4.2

What are the two driving forces for all chemical reactions and physical processes?

Q11.4.3

Does entropy determine the spontaneity of a reaction? Does enthalpy determine the spontaneity of a reaction?

Q11.4.4

How does an increase in temperature affect the entropy of a system?

Q11.4.5

Which system has the greater entropy in each of the following?

a. solid sodium chloride or a sodium chloride solution
b. bromine liquid or bromine vapor
c. 25 g of water at 80°C or 25 g of water at 50°C
d. liquid mercury or solid mercury

Q11.4.6

How does the entropy of a system change for each of the following processes?

a. A solid melts.
b. A liquid freezes.
c. A liquid boils.
d. A vapor condenses to a liquid.
e. Sugar dissolves in water.
f. A solid sublimes.

11.5: Spontaneous Reactions and Free Energy

(click here for solutions)

Q11.5.1

What is true about the relative amounts of reactants and products at the end of a spontaneous reaction?

Q11.5.2

Can a proposed reaction be spontaneous and yet still not be observed to occur? Explain.

Q11.5.3

The forward reaction is spontaneous for a particular reversible reaction. What can you conclude about the spontaneity of the reverse reaction?

Q11.5.4

Explain how free energy is used to determine whether or not a reaction is spontaneous.

Q11.5.5

Under what conditions of enthalpy and entropy change is a reaction always spontaneous? Under what conditions is a reaction never spontaneous?

Q11.5.6

If the enthalpy change is unfavorable (endothermic), but the entropy change is favorable (increasing), would a high temperature or a low temperature be more likely to lead to a spontaneous reaction?

Q11.5.7

If the enthalpy change is favorable and the entropy change is favorable, would the reaction be spontaneous at high temperatures, low temperatures, or all temperatures?

11.6: Rates of Reactions

(click here for solutions)

Q11.6.1

In what unit is the rate of a chemical reaction typically expressed?

Q11.6.2

A 2.50 M solution undergoes a chemical reaction. After 3.00 minutes, the concentration of the solution is 2.15 M. What is the rate of change in M/s?

Q11.6.3

Substance A disappears at a rate of 0.0250 M/s. If the initial concentration is 4.00 M, what is the concentration after one minute?

Q11.6.4

The concentration of product B increases from 0 to 1.75 M in 45 seconds. What is the rate of formation of B?

Q11.6.5

The concentration of product B increases from 0.50 M to 1.25 M in 2.5 seconds. What is the rate of formation of B?

Q11.6.6

Reactant B goes from 2.25 M to 1.50 M in 0.85 seconds. What is the rate of change of B?

Q11.6.7

Does every collision between reacting particles lead to the formation of products? Explain.

Q11.6.8

What two conditions must be met in order for a collision to be effective?

Q11.6.9

Explain why the activation energy of a reaction is sometimes referred to as a barrier.

Q11.6.10

Why is it difficult to study activated complexes?

Q11.6.11

Explain how reaction rates can be affected by

a. changes in concentration.
b. changes in pressure.
c. increased surface area.
d. changes in temperatures.

Q11.6.12

What is the effect of a catalyst on the rate of a reaction?

Q11.6.13

Explain how the presence of a catalyst affects the activation energy of a reaction.

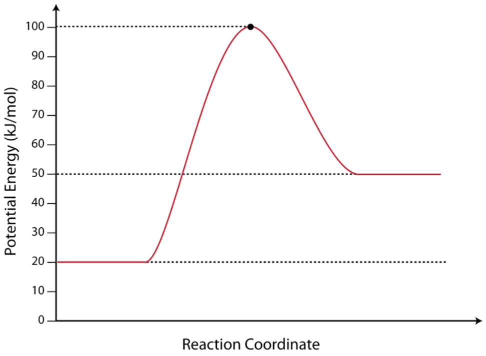
Q11.6.14

Zinc metal reacts with hydrochloric acid. Which one would result in the highest rate of reaction?

a. A solid piece of zinc in 1 M HCl
b. A solid piece of zinc in 3 M HCl
c. Zinc powder in 1 M HCl
d. Zinc powder in 3 M HCl

Q11.6.15

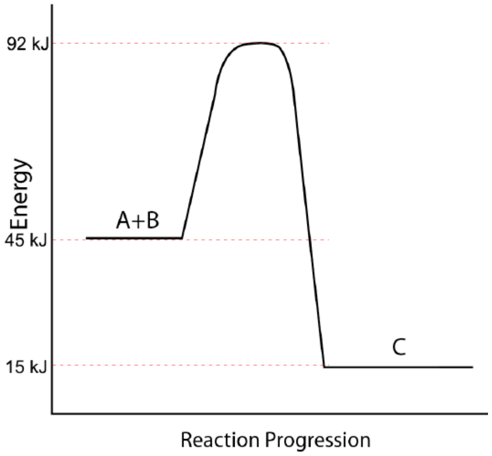
Use the potential energy diagram below to answer the following questions.



a. What is the potential energy of the reactants?
b. What is the potential energy of the products?
c. What is the heat of reaction ($\Delta H = E_{products} - E_{reactants}$)?
d. What is the potential energy of the activated complex?
e. What is the activation energy for the reaction?
f. Is the reaction endothermic or exothermic?
g. Which of the values in a-e above would be changed by the use of a catalyst in the reaction?
h. What is the activation energy of the reverse reaction?
i. What is the heat of reaction ($\Delta H = E_{products} - E_{reactants}$) of the reverse reaction?

Q11.6.16

Use the potential energy diagram below to answer the following questions.



a. What is the potential energy of the reactants?
b. What is the potential energy of the products?
c. What is the heat of reaction (ΔH = $E_{products}$ − $E_{reactants}$)?
d. What is the potential energy of the activated complex?
e. What is the activation energy for the reaction?
f. Is the reaction endothermic or exothermic?
g. Which of the values in a-e above would be changed by the use of a catalyst in the reaction?
h. What is the activation energy of the reverse reaction?
i. What is the heat of reaction (ΔH = $E_{products}$ − $E_{reactants}$) of the reverse reaction?

Answers

11.1: Oxidation Numbers

Q11.1.1

Any element by itself, either monatomic or diatomic, without a charge. The oxidation number of any free element is zero.

Q11.1.2

Sulfur can have an oxidation number up to +6 or as low as -2.

Q11.1.3

a. $\overset{+1}{\mathrm{K}}\overset{+7}{\mathrm{Mn}}\overset{-2}{\mathrm{O}}_4$
  - Rule 2: K forms an ion with a 1+ charge so its oxidation number is +1.
  - Rule 3: O has an oxidation number of -2.
  - Rule 6 for Mn: $$\begin{aligned} &1+x+4(-2)=0\\ &1+x+-8=0\\ &x-7=0\\ &x=7 \end{aligned}$$

b. $\overset{-2}{\mathrm{O}}\overset{+1}{\mathrm{Cl}}_2$
  - Rule 3: O has an oxidation number of -2.
  - Rule 6 for Cl: $$\begin{aligned} &-2+2x=0\\ &2x=-2\\ &x=1 \end{aligned}$$

c. $\overset{+1}{\mathrm{H}}_2\overset{+3}{\mathrm{C}}_2\overset{-2}{\mathrm{O}}_4$
  - Rule 3: O has an oxidation number of -2.
  - Rule 4: H has an oxidation number of \{+1\).

- Rule 6 for C: $2(+1)+2x+4(-2)=0$
  $2+2x-8=0$
  $2x-6=0$
  $2x=6$
  $x=3$

d. $\overset{+1}{\mathrm{Li}}_3\overset{+5}{\mathrm{P}}\overset{-2}{\mathrm{O}}_4$

- Rule 2: Li forms an ion with a 1+ charge so its oxidation number is +1.
- Rule 3: O has an oxidation number of -2.
- Rule 6 for P: $3(+1)+x+4(-2)=0$
  $3+x+-8=0$
  $x-5=0$
  $x=5$

e. $\overset{+1}{\mathrm{Na}}\overset{+1}{\mathrm{Cl}}\overset{-2}{\mathrm{O}}$

- Rule 2: Na forms an ion with a 1+ charge so its oxidation number is +1.
- Rule 3: O has an oxidation number of -2.
- Rule 6 for Cl: $1+x+-2=0$
  $x-1=0$
  $x=1$

f. $\overset{0}{\mathrm{Br}}_2$

- Rule 1: Oxidation number of a free element is zero.

g. $\overset{+3}{\mathrm{Cl}}\overset{-1}{\mathrm{F}}_3$

- Rule 5: F has an oxidation number of -1.
- Rule 6 for Cl: $x+3(-1)=0$
  $x-3=0$
  $x=3$

h. $\overset{+2}{\mathrm{Ca}}\overset{-1}{\mathrm{Cl}}_2$

- Rule 2: Ca forms an ion with a 2+ charge so its oxidation number is +2.
- Rule 6 for Cl: $2+2(x)=0$
  $2x=-2$
  $x=-1$

i. $\overset{+1}{\mathrm{K}}_2\overset{-2}{\mathrm{O}}$

- Rule 2: K forms an ion with a 1+ charge so its oxidation number is +1.
- Rule 3: O has an oxidation number of -2.

Q11.1.4

Determine the oxidation number of each of the atoms in the following ions.

a. $\overset{+3}{\mathrm{N}}\overset{-2}{\mathrm{O}}_2^-$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for N: $x+2(-2)=-1$
  $x-4=-1$
  $x=3$

b. $\overset{+5}{\mathrm{N}}\overset{-2}{\mathrm{O}}_3^-$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for N: $x+3(-2)=-1$
  $x-6=-1$
  $x=5$

c. $\overset{+6}{\mathrm{Cr}}_2\overset{-2}{\mathrm{O}}{}_7^{2-}$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for Cr: $2x + 7(-2) = -2$
  $2x - 14 = -2$
  $2x = 12$
  $x = 6$

d. $\overset{+5}{\mathrm{Br}}\overset{-2}{\mathrm{O}}{}_3^-$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for Br: $x + 3(-2) = -1$
  $x - 6 = -1$
  $x = 5$

e. $\overset{+5}{\mathrm{Cl}}\overset{-2}{\mathrm{O}}{}_3^-$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for Cl: $x + 3(-2) = -1$
  $x - 6 = -1$
  $x = 5$

f. $\overset{+3}{\mathrm{B}}\overset{-2}{\mathrm{O}}{}_3^-$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for B: $x + 3(-2) = -3$
  $x - 6 = -3$
  $x = 3$

g. $\overset{+4}{\mathrm{C}}\overset{-2}{\mathrm{O}}{}_3^{2-}$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for C: $x + 3(-2) = -2$
  $x - 6 = -2$
  $x = 4$

h. $\overset{-3}{\mathrm{N}}\overset{+1}{\mathrm{H}}{}_4^+$

- Rule 2: H has an oxidation number of $+1$.
- Rule 7 for N: $x + 4(+1) = +1$
  $x + 4 = 1$
  $x = -3$

i. $\overset{+6}{\mathrm{Cr}}\overset{-2}{\mathrm{O}}{}_4^{2-}$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for Cr: $x + 4(-2) = -2$
  $x - 8 = -2$
  $x = 6$

j. $\overset{+2}{\mathrm{S}}_2\overset{-2}{\mathrm{O}}{}_3^{2-}$

- Rule 2: O has an oxidation number of -2.
- Rule 7 for S: $2x + 3(-2) = -2$
  $2x - 6 = -2$
  $2x = 4$
  $x = 2$

11.2: The Nature of Oxidation and Reduction

Q11.2.1

Oxidation and reduction are opposite processes. Electrons are lost by one substance and gained by another so they have a transfer of electrons.

Q11.2.2

The oxidizing agent is reduced.

Q11.2.3

The reducing agent is oxidized.

Q11.2.4

a. oxidation
b. reduction
c. reduction
d. oxidation
e. oxidation
f. reduction
g. reduction
h. oxidation

Q11.2.5

a. $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
   1. $2\overset{+1}{K}\overset{+5}{Cl}\overset{-2}{O_3} \rightarrow 2\overset{+1}{K}\overset{-1}{Cl} + 3\overset{0}{O_2}$
   2. redox
   3. oxygen is being oxidized, chlorine is being reduced
   4. $KClO_3$ is the oxidizing and reducing agent (specifically, the oxygen is the reducing agent and the chlorine is the oxidizing agent)

b. $H_2(g) + CuO(s) \rightarrow Cu(s) + H_2O(l)$
   1. $\overset{0}{H_2} + \overset{+2}{Cu}\overset{-2}{O} \rightarrow \overset{0}{Cu} + \overset{+1}{H_2}\overset{-2}{O}$
   2. redox
   3. hydrogen is being oxidized, copper is being reduced
   4. CuO is the oxidizing agent, H_2 is the reducing agent

c. $2Al(s) + 3Mg(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3Mg(s)$
   1. $overset02Al + 3\overset{+2}{Mg}(\overset{+5}{N}\overset{-2}{O_3})_2 \rightarrow 2\overset{+3}{Al}(\overset{+5}{N}\overset{-2}{O_3})_3 + 3\overset{0}{Mg}$
   2. redox
   3. aluminum is being oxidized, magnesium is being reduced
   4. $Mg(NO_3)_2$ is the oxidizing agent, Al is the reducing agent

d. $2HNO_3(aq) + 6HI(aq) \rightarrow 2NO(g) + 3I_2(s) + 4H_2O(l)$
   1. $2\overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3} + 6\overset{+1}{H}\overset{-1}{I} \rightarrow 2\overset{+2}{N}\overset{-2}{O} + 3\overset{0}{I_2} + 4\overset{+1}{H_2}\overset{-2}{O}$
   2. redox
   3. iodine is being oxidized, nitrogen is being reduced
   4. HNO_3 is the oxidizing agent, HI is the reducing agent

e. $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   1. $\overset{+1}{Ag}\overset{+5}{N}\overset{-2}{O_3} + \overset{+1}{Na}\overset{-1}{Cl} \rightarrow \overset{+1}{Ag}\overset{-1}{Cl} + \overset{+1}{Na}\overset{+5}{N}\overset{-2}{O_3}$
   2. not redox

f. $2FeCl_3(aq) + H_2S(g) \rightarrow 2FeCl_2(aq) + 2HCl(aq) + S(s)$
   1. $2\overset{+3}{Fe}\overset{-1}{Cl_3} + \overset{+1}{H_2}\overset{-2}{S} \rightarrow 2\overset{+2}{Fe}\overset{-1}{Cl_2} + 2\overset{+1}{H}\overset{-1}{Cl} + \overset{0}{S}$
   2. redox
   3. sulfur is being oxidized, iron is being reduced

4. $FeCl_3$ is the oxidizing agent, H_2S is the reducing agent

Q11.2.6

a. Oxidation due to loss of hydrogen.
b. Reduction due to loss of oxygen.
c. Reduction due to loss of oxygen or gain of hydrogen.
d. Reduction due to gain of hydrogen.
e. Reduction due to gain of hydrogen.

11.3: Types of Chemical Reactions

Q11.3.1

a. single replacement (single displacement), redox
b. combination (synthesis), redox
c. combustion, redox
d. decomposition, redox
e. double replacement (double displacement), not redox

Q11.3.2

a. $2Na + Cl_2 \rightarrow 2NaCl$, combination (synthesis), redox
b. $Na_3PO_4 + 3KOH \rightarrow 3NaOH + K_3PO_4$, double replacement (double displacement), not redox
c. $P_4 + 3O_2 \rightarrow 2P_2O_3$, combination (synthesis), redox
d. $N_2 + 3H_2 \rightarrow 2NH_3$, combination (synthesis), redox
e. $2Al + 6HCl \rightarrow 3H_2 + 2AlCl_3$, single replacement (single displacement), redox
f. $2H_2O_2 \rightarrow 2H_2O + O_2$, decomposition, redox
g. $2NH_3 + 3CuO \rightarrow 3Cu + N_2 + 3H_2O$, decomposition, redox
h. $NH_4NO_3 \rightarrow N_2O + 2H_2O$, decomposition, not redox

Q11.3.3

Combustion produces CO_2 and H_2O.

Q11.3.4

a. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
b. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
c. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$
d. $CH_3CH_2OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
e. $C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O$

11.4: Entropy

Q11.4.1

endothermic

Q11.4.2

enthalpy and entropy

Q11.4.3

Neither entropy nor enthalpy determine the spontaneity of a reaction but both contribute to determining the spontaneity.

Q11.4.4

Entropy increases with temperature.

Q11.4.5

Which system has the greater entropy in each of the following?

a. solution

b. vapor
c. 80°C
d. liquid

Q11.4.6

How does the entropy of a system change for each of the following processes?

a. increases
b. decreases
c. increasese
d. decreases
e. increases
f. increases

11.5: Spontaneous Reactions and Free Energy

Q11.5.1

There are more products than reactants.

Q11.5.2

Yes, because it is a very slow reaction.

Q11.5.3

The reverse reaction is not spontaneous.

Q11.5.4

Free energy (ΔG) is negative for spontaneous reactions and positive for non-spontaneous reactions.

Q11.5.5

ΔG is negative (spontaneous) at all temperatures when ΔH is negative and ΔS is positive.

ΔG is positive (non-spontaneous) at all temperatures when ΔH is positive and ΔS is negative.

Q11.5.6

high temperature

Q11.5.7

all temperatures

11.6: Rates of Reactions

Q11.6.1

molarity per second, $\frac{M}{s}$ (may be written as Ms^{-1})

Q11.6.2

$$\begin{aligned} rate &= \frac{\Delta[A]}{\Delta t} \\ &= \frac{2.15-2.50\ M}{180.\ s} \\ &= \frac{-0.35\ M}{180.\ s} \\ &= -0.0019\ \frac{M}{s} \end{aligned}$$

Q11.6.3

$$60\ s\left(\frac{0.0250\ M}{s}\right) = 1.50\ M \text{ lost}$$
$$initial + change = final$$
$$4.00\ M + (-1.50\ M) = 2.50\ M$$

Q11.6.4

$$\begin{aligned} rate &= \frac{\Delta[B]}{\Delta t} \\ &= \frac{1.75-0\ M}{45\ s} \\ &= 0.039\ \frac{M}{s} \end{aligned}$$

Q11.6.5

$$\begin{aligned} rate &= \frac{\Delta[B]}{\Delta t} \\ &= \frac{1.25-0.50\ M}{2.5\ s} \\ &= \frac{0.75\ M}{2.5\ s} \\ &= 0.30\ \frac{M}{s} \end{aligned}$$

Q11.6.6

$$\begin{aligned} rate &= \frac{\Delta[B]}{\Delta t} \\ &= \frac{1.50-2.25\ M}{0.85\ s} \\ &= \frac{-0.75\ M}{0.85\ s} \\ &= -0.88\ \frac{M}{s} \end{aligned}$$

Q11.6.7

No, because collisions have to occur at the correct orientation with sufficient energy.

Q11.6.8

Correct orientation and sufficient energy.

Q11.6.9

Because it is the energy requirement that must be met before the reaction can occur. When there is not enough energy, the reaction is blocked from occurring.

Q11.6.10

They are very short-lived.

Q11.6.11

a. Increasing concentration increases the probability of an effective collision so the reaction rate will increase. The reverse is true for a decrease in concentration.
b. Change in pressure of a substance involved in the reaction will have the same effect as a change in concentration.
c. Increasing the surface area creates more places for reactants to interact which will increase the probability of an effective collision so the reaction rate will increase.
d. Increasing the temperature will increase the energy of collisions so a greater number of collisions will have sufficient energy to overcome the activation barrier which will result in an increase in the reaction rate. Decreasing the temperature will have the reverse effect.

Q11.6.12

Catalysts increase the rate of a reaction by lowering the activation energy.

Q11.6.13

Catalysts provide an alternate mechanism or "path" for the reaction to occur. This new mechanism has a lower activation energy so more collisions have enough energy to overcome the barrier which will increase the reaction rate.

Q11.6.14

d. The powder has a higher surface area than a solid piece of zinc and the higher concentration will result in a higher reaction rate.

Q11.6.15

a. 20 kJ/mol

b. 50 kJ/mol
c. 30 kJ/mol
d. 100 kJ/mol
e. 80 kJ/mol
f. endothermic
g. d (energy of the activated complex) and e (activation energy)
h. 50 kJ/mol
i. −30 kJ/mol

Q11.6.16

a. 45 kJ/mol
b. 15 kJ/mol
c. −30 kJ/mol
d. 92 kJ/mol
e. 47 kJ/mol
f. exothermic
g. d (energy of the activated complex) and e (activation energy)
h. 77 kJ/mol
i. 30 kJ/mol

CHAPTER OVERVIEW

12: Organic Reactions

Organic reactions are chemical reactions involving organic compounds. The basic organic chemistry reaction types are addition reactions, elimination reactions, substitution reactions, pericyclic reactions, rearrangement reactions, photochemical reactions and redox reactions.

12.1: Organic Reactions

Learning Outcomes

- Identify and describe substitution, elimination, addition, hydrolysis, and condensation reactions.
- Predict products of each reaction type.
- Predict products of hydration reactions based on Markovnikov's rule.
- Explain why condensation reactions are also called dehydration reactions.
- Define "polymer".

Organic reactions require the breaking of strong covalent bonds, which takes a considerable input of energy. In order for relatively stable organic molecules to react at a reasonable rate, they often must be modified with the use of highly reactive materials or in the presence of a catalyst. In this lesson, you will learn about several general categories of organic reactions.

Substitution Reactions

A **substitution reaction**, which is the same as a single replacement reaction in inorganic reactions, *is a reaction in which one or more atoms in a molecule are replaced with another atom or group of atoms.* Alkyl halides are formed by the substitution of a halogen atom for a hydrogen atom. When methane reacts with chlorine gas, ultraviolet light can act as a catalyst for the reaction.

$$\mathrm{CH_4}\,(g) + \mathrm{Cl_2}\,(g) \overset{\text{UV light}}{\rightarrow} \mathrm{CH_3Cl}\,(g) + \mathrm{HCl}\,(g) \tag{12.1.1}$$

The reaction produces chloromethane and hydrogen chloride. When the mixture is allowed to react for longer periods of time, further substitution reactions may occur.

$$\mathrm{CH_3Cl}\,(g) + \mathrm{Cl_2}\,(g) \overset{\text{UV light}}{\rightarrow} \mathrm{CH_2Cl_2}\,(g) + \mathrm{HCl}\,(g) \tag{12.1.2}$$

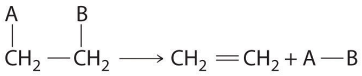
The product above is dichloromethane. Further substitution products trichloromethane and tetrachloromethane, commonly called carbon tetrachloride. A mixture of products occurs in the reaction, with the relative amounts dependent upon the time that the reaction is allowed to proceed. Chlorofluorocarbons are produced by reacting chloroalkanes with HF, because the fluorine atom makes a stronger bond to the carbon atom than chlorine does.

$$\mathrm{CCl_4}\,(g) + \mathrm{HF}\,(g) \overset{\mathrm{SbF_5}}{\rightarrow} \mathrm{CCl_3F}\,(g) + \mathrm{HCl}\,(g) \tag{12.1.3}$$

The fluorine atom substitutes for a chlorine atom in the reaction.

Elimination Reactions

An **elimination reaction** involves the removal of adjacent atoms from a molecule. This results in the formation of a multiple bond and the release of a small molecule, so they are called elimination reactions. They have the general form


$$\mathrm{CH_2(A)-CH_2(B)} \longrightarrow \mathrm{CH_2=CH_2} + \mathrm{A-B}$$

A typical example is the conversion of ethyl chloride to ethylene:

$$CH_3CH_2Cl \rightarrow CH_2 = CH_2 + HCl \tag{12.1.4}$$

Much of the approximately 26 million tons of ethylene produced per year in the United States is used to synthesize plastics, such as polyethylene. In the above reaction, the A–B molecule eliminated is HCl, whose components are eliminated as H^+ from the carbon atom on the left and Cl^- from the carbon on the right. When an acid is produced, as occurs here, the reaction is generally carried out in the presence of a base (such as NaOH) to neutralize the acid. Other elimination reactions will produce H_2, X_2 (where X = halogen), or H_2O. These reactions are often referred to by more descriptive terms such as dehydrogenation (removing hydrogen) or dechlorination (removing chlorine).

Addition Reactions

An **addition reaction** *is a reaction in which an atom or molecule is added to an unsaturated molecule, making a single product.* An addition reaction can often be thought of as adding a molecule across the double bond of an alkene or across the triple bond of an alkyne. Knowing that "ation" means to add, the specific names of these reactions, such as hydrogenation, hydration, or chlorination, should make sense. Note that hydrogenation (adding H_2) and hydration (adding H_2O) are very different processes.

One type of addition reaction is called hydrogenation. **Hydrogenation** *is a reaction that occurs when molecular hydrogen is added to an alkene to produce an alkane* or *hydrogen is added to an alkyne to produce an alkene or alkane.* The reaction is typically performed with the use of a transition metal catalyst. For example, ethene reacts with hydrogen to form ethane.

$$\mathrm{CH_2{=}CH_2}\,(g) + \mathrm{H_2}\,(g) \overset{\mathrm{Pt}}{\rightarrow} \mathrm{CH_3CH_3}\,(g) \tag{12.1.5}$$

Note that the hydrogenation reaction is also a redox reaction. Ethene is reduced, because the oxidation numbers of the carbon atoms change from -2 to -3 as a result of the reaction.

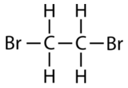
Vegetable oils consist of long carbon chains with carboxyl groups on the end; these molecules are referred to as fatty acids. The carbon chains of the fatty acids in vegetable oils are unsaturated, usually containing multiple double bonds. When hydrogen gas is blown through a sample of the oil, hydrogen atoms add across the double bonds. This conversion changes the substance from a liquid oil into a solid fat. The "hydrogenated" on a food product is an indication that oil (liquid) has been converted into fat (solid) by this process. Margarine is manufactured from unsaturated vegetable oil in this way by hydrogenating some of the double bonds making it a "partially hydrogenated vegetable oil".

Addition reactions are also useful ways to introduce a new functional group into an organic molecule. Alkyl halides can be produced from an alkene by the addition of either the elemental halogen or the hydrogen halide. A monosubstituted alkyl halide can be produced by the addition of a hydrogen halide to an alkene. Shown below is the formation of chloroethane.

$$\mathrm{CH_2{=}CH_2}\,(g) + \mathrm{HCl}\,(g) \rightarrow \mathrm{CH_3CH_2Cl}\,(g) \tag{12.1.6}$$

When the reactant is the diatomic halogen, the product is a disubstituted alkyl halide as in the addition of bromine to ethene.

$$\mathrm{CH_2{=}CH_2}\,(g) + \mathrm{Br_2}\,(l) \rightarrow \mathrm{CH_2BrCH_2Br}\,(g) \tag{12.1.7}$$



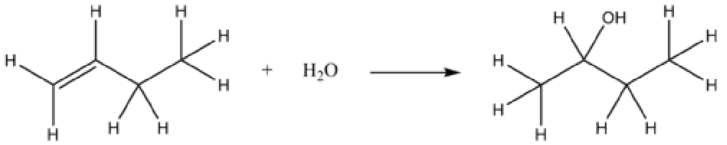
1,2-dibromoethane

The addition of bromine to an unknown organic compound can be used as a test for unsaturation in the compound. Bromine has a distinctive brownish-orange color, while most bromoalkanes are colorless. When bromine is slowly added to a solution of the compound, the orange color will fade if it undergoes an addition reaction to produce an alkyl halide. If the orange color remains, then the original compound was already saturated, and no reaction occurred.

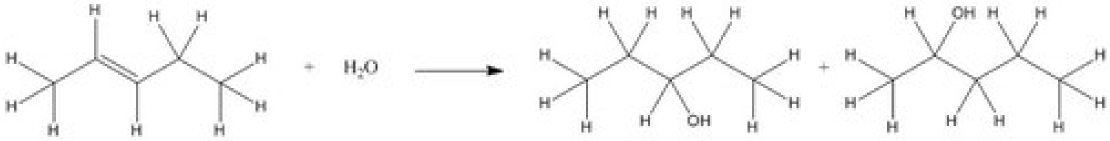
A **hydration reaction** *is a reaction in which water is added to an alkene*. Hydration reactions can take place when the alkene and water are heated to near $100^{\circ}\mathrm{C}$ in the presence of a strong acid, which acts as a catalyst. Shown below is the hydration of ethene to produce ethanol.

$$\mathrm{CH_2{=}CH_2}\,(g) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{CH_3CH_2OH}\,(g) \tag{12.1.8}$$

Unlike addition reactions involving H_2 or Br_2, the addition of water can have two possible products because an $-\mathrm{H}$ and an $-\mathrm{OH}$ are being added to the carbons in the double bond. **Markovnikov's rule** helps predict the main product in a hydration reaction. The rule states that the hydrogen atom from water will add to the carbon that originally had more hydrogen atoms. For example, look at the hydration of 1-butene.



Note that the first carbon in 1-butene started with two hydrogen atoms and the second carbon started with one hydrogen. Therefore, the hydrogen from water adds to the first carbon and the $-\mathrm{OH}$ group adds to the second carbon. When there are equal numbers of hydrogen atoms on both carbons in a double bond then the two products will form in approximately equal amounts. For example, the hydration of 2-pentene results in two products. In the first product, the $-\mathrm{OH}$ group is on the third carbon and in the second product, the $-\mathrm{OH}$ group is on the second carbon. While these two molecules will have similar properties, there will be differences.



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. 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 second product (see figure below).

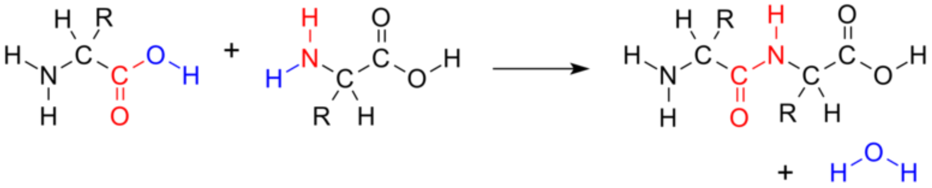
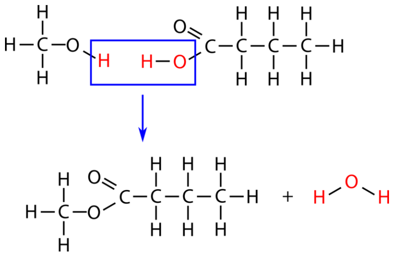


Figure 12.1.1: Amino acids join together to form a molecule called a dipeptide. The $-\mathrm{OH}$ from the carboxyl group of one amino acid combines with a hydrogen atom from the amine group of the other amino group to produce water (blue).

This reaction forms a molecule called a dipeptide, and the resulting carbon-nitrogen covalent bond is often called a peptide bond. When repeated numerous times, a long molecule called a protein is eventually produced.

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 can be 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. Esterification reactions are reversible.

$$\mathrm{CH_3CH_2CH_2COOH + HOCH_3 \overset{H_2SO_4}{\rightarrow} CH_3CH_2CH_2COOCH_3 + H_2O} \tag{12.1.9}$$



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.

$$\underset{\text{ethyl stearate}}{\mathrm{C_{17}H_{35}COOC_2H_5}} \quad + \quad \mathrm{NaOH} \quad \rightarrow \quad \underset{\text{sodium stearate (soap)}}{\mathrm{C_{17}H_{35}COO{-}Na^+}} \quad + \quad \mathrm{C_2H_5OH} \tag{12.1.10}$$

Hydrolysis Reactions

Hydrolysis is the reverse of condensation. "Hydro" indicates that water is involved and "lysis" means to break apart. In a hydrolysis reaction, water is added and the molecule breaks apart, usually at a $\mathrm{C{-}O{-}C}$ linkage.

Polymerization

Polymers are very different than 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 to one another in a repeating pattern*. The *small molecules that make up the polymer are called* **monomers**. Polymers are generally formed by either addition or condensation reactions. Teflon (see figure below) is a non-reactive, non-stick coating used on cookware as well as in containers and pipes for reactive or corrosive chemicals.

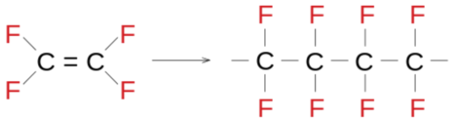


Figure 12.1.2: Polytetrafluoroethylene (also known as Teflon) is formed from the polymerization of tetrafluoroethylene.

12.2: Organic Reactions (Exercises)

These are homework exercises to accompany Chapter 12 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Answers are below the questions.

Questions

12.1: Organic Reactions

(click here for solutions)

Q12.1.1

Explain the difference between hydrolysis and condensation reactions.

Q12.1.2

What is being added in a hydrogenation reaction? Hydration? Bromination?

Q12.1.3

What is the product of the oxidation of a primary alcohol? A secondary alcohol?

Q12.1.4

Why is a condensation reaction also called a dehydration reaction?

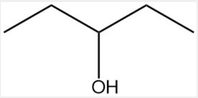
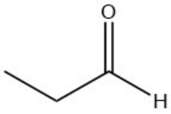
Q12.1.5

Draw the structure for the primary product of each reaction..

a. $CH_3CH{=}CH_2 + Br_2 \rightarrow$
b. $CH_3CH_2CH{=}CH_2 + H_2 \rightarrow$
c. $CH_3CH{=}CHCH_3 + H_2O \rightarrow$
d. $CH_3CH_2CH{=}CH_2 + H_2O \rightarrow$
e. $CH_3CH_2CH{=}CHCH_3 + H_2O \rightarrow$

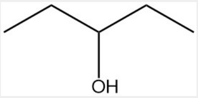
Q12.1.6

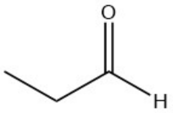
Draw the structure of the compound that would be formed from the single oxidation of each compound.

a.	$H_3C—OH$
b.	
c.	

Q12.1.7

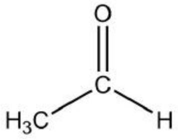
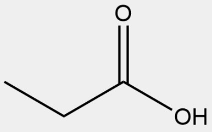
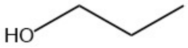
Draw the structure of the compound that would be formed from the complete oxidation of each compound.

a.	$H_3C—OH$
b.	

c.	

Q12.1.8

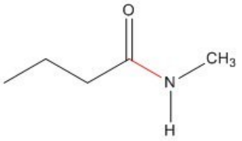
Draw the structure of the compound that would be formed from a single reduction of each compound.

a.	
b.	
c.	

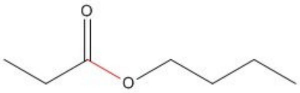
Q12.1.9

Draw the product(s) for the hydrolysis of the given reactant. The red bond in the reactant is the one that is broken.

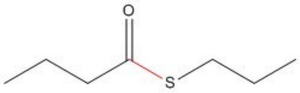
a.



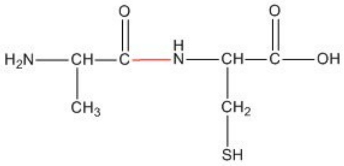
b.



c.



d.



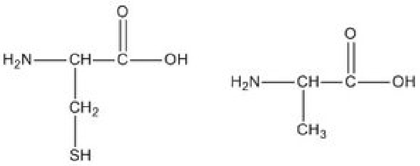
$H_2N-CH(CH_3)-C(=O)-NH-CH(CH_2SH)-C(=O)-OH$

Q12.1.10

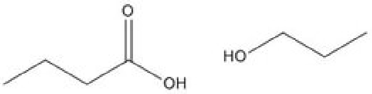
Draw the product(s) for the condensation of the given reactants. The layout of the molecules indicates where the bond will form between them. If there are three molecules, two condensations will occur to produce one final product.

a.

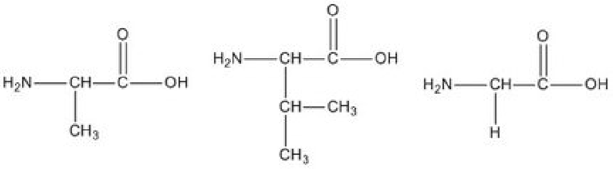
b.



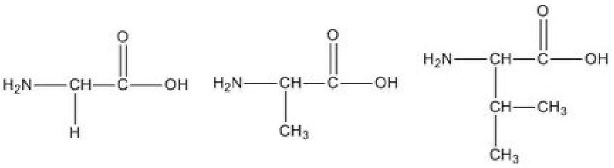
c.



d.



e.



Q12.1.11

Compare the reactants and products in part d and e of 12.1.10. Are the reactants the same or different? Are the products the same or different? Why?

Answers

12.1: Organic Reactions

Q12.1.1

Hydrolysis is a molecule breaking apart after the addition of water. Condensation is two molecules coming together and producing H_2O.

Q12.1.2

hydrogenation = adding hydrogen (H_2)

hydration = adding water

bromination = adding bromine (Br_2)

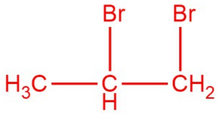
Q12.1.3

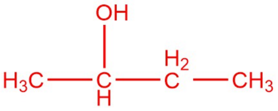
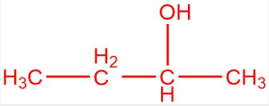
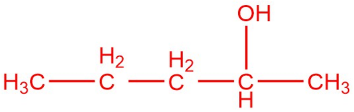
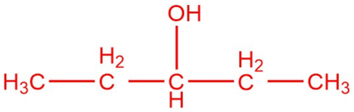
A primary alcohol oxidizes to an aldehyde and then to a carboxylic acid. A secondary alcohol oxidizes to a ketone.

Q12.1.4

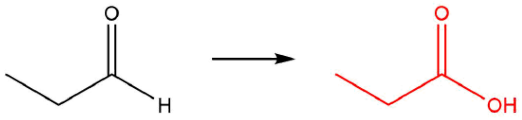
Because water is released during a condensation reaction.

Q12.1.5

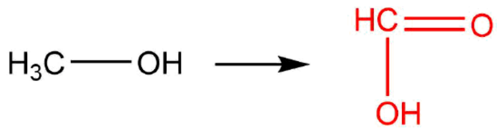
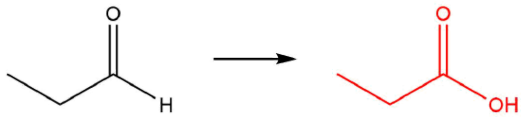
a.	 Each bromine atom is added to one of the carbons in the double bond.

b.	$CH_3CH_2CH_2CH_3$ Each hydrogen atom is added to one of the carbons in the double bonds.
c.	 $H_3C{-}CH(OH){-}CH_2{-}CH_3$ Water is added as H and OH. The carbons in the double bond each have the same number of hydrogen atoms so Markovnikov's rule does <u>not</u> apply. The molecule is symmetrical so adding the OH group to either carbon results in the same molecule.
d.	 $H_3C{-}CH_2{-}CH(OH){-}CH_3$ Water is added as H and OH. The carbon atoms in the double bond have different number of hydrogen atoms so Markovnikov's rule does apply. The H is added to the carbon with more hydrogen atoms and the OH is added to the carbon with fewer hydrogen atoms.
e.	 $H_3C{-}CH_2{-}CH_2{-}CH(OH){-}CH_3$ and  $H_3C{-}CH_2{-}CH(OH){-}CH_2{-}CH_3$ Water is added as H and OH. The carbon atoms in the double bond have the same number of hydrogen atoms on each so Markovnikov's rule does <u>not</u> apply. There are two possible products because adding the OH to the carbon on the left side of the double bond results in a different molecule than adding the OH to the carbon on the right side of the double bond. This differs from c because the molecule is not symmetrical.

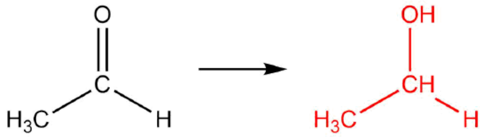
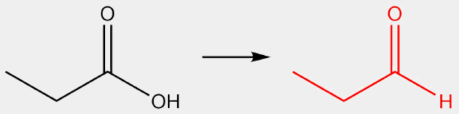
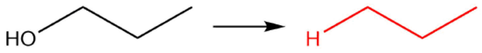
Q12.1.6

a.	$H_3C{-}OH \longrightarrow H_2C{=}O$
b.	
c.	

Q12.1.7

a.	
b.	
c.	

Q12.1.8

a.	
b.	
c.	

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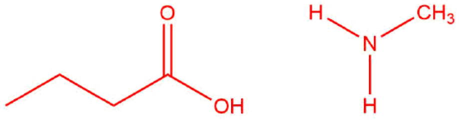
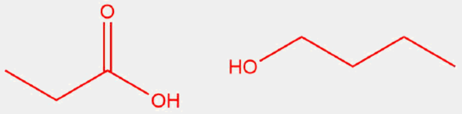
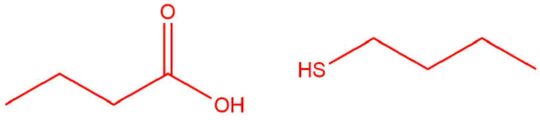
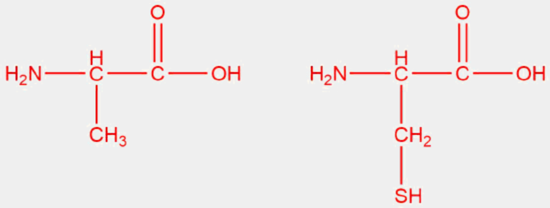
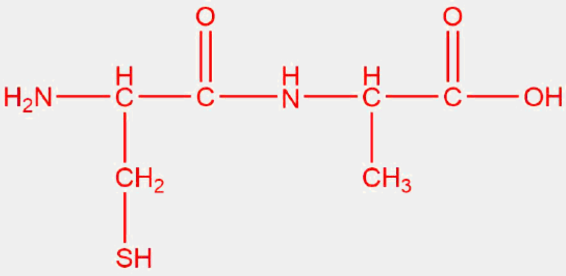
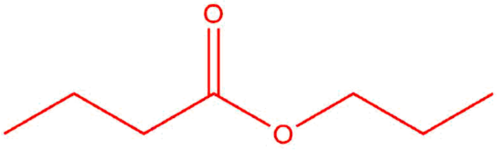
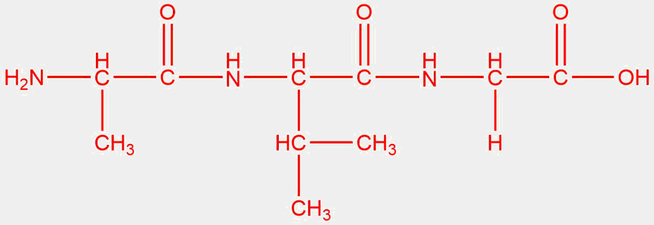
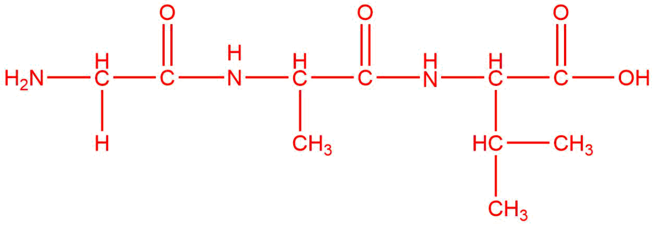
a.	
b.	
c.	
d.	

Q12.1.10

a.	

b.	
c.	
d.	
e.	

Q12.1.11

The reactants are the same, but the products are different because they are combined in a different order.

CHAPTER OVERVIEW

13: Amino Acids and Proteins

Amino acids are molecules containing an amine group(NH_2), a carboxylic acid group(R-C=O-OH) and a side-chain(usually denoted as R) that varies between different amino acids. They are particularly important in biochemistry, where the term usually refers to alpha-amino acids. Proteins are biochemical compounds consisting of one or more polypeptides typically folded into a globular or fibrous form in a biologically functional way.

13.1: Amino Acids

Learning Outcomes

- Identify structural components of an amino acid.
- Define zwitterion and isoelectric point.
- Determine the charge on an amino acid when it is not at the isoelectric point.
- Label amino acids as polar and nonpolar and as acidic, basic, or neutral.

Athletics are very competitive these days at all levels, from school sports to the pros. Everybody is looking for that edge that will make them faster, stronger, more physically fit. One approach taken by many athletes is the use of amino acid supplements. The theory is that the increase in amino acids in the diet will lead to increased protein for muscles. However, the only real benefit comes to the people who make and sell the pills. Studies have not showed any advantage obtained by the athletes themselves. You're much better off just maintaining a healthy diet.

Amino Acids

An **amino acid** is a compound that contains both an amine group $(-\mathrm{NH_2})$ and a carboxyl group $(-\mathrm{COOH})$ in the same molecule. While any number of amino acids can possibly be imagined, biochemists generally reserve the term for a group of 20 amino acids which are formed and used by living organisms. The figure below shows the general structure of an amino acid. Either structure is considered correct for an amino acid.

Basic structure of an amino acid

Figure 13.1.1: An amino acid is an organic molecule that contains an amine group, a carbonyl group, and a side chain (R), all bonded to a central carbon atom. Amino acids can be shown with or without charges. These are equivalent structures.

The amine and carboxyl groups of an amino acid are both covalently bonded to a central carbon atom. That carbon atom is also bonded to a hydrogen atom and an R group. It is this R group which varies from one amino acid to another and is called the amino acid side chain.

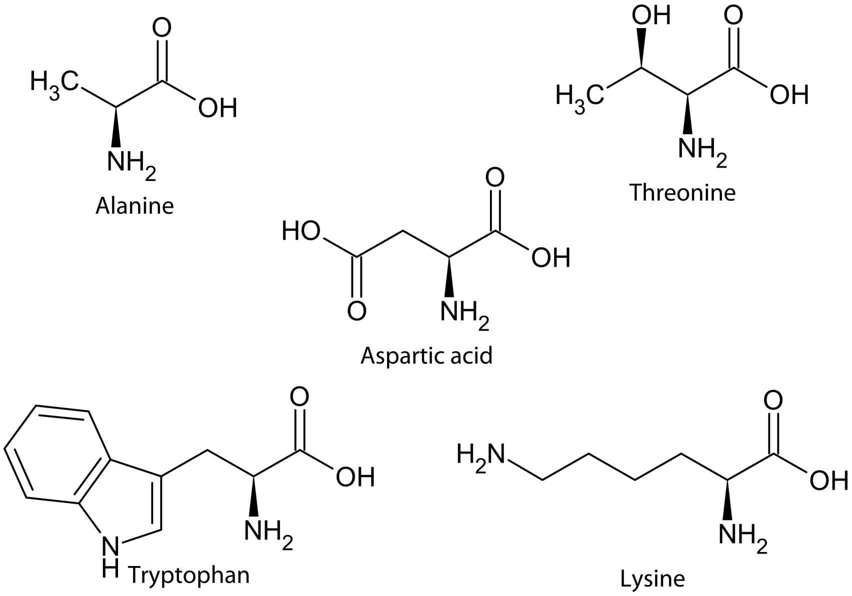
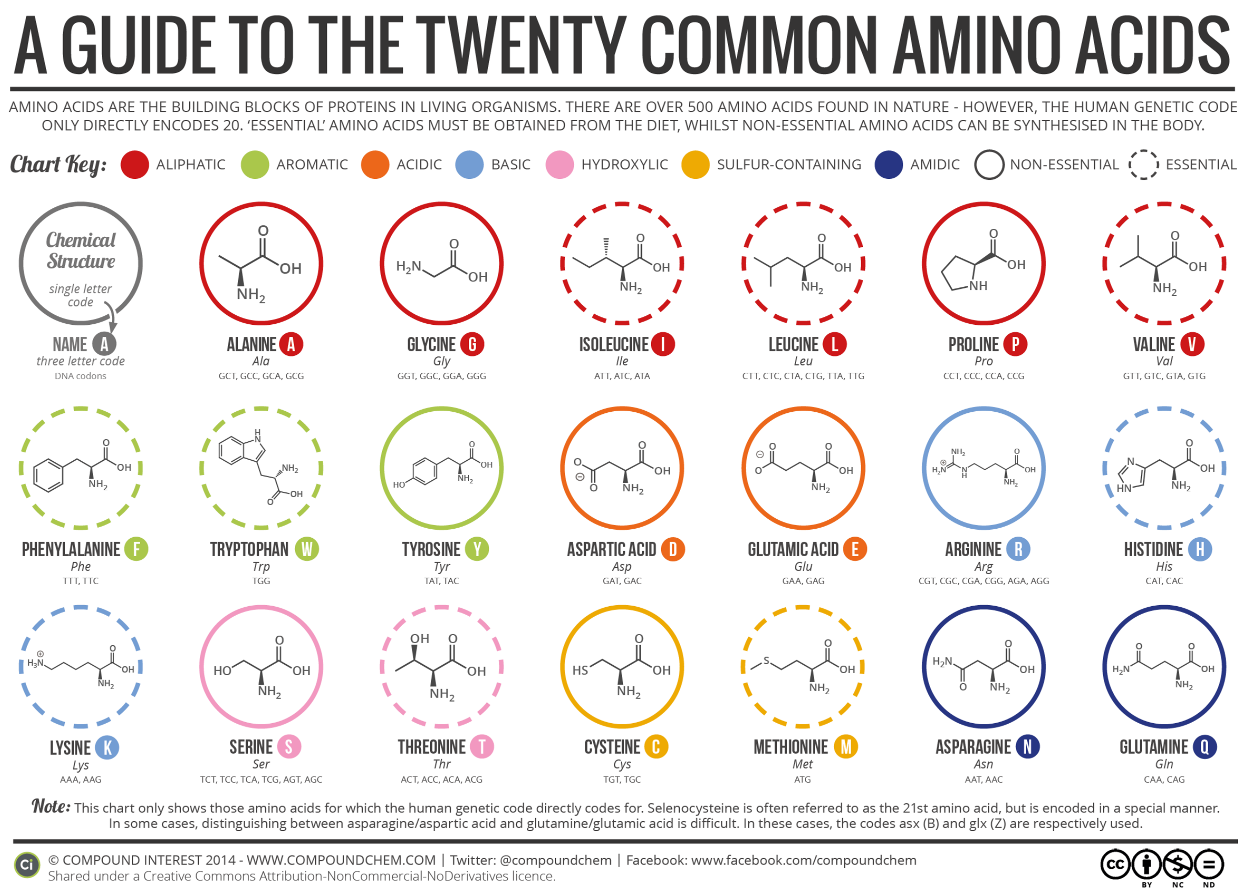


Figure 13.1.2: Five of the twenty biologically relevant amino acids, each having a distinctive side chain (R). Alanine's side chain is nonpolar, while threonine's is polar. Tryptophan is one of several amino acids whose side chain is aromatic. Aspartic acid has an acidic side chain, while lysine has a basic side chain.

The nature of the side chains accounts for the variability in physical and chemical properties of the different amino acids. Each amino acid is grouped based on the properties of the side chain. The groups are designated as polar (hydroxylic, sulfur-containing, amidic) , nonpolar (aliphatic and aromatic), acidic, or basic.



AMINO ACIDS ARE THE BUILDING BLOCKS OF PROTEINS IN LIVING ORGANISMS. THERE ARE OVER 500 AMINO ACIDS FOUND IN NATURE - HOWEVER, THE HUMAN GENETIC CODE ONLY DIRECTLY ENCODES 20. 'ESSENTIAL' AMINO ACIDS MUST BE OBTAINED FROM THE DIET, WHILST NON-ESSENTIAL AMINO ACIDS CAN BE SYNTHESISED IN THE BODY.

Note: This chart only shows those amino acids for which the human genetic code directly codes for. Selenocysteine is often referred to as the 21st amino acid, but is encoded in a special manner. In some cases, distinguishing between asparagine/aspartic acid and glutamine/glutamic acid is difficult. In these cases, the codes asx (B) and glx (Z) are respectively used.

Figure 13.1.3: Twenty amino acids.

In addition to the full name of the amino acid, there are also one-letter and three-letter abbreviations for each. These abbreviations are especially helpful when listing the amino acids in a protein (a chain of many amino acids that will be discussed later).

Rules for classifying amino acids

The following rules (along with two exceptions) can help you classify amino acids as nonpolar, polar acidic (sometimes called acidic), polar basic (sometimes called basic), or polar neutral. We will look at two exceptions but note that the transition from nonpolar to polar neutral is a gradual transition (like the colors of a rainbow) so you may see variations in how amino acids are classified if you look at other sources.

1. **Nonpolar** amino acids (there are 9) contain aliphatic (hydrocarbon) chains or aromatic rings.
2. **Polar acidic** amino acids (2) contain a carboxylic acid (or carboxylate) group in the side chain (R group). This is in addition to the one in the backbone of the amino acid.
3. **Polar basic** amino acids (3) contain an amine (may be neutral or charged) group in the side chain (R group). This is in addition to the one in the backbone of the amino acid.
4. **Polar neutral** amino acids (6) contain a hydroxyl (-OH), sulfur, or amide in the R group).

There are two important exceptions to the above rules.

1. **Tyrosine** has an aromatic group and an -OH group and is considered **polar neutral**.
2. **Methionine** contains a sulfur but as a part of carbon chain. Sulfur has the same electronegativity as carbon, so it is considered **nonpolar**.

Zwitterion

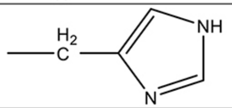
Amino acids are typically drawn either with no charges or with a plus and minus charge (see figure 13.1.1). When an amino acid contains both a plus and a minus charge in the "backbone", it is called a **zwitterion** and has an overall neutral charge. The zwitterion of an amino acid exists at a pH equal to the isoelectric point. Each amino acid has its own pI value based on the properties of the amino acid. At pH values above or below the isoelectric point, the molecule will have a net charge which depends on its pI value as well as the pH of the solution in which the amino acid is found.

pH < pI

When pH is less than pI, there is an excess amount of H^+ in solution. The excess H^+ is attracted to the negatively charged carboxylate ion resulting in its protonation. The carbohydrate ion is protonated, making it neutral, leaving only a positive charge on the amine group. Overall, the amino acid will have a charge of $+1$.

pH > pI

When pH is greater than pI, there is an excess amount of OH^- in solution. The excess OH^- is attracted to the positively charged amine group resulting in the removal of an H^+ ion to form (\ce{H_2O}\). The amine group has a neutral charge leaving only a negative charge on the carboxylate group. Overall, the amino acid will have a charge of -1.

Name (abbreviations)	Side chain	pI
alanine (ala, A)	$-CH_3$	6.11
arginine (arg, R)	$-\overset{H_2}{C}-\overset{H_2}{C}-\overset{H_2}{C}-\overset{H}{N}-\overset{NH}{\overset{\|}{C}}-NH_2$	10.76
asparagine (asn, N)	$-\overset{H_2}{C}-\overset{O}{\overset{\|}{C}}-NH_2$	5.41
aspartic acid, aspartate (asp, D)	$-\overset{H_2}{C}-\overset{O}{\overset{\|}{C}}-OH$	2.87
cysteine (cys, C)	$-\overset{H_2}{C}-SH$	5.02
glutamine (gln, Q)	$-\overset{H_2}{C}-\overset{H_2}{C}-\overset{O}{\overset{\|}{C}}-NH_2$	5.65
glutamic acid, glutamate (glu, E)	$-\overset{H_2}{C}-\overset{H_2}{C}-\overset{O}{\overset{\|}{C}}-OH$	3.08
glycine (gly, G)	$-H$	6.06
histidine (his, H)		7.64
isoleucine (ile, I)	$-\underset{H}{\overset{CH_3}{\overset{\|}{C}}}-\overset{H_2}{C}-CH_3$	6.04
leucine (leu, L)	$-\overset{H_2}{C}-\overset{CH_3}{\overset{\|}{CH}}-CH_3$	6.04
lysine (lys, K)	$-\overset{H_2}{C}-\overset{H_2}{C}-\overset{H_2}{C}-\overset{H_2}{C}-NH_2$	9.47

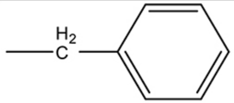
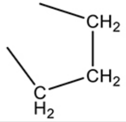
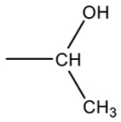
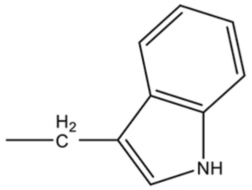
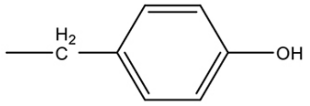
Name (abbreviations)	Side chain	pI
methionine (met, M)	$-\overset{H_2}{C}-\overset{H_2}{C}-S-CH_3$	5.74
phenylalanine (phe, F)		5.91
proline (pro, P)		6.30
serine (ser, S)	$-\overset{H_2}{C}-OH$	5.68
threonine (thr, T)		5.60
tryptophan (trp, W)		5.88
tyrosine (tyr, Y)		5.63
valine (val, V)	$-\overset{CH_3}{\overset{\|}{CH}}-CH_3$	6.02

Figure 13.1.4: Amino acid side chains and pI values.

Example 13.1.1

a. Identify the amino acid pictured below.
b. Find the pI value for the amino acid.
c. Determine how the amino acid will exist at pH = 3.52
d. Determine how the amino acid will exist at pH = 9.34
e. Determine how the amino acid will exist at pH = 5.02

Solution:

a. Look at the side chain to identify the amino acid. The side chain contains $-CH_2SH$ which matches the structure of cysteine.

b. The pI values for amino acids are found in the table of amino acids. For cysteine, pI = 5.02.

c. At pH = 3.52, the H^+ concentration is high (low pH = more acidic = more H^+). Therefore the H^+ will add to the carboxylate ion and neutralize the negative charge. The amino acid will have a positive charge on the amine group left and will have an overall charge of $+1$.

d. At pH = 9.34, the OH^- concentration is high (high pH = more basic = less H^+ = more OH^-). Therefore the OH^- will be attracted to the positively charged amine group and will "steal" an H^+ from it. As a result, the only remaining charge will be on the carboxylate ion so the amino acid will have a -1 charge.

e. At pH = 5.02, the pH = pI so the amino acid will exist as the zwitterion with both the positive and negative charges as shown above.

Contributors and Attributions

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

13.2: Peptides

Learning Outcomes

- Define polypeptide.
- Identify amide bond.
- Predict product of condensation of amino acids.
- Name polypeptides given the abbreviation of the amino acids.

Peptide cells in our bodies have an intricate mechanism for the manufacture of proteins. Humans have to use other techniques in order to synthesize the same proteins in a lab. The chemistry of peptide synthesis is complicated. Both active groups on an amino acid can react and the amino acid sequence must be a specific one in order for the protein to function. Robert Merrifield developed the first synthetic approach for making proteins in the lab, a manual approach which was lengthy and tedious (and, he won the Nobel Prize in Chemistry in 1984 for his work). Today, however, automated systems can crank out a peptide in a very short period of time.

Peptides

A **peptide** is a combination of amino acids in which the amine group of one amino acid has undergone a reaction with the carboxyl group of another amino acid. The reaction is a condensation reaction, forming an amide group $(\mathrm{CO{-}N})$, shown below.

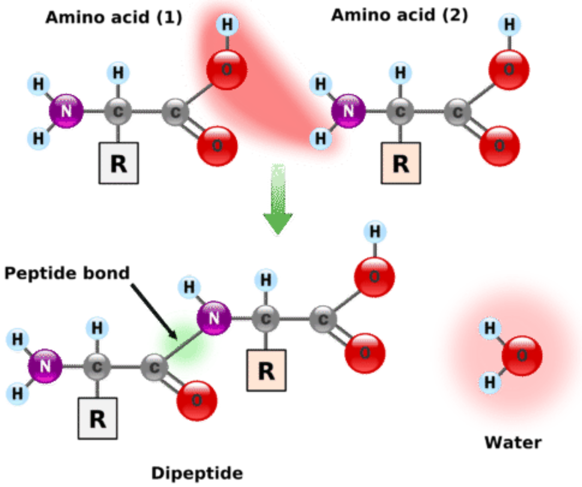
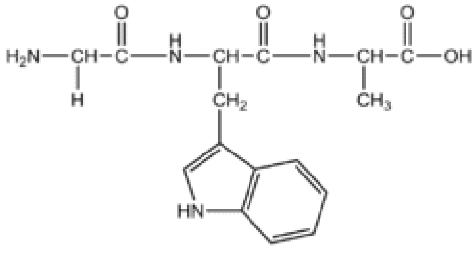


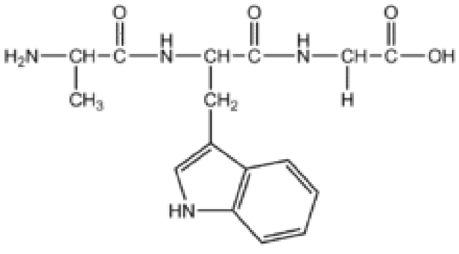
Figure 13.2.1: Amino acids undergo condensation to form a molecule called a dipeptide. The $\mathrm{C{-}N}$ bond is called a peptide bond.

A **peptide bond** is the amide bond that occurs between the amine nitrogen of one amino acid and the carboxyl carbon of another amino acid. The resulting molecule is called a dipeptide. Notice that the particular side chains of each amino acid are irrelevant since the R groups are not involved in the peptide bond.

The dipeptide has a free amine group on one end of the molecule (known as the N-terminus) and a free carboxyl group on the other end (known as the C-terminus). Each is capable of extending the chain through the formation of another peptide bond. The particular sequence of amino acids in a longer chain is called an amino acid sequence. By convention, the amino acid sequence is listed in the order such that the free amine group is on the left end of the molecule and the free carboxyl group is on the right end of the molecule. For example, suppose that a sequence of the amino acids glycine, tryptophan, and alanine is formed with the free amine group as part of the glycine and the free carboxyl group as part of the alanine. The amino acid sequence can be easily written using the abbreviations as Gly-Trp-Ala. This is a different sequence from Ala-Trp-Gly because the free amine and carboxyl groups would be on different amino acids in that case.



Gly-Trp-Ala



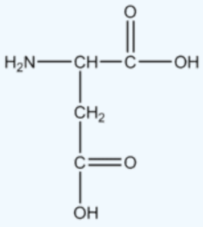
Ala-Trp-Gly

Example 13.2.1

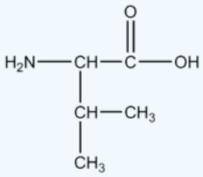
Draw the polypeptide Asp-Val-Ser.

Solution

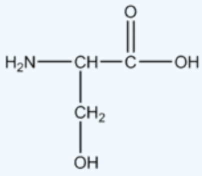
1. Identify the structures of each of the three given amino acids and draw them in the same order as given in the name.



Asp

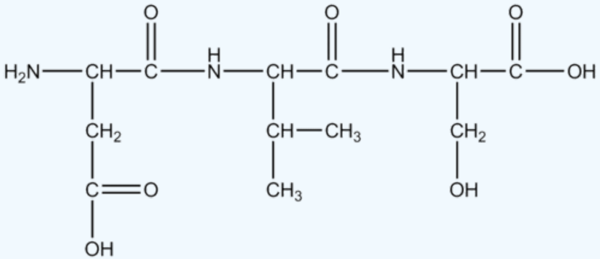


Val



Ser

2. Leaving the order the same, connect the amino acids to one another by forming peptide bonds. Note that the order given in the name is the same way the amino acids are connected in the molecule. The first one listed is always the N-terminus of the polypeptide.



Asp-Val-Ser

Example 13.2.2

List all of the possible polypeptides that can be formed from cysteine (Cys), leucine (Leu), and arginine (Arg).

Solution

Although there are only three amino acids, the order in which they are bonded changes the identity, properties, and function of the resulting polypeptide. There are six possible polypeptides formed from these three amino acids.

Cys-Leu-Arg

Cys-Arg-Leu

Leu-Cys-Arg

Leu-Arg-Cys

Arg-Cys-Leu

Arg-Leu-Cys

Contributors and Attributions

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

13.3: Protein Structure

Learning Outcomes

- Describe the four levels of protein structure.
- Identify the two types of secondary structure in proteins.
- Describe the interactive forces in each level of protein structure.
- Distinguish between globular and fibrous proteins.
- Define denaturation of proteins.
- Identify ways in which proteins are denatured.

Hemoglobin is a complex protein which has a quaternary structure and contains iron. There are four subunits in the hemoglobin molecule - two alpha subunits and two beta subunits. Each subunit contains one iron ion, whose oxidation state changes from $+2$ to $+3$ and back again, depending upon the environment around the iron. When oxygen binds to the iron, the three-dimensional shape of the molecule changes. Upon release of the oxygen to the cells, the shape changes again.

With hemoglobin of normal structure, this shift in conformation does not present any problems. However, individuals with hemoglobin S do experience serious complications. This hemoglobin has one amino acid in the two beta chains that is different from the amino acid at that point in the primary structure of normal hemoglobin. The result of this one structural change is aggregation of the individual protein molecules when oxygen is released. Adjacent hemoglobin molecules come in contact with one another and clump up, causing the red cells to deform and break.

This abnormality, known as sickle cell, is genetic in nature. A person may inherit the gene from one parent and have sickle cell trait (only some of the hemoglobin is hemoglobin S), which is usually not life-threatening. Inheriting the gene from both parents will result in sickle cell disease, a very serious condition.

Proteins

A **polypeptide** is a sequence of amino acids between ten and one hundred in length. A **protein** is a peptide that is greater than one hundred amino acids in length. Proteins are very prevalent in living organisms. Hair, skin, nails, muscles, and the hemoglobin in red blood cells are some of the important parts of your body that are made of different proteins. The wide array of chemical and physiological properties of proteins is a function of their amino acid sequences. Since proteins generally consist of one hundred or more amino acids, the number of amino acid sequences that are possible is virtually limitless.

The three-dimensional structure of a protein is very critical to its function. This structure can be broken down into four levels. The **primary structure** is the amino acid sequence of the protein. The amino acid sequence of a given protein is unique and defines the function of that protein. Peptide bonds form the connections between the amino acids.

The **secondary structure** is a highly irregular sub-structure of the protein. The two most common types of protein secondary structure are the alpha helix (see figure below) and the beta sheet (see figure below). An alpha helix consists of amino acids that adopt a spiral shape. A beta pleated sheet (like a fan-folded paper) is alternating rows of amino acids that line up in a side-by-side fashion. In both cases, the secondary structures are stabilized by extensive hydrogen bonding between the side chains. The interaction of the various side chains in the amino acid, specifically the hydrogen bonding, leads to the adoption of a particular secondary structure. The hydrogen bonding occurs between amino acids that are close to each other in the primary structure.

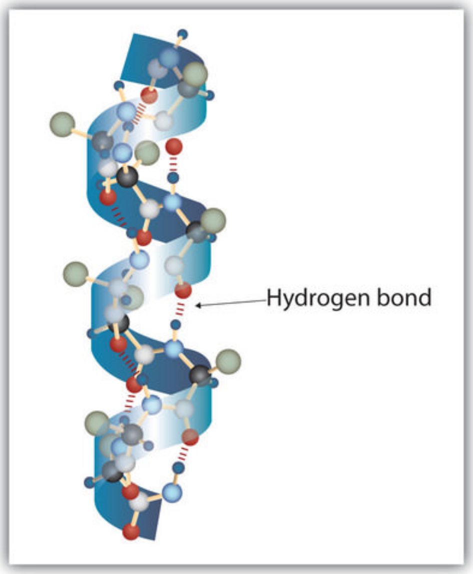


Figure 13.3.1: Alpha helix.

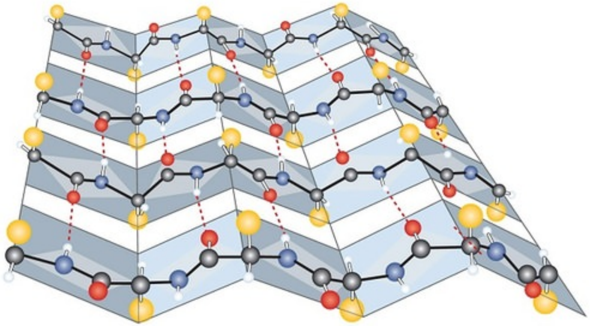
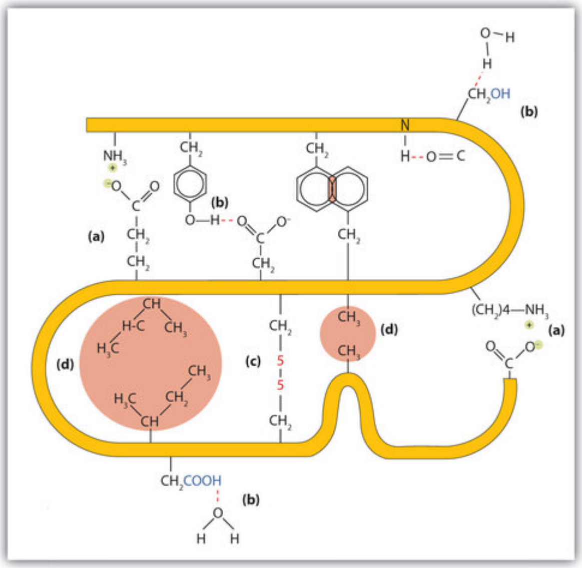


Figure 13.3.2: Beta pleated sheet.

The **tertiary structure** is the overall three-dimensional structure of the protein. A typical protein consists of several sections of a specific secondary structure (alpha helix or beta sheet) along with other areas in which a more random structure occurs. These areas combine to produce the tertiary structure. The tertiary structure is stabilized by forces similar to the intermolecular forces previously seen between molecules. These attractive forces include London dispersion forces (hydrophobic), hydrogen bonding, dipole-dipole forces, ion-dipole interactions, ionic salt bridges, and disulfide bonds (see figure below).



Some protein molecules consist of multiple protein subunits. The **quaternary structure** of a protein refers to the specific interaction and orientation of the subunits of that protein. The quaternary structure is a result of the same types of interactions as seen in tertiary structure but between different subunits. Quaternary structures can have different numbers of subunits. For example, hemoglobin contains four subunits while insulin contains two subunits.

Hemoglobin is a very large protein found in red blood cells and whose function is to bind and carry oxygen throughout the bloodstream. Hemoglobin consists of four subunits - two α subunits (yellow) and two β subunits (gray) - which then come together in a specific and defined way through interactions of the side chains (see figure below). Hemoglobin also contains four iron atoms, located in the middle of each of the four subunits. The iron atoms are part of a structure called a porphyrin, shown in red in the figure.

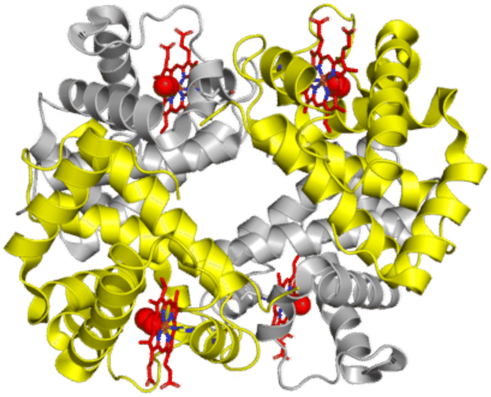


Figure 13.3.3: Hemoglobin.

Some proteins consist of only one subunit and thus do not have a quaternary structure. The figure below diagrams the interaction of the four levels of protein structure.

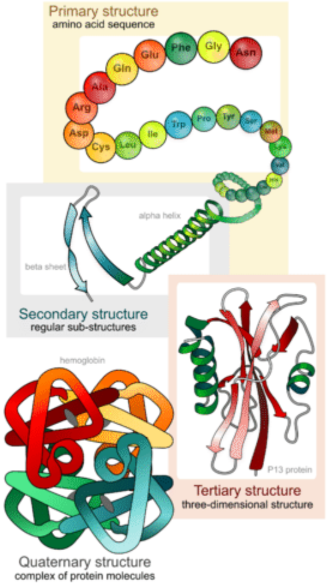


Figure 13.3.4: The four levels of protein structure.

Globular and Fibrous Proteins

Once proteins form and have developed all levels of their structure, they can be classified as either fibrous or globular. These classifications give the basic shape of the entire protein molecule. While many proteins are globular proteins (see figure below), keratin proteins are fibrous (see figure below) and make up the hair, nails, and the outer layer of skin.

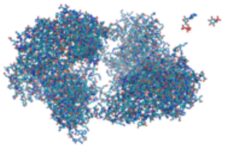


Figure 13.3.5: Globular protein.

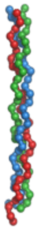


Figure 13.3.6: Fibrous protein.

Denaturation of Proteins

The highly organized structures of proteins are truly masterworks of chemical architecture. But highly organized structures tend to have a certain delicacy, and this is true of proteins. Denaturation is the term used for any change in the three-dimensional structure of a protein that renders it incapable of performing its assigned function. A denatured protein cannot do its job because there is a change in the secondary, tertiary, or quaternary structure (see figure below). A wide variety of reagents and conditions, such as heat, organic compounds, pH changes, and heavy metal ions can cause protein denaturation. Anyone who has fried an egg has observed denaturation. The clear egg white turns opaque as the albumin denatures and coagulates.

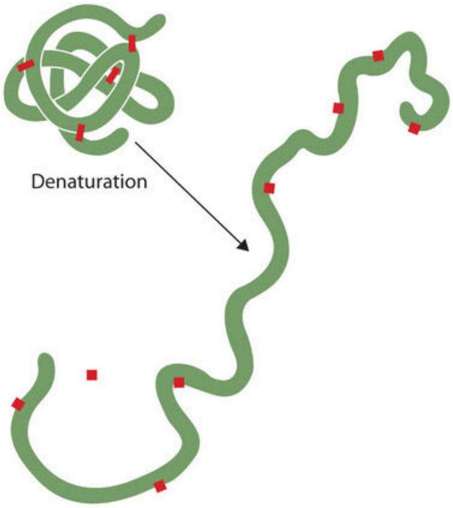


Figure 13.3.7: Denaturation of a protein.

The primary structures of proteins are quite sturdy. In general, fairly vigorous conditions are needed to hydrolyze peptide bonds. At the secondary through quaternary levels, however, proteins are quite vulnerable to attack, though they vary in their vulnerability to denaturation. The delicately folded globular proteins are much easier to denature than are the tough, fibrous proteins of hair and skin.

There are a variety of ways to denature proteins including those below.

- Heat above $50^{o}C$
- Strong acids
- Strong bases
- Ionic compounds (i.e. $NaCl$)
- Reducing agents
- Detergents
- Heavy metal ions
- Agitation

If you have ever had a hair permanent or chemically straightened your hair, the process involved the denaturation of proteins. The reducing agent (usually an ammonium compound) breaks the disulfide bonds in the hair. The hair is then curled or straightened which aligns the amino acids in a different pattern. An oxidizing agent is applied and the disulfide bonds reform between different amino acids. The change is permanent for the hair that you have at the time but new hair growing in will have the structure of the original proteins and your hair is back to its normal state.

Contributors and Attributions

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

13.4: Amino Acids and Proteins (Exercises)

These are homework exercises to accompany Chapter 13 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health.

Questions

13.1: Amino Acids

(click here for solutions)

Q13.1.1

Read the material at http://hyperphysics.phy-astr.gsu.edu/hbase/organic/essam.html and answer the following questions:

a. What are essential amino acids?
b. What are nonessential amino acids?
c. What happens if you are deficient in an amino acid?

Q13.1.2

Draw the functional groups present in all amino acids.

Q13.1.3

Complete the following for threonine, lysine, and tyrosine.

a. Draw the amino acid.
b. Circle the side chain.
c. Identify whether it is polar, nonpolar, acidic, or basic.
d. At what pH will it exist as a zwitterion?
e. What is the range of pH values when it will be positively charged?
f. What is the range of pH values when it will be negatively charged?

13.2: Peptides

(click here for solutions)

Q13.2.1

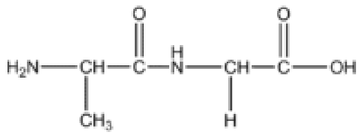
Draw the two dipeptides formed from each pair of amino acids.

a. tyrosine and lysine
b. threonine and gluatmine
c. alanine and histidine

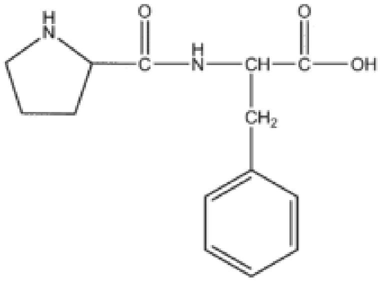
Q13.2.2

Draw and give the full names of the amino acids in the following dipeptides.

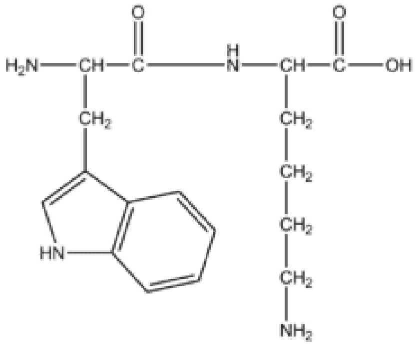
a.



b.



c.



Q13.2.3

List of all of the possible polypeptides that can be formed from threonine, alanine, and phenylalanine (use three character abbreviations for each amino acid).

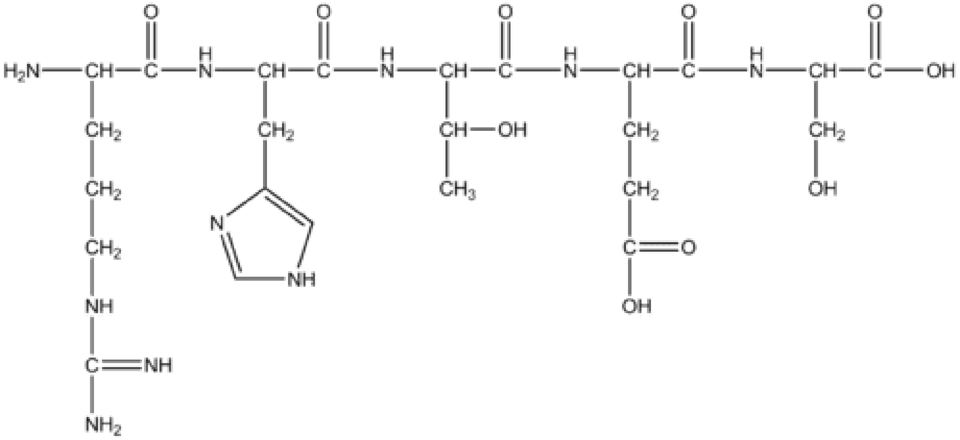
Q13.2.4

Draw the following polypeptides.

a. Ser-Tyr-Gln
b. Lys-Met-Gly

Q13.2.5

Identify each of the amino acids in the polypeptide and then name it using the three character abbreviations.



13.3: Protein Structure

(click here for solutions)

Q13.3.1

Describe the four levels of protein structure.

Q13.3.2

What levels of structure involve hydrogen bonding?

Q13.3.3

What types of structure is the result of interactions between amino acids that are far apart in the primary structure?

Q13.3.4

What types of interactions hold the secondary structure together?

Q13.3.5

What types of interactions hold the tertiary structure together?

Q13.3.6

What levels of structure are affected by denaturation?

Q13.3.7

A protein has one subunit. Would it have a quaternary structure?

Answers

13.1: Amino Acids

Q13.1.1

a. Essential amino acids are those you get from your diet.
b. Nonessential amino acids are produced in the body.
c. Illness and/or degradation of body's proteins.

Q13.1.2

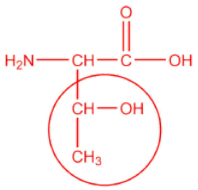
amine and carboxylic acid

Q13.1.3

Complete the following for threonine, lysine, and tyrosine.

threonine

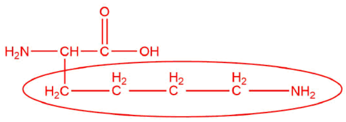
a.



b.
c. polar
d. 5.60
e. < 5.60
f. > 5.60

lysine

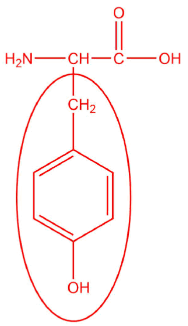
a.



b.
c. basic
d. 9.47
e. < 9.47
f. > 9.47

tyrosine

a.



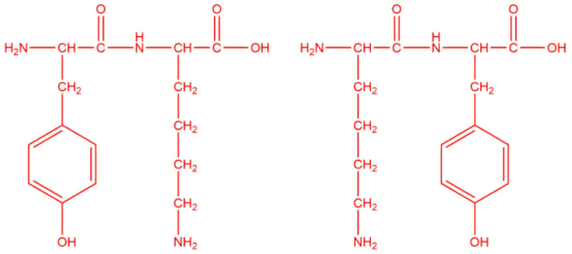
b.
c. polar
d. 5.63
e. < 5.63
f. > 5.63

13.2: Peptides

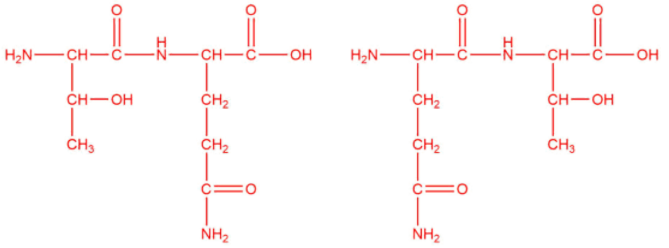
Q13.2.1

Draw the two dipeptides formed from each pair of amino acids.

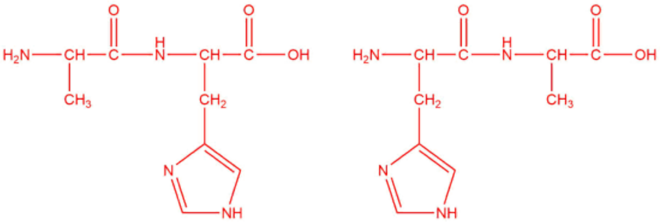
a.



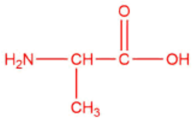
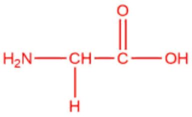
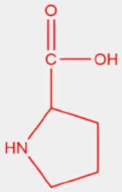
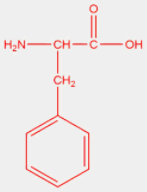
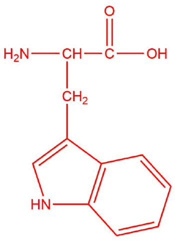
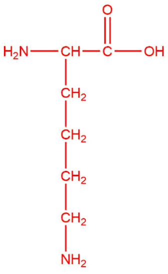
b.



c.



Q13.2.2

a.	 alanine	 glycine
b.	 proline	 phenylalanine
c.	 tryptophan	 lysine

Q13.2.3

Thr-Ala-Phe

Thr-Phe-Ala

Ala-Thr-Phe

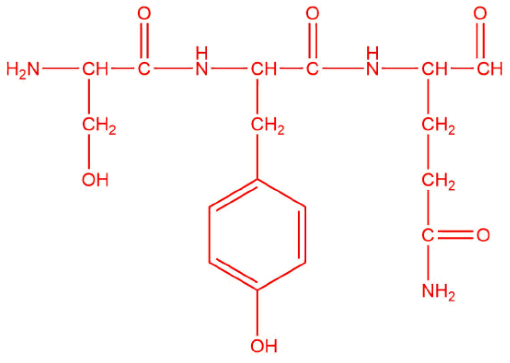
Ala-Phe-Thr

Phe-Ala-Thr

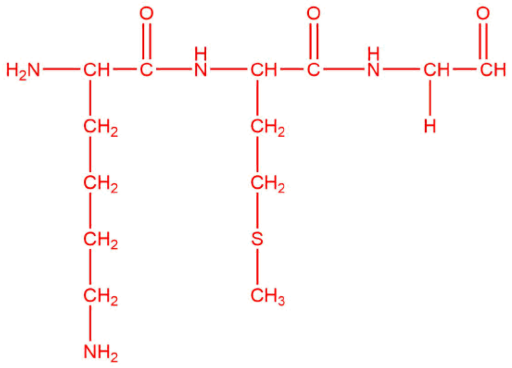
Phe-Thr-Ala

Q13.2.4

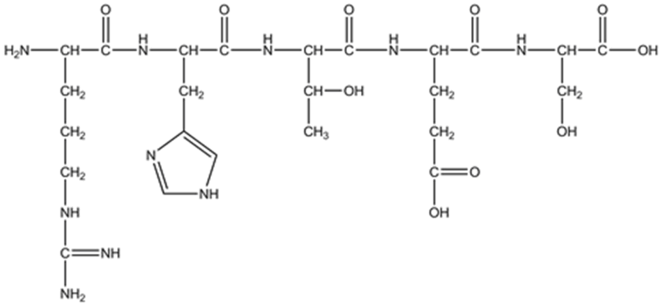
a.



b.



Q13.2.5



arginine histidine threonine glutamic acid serine

Arg-His-Thr-Glu-Ser

13.3: Protein Structure

Q13.3.1

Primary - sequence of amino acids

Secondary - alpha helix and Beta-pleated sheets held together by hydrogen bonds

Tertiary - third level of structure of protein often forming globular or fibrous structure, held together by variety of attractive forces

Quaternary - complex of multiple proteins held together to function as one, held together by variety of attractive forces (same as tertiary)

Q13.3.2

secondary, tertiary, and quaternary structures

Q13.3.3

tertiary structures

Q13.3.4

hydrogen bonds

Q13.3.5

London dispersion forces, hydrogen bonds, dipole-dipole forces, ion-dipole interactions, salt bridges, and disulfide bonds

Q13.3.6

secondary, tertiary, and quaternary

Q13.3.7

No, a quaternary structure must have multiple subunits.

CHAPTER OVERVIEW

14: Biological Molecules

Biomolecules include large macromolecules (or polyanions) such as proteins, carbohydrates, lipids, and nucleic acids, as well as small molecules such as primary metabolites, secondary metabolites, and natural products.

14.1: Enzymes

Learning Outcomes

- Explain the role of an enzyme in the body.
- Define active site, substrate, and allosteric site.
- Distinguish between competitive and noncompetitive inhibitors.
- Describe the lock and key vs. induced-fit model of enzymes.
- Provide the characteristics of a cofactor and a coenzyme.
- Describe how substrate concentration, pH, and temperature affect enzyme activity.
- Interpret graphs of reaction rate vs. reaction conditions.

The first enzyme to be isolated was discovered in 1926 by American chemist James Sumner, who crystallized the protein. The enzyme was urease, which catalyzes the hydrolytic decomposition of urea, a component of urine, into ammonia and carbon dioxide.

$$H_2NCON_2\,(aq) + H_2O\,(l) \overset{\text{urease}}{\rightarrow} 2NH_3\,(g) + CO_2\,(g) \tag{14.1.1}$$

His discovery was ridiculed at first because nobody believed that enzymes would behave the same way that other chemicals did. Sumner was eventually proven right and won the Nobel Prize in Chemistry in 1946.

Enzymes and Biochemical Reactions

Most chemical reactions within organisms would be impossible under the conditions in cells. For example, the body temperature of most organisms is too low for reactions to occur quickly enough to carry out life processes. Reactants may also be present in such low concentrations that it is unlikely they will meet and collide. Therefore, the rate of most biochemical reactions must be increased by a catalyst. A **catalyst** is a chemical that speeds up chemical reactions. In organisms, catalysts are called **enzymes**. Essentially, enzymes are biological catalysts.

Like other catalysts, enzymes are not reactants in the reactions they control. They help the reactants interact but are not used up in the reactions. Instead, they may be used over and over again. Unlike other catalysts, enzymes are usually highly specific for particular chemical reactions. They generally catalyze only one or a few types of reactions.

Enzymes are extremely efficient in speeding up reactions. They can catalyze up to several million reactions per second. As a result, the difference in rates of biochemical reactions with and without enzymes may be enormous. A typical biochemical reaction might take hours or even days to occur under normal cellular conditions without an enzyme, but less than a second with an enzyme.

Figure 14.1.1 diagrams a typical enzymatic reaction. A **substrate** is the molecule or molecules on which the enzyme acts. In the urease catalyzed reaction, urea is the substrate.

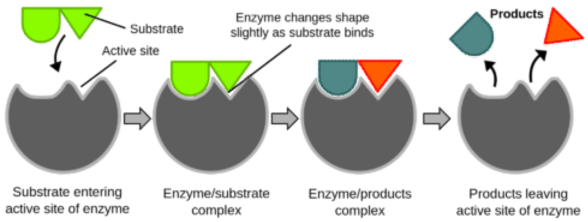


Figure 14.1.1: The sequence of steps for a substrate binding to an enzyme in its active site, reacting, then being released as products.

The first step in the reaction is that the substrate binds to a specific part of the enzyme molecule, known as the active site. The binding of the substrate is dictated by the shape of each molecule. Side chains on the enzyme interact with the substrate in a specific way, resulting in the making and breaking of bonds. The **active site** is the place on an enzyme where the substrate binds. An enzyme folds in such a way that it typically has one active site, usually a pocket or crevice formed by the folding pattern of the protein. Because the active site of an enzyme has such a unique shape, only one particular substrate is capable of binding to that enzyme. In other words, each enzyme catalyzes only one chemical reaction with only one substrate. Once the enzyme/substrate complex is formed, the reaction occurs and the substrate is transformed into products. Finally, the product molecule or molecules are released from the active site. Note that the enzyme is left unaffected by the reaction and is now capable of catalyzing the reaction of another substrate molecule.

For many enzymes, the active site follows a **lock and key** (A in the figure below) model where the substrate fits exactly into the active site. The enzyme and substrate must be a perfect match so the enzyme only functions as a catalyst for one reaction. Other enzymes have an **induced fit** (B in the figure below) model. In an induced fit model, the active site can make minor adjustments to accommodate the substrate. This results in an enzyme that is capable of interacting with a small group of similar substrates. Look at the shape of the active site compared to the shape of the substrate in B of the figure below. The active site adjusts to accommodate the substrate.

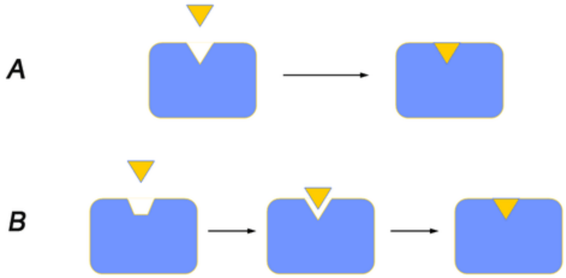


Figure 14.1.2: (A) Lock and key enzyme model and (B) induced fit enzyme model.

Inhibitors

An **inhibitor** is a molecule which interferes with the function of an enzyme, either by slowing or stopping the chemical reaction. Inhibitors can work in a variety of ways, but one of the most common is illustrated in the figure below.

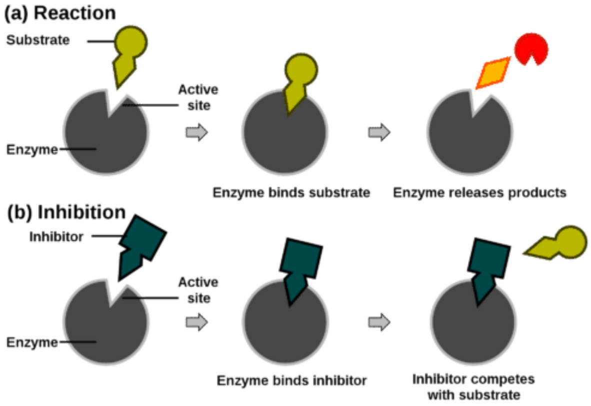


Figure 14.1.3: A competitive inhibitor is a molecule that binds to the active site of an enzyme without reacting, thus preventing the substrate from binding.

A **competitive inhibitor** binds competitively at the active site and blocks the substrate from binding. Since no reaction occurs with the inhibitor, the enzyme is prevented from catalyzing the reaction.

A **non-competitive inhibitor** does not bind at the active site. It attaches at an **allosteric site**, which is some other site on the enzyme, and changes the shape of the protein. The allosteric site is any site on the enzyme that is not the active site. The attachment of the non-competitive inhibitor to the allosteric site results in a shift in three-dimensional structure that alters the shape of the active site so that the substrate will no longer fit in the active site properly (see figure below).

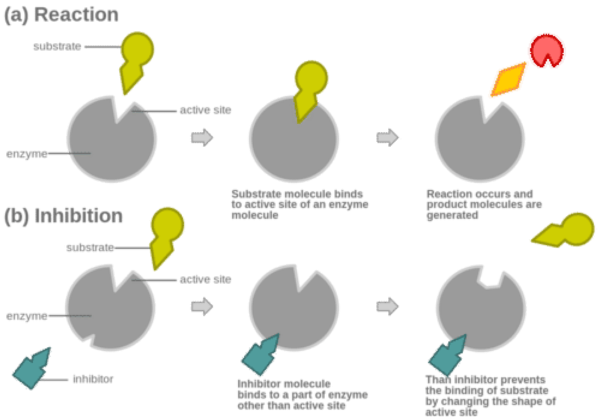


Figure 14.1.4: Non-competitive inhibition.

Cofactors and Coenzymes

Some enzymes require the presence of another substrate as a "helper" molecule in order to function properly. **Cofactors** and **coenzymes** serve in this role. Cofactors are inorganic species and coenzymes are small organic molecules. Many vitamins, such as B vitamins, are coenzymes. Some metal ions which function as cofactors for various enzymes include zinc, magnesium, potassium, and iron.

Catalytic Activity of Enzymes

Enzymes generally lower activation energy by reducing the energy needed for reactants to come together and react. One way that enzymes act is to bring reactants (substrates) together so they don't have to expend energy moving about until they collide at random. Enzymes bind both reactant molecules (substrates), tightly and specifically, at a site on the enzyme's active site. Enzymes can also bring molecules to the active site to break them apart. For example, sucrase is the enzyme for the breakdown of sucrose which enters the active site of the enzyme and helps weaken the interactions between the fructose and glucose that make up sucrose. Sucrase is specific to the breakdown of sucrose as are most enzymes. The active site is specific for the reactants of the biochemical reaction the enzyme catalyzes. Similar to puzzle pieces fitting together, the active site can only bind certain substrates. The activities of enzymes also depend on the temperature, concentration, and the pH of the surroudings.

Concentration

As with most reactions, the concentration of the reactant(s) affects the reaction rate. This is also true in enzyme concentration. When either substrate or enzyme concentration is low, the rate of the reaction will be slower than where there are higher concentrations. The two species must interact for a reaction to occur and higher concentrations of one or both will result in more effective interactions between the two.

However, continuing to increase the substrate's concentration will not always increase the reaction rate. This is because at some point, all of the enzymes will be occupied and unavailable to bind with another substrate molecule until the substrate forms a product molecule and is released from the enzyme.

pH

Some enzymes work best at acidic pHs, while others work best in neutral environments. For example, digestive enzymes secreted in the acidic environment (low pH) of the stomach help break down proteins into smaller molecules. The main digestive enzyme in the stomach is **pepsin**, which works best at a pH of about 1.5. These enzymes would not work optimally at other pHs. Trypsin is another enzyme in the digestive system, which breaks protein chains in food into smaller particles. **Trypsin** works in the small intestine, which is not an acidic environment. Trypsin's optimum pH is about 8.

Different reactions and different enzymes will achieve their maximum rate at certain pH values. As shown in the figure below, the enzyme achieves a maximum reaction rate at a pH of 4. Notice that the reaction will continue at lower and higher pH values because the enzyme will still function at other pH values but will not be as effective. At very high or very low pH values, denaturation will occur because an enzyme is just a protein with a specific function.

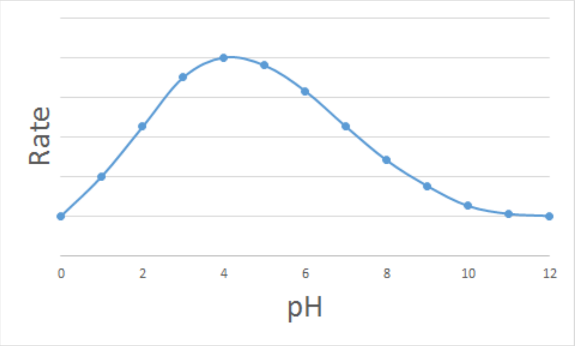


Figure 14.1.5: Relationship between rate and pH.

Temperature

As with pH, reactions also have an ideal temperature where the enzyme functions most effectively. It will still function at higher and lower temperatures, but the rate will be less. For many biological reactions, the ideal temperature is at physiological conditions which is around 37^{o}C which is normal body temperature. Many enzymes lose function at lower and higher temperatures. At higher temperatures, an enzyme's shape deteriorates. Only when the temperature comes back to normal does the enzyme regain its shape and normal activity unless the temperature was so high that it caused irreversible damage.

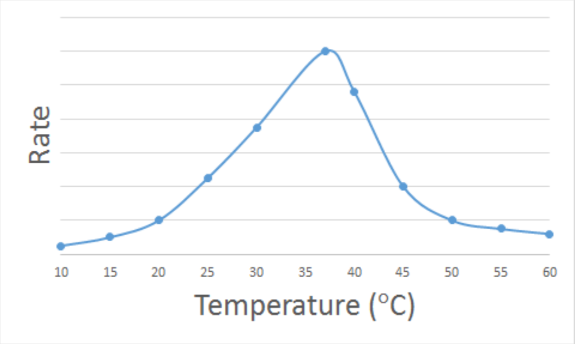


Figure 14.1.6: Relationship between temperature and rate.

Supplemental Resources

- Enzymes: https://youtu.be/E90D4BmaVJM

Contributors and Attributions

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

14.2: Lipids and Triglycerides

Learning Outcomes

- Identify the basic structural features of fatty acids.
- Distinguish between saturated and unsaturated fatty acids.
- Define essential fatty acids.
- Identify the structural features of a triglyceride.
- Label the type of bond formed between fatty acids and glycerol in triglyceride.

There is a lot of interest these days on healthy diets as well as concerns about heart problems. There is also a strong market for the sales of omega-3 fatty acids, which are said to help lower fat levels in blood. But too many people rely on the supplements to help their hearts and don't understand the chemistry behind it all. Yes, taking omega-3 fatty acids will give you some of the fatty acids your body requires. No, this is not a substitute for eating a healthy diet and exercising. You can't sit in front of the TV set, eating your large pizza, and expect these pills to keep you health. You've got to do things the hard way - eat your vegetables and get some exercise.

Fatty Acids

A **lipid** is an organic compound such as fat or oil. Organisms use lipids to store energy, but lipids have other important roles as well. Lipids consist of repeating units called fatty acids. **Fatty acids** are organic compounds that have the general formula $\mathrm{CH_3(CH_2)_nCOOH}$, where n usually ranges from 2 to 28 and is always an even number. There are two types of fatty acids: saturated fatty acids and unsaturated fatty acids.

Saturated Fatty Acids

In **saturated fatty acids**, carbon atoms are bonded to as many hydrogen atoms as possible. This causes the molecules to form straight chains, as shown in the figure below. The straight chains can be packed together very tightly, allowing them to store energy in a compact form. This explains why saturated fatty acids are solids at room temperature. Animals use saturated fatty acids to store energy.

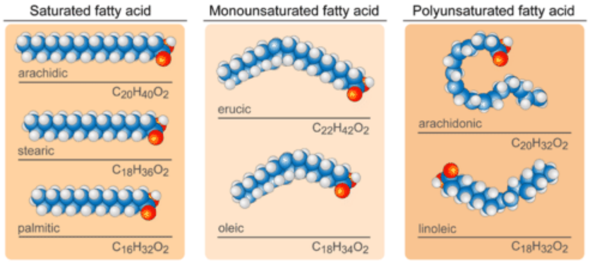


Figure 14.2.1: Structures of saturated and unsaturated fatty acids.

Unsaturated Fatty Acids

In **unsaturated fatty acids**, some carbon atoms are not bonded to as many hydrogen atoms as possible due to the presence of one or more double bonds in the carbon chain. Instead, they are bonded to other groups of atoms. Wherever carbon binds with these other groups of atoms, it causes chains to bend (see figure above). The bent chains cannot be packed together very tightly, so unsaturated fatty acids are liquids at room temperature. Plants use unsaturated fatty acids to store energy.

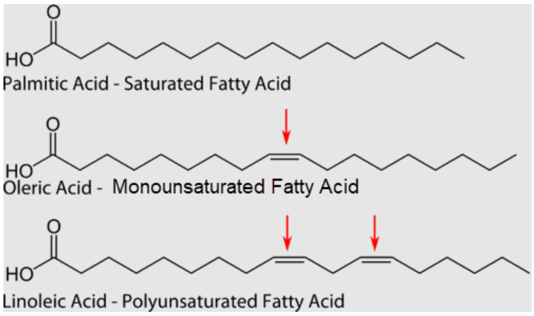


Figure 14.2.2: Saturated fatty acids have only single bonds while monounsaturated fats have one double bond and polyunsaturated fats have more than one double bond.

Lipids and Diet

Unsaturated fat is generally considered to be healthier because it contains fewer calories than an equivalent amount of saturated fat. Additionally, high consumption of saturated fats is linked to an increased risk of cardiovascular disease. Some examples of foods with high concentrations of saturated fats include butter, cheese, lard, and some fatty meats. Foods with higher concentrations of unsaturated fats include nuts, avocado, and vegetable oils such as canola oil and olive oil.

Humans need lipids for many vital functions, such as storing energy and forming cell membranes. Lipids can also supply cells with energy. In fact, a gram of lipids supplies more than twice as much energy as a gram of carbohydrates or proteins. Lipids are necessary in the diet for most of these functions. Although the human body can manufacture most of the lipids it needs, there are others, called essential fatty acids, that must be consumed in food. Essential fatty acids include omega-3 and omega-6 fatty acids. Both of these fatty acids are needed for important biological processes, not just for energy.

Although some lipids in the diet are essential, excess dietary lipids can be harmful. Because lipids are very high in energy, eating too many may lead to unhealthy weight gain. A high-fat diet may also increase lipid levels in the blood. This, in turn, can increase the risk for health problems such as cardiovascular disease. The dietary lipids of most concern are saturated fatty acids, trans fats, and cholesterol. For example, cholesterol is the lipid mainly responsible for narrowing arteries and causing the disease atherosclerosis.

Types of Lipids

Lipids may consist of fatty acids alone, or they may contain other molecules as well. For example, some lipids contain alcohol or phosphate groups. They include

1. triglycerides: the main form of stored energy in animals.
2. phospholipids: the major components of cell membranes.
3. steroids: serve as chemical messengers and have other roles.

Triglycerides

One type of lipid is called a **triglyceride**, an ester derived from glycerol combined with three fatty acid molecules.

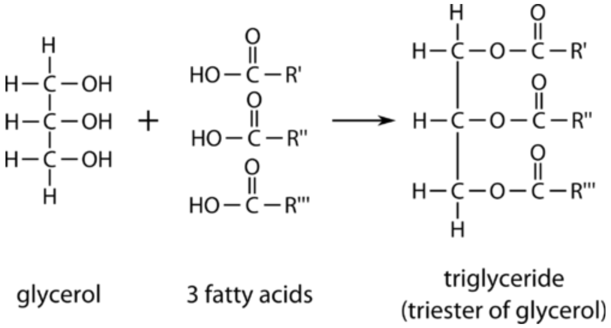


Figure 14.2.3: Triglyceride components.

Glycerol is a triol, an alcohol which contains three hydroxyl functional groups. A fatty acid is a long carbon chain, generally from 12 to 24 carbons in length, with an attached carboxyl group. Each of the three fatty acid molecules undergoes an esterification with one of the hydroxyl groups of the glycerol molecule. The result is a large triester molecule referred to as a triglyceride.

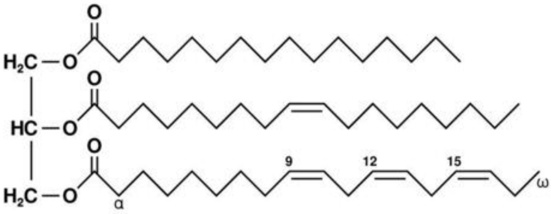


Figure 14.2.4: A triglyceride molecule can be formed from any combination of fatty acids.

Triglycerides function as a long-term storage form of energy in the human bods. Because of the long carbon chains, triglycerides are nearly nonpolar molecules and thus do not dissolve readily in polar solvents such as water. Instead, oils and fats are soluble in nonpolar organic solvents such as hexane and ethers.

Contributors and Attributions

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

14.3: Phospholipids in Cell Membranes

Learning Outcomes

- Describe the structure of a phospholipid.
- Identify the polar (hydrophilic) and nonpolar (hydrophobic) regions of a phospholipid.
- Explain how the phospholipid molecules form the bilayer of the cell membrane.

When you go to the dentist to get a tooth pulled, you really do not want to feel any pain. The dentist injects an anesthetic into your gum and it eventually becomes numb. One theory as to why anesthetics work deals with the movement of ions across the cell membrane. The anesthetic gets into the membrane structure and causes shifts in how ions move across the membrane. If ion movement is disrupted, nerve impulses will not be transmitted and you will not sense pain - at least not until the anesthetic wears off.

Phospholipids

A **phospholipid** is a lipid that contains a phosphate group and is a major component of cell membranes. A phospholipid consists of a hydrophilic (water-loving) head and hydrophobic (water-fearing) tail (see figure below). The phospholipid is essentially a triglyceride in which a fatty acid has been replaced by a phosphate group of some sort.

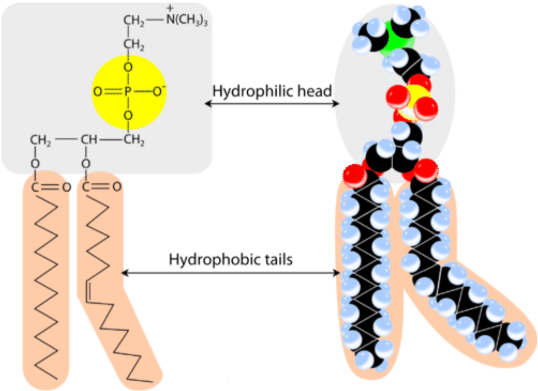


Figure 14.3.1: A phospholipid consists of a head and a tail. The "head" of the molecule contains the phosphate group and is hydrophilic, meaning that it will dissolve in water. The "tail" of the molecule is made up of two fatty acids, which are hydrophobic and do not dissolve in water.

Following the rule of "like dissolves like", the hydrophilic head of the phospholipid molecule dissolves readily in water. The long fatty acid chains of a phospholipid are nonpolar and thus avoid water because of their insolubility. In water, phospholipids spontaneously form a double layer called a lipid bilayer in which the hydrophobic tails of phospholipid molecules are sandwiched between two layers of hydrophilic heads (see figure below). In this way, only the heads of the molecules are exposed to the water, while the hydrophobic tails interact only with each other.

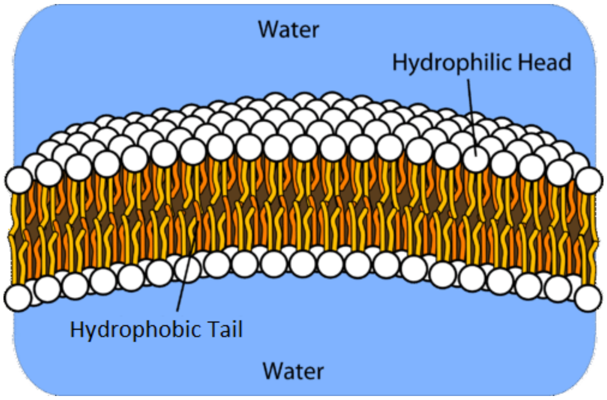


Figure 14.3.2: In a water solution, phospholipids form a bilayer where the hydrophobic tails point towards each other on the interior and only the hydrophilic heads are exposed to the water.

Phospholipid bilayers are critical components of cell membranes. The lipid bilayer acts as a barrier to the passage of molecules and ions into and out of the cell. However, an important function of the cell membrane is to allow selective passage of certain substances into and out of cells. This is accomplished by the embedding of various protein molecules in and through the lipid bilayer (see figure below). These proteins form channels through which certain specific ions and molecules are able to move. Many membrane proteins also contain attached carbohydrates on the outside of the lipid bilayer, allowing it to form hydrogen bonds with water.

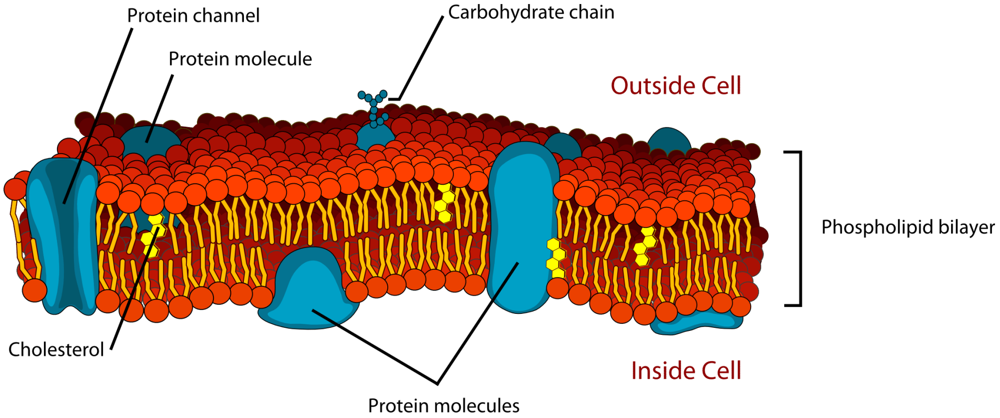


Figure 14.3.3: The phospholipid bilayer of a cell membrane contains embedded protein molecules which allow for selective passage of ions and molecules through the membrane.

Contributors and Attributions

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

14.4: Biological Molecules (Exercises)

These are homework exercises to accompany Chapter 14 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Answers are below the questions.

Questions

14.1: Enzymes

(click here for solutions)

Q14.1.1

What do enzymes do in the body?

Q14.1.2

Describe each of the following.

a. active site
b. substrate
c. allosteric site
d. inhibitor

Q14.1.3

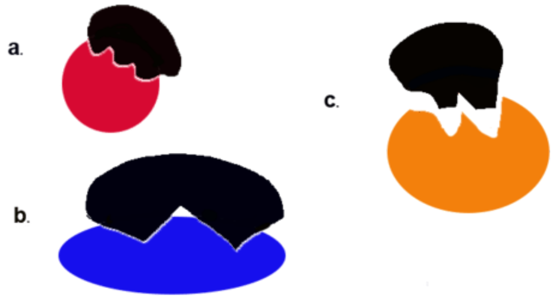
How are competitive and noncompetitive inhibitors similar and different from one another?

Q14.1.4

An inhibitor interacts with the enzyme at an allosteric site. Is it competitive or noncompetitive?

Q14.1.5

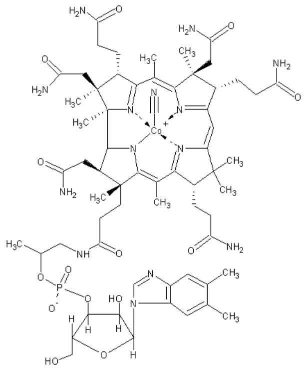
Describe each of the following enzyme-substrate pairs as using the lock-and-key model or the induced fit model.



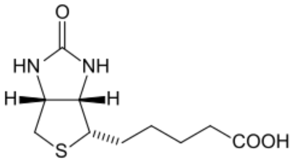
Q14.1.6

Each of the following behaves as a cofactor or a coenzyme. Identify each.

a. Zn^{2+}
b. Vitamin B12
c. biotin
d. Fe^{3+}



Vitamin B12



biotin

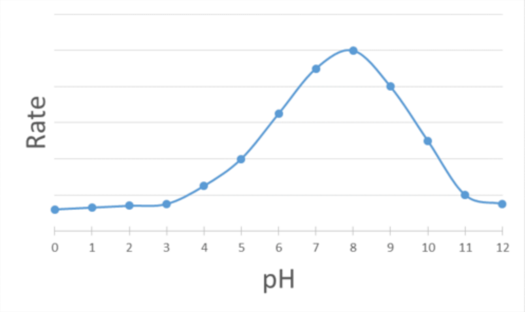
Q14.1.7

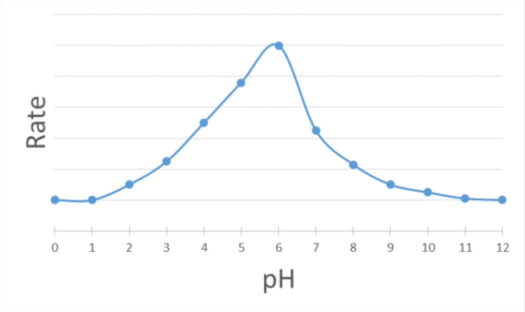
Identify each statement as true or false. Correct the false statement(s).

a. Enzyme activity increases with temperature.
b. Enzyme activity depends on pH.
c. Enzymes are consumed in a chemical reaction.
d. Increasing the concentration will increase the rate of a reaction until all of the enzyme is "occupied".
e. Enzymes can only function at their ideal temperature and pH.

Q14.1.8

The enzyme activity is graphed with respect to the pH of the mixture. Determine the pH at which each enzyme is most effective.

a. 

b. 

14.2: Lipids and Triglycerides

(click here for solutions)

Q14.2.1

What is the functional group in a fatty acid?

Q14.2.2

What is the difference between a fat and an oil?

Q14.2.3

Butter is a fat that is a solid at room temperature. What type of fatty acid does butter contain? How do you know?

Q14.2.4

Describe the difference between saturated and unsaturated fatty acids.

Q14.2.5

Explain why molecules of saturated and unsaturated fatty acids have different shapes.

Q14.2.6

Draw each structure.

a. Saturated fatty acid with 18 carbon atoms.
b. Monounsaturated fatty acid with 14 carbon atoms.
c. Polyunsaturated fatty acid with 14 carbon atoms.
d. Monounsaturated with 16 carbon atoms.
e. Polyunsaturated fatty acid with 18 carbon atoms and three double bonds.

Q14.2.7

Where does the body get essential fatty acids?

Q14.2.8

What molecules react to form a triglyceride?

Q14.2.9

Draw a triglyceride formed from three identical fatty acids.

Q14.2.10

Draw a triglyceride formed from three different fatty acids.

14.3: Phospholipids in Cell Membranes

(click here for solutions)

Q14.3.1

What is a phospholipid?

Q14.3.2

Which part of the phospholipid molecule is water-soluble?

Q14.3.3

Which part is not water-soluble?

Q14.3.4

What is the purpose of a semipermeable membrane like the cell membrane?

Answers

14.1: Enzymes

Q14.1.1

Enzymes act as catalysts for biological reactions.

Q14.1.2

Describe each of the following.

a. The active site is where the reaction occurs in the enzyme.
b. The substrate is the molecule that interacts with the enzyme.

c. An allosteric site is on the enzyme away from the active site. Inhibitors can interactive with the enzyme at the allosteric site.
d. An inhibitor is a molecule that interacts with the molecule to slow or stop a reaction.

Q14.1.3

Both competitive and noncompetitive inhibitors slow or stop a reaction. Competitive inhibitors bind with the active site and non-competitive inhibitors bind with an allosteric site.

Q14.1.4

noncompeititve

Q14.1.5

a. lock & key
b. lock & key
c. induced fit

Q14.1.6

a. cofactor
b. coenzyme
c. coenzyme
d. cofactor

Q14.1.7

Identify each statement as true or false. Correct the false statement(s).

a. Enzyme activity increases with temperature and then decreases with increasing temperature beyond a peak temperature.
b. True
c. Enzymes are not consumed in a chemical reaction.
d. True
e. Enzymes
f.
g.
~~can only~~ function most effectively at their ideal temperature and pH, but can function at a range of temperature and pH values.

Q14.1.8

a. pH = 8
b. pH = 6

14.2: Lipids and Triglycerides

Q14.2.1

carboxylic acid

Q14.2.2

Fats are solids at room temperatures while oils are liquids.

Q14.2.3

Saturated fatty acid.

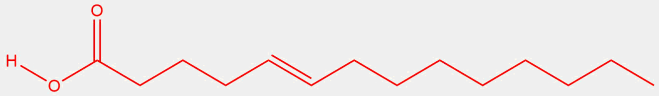
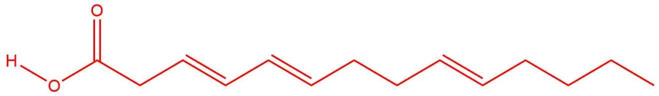
Q14.2.4

Saturated fatty acids contain only single bonds between carbon atoms while unsaturated fatty acids contain one or more double bonds.

Q14.2.5

Saturated fatty acids have a straight chain of carbon atoms while unsaturated fatty acids have a bend at every double bond.

Q14.2.6

a.	
b.	Answers will vary. The double bond may be located between any two carbon atoms. 
c.	Answers will vary. There should be multiple double bonds and can be located between any pairs of carbon atoms. 
d.	Answers will vary. The double bond may be located between any two carbon atoms. 
e.	Answers will vary. There should be three double bonds and can be located between any pairs of carbon atoms. 

Q14.2.7

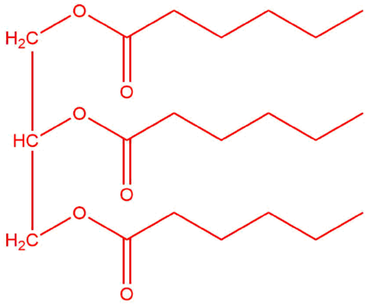
Essential fatty acids come from the food we eat.

Q14.2.8

Glycerol and three fatty acids from a triglyceride.

Q14.2.9

Answers will vary but each "tail" should have the same number of carbon atoms.



Q14.2.10

Answers will vary but each "tail" should have a different number of carbon atoms.

14.3: Phospholipids in Cell Membranes

Q14.3.1

A phospholipid is a triglyceride where one of the fatty acid tails is replaced by a phosphate group.

Q14.3.2

The phosphate end is water-soluble.

Q14.3.3

The hydrocarbon tails are not water-soluble.

Q14.3.4

To control the flow of substances in and out of the cell.

CHAPTER OVERVIEW

15: Metabolic Cycles

Metabolic pathways linked series of chemical reactions occurring within a cell. The reactants, products, and intermediates of an enzymatic reaction are known as metabolites, which are modified by a sequence of chemical reactions catalyzed by enzymes. In a metabolic pathway, the product of one enzyme acts as the substrate for the next. These enzymes often require dietary minerals, vitamins, and other cofactors to function. There are two types of metabolic pathways that are characterized by their ability to either synthesize molecules with the utilization of energy (anabolic pathway) or break down of complex molecules by releasing energy in the process (catabolic pathway). The two pathways complement each other in that the energy released from one is used up by the other.

15.1: Glycolysis

Learning Outcomes

- List the three stages of cellular respiration.
- Explain the purpose of glycolysis.
- Describe the use and formation of ATP during glycolysis.
- Name the products of glycolysis.
- Distinguish between aerobic and anaerobic.

How do you slice a molecule of glucose in half? With sharp knives? Not really. But you lyse it with enzymes during a process named glycolysis. Glucose is sliced right in half from a 6-carbon molecule to two 3-carbon molecules. This is the first step and an extremely important part of cellular respiration. It happens all the time, both with and without oxygen. And in the process, transfers some energy to ATP.

Glycolysis: A Universal and Ancient Pathway for Making ATP

When was the last time you enjoyed yogurt on your breakfast cereal, or had a tetanus shot? These experiences may appear unconnected, but both relate to bacteria which do not use oxygen to make ATP. In fact, tetanus bacteria cannot survive if oxygen is present. However, *Lactobacillus acidophilus* (bacteria which make yogurt) and *Clostridium tetani* (bacteria which cause tetanus or lockjaw) share with nearly all organisms the first stage of cellular respiration, **glycolysis** (see figure below). Because glycolysis is universal, whereas aerobic (oxygen-requiring) cellular respiration is not, most biologists consider it to be the most fundamental and primitive pathway for making ATP.

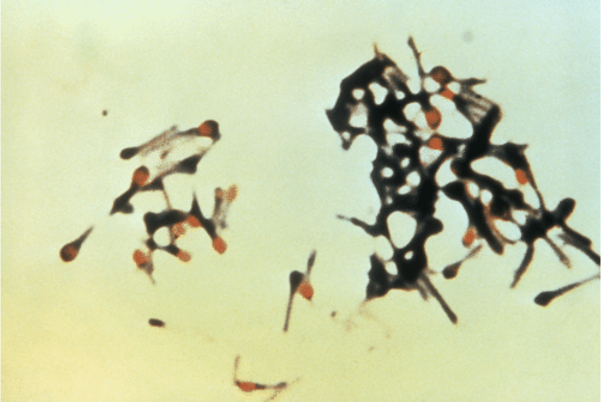


Figure 15.1.1: Clostridium tetani *bacteria are obligate anaerobes, which cannot grow in the presence of oxygen and use a variation of glycolysis to make ATP. Because they can grow in deep puncture wounds and secrete a toxin, which can cause muscle spasms, seizures, and death, most people receive tetanus vaccinations at least every ten years throughout life.*

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O \tag{15.1.1}$$

Like photosynthesis, the process represented by this equation is actually many small, individual chemical reactions. We will divide the reactions of cellular respiration into three stages: glycolysis, the **Krebs Cycle** (also known as the citric acid cycle), and the **electron transport chain** (see figure below). In this concept, Stage 1, glycolysis, the oldest and most widespread pathway for making ATP, is discussed. Before diving into the details, we must not that this first stage of cellular respiration is unique among the three stages: it does not require oxygen, and it does not take place in the mitochondrion. The chemical reactions of glycolysis occur without oxygen in the cytosol of the cell (see figure below).

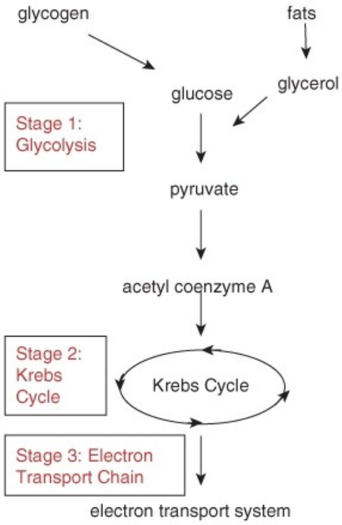


Figure 15.1.2: The many steps in the process of aerobic cellular respiration can be divided into three stages. The first stage, glycolysis, produces ATP without oxygen. Because this part of the cellular respiration pathway is universal, biologists consider it the oldest segment. Note that **glycogen** and fats can also enter the glycolysis pathway. The second stage is the Krebs Cycle, and the third stage is the electron transport chain. It is during the third stage that chemiosmosis produces numerous ATP molecules.

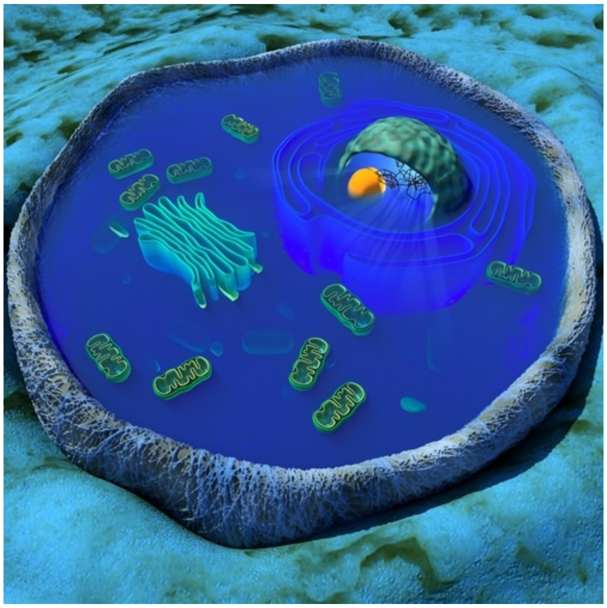


Figure 15.1.3: Glycolysis, unlike the latter two stages of cellular respiration, takes place without oxygen in the cytosol (blue) of the cell. For many organisms, aerobic respiration continues with the Krebs Cycle and the electron transport chain in the mitochondria (green). To enter the mitochondria, glucose must first be lysed into smaller molecules.

The name for Stage 1 clearly indicates what happens during that stage: *glyco-* refers to glucose, and *-lysis* means "splitting". In glycolysis, within the cytosol of the cell, a minimum of eight different enzymes break apart glucose into two 3-carbon molecules. The energy released in breaking those bonds is transferred to carrier molecules, ATP and NADH. **NADH** temporarily holds small amounts of energy which can be used to later build ATP. The 3-carbon product of glycolysis is **pyruvate**, or pyruvic acid (see figure below). (The sole difference between them is actually a sole hydrogen atom. Pyruvic acid: $\mathrm{CH_3COCOOH}$, pyruvate: $\mathrm{CH_3COCOO^-}$). Overall, glycolysis can be represented as:

$$\mathrm{C_6H_{12}O_6 + 2NAD^+ + 2P_i + 2ADP \rightarrow 2\ pyruvate\ + 2NADH + 2ATP} \tag{15.1.2}$$

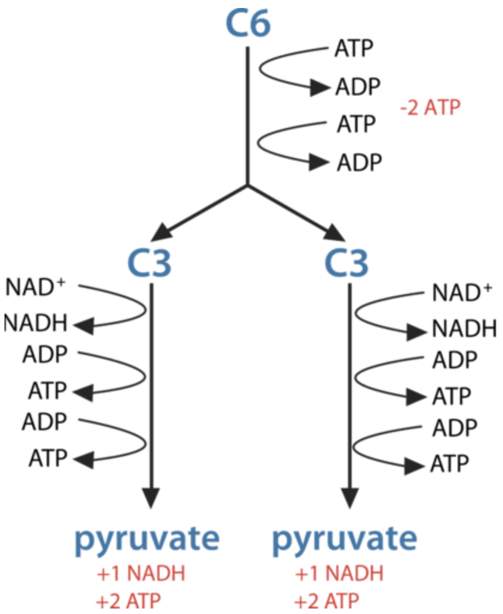


Figure 15.1.4: In glycolysis, glucose (C6) is split into two 3-carbon (C3) pyruvate molecules. This releases energy, which is transferred to ATP. How many ATP molecules are made during this stage of cellular respiration?

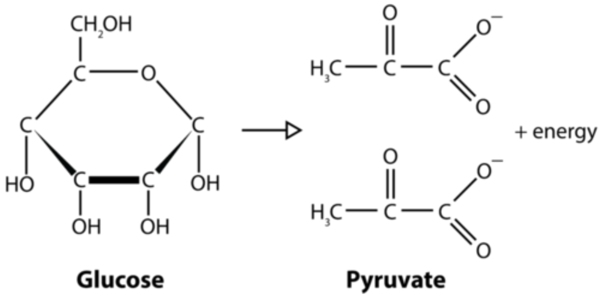


Figure 15.1.5: Glycolysis breaks the 6-carbon molecule glucose into two 3-carbon pyruvate molecules, releasing some of the chemical energy which had been stored in glucose.

However, even this equation is deceiving. Just the splitting of glucose requires many steps, each transferring or capturing small amounts of energy. Individual steps appear in the figure below. Studying the pathway in detail reveals that cells must "spend" or "invest" two ATP in order to begin the process of breaking glucose apart. Note that the phosphates produced by breaking apart ATP join with glucose, making it unstable and more likely to break apart. Later steps harness the energy released when glucose splits, and use it to build "hot hydrogens" (NAD^+ is reduced to NADH) and ATP (ADP + $P_i \rightarrow$ ATP). If you count the ATP produced, you will find a net yield of two ATP per glucose (4 produced $-$ 2 spent). Remember to double the second set of reactions to account for the two 3-carbon molecules which follow that pathway! The "hot hydrogens" can power other metabolic pathways, or in many organisms, provide energy for further ATP synthesis.

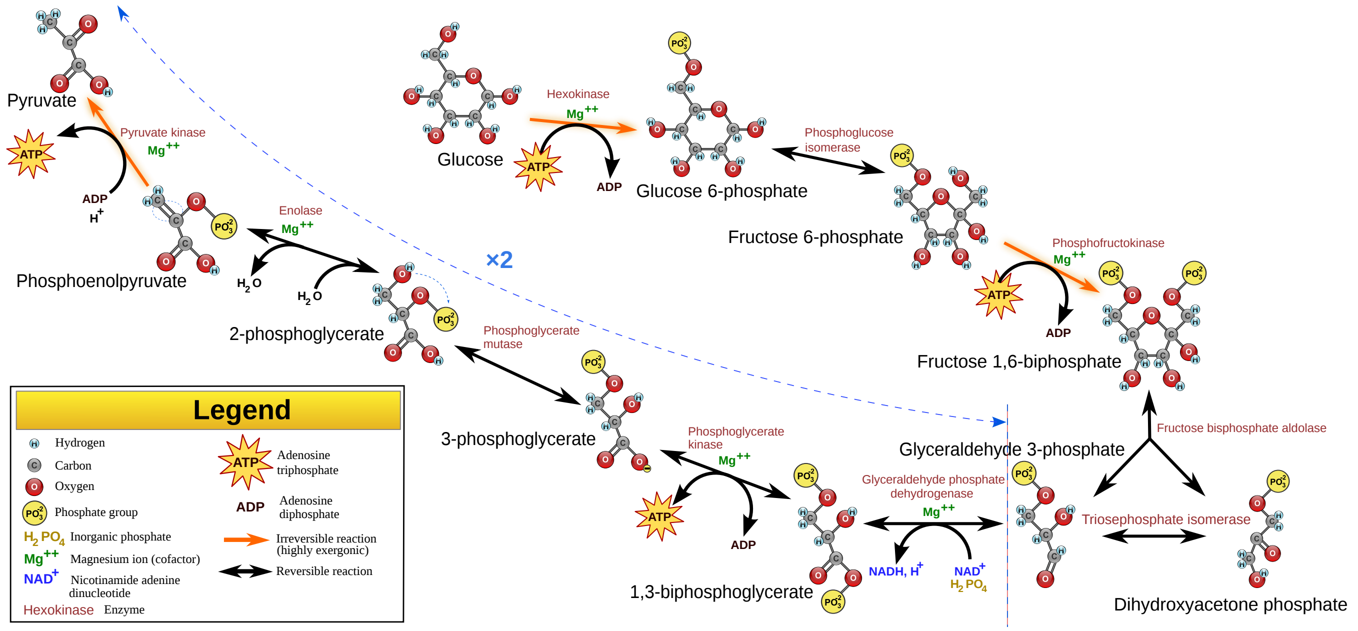


Figure 15.1.6: *Glycolysis pathway. from Wikipedia (CCBY-SA 3.0; YassineMrabet).*

To summarize: In the cytosol of the cell, glycolysis transfers some of the chemical energy stored in one molecule of glucose to two molecules of ATP and two NADH. This makes (some of) the energy in glucose, a universal fuel molecule for cells, available to use in cellular work - moving organelles, transporting molecules across membranes, or building large organic molecules.

Although glycolysis is universal, pathways leading away from glycolysis vary among species depending on the availability of oxygen. If oxygen is unavailable, pyruvate may be converted to lactic acid or ethanol and carbon dioxide in order to regenerate NAD^+, called anaerobic respiration. **Anaerobic respiration** is also called **fermentation**, which will be discussed in another concept.

If oxygen is present, pyruvate enters the mitochondria for further breakdown, releasing far more energy and producing many additional molecules of ATP in the latter two stages of **aerobic respiration** - the Krebs Cycle and electron transport chain.

Supplemental Resources

- Glycolysis: http://johnkyrk.com/glycolysis.html

Contributors and Attributions

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

15.2: The Citric Acid Cycle

Learning Outcomes

- Describe the citric acid cycle (Krebs Cycle).
- Name the products of the citric acid cycle.
- Identify the energy carrier molecules produced in the citric acid cycle.
- Describe what happens to pyruvate before it enters the citric acid cycle.

Aerobic Respiration

Enticing clues - volcanic gases, vast iron ore sediments, and bubbles of ancient air trapped in amber - suggest dramatic changes during the history of earth's atmosphere. Correlating these clues with the fossil record leads to two major conclusions: that early life evolved in the absence of oxygen, and that oxygen first appeared between 2 and 3 billion years ago (see figure below) because of photosynthesis by the blue green bacteria, cyanobacteria. The chemistry of cellular respiration reflects this history. Its first stage, glycolysis, is universal and does not use oxygen.

Absolutely dependent on oxygen gas, we find it difficult to imagine that its appearance must have been disastrous for the anaerobic organisms that evolved in its absence. But oxygen is highly reactive, and at first, its effect on evolution was so negative that some have named this period the "oxygen catastrophe". However, as oxygen gradually formed a protective **ozone layer**, life rebounded. After the first organisms evolved to use oxygen to their advantage, the diversity of aerobic organisms exploded. According to the Theory of Endosymbiosis, engulfing of some of these aerobic bacteria led to eukaryotic cells with mitochondria, and **multicellularity**, the evolution of multicellular eukaryotic organisms, followed. Today, we live in an atmosphere which is 21% oxygen, and most of life follows glycolysis with the last two, aerobic stages of cellular respiration.

Recall the purpose of cellular respiration: to release energy from glucose to make ATP, the universal molecule of energy for cellular work. The following equation describes the overall process, although it summarizes many individual chemical reactions.

$$6\,\mathrm{O_2} + \underbrace{\mathrm{C_6H_{12}O_6}}_{\substack{\text{stored chemical}\\ \text{energy, deliverable}}} + 398\,\mathrm{Pi} \xrightarrow{\text{mitochondria}} 38\,\mathrm{ATP} + 6\,\mathrm{O_2} + 6\,\mathrm{H_2O} \tag{15.2.1}$$

Once again, the first stage of this process, glycolysis, is ancient, universal, and anaerobic. In the cytoplasm of most cells, glycolysis breaks each 6-carbon molecule of glucose into two 3-carbon molecules of **pyruvate**. Chemical energy, which had been stored in the now broken bonds, is transferred to 2 ATP and 2 NADH molecules.

The fate of pyruvate depends on the species and the presence or absence of oxygen. If oxygen is present to drive subsequent reaction, pyruvate enters the mitochondria, where the **citric acid cycle** (also known as the **Krebs Cycle**) (Stage 2) and **electron transport chain** (Stage 3) break it down and oxidize it completely to $\mathrm{CO_2}$ and $\mathrm{H_2O}$. The energy released builds many more ATP molecules, though of course some is lost as heat. Let's explore the details of how mitochondria use oxygen to make more ATP from glucose by aerobic respiration.

The Citric Acid Cycle: Capturing Energy from Pyruvate

Aerobic respiration begins with the entry of the product of glycolysis, pyruvate, into the mitochondria. For each initial glucose molecule, two pyruvate molecules will enter the mitochondria. Pyruvate, however, is not the molecule that enters the citric acid cycle. Prior to entry into this cycle, pyruvate must be converted into a 2-carbon acetyl-CoenzymeA (acetyl-CoA) unit.

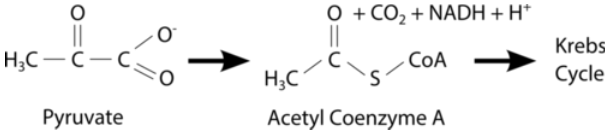


Figure 15.2.1: After glycolysis, two 3-carbon pyruvates enter the mitochondria, where they are converted to two 2-carbon acetyl-CoenzymeA (CoA) molecules. Acetyl-CoA then enters the Krebs Cycle. Note that the carbons removed become carbon dioxide, accounting for two of the six such end products of glucose oxidation. The energy released by this breakdown is carried by NADH.

The conversion of pyruvate into acetyl-CoA is referred to as the pyruvate dehydrogenase reaction. It is catalyzed by the pyruvate dehydrogenase complex. This process produces one NADH electron carrier while releasing a $\mathrm{CO_2}$ molecule. This step is also known as the link reaction or transition step, as it links glycolysis and the citric acid cycle. Of course, as two pyruvates result from glycolysis, two acetyl-CoAs are produced as are 2 NADH molecules.

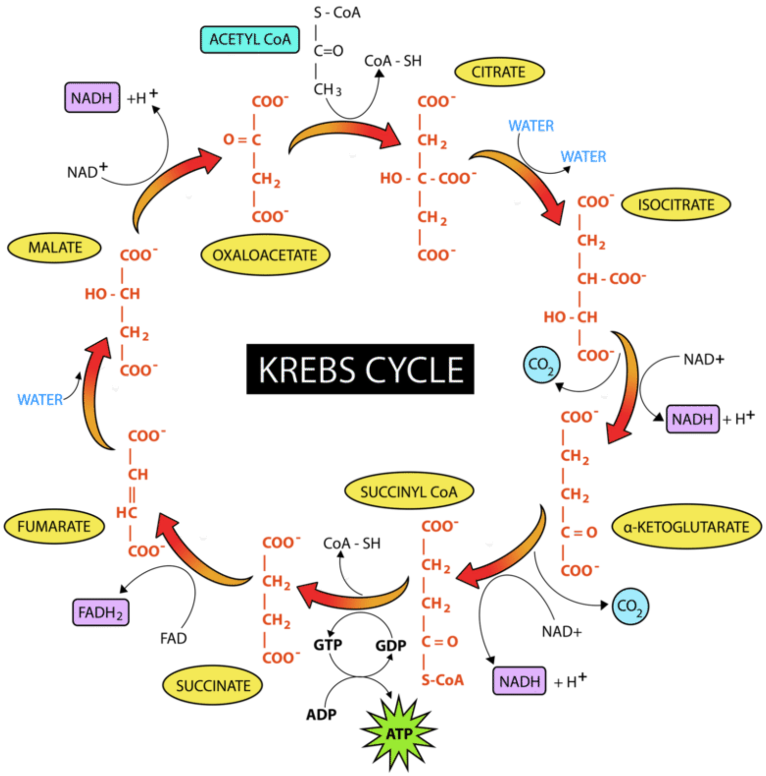


Figure 15.2.2: The Krebs Cycle completes the breakdown of glucose begun in glycolysis. If oxygen is present, pyruvate enters the mitochondria and is converted to acetyl-CoA. Acetyl-CoA enters the cycle by combining with 4-carbon oxaloacetate. Study the diagram to confirm that each turn of the cycle (two for each glucose) stores energy in 3 NADH $+H^+$, one $FADH_2$, and one ATP (from GTP), and releases 2 CO_2. The Krebs Cycle is also known as the citric acid cycle or the tricarboxylic acid cycle (TCA cycle).

1. Within the mitochondria, each pyruvate is broken apart and combined with a coenzyme known as CoA to form a 2-carbon molecule, acetyl-CoA, which can enter the Krebs Cycle. A single atom of carbon (per pyruvate) is "lost" as carbon dioxide. The energy released in this breakdown is captured in two NADH molecules. See the figure above. Fatty acids can also break down in to acetyl-CoA. By this means, lipids, like fats, can be "burned" to make ATP using the citric acid cycle.
2. The Krebs Cycle (see figure above) begins by combining each acetyl-CoA with a four-carbon carrier molecule to make a 6-carbon molecule of citric acid (or citrate, its ionized form).
3. The cycle carries citric acid through a series of chemical reactions which gradually release energy and capture it in several carrier molecules. For each acetyl-CoA which enters the cycle, 3 NAD^+ are reduced to NADH, one molecule of FAD (another temporary energy carrier) is reduced to \(**FADH_2\)**, and one molecule of ATP (actually a precursor, GTP, guanine triphosphate) is produced. Study the figure above to locate each of these energy-capturing events.
4. Note what happens to carbon atoms (black dots in the figure above). For each 2-carbon acetyl-CoA which enters the cycle, two molecules of carbon dioxide are released, completing the breakdown of the original 6-carbon glucose molecule. The final step regenerates the original 4-carbon molecule which began the cycle, so that another acetyl-CoA can enter the cycle.

In summary, the citric acid cycle completes the breakdown of glucose which began with glycolysis. Its chemical reactions oxidize all six of the original carbon atoms to CO_2, and capture the energy released in 2 ATP, 6 NADH, and 2 $FADH_2$. These energy carriers join the 2 ATP and 2 NADH produced in glycolysis and the 2 NADH produced in the conversion of 2 pyruvates to 2 acetyl-CoA molecules.

At the conclusion of the citric acid cycle, glucose is completely broken down, yet only four ATP have been produced. Moreover, although oxygen is required to drive the citric acid cycle, the cycle's chemical reactions do not themselves consume O_2. The conclusion of cellular respiration, stage 3, produces the majority of the ATP.

Supplemental Resources

- Citric Acid Cycle (aka Krebs Cycle): virtuallabs.stanford.edu/other/biochem/TCA.swf
- Krebs Cycle (aka Citric Acid Cycle): http://johnkyrk.com/krebs.html

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15.3: Lactic Acid Fermentation

Learning Outcomes

- Describe lactic acid fermentation.
- Describe how bacteria, including those we employ to make yogurt, make ATP in the absence of oxygen.
- Discuss how your muscles continue to work for you even when your respiratory and cardiovascular system can no longer keep up a continuous supply of oxygen.

Short spurts of sprinting are sustained by fermentation in muscle cells. This produces just enough ATP to allow these short bursts of increased activity.

Lactic Acid Fermentation: Muscle Cells and Yogurt

For chicken or turkey dinners, do you prefer light meat or dark? Do you consider yourself a sprinter or a long distance runner? What is the biological difference between light meat and dark meat? Or between the two types of runners? Would you believe it has something to do with muscle color?

Figure 15.3.1**:** *Light meat or dark? Sprinting or endurance? Muscle cells know two ways of making ATP - aerobic and anaerobic respiration.*

Are Drumsticks and Athletic Prowess Related?

Muscle color reflects its specialization for aerobic or anaerobic metabolism. Although humans are **obligate aerobes** (an organism which requires oxygen for cellular respiration), our muscle cells have not given up on ancient pathways which allow them to keep producing ATP quickly when oxygen runs low. The difference is more pronounced in chickens and grouse (see figure below), which stand around all day on their legs. For long periods of time, they carry out aerobic respiration in their "specialized-for-endurance" red muscles. If you are familiar with grouse, you know that these birds "flush" with great speed over short distances. Such "sprinting" flight depends on anaerobic respiration in the white cells of breast and wing muscle, allowing rapid production of ATP in low oxygen situations.

Figure 15.3.2**:** *Ruffed grouse use anaerobic respiration (lactic acid fermentation) in wing and breast muscles for quick bursts of speed to escape from predators*.

No human muscle is all red or all white, but chances are, if you excel at sprinting short distances or at a sport such as weight lifting, you have more white glycolytic fibers in your leg muscles, allowing anaerobic respiration. If you run marathons, you probably have more red oxidative fibers, performing aerobic respiration.

Lactic Acid Fermentation

You may have not been aware that your muscle cells can ferment. **Fermentation** is the process of producing ATP in the absence of oxygen, through glycolysis alone. Recall that glycolysis breaks a glucose molecule into two pyruvate molecules, producing a net gain of two ATP and two NADH molecules. **Lactic acid fermentation** is the type of anaerobic respiration carried out by yogurt bacteria (*Lactobacillus* and others) and by your own muscle cells when you work them hard and fast.

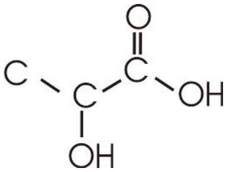

Figure 15.3.3*: Lactic acid,* $C_3H_6O_3$.

Lactic acid fermentation converts the 3-carbon pyruvate to the 3-carbon lactic acid $\left(C_3H_6O_3\right)$ (see figure below) and regenerates NAD$^+$ in the process, allowing glycolysis to continue to make ATP in low-oxygen conditions. Since there is a limited supply of NAD$^+$ available in any given cell, this electron acceptor must be regenerated to allow ATP production to continue. To achieve this, NADH donates its extra electrons to the pyruvate molecules, regenerating NAD$^+$. Lactic acid is formed by the reduction of pyruvate.

$$\mathrm{C_3H_3O_3\ (pyruvate)\ + NADH \rightarrow C_3H_6O_3\ (lactic\ acid)\ + NAD^+} \tag{15.3.1}$$

actic acid fermentation converts pyruvate to lactic acid, and regenerates NAD$^+$ from $NADH$.

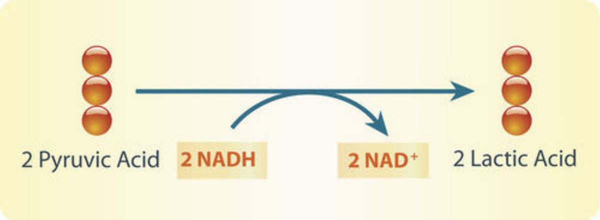

Figure 15.3.4*: Lactic acid fermentation makes ATP in the absence of oxygen by converting glucose to lactic acid (through a pyruvate intermediate). Making lactic acid from pyruvate oxidizes NADH, regenerating NAD$^+$ so that glycolysis can continue to make more ATP rapidly. Each circle represents a carbon atom.*

For *Lactobacillus* bacteria, the acid resulting from fermentation kills bacterial competitors in buttermilk, yogurt, and some cottage cheese. The benefits extend to humans who enjoy these foods, as well (Figure 15.3.5).

Figure 15.3.5*:* Lactobacillus *bacteria use the same type of anaerobic respiration as our muscle cells. Lactic acid reduces competition from other bacteria and flavors yogurt.*

You may have notice this type of fermentation in your own muscles, because muscle fatigue and pain are associated with lactic acid. Lactic acid accumulates in your muscle cells as fermentation proceeds during times of strenuous exercise. During these times, your respiratory and cardiovascular systems cannot transport oxygen to your muscle cells, especially those in your legs, fast enough to maintain aerobic respiration. To allow the continuous production of some ATP, your muscle cells use lactic acid fermentation.

Supplemental Resources

- Fermentation: http://www.tempeh.info/fermentation/fermentation.php
- Lactic Acid Fermentation: http://www.tempeh.info/fermentation/...rmentation.php

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15.4: The Electron Transport Chain

Learning Outcomes

- Summarize the electron transport chain.
- Recognize that electron transport chain is the third and final stage of aerobic cellular respiration.
- Identify the products of the citric acid cycle.

What do trains, trucks, boats, and planes all have in common? They are ways to transport. And they all use a lot of energy. To make ATP, energy must be "transported" - first from glucose to NADH, and then somehow passed to ATP. How is this done? With an electron transport chain, the third stage of aerobic respiration. This third stage uses energy to make energy.

The Electron Transport Chain: ATP for Life in the Fast Lane

At the end of the Krebs Cycle, energy from the chemical bonds of glucose is stored in diverse energy carrier molecules: four ATPs, but also two $FADH_2$ and ten NADH molecules. The primary task of the last stage of cellular respiration, the **electron transport chain**, is to transfer energy from the electron carriers to even more ATP molecules, the "batteries" which power work within the cell.

Pathways for making ATP in stage 3 of aerobic respiration closely resemble the electron transport chains used in photosynthesis. In both electron transport chains, energy carrier molecules are arranged in sequence within a membrane so that energy-carrying electrons cascade from one to another, losing a little energy in each step. In both photosynthesis and aerobic respiration, the energy lost is harnessed to pump hydrogen ions into a compartment, creating an **electrochemical gradient** or **chemiosmotic gradient** across the enclosing membrane. And in both processes, the energy stored in the chemiosmotic gradient is used with **ATP synthase** to build ATP.

For aerobic respiration, the electron transport chain or "respiratory chain" is embedded in the inner membrane of the mitochondria (see figure below). The $FADH_2$ and NADH molecules produced in glycolysis and the Krebs Cycle, donate high-energy electrons to energy carrier molecules within the membrane. As they pass from one carrier to another, the energy they lose is used to pump hydrogen ions into the mitochondrial intermembrane space, creating an electrochemical gradient. Hydrogen ions flow "down" the gradient - from outer to inner compartment - through the ion channel/enzyme ATP synthase, which transfers their energy to ATP. Note the paradox that it requires energy to create and maintain a concentration gradient of hydrogen ions that are then used by ATP synthase to create stored energy (ATP). In broad terms, it takes energy to make energy. Coupling the electron transport chain to ATP synthesis with a hydrogen ion gradient is **chemiosmosis**, first described by Nobel laureate Peter D. Mitchell. This process, the use of energy to phosphorylate ADP and produce ATP is also known as **oxidative phosphorylation**.

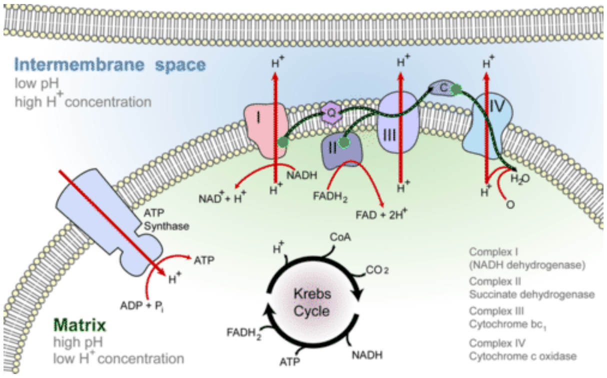

Figure 15.4.1: The third stage of cellular respiration uses the energy stored during the earlier stages in NADH and $FADH_2$ to make ATP. Electron transport chains embedded in the mitochondrial inner membrane capture high-energy electrons from the carrier molecules and use them to concentrate hydrogen ions in the intermembrane space. Hydrogen ions flow down their electrochemical gradient back into the matrix through ATP synthase channels which capture their energy to convert ADP to ATP. Notice that the process regenerates NAD^+, supplying the electron acceptor molecule needed in glycolysis. (CC BY-NC 3.0; Mariana Ruiz Villarreal (LadyofHats) for the CK-12 Foundation).

After passing through the electron transport chain, low-energy electrons and low-energy hydrogen ions combine with oxygen to form water. Thus, oxygen's role is to drive the entire set of ATP-producing reactions within the mitochondrion by accepting "spent" hydrogens. Oxygen is the final electron acceptor, no part of the process - from the Krebs Cycle through the electron transport chain- can happen without oxygen.

The electron transport chain can convert the energy from one glucose molecule's worth of $FADH_2$ and $NADH$ + H^+ into as many as 34 ATP. When the four ATP produced in glycolysis and the Krebs Cycle are added, the total of 38 ATP fits the overall equation for aerobic cellular respiration:

$$6\,O_2 + \underbrace{C_6H_{12}O_6}_{\text{stored chemical energy}} + 38\,ADP + 39\,P_i \rightarrow \underbrace{38\,ATP}_{\text{stored chemical energy}} + 6\,CO_2 + 6\,H_2O \tag{15.4.1}$$

Aerobic respiration is complete. If oxygen is available, cellular respiration transfers the energy from one molecule of glucose to 38 molecules of ATP, releasing carbon dioxide and water as waste. "Deliverable" food energy has become energy which can be used for work within the cell - transport within the cell, pumping ions and molecules across membranes, and building large organic molecules. Can you see how this could lead to "life in the fast lane" compared to anaerobic respiration (glycolysis alone)?

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15.5: Metabolic Cycles (Exercises)

These are homework exercises to accompany Chapter 15 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health.

15.1: Glycolysis

Q15.1.1

What are the three stages of cellular respiration?

Q15.1.2

What is the purpose of glycolysis?

Q15.1.3

What is the output of glycolysis from a single glucose molecule?

Q15.1.4

How many molecules of ATP are "invested" in glycolysis? How many are produced?

Q15.1.5

Define aerobic and anaerobic.

15.2: The Citric Acid Cycle

Q15.2.1

Where does the Krebs cycle occur in the cell?

Q15.2.2

What happens to the pyruvate produced during glycolysis?

Q15.2.3

How many reactions does it take to complete the cycle?

Q15.2.4

How many "turns" of the citric acid cycle must occur for each molecule of glucose entering glycolysis?

Q15.2.5

What is he output of the citric acid cycle?

Q15.2.6

Trace the six carbon atoms originally from acetyl-CoA through the Krebs Cycle. Trace the flow of energy from the pyruvates produced in glycolysis through the Krebs Cycle.

Q15.2.7

How many energy carriers are produced during the Krebs cycle per acetyl-CoA?

15.3: Lactic Acid Fermentation

Q15.3.1

What is fermentation?

Q15.3.2

Define lactic acid fermentation.

Q15.3.3

Identify yourself as a “sprinter” or an “endurance runner” and predict the type of muscle fiber (red or white) which predominates in your body. Explain your reasoning.

Q15.3.4

What is the chemical equation of lactic acid fermentation?

15.4: The Electron Transport Chain

Q15.4.1

What molecules "feed" the electron transport chain?

Q15.4.2

What is the primary product of the electron transport chain?

Q15.4.3

Where do the reactions of the electron transport chain occur?

15.6: Homeostasis

Learning Outcomes

- Describe how homeostasis and equilibrium are different.

Remove one stone and the whole arch collapses. The same is true for the human body. All the systems work together to maintain stability or homeostasis. Disrupt one system, and the whole body may be affected.

Homeostasis

All of the organs and organ systems of the human body work together like a well-oiled machine. This is because they are closely regulated by the nervous and endocrine systems. The **nervous system** controls virtually all body activities, and the **endocrine system** secretes **hormones** that regulate these activities. Functioning together, the organ systems supply body cells with all the substances they need and eliminate their wastes. They also keep temperature, pH, and other conditions at just the right levels to support life processes.

Maintaining Homeostasis

The process in which organ systems work to maintain a stable internal environment is called **homeostasis**. Keeping a stable internal environment requires constant adjustments. Here are just three of the many ways that human organ systems help the body maintain homeostasis:

- Respiratory system: A high concentration of carbon dioxide in the blood triggers faster breathing. The lungs exhale more frequently, which removes carbon dioxide from the body more quickly.
- Excretory system: A low level of water in the blood triggers retention of water by the kidneys. The kidneys produce more concentration urine, so less water is lost from the body.
- Endocrine system: A high concentration of sugar in the blood triggers secretion of insulin by an endocrine gland called the pancreas. Insulin is a hormone that helps cells absorb sugar from the blood.

So how does your body maintain homeostasis? The regulation of your internal environment is done primarily through negative feedback. **Negative feedback** is a response to a stimulus that keeps a variable close to a set value (see figure below). Essentially, it "shuts off" or "turns off" a system when it varies from a set value.

For example, your body has an internal thermostat. During a winter day, in your house a thermostat senses the temperature in a room and responds by turning on or off the heater. Your body acts in much the same way. When body temperature rises, receptors in the skin and the brain sense the temperature change. The temperature change triggers a command from the brain. This command can cause several responses. If you are too hot, the skin makes sweat and blood vessels near the skin surface dilate. This response helps decrease body temperature.

Another example of negative feedback has to do with blood glucose levels. When glucose (sugar) levels in the blood are too high, the pancreas secretes insulin to stimulate the absorption of glucose and the conversion of glucose into glycogen, which is stored in the liver. As blood glucose levels decrease, less insulin is produced. When glucose levels are too low, another **hormone** called glucagon is produced, which causes the liver to convert glycogen back to glucose.

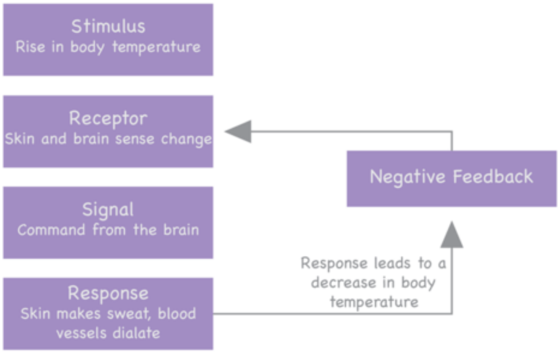

Figure 15.6.1: Feedback Regulation. If a raise in body temperature (stimulus) is detected (receptor), a signal will cause the brain to maintain homeostasis (response). Once the body temperature returns to normal, negative feedback will cause the response to end. This sequence of stimulus-receptor-signal-response is used throughout the body to maintain homeostasis.

Positive Feedback

Some processes in the body are regulated by positive feedback. **Positive feedback** is when a response to an event increases the likelihood of the event to continue. An example of positive feedback is milk production in nursing mothers. As the baby drinks her mother's milk, the hormone prolactin, a chemical signal, is released. The more the baby suckles, the more prolactin is released, which causes more milk to be produced. Other examples of positive feedback include contractions during childbirth. When constrictions in the uterus push a baby into the birth canal, additional contractions occur.

Failure of Homeostasis

Many homeostatic mechanisms such as these work continuously to maintain stable conditions in the human body. Sometimes, however, the mechanisms fail. When they do, cells may not get everything they need, or toxic wastes may accumulate in the body. If homeostasis is not restored, the imbalance may lead to disease or even death.

Homeostasis vs. Equilibrium

Homeostasis requires an input of energy to maintain a specific condition necessary for life. Disturbances to homeostasis must be responded to in order to avoid death or disease. For example, a body needs to maintain a certain internal temperature. Go outside in cold weather - body shivers to maintain its body temperature.

Dynamic equilibrium is maintaining a specific condition that minimizes the system's energy, depending on the circumstances. A disturbance to an equilibrium is responded to in order to shift the process to reestablish an equilibrium. For example, if a warm object (say a metal bowl) is placed outside in cold weather - the transfer of heat occurs and the temperature of the bowl equilibrates to the outside temperature. If this happened to a person, it would not be good.

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Index

Index

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    - 9: Equilibrium Applications - *CK-12 License*
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        - 9.2: Buffers - *CK-12 License*
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  - 15.2: The Citric Acid Cycle - *CK-12 License*
  - 15.3: Lactic Acid Fermentation - *CK-12 License*
  - 15.4: The Electron Transport Chain - *CK-12 License*
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