CHEM 309: APPLIED CHEMISTRY FOR THE HEALTH SCIENCES



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CHE 309: Applied Chemistry for the Health Sciences

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Detailed Licensing

Detailed Licensing

Licensing

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CHAPTER OVERVIEW

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1.1: Expressing Numbers - Significant Figures

Learning Objectives

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

Scientists have established certain conventions for communicating the degree of **precision** of a measurement, which is dependent on the measuring device used. Imagine, for example, that you are using a meterstick to measure the width of a table. The centimeters (cm) marked on the meterstick, tell you how many *centimeters* wide the table is. Many metersticks also have markings for millimeters (mm), so we can measure the table to the nearest *millimeter*. Most metersticks do not have any smaller (or more precise) markings indicated, so you cannot report the measured width of the table any more precise than to the nearest millimeter. However, you can *estimate* one past the smallest marking, in this case the millimeter, to the next decimal place in the measurement (Figure **1.1.1**).

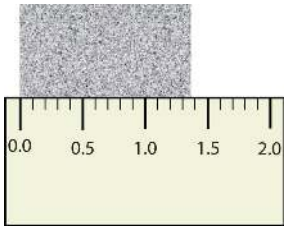


Figure **1.1.1**: Measuring the length of an Object to the correct number of digits as 1.35 cm. Three digits should be reported for the length of this object?

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

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

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

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

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

1. All nonzero digits are significant. In 1,357 mm, all the digits are significant.
2. *Sandwiched (or embedded) zeros*, those between significant digits, are significant. Thus, 405 g has *three* significant figures.
3. *Leading zeros*, which are zeros at the beginning of a decimal number less than 1, are not significant. In 0.000458 mL, the first four digits are leading zeros and are not significant. The zeros serve only to put the digits 4, 5, and 8 in the correct decimal positions. This number has three significant figures.
4. *Trailing zeros*, which are zeros at the end of a number, are significant only if the number has a decimal point. In 1,500 m, the two trailing zeros are not significant because the number is written without a decimal point, therefore the number has two

significant figures. However, in 1,500.00 m, all six digits are significant because the number has a decimal point.

5. When determining significant figures for numbers in scientific notation, the power of 10 is not included. So 7.0 X 10^{-5} g is a measurement of mass in 2 significant figures,

Example 1.1.1

How many significant figures does each number have?

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

Answer a

four (by rules 1 and 4)

Answer b

seven (by rules 1 and 2)

Answer c

nine (by rules 1, 2, and 4)

Answer d

four (by rules 1 and 3)

Exercise 1.1.1

How many significant figures does each number have?

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

Answer a

five

Answer b

four

Answer c

four

Answer d

one

Answer e

one

Rounding Off Numbers

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know that, round to that many digits, starting from the left. **If the number immediately to the right of the last significant digit is less**

than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

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

Table 1.5.1: Rounding examples

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

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

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

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

Combining Numbers

For **addition or subtraction**, the rule is to stack all the numbers with their decimal points aligned. The answer should have the same number of decimal places as the quantity with the fewest decimal places. Consider the following:

```
    56.789
+   102.2
+ 1,300.099
= 1,459.088
      ↑ Limit to this column
```

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

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

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

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

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

✔ Example 1.1.2

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

a. 23.096×90.300
b. 125×9.000
c. $1{,}027 + 610.0 + 363.06$

Answer a

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

Answer b

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

Answer c

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

? Exercise 1.1.2

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

a. $217 \div 903$
b. $13.77 + 908.226 + 515$
c. $255.0 - 99$
d. 0.00666×321

Answer a

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

Answer b

$1437 = 1.437 \times 10^3$

Answer c

$156 = 1.56 \times 10^2$

Answer d

$2.14 = 2.14 \times 10^0$

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

Concept Review Exercises

1. Explain why the concept of significant figures is important in scientific measurements.
2. State the rules for determining the significant figures in a measurement.
3. When do you round a number up, and when do you not round a number up?

Answers

1. Significant figures represent all the known digits of a measurement plus the first estimated one. It gives information about how precise the measuring device and measurement is.
2. All nonzero digits are significant; zeros between nonzero digits are significant; zeros at the end of a nondecimal number or the beginning of a decimal number are not significant; zeros at the end of a decimal number are significant.
3. Round up only if the first digit dropped is 5 or higher.

Key Takeaways

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

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1.2: Expressing Numbers - Scientific Notation

Learning Objectives

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

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

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

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

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

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

Recall that:

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

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

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

Example 1.2.1: Scientific Notation

Express each number in scientific notation.

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

Answer a

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

Answer b

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

Answer c

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

? Exercise 1.2.1

Express each number in scientific notation.

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

Answer a

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

Answer b

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

Answer c

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

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

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

✓ Example 1.2.2

Express each number in standard, or conventional notation.

a. 5.27×10^4
b. 1.0008×10^6

Answer a

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

Answer b

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

✓ Exercise 1.2.2

Express each number in standard, or conventional notation.

a. 6.98×10^8
b. 1.005×10^2

Answer a

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

Answer b

Moving the decimal point two places to the right gives 100.5

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

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

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

Recall that:

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

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

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

Note:

When writing numbers in scientific notation, the convention is to have only one digit before the decimal point.

- Numbers that are greater than one have a positive power in scientific notation. If the decimal point is moved to the left 6 places, the power (6) of 10 is positive. $1{,}500{,}000 = 1.5 \times 10^{6}$
- Numbers that are less than one have a negative power in scientific notation. If the decimal point is moved to the right 3 places, the power (3) of 10 is negative. $0.006 = 6 \times 10^{-3}$

Example 1.2.3

Express each number in scientific notation.

a. 0.000006567
b. −0.0004004
c. 0.000000000000123

Answer a

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

Answer b

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

Answer c

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

Exercise 1.2.3

Express each number in scientific notation.

a. 0.000355
b. 0.314159
c. −0.051204

Answer a

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

Answer b

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

Answer c

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

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

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

Note

Changing a number in scientific notation to standard form:

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

Example 1.2.4

Change the number in scientific notation to standard form.

a. 6.22×10^{-2}
b. 9.9×10^{-9}

Answer a

0.0622

Answer b

0.0000000099

Exercise 1.2.4

Change the number in scientific notation to standard form.

a. 9.98×10^{-5}
b. 5.109×10^{-8}

Answer a

0.0000998

Answer b

0.00000005109

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

Concept Review Exercises

1. Why it is easier to use scientific notation to express very large or very small numbers?
2. What is the relationship between how many places a decimal point moves and the power of 10 used in changing a conventional number into scientific notation?

Answers

1. Scientific notation is more convenient than listing a large number of zeros.
2. The number of places the decimal point moves equals the power of 10. The power of 10 is positive if the decimal point moves to the left and negative if the decimal point moves to the right.

Key Takeaway

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

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1.3: Converting Units-Dimensional Analysis

Learning Objectives

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

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

Conversion Factors

Here is a simple example. How many centimeters are there in 3.55 m?

The relationship between cm and m is shown below:

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

From this relationship we can derive two conversion factors:

$$\left(\frac{100\ \text{cm}}{1\ \text{m}}\right) \quad \text{and} \quad \left(\frac{1\ \text{m}}{100\ \text{cm}}\right)$$

Conversion Factor # 1 Conversion Factor # 2

A fraction that has equivalent quantities in the numerator and the denominator but expressed in *different units* is called a **conversion factor**.

To solve the problem we use one of the two conversion factors. We first write the quantity we are given, 3.55 m. Then we multiply this quantity by conversion factor # 1.

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

Conversion factor # 1 is chosen because m, the abbreviation for meters, occurs in both the numerator *and* the denominator of our expression, they cancel out. The final step is to perform the calculation that remains once the units have been canceled. We find that 3.55 m equals 355 cm.

$$3.55\ \cancel{\text{m}} \quad \text{X} \quad \left(\frac{100\ \text{cm}}{1\ \cancel{\text{m}}}\right) \quad = \quad 355\ \text{cm}$$

A generalized description of this process is as follows:

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

If you can master the technique of applying conversion factors, you will be able to solve a large variety of problems. Figure **1.3.1** shows a **concept map** for constructing a proper conversion.

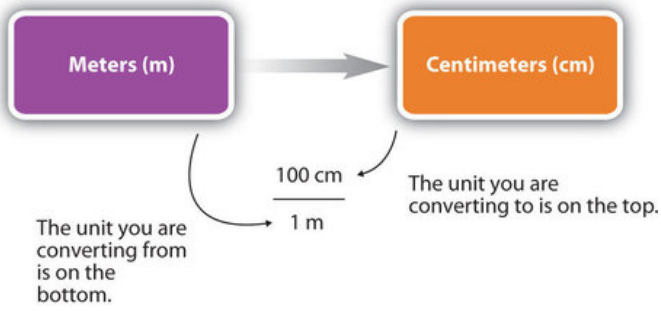


Figure **1.3.1**: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Meters is converted to centimeters. In the fraction showing 100 cm over 1 m, the unit you convert to is on top, and the unit you convert from is on bottom.

Significant Figures in Conversions

How do conversion factors affect the determination of significant figures? Numbers in conversion factors based on prefix changes, such as meters to centimeters, are *not* considered in the determination of significant figures in a calculation because the numbers in conversion factors are exact. **Exact numbers** are *defined* or *counted* numbers, (not measured numbers), and can be considered as having an infinite number of significant figures. (In other words, 1 m is exactly 100 cm, by the definition of centi-.) Counted numbers are also exact. If there are 16 students in a classroom, the number 16 is exact. In contrast, conversion factors that come from measurements (such as density, as we will see shortly) are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

Example 1.3.1

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

Solution

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

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

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



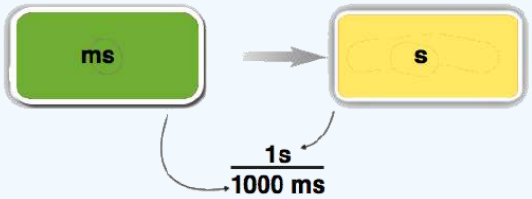
$$4.7\,\cancel{\text{L}} \times \frac{1{,}000\text{ mL}}{1\,\cancel{\text{L}}} = 4{,}700\text{ mL}$$

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

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

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

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



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

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

Exercise 1.3.1

Perform each conversion.

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

Answer a

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

Answer b

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

Concept Review Exercises

1. How do you determine which quantity in a conversion factor goes in the denominator of the fraction?
2. State the guidelines for determining significant figures when using a conversion factor.

Answers

1. The unit you want to cancel from the numerator goes in the denominator of the conversion factor.
2. Exact numbers that appear in many conversion factors do not affect the number of significant figures; otherwise, the normal rules of multiplication and division for significant figures apply.

Key Takeaway

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

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1.4: States of Matter - Solid, Liquid, and Gas

Learning Objectives

- To describe the solid, liquid and gas phases.

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

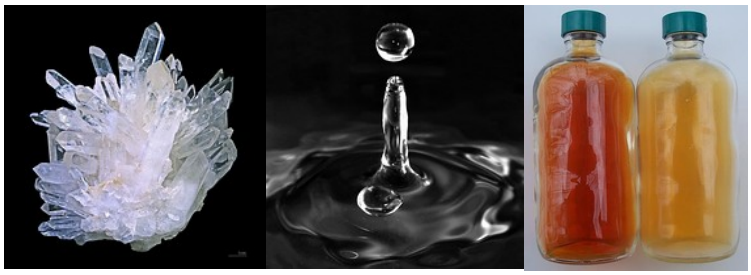


Figure **1.4.1**: Matter is usually classified into three classical states. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances exist as gases at room temperature such as oxygen and carbon dioxide. Others, like water and mercury metal, exist as liquids at room temperature. Most metals exist as solids at room temperature. Figure **1.4.2** shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

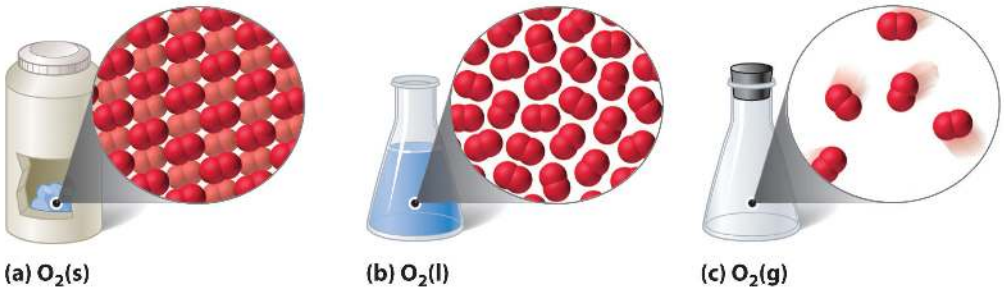


Figure **1.4.2**: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

Plasma: A Fourth State of Matter

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



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

Solids

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

Solids are defined by the following characteristics:

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

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

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure 1.4.3. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

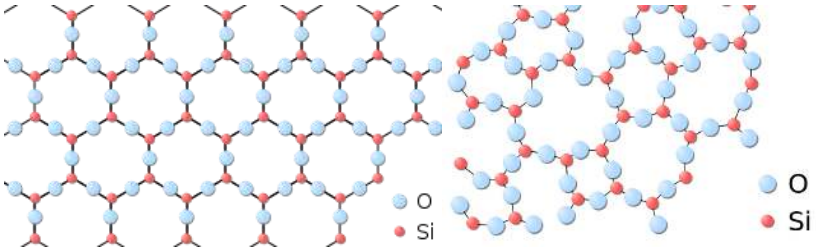


Figure 1.4.3: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that in the crystal, each silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

Liquids

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

Liquids have the following characteristics:

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

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature.

Gases

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

- No definite shape (takes the shape of its container)
- No definite volume

- Particles move in random motion with little or no attraction to each other
- Highly compressible

Table 1.4.1: Characteristics of the Three States of Matter

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

Example 1.4.1

What state or states of matter does each statement describe?

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

Solution

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

Exercise 1.4.1

What state or states of matter does each statement describe?

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

Answer a:

solid

Answer b:

gas

Answer c:

solid

Summary

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

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1.5: Energy

Learning Objectives

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

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we could not turn on lights, make our lunch, nor travel to school. In fact, without energy, we could not even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we are awake or asleep.

Ability to Do Work or Produce Heat

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

When scientists speak of ***heat***, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot—the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents (Figure 1.5.1).



Figure 1.5.1: Energy is transferred as heat from the hot stove element to the pot until the pot and its contents become just as hot as the element. The energy that is transferred into the pot as heat is then used to cook the food.

Heat is one way in which energy can be transferred. Energy can also be transferred as **work**. The scientific definition of work is *force (any push or pull) applied over a distance*. Whenever you push an object and cause it to move, you have done work, and transferred some energy to the object. So far, we've talked about the two ways in which energy can be transferred from one place, or object, to another. Energy can be transferred as heat, and energy can be transferred as work.

Kinetic Energy

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

Kinetic energy is very common, and it's easy to spot examples of it in the world around you. Sometimes we even try to capture kinetic energy and use it to power things like our home appliances. If you are from California, you might have driven through the Tehachapi Pass near Mojave or the Montezuma Hills in Solano County and seen the windmills lining the slopes of the mountains (Figure 1.5.2). These are two of the larger wind farms in North America. As wind rushes along the hills, the kinetic energy of the moving air particles turns the windmills, trapping the wind's kinetic energy so that people can use it in their houses and offices.



Figure 1.5.2: A wind farm in Solano County harnesses the kinetic energy of the wind. (CC BY-SA 3.0 Unported; BDS2006 at Wikipedia)

Potential Energy

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

Any stored energy is potential energy. Unfortunately, there are a lot of different ways in which energy can be stored, and that can make potential energy very difficult to recognize. In general, an object has potential energy because of its *position relative to another object*. For example, when you hold a rock above the earth, it has potential energy because of its position relative to the ground. You can tell that this is *potential energy* because the energy is *stored* for as long as you hold the rock in the air. Once you drop the rock, though, the stored energy is converted to kinetic energy as the rock falls. A ball at the top of a hill stores potential energy until it is allowed to roll to the bottom. When you hold two magnets next to each other, they store potential energy too. For some examples of potential energy, it is harder to see how "position" is involved.

Chemical Energy

In chemistry, we are often interested in what is called **chemical potential energy**. Chemical potential energy is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position?

As you learned earlier, the world, and all of the chemicals in it are made up of atoms and molecules. These store potential energy that is dependent on their positions relative to one another. Of course, you cannot see atoms and molecules. However, scientists do know a lot about ways in which atoms and molecules interact. This allows them to figure out how much potential energy is stored in a specific quantity (like a cup or a gallon) of a particular chemical. *Different chemicals have different amounts of potential energy* because they are made up of different atoms, and those atoms have different positions relative to one another.

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

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

Units of Energy

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

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

$$\textbf{1 kilocalorie} = \textbf{1 Calorie} = \textbf{1000 calories} \tag{1.5.1}$$

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

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

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

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

$$\textbf{400 Cal} = \textbf{400 kcal} \times \frac{\textbf{4.184 kJ}}{\textbf{1 kcal}} = \textbf{1.67} \times \textbf{10}^{\textbf{3}}\ \textbf{kJ} \tag{1.5.2}$$

Summary

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

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1.6: Measurements and Units

Learning Objectives

- Express quantities properly, using a number and a unit.
- State the different measurement systems used in chemistry.

Measurements

A coffee maker's instructions tell you to fill the coffeepot with 4 cups of water and to use 3 scoops of coffee. When you follow these instructions, you are measuring. When the nurse checks your temperature, height, weight, and blood pressure at the doctor's office, (Figure 1.6.1), the nurse is also measuring.



Figure 1.6.1: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GFDL; Pia von Lützau).

Chemists measure the properties of matter and express these measurements as quantities. A quantity is an amount of something and consists of a number and a unit. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5. If you ask a friend how far they walk from home to school, and the friend answers "12" without specifying a unit, you do not know whether your friend walks 12 kilometers, 12 miles, 12 furlongs, or 12 yards. *Both a number and a unit must be included to express a quantity properly.*

Example 1.6.1

Identify the number and the unit in each quantity.

1. one dozen eggs
2. 2.54 centimeters
3. a box of pencils
4. 88 meters per second

Solution

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

Units of Measurements

To understand chemistry, we need a clear understanding of the units chemists work with and the rules they follow for expressing numbers. How long is a yard? It depends on whom you ask and when you asked the question. Today we have a standard definition

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

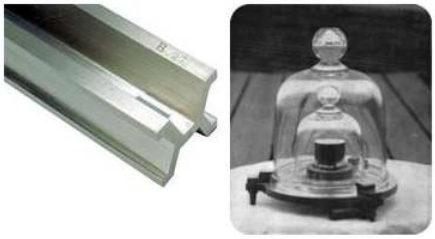


Figure **1.6.2**: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science. 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. Five of the base units are shown in Table **1.6.1**.

Table **1.6.1**: SI Base Units of Measurement

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

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

Unfortunately, The Metric System is Not Ubiquitous

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

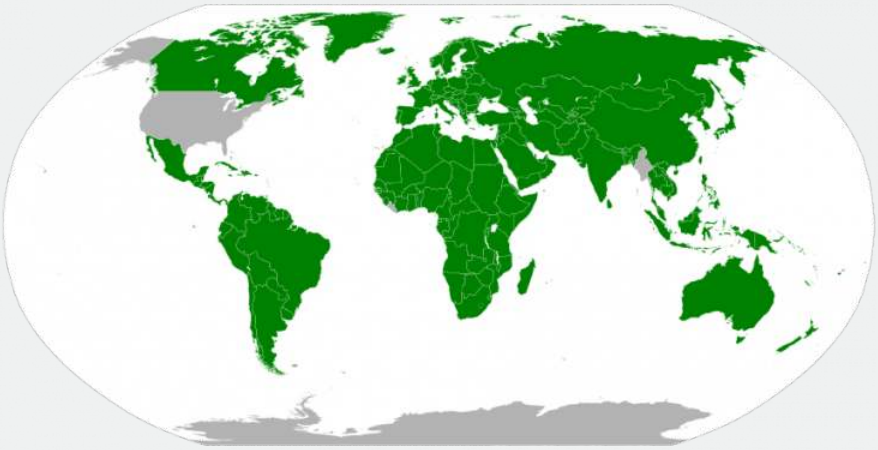


Figure 1.6.1: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Summary

- Identify a quantity properly with a number and a unit.
- The basic units in the SI system such as mass, length, temperature, time, and amount of substance are commonly used in chemistry.

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1.7: Temperature - Motion of Molecules and Atoms

Learning Objectives

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

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



Figure **1.7.1**: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

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



Figure **1.7.2**: Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right).

The Fahrenheit Scale

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

The Celsius Scale

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

The Kelvin Scale

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

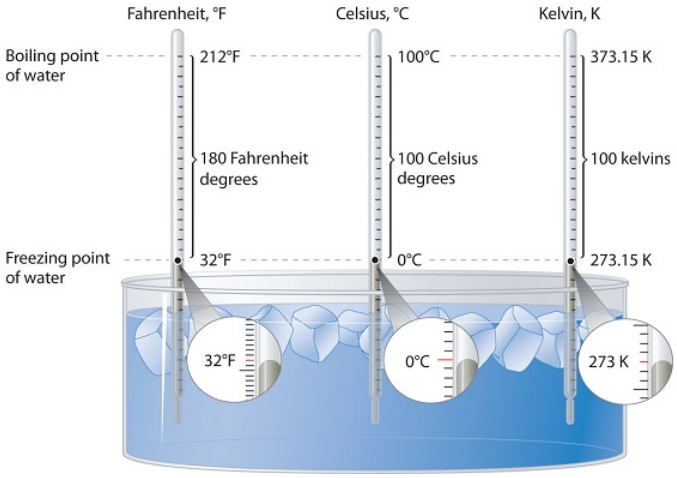


Figure 1.7.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

Converting between Scales

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

$$T\,(\text{in °C}) + 273.15 = T\,(\text{in K}) \tag{1.7.1}$$

$$T\ \ (\text{in K}) - 273.15 = T\,(\text{in °C}) \tag{1.7.2}$$

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

$$°C = \frac{(°F - 32)}{1.8} \tag{1.7.3}$$

$$°F = 1.8 \times (°C) + 32 \tag{1.7.4}$$

Example 1.7.1: Temperature Conversions

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

Solution

Converting from Fahrenheit to Celsius requires the use of Equation 1.7.3:

(103.5°F - 32) / 1.8 = 39.7 °C

Converting from Celsius to Kelvin requires the use of Equation 1.7.1:

39.7 °C + 273.15 = 312.9 K

Exercise 1.7.1

Convert each temperature to °C and °F.

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

Answer (a)

5527 °C, 9981. °F

Answer (b)

2807 °C, 5084 °F

Answer (c)

-195.79 °C, -320.42 °F

Summary

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

Contributions & Attributions

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1.8: Density and Specific Gravity

Learning Objectives

- Define density and specific gravity.
- Perform calculations involving both density and specific gravity.

After trees are cut, logging companies often move these materials down a river to a sawmill where they can be shaped into building materials or other products. The logs float on the water because they are less dense than the water they are in. Knowledge of density is important in the characterization and separation of materials. Information about density allows us to make predictions about the behavior of matter.

Density

A golf ball and a table tennis ball are about the same size. However, the golf ball is much heavier than the table tennis ball. Now imagine a similar size ball made out of lead. That would be very heavy indeed! What are we comparing? By comparing the mass of an object relative to its size, we are studying a property called **density**. Density is the ratio of the mass of an object to its volume.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{1.8.1}$$

$$D = \frac{m}{V} \tag{1.8.2}$$

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. Common units for density include g/mL, g/cm^3, g/L, kg/L, and even kg/m^3. Densities for some common substances are listed in Table 1.8.1.

Table 1.8.1: Densities of Some Common Substances

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084
Ice (0°C)	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water (4°C)	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.92		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

The SI units of density are kilograms per cubic meter $\left(\text{kg/m}^3\right)$, since the kg and the m are the SI units for mass and length respectively. In everyday usage in a laboratory, this unit is awkwardly large. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter $\left(\text{g/cm}^3\right)$. Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Water has a density of 1.0 g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L.



Figure 1.8.1: Separatory Funnel containing oil and colored water to display density differences. (CC BY-SA 3.0; PRHaney vai Wikipedia).

If a substance's relative density is less than one then it is less dense than water. If greater than 1, then the substance is denser than water. If the relative density is exactly 1 then the densities are equal. For example, an ice cube, with a relative density of about 0.91 g/mL, will float on water and a substance with a relative density greater than 1 will sink. Corn syrup, being denser, would sink if placed in water.

Since most materials expand as temperature increases, the density of a substance is temperature dependent and usually decreases as temperature increases. For example, cooking oil becomes less dense when heated in a pan.

Density as a Conversion Factor

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

$$13.6\ \text{g mercury} = 1\ \text{mL mercury}$$

This relationship can be used to construct two conversion factors:

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

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

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

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

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

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

Example 1.8.1

An $18.2\ \text{g}$ sample of zinc metal has a volume of $2.55\ \text{cm}^3$. Calculate the density of zinc.

Solution

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

Known

- Mass $= 18.2\ \text{g}$
- Volume $= 2.55\ \text{cm}^3$

Unknown

- Density $=?\ \text{g/cm}^3$

Use Equation 1.8.1 to solve the problem.

Step 2: Calculate

$$D = \frac{m}{V} = \frac{18.2\ \text{g}}{2.55\ \text{cm}^3} = 7.14\ \text{g/cm}^3 \tag{1.8.3}$$

Step 3: Think about your result.

$1\ \text{cm}^3$ of zinc has a mass of 7.14 grams. Metals are expected to have a density greater than that of water and zinc's density falls within the range of the other metals listed above.

Since density values are known for many substances, density can be used to determine an unknown mass or an unknown volume. Dimensional analysis will be used to ensure that units cancel appropriately.

Example 1.8.2

What is the mass of $2.49\ \text{cm}^3$ of aluminum?

Solution

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

Known

- Density $= 2.70\ \text{g/cm}^3$
- Volume $= 2.49\ \text{cm}^3$

Unknown

- Mass $=?\ \text{g}$

Use the equation for density, $D = \frac{m}{V}$, and dimensional analysis to solve each problem.

Step 2: Calculate

$$2.49\ \text{cm}^3 \times \frac{2.70\ \text{g}}{1\ \text{cm}^3} = 6.72\ \text{g} \tag{1.8.4}$$

In this problem, the mass is equal to the density multiplied by the volume.

Step 3: Think about your results.

Because a mass of $1\ \text{cm}^3$ of aluminum is $2.70\ \text{g}$, the mass of about $2.5\ \text{cm}^3$ should be about 2.5 times larger.

Example 1.8.3

What is the volume of $50.0\ \text{g}$ of aluminum?

Solution

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

Known

- Density $= 2.70\ \text{g/cm}^3$
- Mass $= 50.0\ \text{g}$

Unknown

- Volume $=?\ \text{cm}^3$

Use the equation for density, $D = \frac{m}{V}$, and dimensional analysis to solve each problem.

Step 2: Calculate

$$50.0\ \text{g} \times \frac{1\ \text{cm}^3}{2.70\ \text{g}} = 18.5\ \text{cm}^3 \tag{1.8.5}$$

In this problem 2, the volume is equal to the mass divided by the density.

Step 3: Think about your results.

The $50\ \text{g}$ of aluminum is substantially more than its density, so that amount should occupy a relatively large volume.

Specific Gravity

Specific gravity is the ratio of the density (mass of a unit volume) of a substance to the density of water,

$$\text{specific gravity} = \frac{\text{Density of a substance}(\cancel{g/mL})}{\text{Density of the water at the same temperature}(\cancel{g/mL})} \tag{1.8.6}$$

The units cancel out in the ratio. Therefore, the specific gravity is a unitless number. The density of water is 1.0 g/mL at room temperature, so the specific gravity is equal to the density of the object expressed without a unit.

✔ Example 1.8.1

At 20.0°C, the density of ethanol is 0.790 g/mL and that of water is 1.00 g/mL. What is the specific gravity of ethanol at 20.0°C?

Solution

$$\text{Specific gravity of alcohol} = \frac{\text{density of alcohol}}{\text{density of water}} = \frac{0.790\ \text{g/mL}}{1.00\ \text{g/mL}} = 0.790$$

The units cancel out and so specific gravity is a unitless quantity.

? Exercise 1.8.1

At 20.0°C, the density of mercury is 13.6 g/mL and that of water is 1.00 g/mL. What is the specific gravity of mercury at 20.0°C?.

Answer

$$\text{Specific gravity of mercury} = \frac{\text{density of mercury}}{\text{density of water}} = \frac{13.6\ \text{g/mL}}{1.00\ \text{g/mL}} = 13.6$$

A *hydrometer* is an instrument used for measuring the specific gravity of liquids based on the concept of buoyancy (Figure **1.8.2**). A hydrometer usually consists of a sealed hollow glass tube with a wider bottom portion for buoyancy, a ballast such as lead or mercury for stability, and a narrow stem with graduations for measuring. The liquid to test is poured into a tall container, often a graduated cylinder, and the hydrometer is gently lowered into the liquid until it floats freely. The point at which the surface of the liquid touches the stem of the hydrometer correlates to relative density. Hydrometers can contain any number of scales along the stem corresponding to properties correlating to the density.



Figure **1.8.2**: A US Navy Aviation Boatswain's Mate tests the specific gravity of JP-5 fuel. (Public Domain; U.S. Navy via Wikipedia)

When substances dissolve in water, the specific gravity of the solution is usually different from pure water. For example, the specific gravity of whole blood for humans is ~1.060. The density of urine varies in the range of 1.005 to 1.030. Both the blood and urine have dissolved substances in water that increase the density from that of pure water. Both high and low density or specific gravity than the normal range of urine indicates medical problems. An increase in the specific gravity of urine indicates that it is due to an increase in the solutes caused by dehydration, diarrhea, or infection. Similarly, a decrease in solute concentration decreases the specific gravity of urine, which indicates medical problems like renal failure.

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

Summary

- Density is the ratio of the mass of an object to its volume.
- Gases are less dense than solids and liquids.
- For liquids and gases, the temperature will affect the density to some extent.
- Specific gravity is the ratio of the density of a substance to the density of water.

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1.9: Dosage Calculations I

Learning Objectives

- Calculate drug dosages using conversion factors.

To Your Health: Dosages

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

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

Drug dosage calculations

Drug dosage calculation is required if the physician's order is different from what is available. The medication may be administered as fluid or tablet.

Example 1.9.1

Epinephrine is used to treat peanut allergies. A 5 year old patient needs an immediate treatment of 0.12 mg of epinephrine. The solution is available in concentration of 0.50 mg/ml. Calculate the volume in ml required for the patient.

Solution

The first step is to understand the meaning of concentration of solution. There are 0.50 mg in 1 mL of solution. From this relationship, we can construct the conversion factor. We use the conversion factor that will cancel out the original unit, *mg*, and introduce the unit we are converting to, which is *mL*.

$$0.12 \cancel{\text{mg}} \times \frac{1 \text{ mL}}{0.50 \cancel{\text{mg}}} = 0.24 \text{ mL}$$

Thus the patient requires 0.24 mL of epinephrine.

Example 1.9.2

A physician ordered 100 mg of Demerol. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?

Solution

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

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

Hence, the nurse should administer 2 tablets.

Example 1.9.3

The physician ordered 20 mg of Valium. Valium is available as 10 mg per tablet. How many tablets should the nurse administer?

Solution

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

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

Contributors and Attributions

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1.10: Metric Prefixes

Learning Objectives

- Describe how prefixes are used in the metric system.
- Identify how the metric prefixes nano, micro, milli-, centi-, deci, and kilo- compare to the base unit.

Many aspects of chemistry use quantitative measurements to describe properties of matter. In this section, we will look at how we deal with very large or very small measured numbers, with prefixes that are used with units of measurements.

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 *μεγας*, meaning "great". Table **1.10.1** lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table **1.10.1**: SI Prefixes

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

Note

*The abbreviation for metric prefix micro is **mc** in the field of medicine.*

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

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

Example 1.10.1: Unit Abbreviations

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

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

Solutions

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

Exercise 1.10.1

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

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

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Example 1.10.2

Convert 81.2 g to kilograms following using the dimensional analysis. Show clearly conversion factors.

Solution

1 kg = 10^3 g

Two conversion factors are possible:

$$\left(\frac{1\ \text{kg}}{10^3\ \text{g}} \right) \text{ and } \left(\frac{10^3\ \text{g}}{1\ \text{kg}} \right)$$

Use the first conversion factor to cancel out the units of grams:

$$81.2\ \text{g} \times \left(\frac{1\ \text{kg}}{1 \times 10^3\ \text{g}} \right) = 0.0812\ \text{kg} = 8.12 \times 10^{-2}\ \text{kg}$$

Problem Solving With Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. Kilo and milli are both metric prefixes. You can do the conversion in two steps.

To convert one metric prefix to another, it is useful to go through the base unit. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence. We first convert the given amount in km to the base unit which is meters. We know that 1 km = 1,000 m.

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

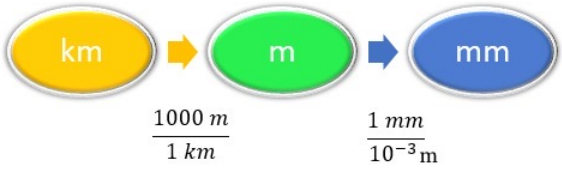
Then we convert meters to mm, remembering that 1 mm = 10^{-3} m.

$$54,700\ \text{m} \times \left(\frac{1\ \text{mm}}{10^{-3}\ \text{m}} \right) = 54,700,000\ \text{mm} = 5.47 \times 10^7\ \text{mm}$$

We have expressed the final answer in scientific notation.

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

Concept Map



Calculation

$$54.7\ \text{km} \times \frac{1,000\ \text{m}}{1\ \text{km}} \times \frac{1\ \text{mm}}{10^{-3}\ \text{m}} = 54,700,000\ \text{mm}$$
$$= 5.47 \times 10^7\ \text{mm}$$

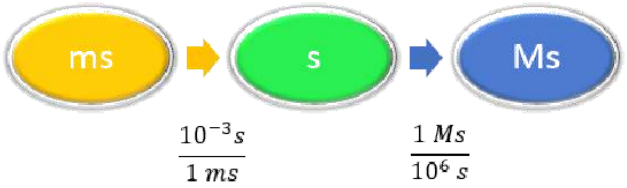
In each step, the previous unit is canceled and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.

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

Example 1.10.3

Convert 58.2 milliseconds (ms) to megaseconds (Ms) in one multi-step calculation.

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 58.2 ms Find: Ms
List other known quantities	1 ms = 10^{-3} s 1 Ms = 10^{6} s
Prepare a concept map	
Calculate	$58.2\ \cancel{ms} \times \frac{10^{-3}\ \cancel{s}}{1\ \cancel{ms}} \times \frac{1\ \text{Ms}}{1,000,000\ \cancel{s}} = 0.0000000582\ \text{Ms}$ $= 5.82 \times 10^{-8}\ \text{Ms}$ Neither conversion factor affects the number of significant figures in the final answer.

Exercise 1.10.2

Convert 43.007 mg to kilograms in one multistep calculation.

Answer

$$43.007\ \cancel{mg} \times \frac{1\ \cancel{g}}{1,000\ \cancel{mg}} \times \frac{1\ \text{kg}}{1,000\ \cancel{g}} = 0.000043007\ \text{kg}$$
$$= 4.3007 \times 10^{-5}\ \text{kg}$$

.

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

Example 1.10.4

Write a concept map (a plan) for how you would convert 1.0×10^{12} *nano*liters (nL) to *kilo*liters (kL).

Solution

Concept Map: Convert the given (nanoliters, nL) to liters; then convert liters to kiloliters.



https://chem.libretexts.org/@go/page/434747

Summary

Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the measured numbers manageable. The SI system is based on multiples of ten.

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- **1.7: Converting Units** by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological.

1.11: Dosage Calculations II

Learning Objectives

- Calculate drug dosages using metric prefixes and conversion factors.

Effects are dose-dependent

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

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

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

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

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

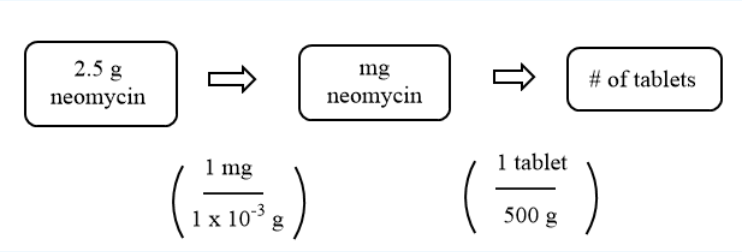
Drug dosage calculation is required if the physician's order is different from what is available. Here we will see examples where the medications are administered as fluids or tablets.

Example 1.11.1

Neomycin is an antibiotic. A patient needs 2.5 grams of Neomycin sulfate. 500 mg tablets are available. How many tablets will you give to the patient?

Solution

It is important to set up a concept map, so as to plan the steps involved:



We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence.

$$2.5\ \text{g} \times \left(\frac{1\ \text{mg}}{1 \times 10^{-3}\ \text{g}}\right) \times \left(\frac{1\ \text{tablet}}{500\ \text{mg}}\right) = 5\ \text{tablets}$$

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

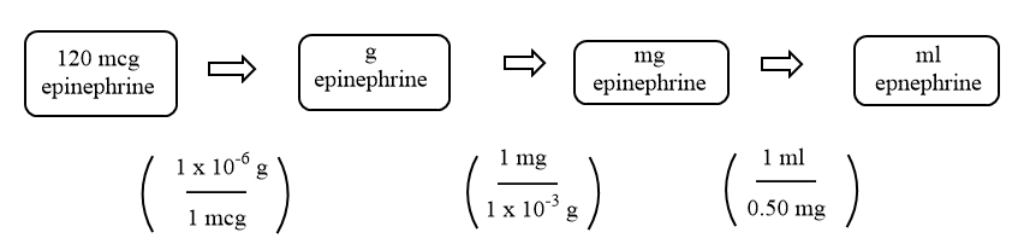
The patient requires 5 tablets of neomycin.

Example 1.11.2

Epinephrine is used to treat peanut allergies. A 5 year old patient needs an immediate treatment of 120 mcg of epinephrine. The solution is available in concentration of 0.50 mg/mL. Calculate the volume (mL) required for the patient?

Solution

It is important to set up a concept map, so as to plan the steps involved:



We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence.

$$120\ \cancel{mcg} \times \left(\frac{1 \times 10^{-6}\ \cancel{g}}{1\ \cancel{mcg}}\right) \times \left(\frac{1\ \cancel{mg}}{1 \times 10^{-3}\ \cancel{g}}\right) \times \left(\frac{1\ mL}{0.50\ \cancel{mg}}\right) = 0.24\ mL$$

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

The 5 year old patient requires 0.24 ml of epinephrine.

Example 1.11.3

The doctor's order is 1.2 g of Folic Acid. Folic Acid is available as 800 mg per tablet. How many tablets should be taken?

Solution

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

conversion factor for each step.

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

Hence, 1.5 tablets should be taken.

Example 1.11.4

The doctor's order is 1 g of Calcium. What is on hand is Calcium as 500 mg per tablet. How many tablets should be taken?

Solution

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

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

Hence, 2 tablets should be taken.

Note

Drug Dosage Calculation based on Body Weight.

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

Example 1.11.5

Demerol is ordered 1.5mg/kg for a patient that is 220 lbs. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?

Solution

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

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

Hence, the nurse administers 3 tablets.

Example 1.11.6

A doctor prescribes amoxicillin 30mg/kg to a child weighing 73.5 lbs. Amoxicillin is available as 500 mg tablets. How many tablets should the nurse administer?

Solution

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

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

Hence, the nurse administers 2 tablets.

Career Focus: Pharmacist

A pharmacist dispenses drugs that have been prescribed by a doctor. Although that may sound straightforward, pharmacists in the United States must hold a doctorate in pharmacy and be licensed by the state in which they work. Most pharmacy programs require four years of education in a specialty pharmacy school. Pharmacists must know a lot of chemistry and biology so they can understand the effects that drugs (which are chemicals, after all) have on the body. Pharmacists can advise physicians on the selection, dosage, interactions, and side effects of drugs. They can also advise patients on the proper use of

their medications, including when and how to take specific drugs properly. Pharmacists can be found in drugstores, hospitals, and other medical facilities. Curiously, an outdated name for pharmacist is *chemist*, which was used when pharmacists formerly did a lot of drug preparation, or *compounding*. In modern times, pharmacists rarely compound their own drugs, but their knowledge of the sciences, including chemistry, helps them provide valuable services in support of everyone's health.



A woman consulting with a pharmacist. (Public Domain; Rhoda Baer via National Cancer Institute, an agency part of the National Institutes of Health.)

Summary

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

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1.12: End of Chapter Problems I

Significant Figures

1. How many significant figures does each number have?

6,798,000

6,000,798

6,000,798.00

0.0006798

2.1828

0.005505

55,050

5

500

2. Solve and write the answer with the appropriate number of significant figures.

23.096 × 90.300

125 × 9.000

1,027 + 610.0 + 363.06

217 ÷ 903

13.77 + 908.226 + 515

255.0 − 99

0.00666 x 321

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

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

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

Scientific Notation and Decimal Form

1. Express each number in scientific notation.

67,000,000,000

1,689

12.6

1,492

102,000,000

101,325

0.000006567

−0.0004004

0.000000000000123

0.000355

0.314159

−0.051204

2. Change the number in scientific notation to standard form.

5.27×10^4

1.0008×10^6

6.98×10^8

1.005×10^2

6.22×10^{-2}

9.9×10^{-9}

9.98×10^{-5}

5.109×10^{-8}

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

Converting Units

1. The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters? [1L =1000 mL]

2. A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms? [1 s = 1000 ms]

3. Perform the conversion 32.08 kg to grams. [1 kg = 1000 g]

States of Matter

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

a. This state has a definite volume, but no definite shape.
b. This state has no definite volume.
c. This state allows the individual particles to move about while remaining in contact.
d. This state has individual particles in a fixed position with regard to each other.
e. This state has individual particles far apart from each other in space.
f. This state has a definite shape.

Energy

1. Classify each of the following energies as kinetic energy or potential energy:

The energy in a chocolate bar.

The energy of rushing water used to turn a turbine or a water wheel.

The energy of a skater gliding on the ice.

The energy in a stretched rubber band.

2. You mix two chemicals in a beaker and notice that as the chemicals react, the beaker becomes noticeably colder. Which chemicals have more chemical potential energy, those present at the start of the reaction or those present at the end of the reaction?

Measurements

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

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

Temperature

1. Convert each temperature to °C.

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

2. Convert each temperature to K.

98.6 ^{o}C

25.0 ^{o}C

0.0 ^{o}C

-40.0 ^{o}C

Density and Specific Gravity

For the following problems assume room temperature of 20°C.

1. Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.7102 g. What is the specific gravity of ethyl alcohol?

2. If you have a 2.130 mL sample of acetic acid with mass 0.002234 kg, what is the density in kg/L? What is the specific gravity of acetic acid?

3. A mercury thermometer for measuring a patient's temperature contains 0.750 g of mercury. What is the volume of this mass of mercury? The density of Hg is 13.6 g/mL. What is the specific gravity of Hg?

4. You have a sample of aluminum that has a volume of 7.88 mL. What is the mass of Al if the density is 2.70 g/mL. What is the specific gravity of Al?

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1.13: End of Chapter Problems II

The following problems on dosage calculations I are to be completed with unit 2.

Dosage Calculations I

Please set up your calculations using the dimensional analysis.

1. Epinephrine is used to treat peanut allergies. A 5 year old patient needs an immediate treatment of 0.12 mg of epinephrine. The solution is available in concentration of 0.50 mg/ml. Calculate the volume in ml required for the patient.

2. A physician ordered 100 mg of Demerol. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?

3. The physician ordered 20 mg of Valium. Valium is available as 10 mg per tablet. How many tablets should the nurse administer?

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

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

6. The tranquilizer valium is sold in 2.0 mL syringes that contain 50.0 mg of the drug per 1.0 mL of liquid. If a physician prescribes 25 mg of this drug, how many mL should be administered?

7. An antibiotic is sold in 3.0 mL ampoules that contain 60.0 mg of the drug. How many mL of the antibiotic should be withdrawn from the ampoule if 45 mg are to be administered to a patient?

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1.14: End of Chapter Problems III

The following problems on metric prefixes and dosage calculations are to completed with unit 3.

Metric Prefixes

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

kiloliter

microsecond

decimeter

nanogram

kilometer

milligram

nanosecond

centiliter

2. Convert 81.2 g to kilograms following using the dimensional analysis. Show clearly conversion factors.

3. Convert 58.2 milliseconds (ms) to megaseconds (Ms).

4. Convert 47.66 mcg to dg using the factor label method.

5. Please complete the following conversions using dimensional analysis.

a. 5.67 mm to m

b. 5.67 nm to m

c. 0.3 cm to m

d. 1.5 km to mm

e. 24 mcg to mg

f. 15 dg to cg

g. 0.712 mg to kg

h. 412 kg to dg

Dosage Calculations II

Please set up your calculations using the dimensional analysis.

1. Neomycin is an antibiotic. A patient needs 2.5 grams of Neomycin sulfate. 500 mg tablets are available. How many tablets will you give to the patient?

2. Epinephrine is used to treat peanut allergies. A 5 year old patient needs an immediate treatment of 120 mcg of epinephrine. The solution is available in concentration of 0.50 mg/mL. Calculate the volume (mL) required for the patient?

3. The doctor's order is 1.2 g of folic acid. Folic Acid is available as 800 mg per tablet. How many tablets should be taken?

4. The doctor's order is 1 g of Calcium. What is on hand is Calcium as 500 mg per tablet. How many tablets should be taken?

5. Demerol is ordered 1.5mg/kg for a patient that is 220 lbs. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?

6. A doctor prescribes amoxicillin 30 mg/kg to a child weighing 73.5 lbs. Amoxicillin is available as 500 mg tablets. How many tablets should the nurse administer?

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

8. A 175 lb patient is to undergo surgery and will be given an intravenous anesthetic. The safe dosage of anesthetic is up 12 mg/kg of body weight. Determine the mg of the dose of anesthetic that should be used. (2.205 lbs = 1 kg)

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

10. The antiviral drug Stavudine is used for the treatment of AIDS. The daily recommended dosage of Stavudine is 1.0 mg/kg of body weight. How many grams of the drug should be administered to a 150 lb patient?

11. Tetracycline is an antibiotic used to mark polar bears. If 25 mg/kg is an effective dose, how much tetracycline is needed in grams to mark a 100 kg polar bear?

12. Ivermectine is used to treat dogs that have intestinal parasites. The effective dosage for this drug is 10.5 mcg/kg of body weight. How much ivermectin should be given to a 9.0 kg dog?

13. Chloroquine is used to treat malaria. Studies have shown that an effective dose for children is 3.5 mg/kg of body weight every 6 hours. If a child weighs 12 kg, how many milligram of this drug should be given in a 24 hour period?

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CHAPTER OVERVIEW

2: Topics in General Chemistry II

2.1: The Structure of Atoms
2.2: The Elements
2.3: Atomic Number and Mass Number
2.4: Atomic Weight
2.5: The Periodic Table
2.6: Valence Electrons and Electron Dot Formula
2.7: The Mole and Mole-Mass Conversions
2.8: Physical and Chemical Properties and Changes
2.9: End of Chapter Problems

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2.1: The Structure of Atoms

Learning Objectives

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

Prior to the 1900's atoms were considered to be indivisible. We know now that atoms not only can be divided but also are composed of three different kinds of particle, *subatomic particles*, with their own properties that are different from the chemical properties of atoms.

Subatomic Particles

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

Finally, additional experiments pointed to the existence of a third particle. Evidence produced in 1932 established the existence of the neutron, a particle with about the same mass as a proton but with no electrical charge, it is *neutral*. We understand now that all atoms can be broken down into subatomic particles: protons, neutrons, and electrons. Table **2.1.1** lists some of their important characteristics and the symbols used to represent each particle.

Table **2.1.1**: Properties of the Subatomic Particles

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

The Nucleus

How are these subatomic particles arranged? in the early 1900's much experimental work happened. A model was proposed to explain the structure of the atom. Protons and neutrons are concentrated in a central region called the nucleus of the atom. Electrons are outside the nucleus and orbit about it because they are attracted to the positive charge in the nucleus. Most of the mass of an atom is in the nucleus, while the orbiting electrons account for an atom’s size. As a result, an atom consists largely of empty space. This description is the “planetary model” of the atom.

It is essentially the same model that we use today to describe atoms but with one important modification. The planetary model suggests that electrons occupy certain specific, circular orbits about the nucleus. We know now that this model is overly simplistic. A better description is that electrons form fuzzy clouds around nuclei. Figure **2.1.1** shows a more modern version of our understanding of atomic structure.

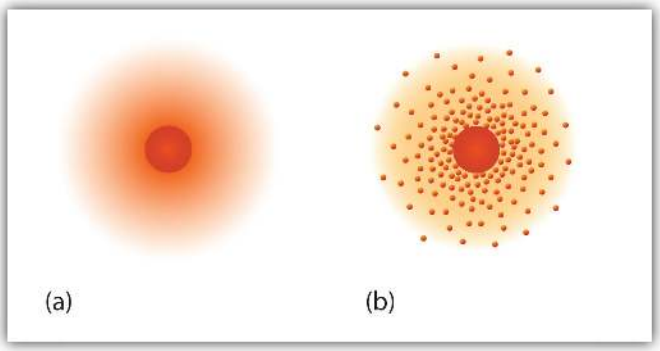


Figure **2.1.1**: A Modern Depiction of Atomic Structure. A more modern understanding of atoms, reflected in these representations of the electron in a hydrogen atom, is that electrons occupy regions of space about the nucleus; they are not in discrete orbits like planets around the sun. (a) The darker the color, the higher the probability that an electron will be at that point. (b) In a two-dimensional cross section of the electron in a hydrogen atom, the more crowded the dots, the higher the probability that an electron will be at that point. In both (a) and (b), the nucleus is in the center of the diagram.

Concept Review Exercises

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

Answers

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

Key Takeaways

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

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2.2: The Elements

Learning Objectives

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

An element is a substance that cannot be broken down into simpler chemical substances. There are about 90 naturally occurring elements known on Earth. Using technology, scientists have been able to create nearly 30 additional elements that do not occur in nature. Today, chemistry recognizes 118 elements—some of which were created an atom at a time. Figure **2.2.1** shows some of the chemical elements.



Figure **2.2.1**: Samples of Elements. Gold is a yellowish solid, iron is a silvery solid, while mercury is a silvery liquid at room temperature. © Thinkstock

Abundance

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

Table **2.2.1**: Elemental Composition of Earth

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

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

On the planet Earth, however, the situation is rather different. Oxygen makes up 46.1% of the mass of Earth's crust (the relatively thin layer of rock forming Earth's surface), mostly in combination with other elements, while silicon makes up 28.2%. Hydrogen, the most abundant element in the universe, makes up only 0.14% of Earth's crust. Table **2.2.1** lists the relative abundances of elements on Earth as a whole and in Earth's crust. Table **2.2.2** lists the relative abundances of elements in the human body. If you compare Table **2.2.1** and Table **2.2.2**, you will find disparities between the percentage of each element in the human body and on Earth. Oxygen has the highest percentage in both cases, but carbon, the element with the second highest percentage in the body, is relatively rare on Earth and does not even appear as a separate entry in Table **2.2.1**; carbon is part of the 0.174% representing "other" elements. How does the human body concentrate so many apparently rare elements?

Table **2.2.2**: Elemental Composition of a Human Body

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

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

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

LOOKING CLOSER: PHOSPHOROUS, THE CHEMICAL BOTTLENECK

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

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

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

Names and Symbols

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

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

The names of the elements can be cumbersome to write in full, especially when combined to form the names of compounds. Therefore, each element name is abbreviated as a one- or two-letter chemical symbol. By convention, the first letter of a chemical symbol is a capital letter, while the second letter (if there is one) is a lowercase letter. The first letter of the symbol is usually the first letter of the element's name, while the second letter is some other letter from the name. Some elements have symbols that derive from earlier, mostly Latin names, so the symbols may not contain any letters from the English name. Table **2.2.3** lists the names and symbols of some of the most familiar elements.

Table **2.2.3**: Element Names and Symbols

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

*The symbol comes from the Latin name of element.

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

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

Example 2.2.1

Write the chemical symbol for each element without consulting Table **2.2.3** "Element Names and Symbols".

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

Answer a

Br

Answer b

B

Answer c

C

Answer d

Ca

Answer e

Au

Exercise 2.2.1

Write the chemical symbol for each element without consulting Table 2.2.3.

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

Answer a

Mn

Answer b

Mg

Answer c

Ne

Answer d

N

Answer e

Ag

Example 2.2.2

What element is represented by each chemical symbol?

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

Answer a

sodium

Answer b

mercury

Answer c

phosphorus

Answer d

potassium

Answer e

iodine

? Exercise 2.2.2

What element is represented by each chemical symbol?

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

Answer a

lead

Answer b

tin

Answer c

uranium

Answer d

oxygen

Answer e

fluorine

Concept Review Exercises

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

Answers

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

Key Takeaways

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

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2.3: Atomic Number and Mass Number

Learning Objectives

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

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

Atomic Number

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

For example, hydrogen has the atomic number 1; all hydrogen atoms have 1 proton in their nuclei. Helium has the atomic number 2; all helium atoms have 2 protons in their nuclei. There is no such thing as a hydrogen atom with 2 protons in its nucleus; a nucleus with 2 protons would be a helium atom. The atomic number *defines* an element. Table **2.3.1** lists some common elements and their atomic numbers. Based on its **atomic number**, you can determine the **number of protons in the nucleus** of an atom. The largest atoms have over 100 protons in their nuclei.

Table **2.3.1**: Some Common Elements and Their Atomic Numbers

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

Example 2.3.1

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

a. aluminum
b. iron
c. carbon

Answer a

According to Table 2.3.1, aluminum has an atomic number of 13. Therefore, every aluminum atom has 13 protons in its nucleus.

Answer b

Iron has an atomic number of 26. Therefore, every iron atom has 26 protons in its nucleus.

Answer c

Carbon has an atomic number of 6. Therefore, every carbon atom has 6 protons in its nucleus.

Exercise 2.3.1

What is the number of protons in the nucleus of each element? Use Table 2.3.1.

a. sodium
b. oxygen
c. chlorine

Answer a

Sodium has 11 protons in its nucleus.

Answer b

Oxygen has 8 protons in its nucleus.

Answer c

Chlorine has 17 protons in its nucleus

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

Example 2.3.2

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

a. sulfur
b. tungsten
c. argon

Answer a

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

Answer b

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

Answer c

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

Exercise 2.3.2

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

a. magnesium
b. potassium
c. iodine

Answer a

Mg has 12 electrons.

Answer b

K has 19 electrons.

Answer c

I has 53 electrons.

Isotopes

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

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. Figure **2.3.1** compares the three isotopes of hydrogen.

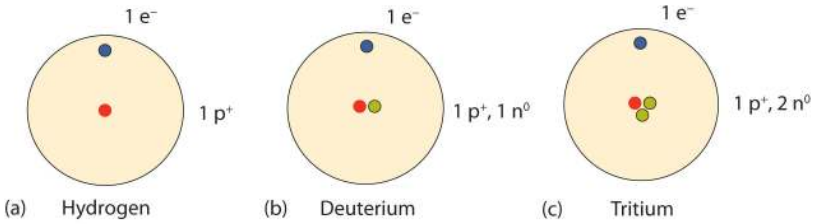


Figure **2.3.1**: Isotopes of Hydrogen. Most hydrogen atoms have only a proton in the nucleus (a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

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

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

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

$$\begin{matrix}\text{mass number} \longrightarrow 56 \\ \text{atomic number} \longrightarrow 26\end{matrix}\text{Fe} \tag{2.3.1}$$

which indicates a particular isotope of iron. The 26 is the atomic number (which is the same for all iron atoms), while the 56 is the mass number of the isotope. To determine the number of neutrons in this isotope, we subtract 26 from 56: 56 − 26 = 30, so there are 30 neutrons in this atom. This way of reprenting the isotope of an element is called the **atomic notation**.

✔ Example 2.3.3

How many protons and neutrons are in each atom?

a. $^{35}_{17}\text{Cl}$
b. $^{127}_{53}\text{I}$

Answer a

In $^{35}_{17}\text{Cl}$ there are 17 protons, and 35 − 17 = 18 neutrons in each nucleus.

Answer b

In $^{127}_{53}\text{I}$ there are 53 protons, and 127 − 53 = 74 neutrons in each nucleus.

? Exercise 2.3.3

How many protons and neutrons are in each atom?

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

Answer a

In $^{197}_{79}\mathrm{Au}$ there are 79 protons, and 197 − 79 = 118 neutrons in each nucleus.

Answer b

In $^{23}_{11}\mathrm{Na}$ there are 11 protons, and 23 − 11 = 12 neutrons in each nucleus.

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

Summary

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

Almost all of the mass of an atom is from the total protons and neutrons contained within a tiny (and therefore very dense) nucleus. The majority of the volume of an atom is the surrounding space in which the electrons reside. A representation of a carbon-12 atom is shown below in Figure **2.3.2**.

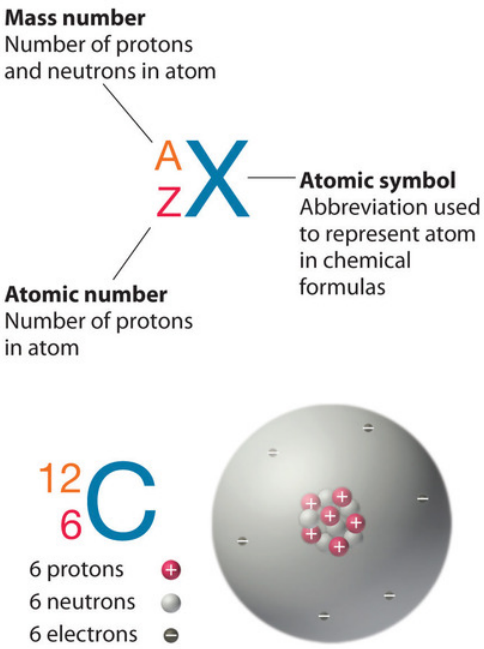


Figure **2.3.2**: Formalism used for identifying specific nuclide (any particular kind of nucleus)

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

Concept Review Exercises

1. Why is the atomic number so important to the identity of an atom?
2. What is the relationship between the number of protons and the number of electrons in an atom?
3. How do isotopes of an element differ from each other?
4. What is the mass number of an element?

Answers

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

Key Takeaways

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

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2.4: Atomic Weight

Learning Objectives

- Define atomic weight and atomic mass unit.

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

Atomic Mass Unit

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

$$1 \text{ amu} = \frac{1}{12} \text{ the mass of } ^{12}\text{C atom} \tag{2.4.1}$$

It is equal to 1.661×10^{-24} g.

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

Atomic Weight is the Weighted Average Mass of Isotopes

As stated above, most elements occur naturally as a mixture of two or more isotopes. Listed below (Table **2.4.1**) are the naturally occurring isotopes of selected elements along with the percent natural abundance of each.

Table **2.4.1**: Atomic Weight and Percent Abundances of Some Natural Isotopes

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic Mass (amu)	Average Atomic Mass (amu)
Hydrogen	$^{1}_{1}H$	99.985	1.0078	1.0079
	$^{2}_{1}H$	0.015	2.0141	
	$^{3}_{1}H$	negligible	3.0160	
Carbon	$^{12}_{6}C$	98.89	12.000	12.011
	$^{13}_{6}C$	1.11	13.003	
	$^{14}_{6}C$	trace	14.003	
Oxygen	$^{16}_{8}O$	99.759	15.995	15.999
	$^{17}_{8}O$	0.037	16.995	
	$^{18}_{8}O$	0.204	17.999	
Chlorine	$^{35}_{17}Cl$	75.77	34.969	35.453
	$^{37}_{17}Cl$	24.23	36.966	
Copper	$^{63}_{29}Cu$	69.17	62.930	63.546
	$^{65}_{29}Cu$	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.

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

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

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

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

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

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

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

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

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

Example 2.4.1

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

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

Solution

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

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

2.4.1 Exercise

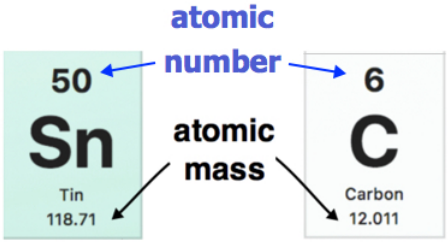
Calculate the atomic weight of copper. Copper exists as a mixture of 2 isotopes. Their respective masses and natural abundance are shown below.

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

Answer

63.546 amu

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



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

Click here to find the Periodic Table.

Concept Review Exercises

1. Define atomic weight. Why is it considered a weighted average?
2. What is an atomic mass unit?

Answers

1. The atomic weight is an average of an element's atomic masses, weighted by the natural abundance of each isotope of that element. It is a weighted average because different isotopes have different masses.
2. An atomic mass unit is 1/12th of the mass of a ^{12}C atom.

Key Takeaway

- Atoms have a mass that is based largely on the number of protons and neutrons in their nucleus.
- The atomic weight of each element in the Periodic Table is the weighted average of the mass of all its isotopes.

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2.5: The Periodic Table

Learning Objectives

- Explain how elements are organized into the periodic table.

The periodic table is one of the cornerstones of chemistry because it organizes all the known elements on the basis of their chemical properties. A modern version is shown in Figure **2.5.1**. Most periodic tables provide additional data (such as atomic weight) in a box that contains each element's symbol. The elements are listed in order of atomic number.

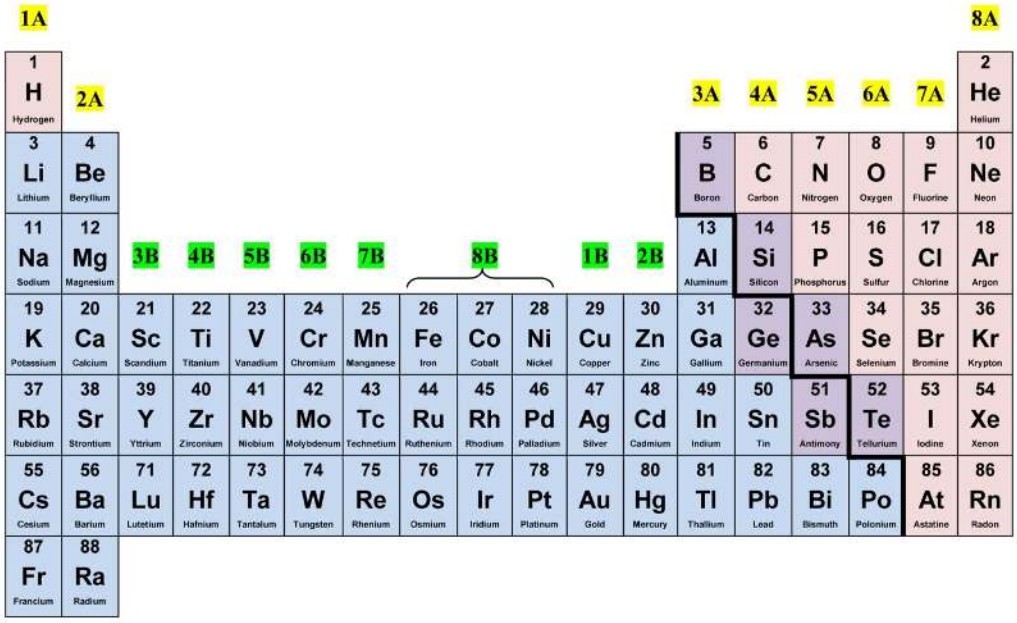


Figure **2.5.1**: The Modern Periodic Table.

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called **groups** (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements-1A), *alkaline earth metals* (the second column of elements-2A), *halogens* (the next-to-last column of elements-7A), and *noble gases* (the last column of elements-8A).

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

Metals, Nonmetals and Metalloids

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a semimetal, as shown in Figure **2.5.2**. A metal is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A nonmetal is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure **2.5.2**, **metals** occupy the left three-fourths of the periodic table, while **nonmetals** (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **semimetals** (or **metalloids**). Elements adjacent to the **bold zigzag** line in the right-hand portion of the periodic table have semimetal properties.

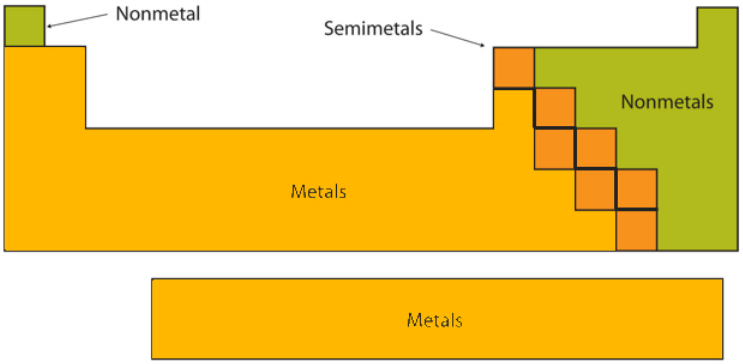


Figure **2.5.2**: Types of Elements. Elements are either metals, nonmetals, or semimetals. Each group is located in a different part of the periodic table.

Example 2.5.1

Based on its position in the periodic table, do you expect selenium (Se) to be a metal, a nonmetal, or a semimetal?

Solution

The atomic number of selenium is 34, which places it in period 4 and group 16. In Figure 2.7.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal. Note, however, that because selenium is close to the metal-nonmetal dividing line, it would not be surprising if selenium were similar to a semimetal in some of its properties.

2.5.1 Exercise

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

Answer

metal

Representative, Transition and Inner-transition

Another way to categorize the elements of the periodic table is shown in Figure **2.5.3**. The first two columns on the left and the last six columns on the right are called the **main group** or **representative** elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the inner transition metals. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

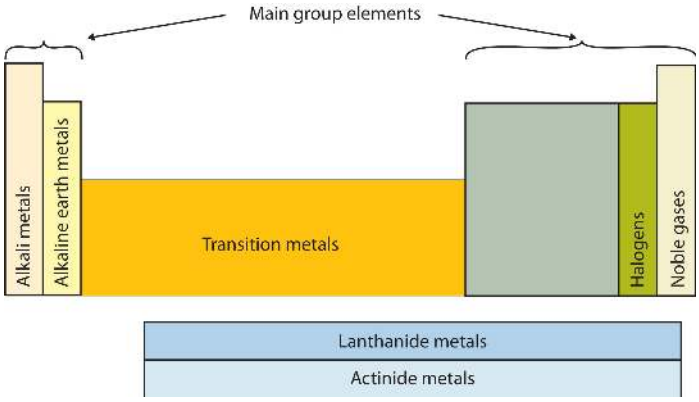


Figure **2.5.3**: Special Names for Sections of the Periodic Table. Some sections of the periodic table have special names. The elements lithium, sodium, potassium, rubidium, cesium, and francium are collectively known as alkali metals.

Alkali metals are the first column. Alkaline earth metals are the second. Halogens are the second to last column. Noble gases are the last column.

To Your Health: Transition Metals in the Body

Most of the elemental composition of the human body consists of main group elements. The first element appearing on the list that is not a main group element is iron, at 0.006 percentage by mass. Because iron has relatively massive atoms, it would appear even lower on a list organized in terms of percent by *atoms* rather than percent by mass.

Iron is a transition metal. Transition metals have interesting chemical properties. The chemistry of iron makes it a key component in the proper functioning of red blood cells.

Red blood cells are cells that transport oxygen from the lungs to cells of the body and then transport carbon dioxide from the cells to the lungs. Without red blood cells, animal respiration as we know it would not exist. The critical part of the red blood cell is a protein called *hemoglobin*. Hemoglobin combines with oxygen and carbon dioxide, transporting these gases from one location to another in the body. Hemoglobin is a relatively large molecule, with a mass of about 65,000 u.

The crucial atom in the hemoglobin protein is iron. Each hemoglobin molecule has four iron atoms, which act as binding sites for oxygen. It is the presence of this particular transition metal in your red blood cells that allows you to use the oxygen you inhale.

Other transition metals have important functions in the body, despite being present in low amounts. Zinc is needed for the body's immune system to function properly, as well as for protein synthesis and tissue and cell growth. Copper is also needed for several proteins to function properly in the body. Manganese is needed for the body to metabolize oxygen properly. Cobalt is a necessary component of vitamin B-12, a vital nutrient. These last three metals are not listed explicitly in Table 2.1.2, so they are present in the body in very small quantities. However, even these small quantities are required for the body to function properly.

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins. For example, the elements of Group 1A are known as the alkali metals, Group 2A are the alkaline earth metals, Group 7A are the halogens, and Group 8A are the noble gases.

Group 1A: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1A, but it is not a metal. The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth's crust, respectively.

Group 2A: The Alkaline Earth Metals

The alkaline earth metals are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 7A: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for "salt forming," which reflects that all the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

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

Group 8A: The Noble Gases

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

To Your Health: Radon

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

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

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

Valence Electrons and Group Number

The number of valence electrons of an element can be determined by the periodic table group (vertical column) in which the element is categorized. With the exception of the transition metals, the units digit of the group number identifies how many valence electrons are associated with a neutral atom of an element listed under that particular column.

Table **2.5.1**. The Group # and the # of Valence Electrons.

Periodic Table Group	Valence Electrons
Group 1A (alkali metals)	1
Group 2 A (alkaline earth metals)	2
Groups 1B-8B (transition metals)	*
Group 3A (boron group)	3
Group 4A (carbon group)	4
Group 5A (nitrogen group)	5
Group 6A (oxygen family)	6
Group 7A (halogens)	7
Group 8A (noble gases)	8**

** The general method for counting valence electrons is generally not useful for transition metals.*

*** Except for helium, which has only two valence electrons.*

Atomic Radius

The periodic table is useful for understanding atomic properties that show periodic trends. One such property is the atomic radius (Figure 2.5.5). The atomic radius is defined as one-half the distance between the nuclei of identical atoms that are bonded together. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be 74 pm. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2} = 37$ pm.

The higher the shell number, the farther from the nucleus the electrons in that shell are likely to be. In other words, the size of an atom is generally determined by the number of the valence electron shell. Therefore, as we go down a column on the periodic table, the atomic radius increases. As we go *across* a period on the periodic table, however, electrons are being added to the *same* valence shell; meanwhile, more protons are being added to the nucleus, so the positive charge of the nucleus is increasing. The increasing positive charge attracts the electrons more strongly, pulling them closer to the nucleus. Consequently, as we go across a period, from left to right, the atomic radius decreases. These trends are seen clearly in Figure 2.5.5

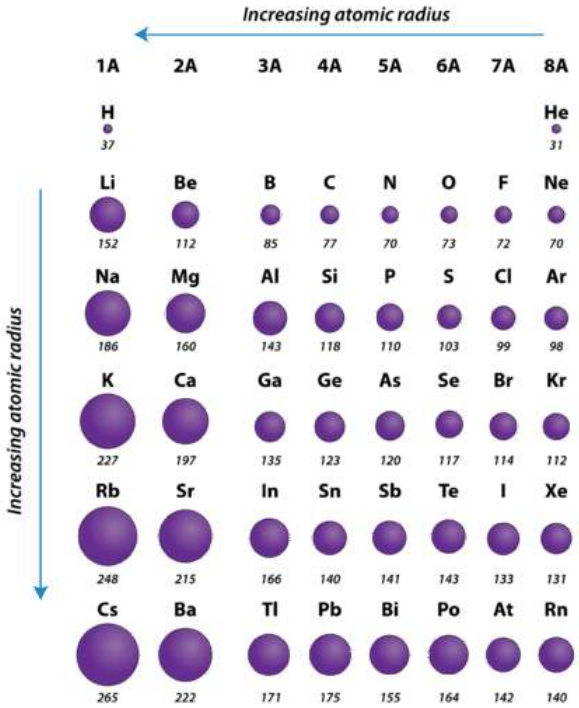


Figure 2.5.5 Trends on the Periodic Table. Atomic radii of the representative elements measured in picometers. The relative sizes of the atoms show several trends with regard to the structure of the periodic table. Atoms become larger going down a group and going from right to left across a period.

Example 2.5.2

Using the periodic table (rather than Figure 2.5.5), which atom is larger?

a. N or Bi
b. Mg or Cl

Answer a

Bi is below N in Group 5A in the periodic table and has electrons in higher-numbered shells, hence we expect that Bi atoms are larger than N atoms.

Answer b

Both Mg and Cl are in period 3 of the periodic table, but Cl lies farther to the right. Therefore we expect Mg atoms to be larger than Cl atoms.

Exercise 2.5.2

Using the periodic table (rather than Figure **2.5.5**), which atom is larger?

a. Li or F
b. Na or K

Answer a

Li and F are on the same period, but F lies farther to the right. Therefore, we expect Li to be larger than F atoms.

Answer b

K lies below Na in Group 1A, hence has more electron shells, making it larger than Na.

Career Focus: Clinical Chemist

Clinical chemistry is the area of chemistry concerned with the analysis of body fluids to determine the health status of the human body. Clinical chemists measure a variety of substances, ranging from simple elements such as sodium and potassium to complex molecules such as proteins and enzymes, in blood, urine, and other body fluids. The absence or presence, or abnormally low or high amounts, of a substance can be a sign of some disease or an indication of health. Many clinical chemists use sophisticated equipment and complex chemical reactions in their work, so they not only need to understand basic chemistry, but also be familiar with special instrumentation and how to interpret test results.

Concept Review Exercises

1. How are the elements organized into the periodic table?
2. Looking at the periodic table, where do the following elements appear?
   a. the metals
   b. the nonmetals
   c. the halogens
   d. the transition metals
   e. the noble gases
3. Describe the trends in atomic radii as related to an element's position on the periodic table.

Answers

1. Elements are organized in order of increasing atomic number.
2. a. the left three-quarters of the periodic table (to the left of the zigzag band)
   b. the right quarter of the periodic table (to the right of the zigzag band)
   c. the next-to-last column of the periodic table
   d. the middle section of the periodic table
   e. the last column of the periodic table
3. As you go across the periodic table, atomic radii decrease; as you go down the periodic table, atomic radii increase.

Key Takeaways

- The chemical elements are arranged in a chart called the periodic table.
- Some characteristics of the elements are related to their position on the periodic table.
- The number of valence electrons of an element can be determined by the group (vertical column) number in the Periodic Table. Elements with the same number of valence electrons have similar chemical properties.

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2.6: Valence Electrons and Electron Dot Formula

Valence Electrons

In any given atom or element there are two types of electrons. There are the core electrons and the valence electrons. Electrons in the outer-most shell, called the valence shell, tend to react (be gained, lost, or shared) during a chemical reaction. You might imagine that, if two atoms bumped into each other, it would be the outer electrons that would interact first.

The number of valence electrons for each main group element can be determined by the column, or group, it occupies on the periodic table.

For example, the elements in the first column (labeled 1A), all have one valence electron.

The second column (2A) has two valence electrons.

We skip the block of ten elements in the middle (transition metals).

The elements in columns 3A, 4A, 5A, 6A, and 7A, and 8A* have three, four, five, six, seven, and eight* valence electrons, respectively.

*Note that helium (He) only has two valence electrons.

Electron Dot Structures of Main Group Elements

Chemists use simple diagrams to show an atom's valence electrons. These diagrams show only valence electrons which are represented as dots. The dots are arranged to the right and left and above and below the symbol, **with no more than two dots on a side**. For example, the representation for sodium atom is as follows:



and the representation for chlorine atom is as follows:



These diagrams are called **Lewis electron dot diagram**s, after Gilbert N. Lewis, the American chemist who introduced them. To write an element's Lewis dot symbol, place the dots representing its valence electrons, one at a time, around the element's chemical symbol.

For the e-dot structures, it does not matter what sides the dots are placed on in Lewis diagrams as long as each side has a maximum of two dots. The following figure shows the electron dot structures of the elements, lithium through neon, which is the entire second period of the periodic table. For the **main group** elements, the **number of valence electrons** is the **same as the group number** listed at the top of the periodic table.



This applies to the main group elements. Groups 1B-8B (transition elements) do not follow the predicted patterns well.

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2.7: The Mole and Mole-Mass Conversions

Learning Objectives

- To define the mole unit.
- To convert quantities between mass units and mole units.

The Mole

We need 2 hydrogen atoms and 1 oxygen atom to make 1 water molecule. If we want to make 2 water molecules, we will need 4 hydrogen atoms and 2 oxygen atoms. If we want to make 5 molecules of water, we need 10 hydrogen atoms and 5 oxygen atoms. The ratio of atoms we will need to make any number of water molecules is the same: 2 hydrogen atoms to 1 oxygen atom as shown in figure 2.7.1.

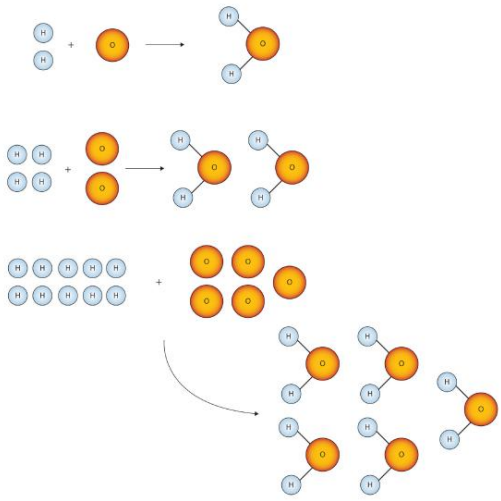


Figure 2.7.1: The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

Chemists use the term **mole** to represent a large number of atoms or molecules. Just as a dozen implies 12 things, a **mole** (abbreviated as ***mol***) represents **6.022×10^{23} things**. The number 6.022×10^{23}, called Avogadro's number after the 19th-century chemist Amedeo Avogadro, is the number we use in chemistry to represent macroscopic amounts of atoms and molecules. Thus, if we have 6.022×10^{23} Na atoms, we say we have 1 mol of Na atoms. If we have 2 mol of Na atoms, we have $2 \times (6.022 \times 10^{23})$ Na atoms, or 1.2044×10^{24} Na atoms. Similarly, if we have 0.5 mol of benzene (C_6H_6) molecules, we have $0.5 \times (6.022 \times 10^{23})$ C_6H_6 molecules, or 3.011×10^{23} C_6H_6 molecules.

> A mole represents a very large number! If 1 mol of quarters were stacked in a column, it could stretch back and forth between Earth and the sun *6.8 billion* times.

Mole to Mass Conversions

With such a large # it is extremely difficult, if not impossible, to organize atoms one at a time. We deal with billions of atoms at a time. **How can we keep track of so many atoms (and molecules) at a time? We do it by using mass rather than by counting individual atoms.** We use atomic weight in grams (molar mass) as a conversion factor in a mole-mass conversion (or its reverse, a mass-mole conversion).

1 mol of Al has a mass of 26.98 g (Example **2.7.1**). Stated mathematically,

$$1 \text{ mol Al} = 26.98 \text{ g Al}$$

We can divide both sides of this expression by either side to get one of two possible conversion factors:

$$\frac{1\ \text{mol Al}}{26.98\ \text{g Al}} \quad \text{and} \quad \frac{26.98\ \text{g Al}}{1\ \text{mol Al}}$$

The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass.

Example 2.7.1

What is the mass of 3.987 mol of Al?

Solution

The first step in a conversion problem is to decide what conversion factor to use. Because we are starting with mole units, we want a conversion factor that will cancel the mole unit and introduce the unit for mass in the numerator. Therefore, we should use the $\frac{26.98\ \text{g Al}}{1\ \text{mol Al}}$ conversion factor. We start with the given quantity and multiply by the conversion factor:



$$3.987\ \text{mol Al} \times \frac{26.98\ \text{g Al}}{1\ \text{mol Al}}$$

Note that the mol units cancel algebraically. (The quantity 3.987 mol is understood to be in the numerator of a fraction that has 1 in the unwritten denominator.) Canceling and solving gives

$$3.987\ \text{mol Al} \times \frac{26.98\ \text{g Al}}{1\ \text{mol Al}} = 107.6\ \text{g Al}$$

Our final answer is expressed to four significant figures.

Exercise 2.7.1

How many moles are present in 100.0 g of Al? (Hint: you will have to use the other conversion factor we obtained for aluminum.)

Answer

$$100.0\ \text{g Al} \times \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} = 3.706\ \text{mol Al}$$

To Your Health: Minerals

For our bodies to function properly, we need to ingest certain substances from our diets. Among our dietary needs are minerals, the noncarbon elements our body uses for a variety of functions, such developing bone or ensuring proper nerve transmission. The US Department of Agriculture has established some recommendations for the RDIs of various minerals. The accompanying table lists the RDIs for minerals, both in mass and moles, assuming a 2,000-calorie daily diet.

Table **2.7.1**: Essential Minerals and their Composition in Humans

Mineral	Male (age 19–30 y)		Female (age 19–30 y)	
Ca	1,000 mg	0.025 mol	1,000 mg	0.025 mol
Cr	35 µg	6.7×10^{-7} mol	25 µg	4.8×10^{-7} mol
Cu	900 µg	1.4×10^{-5} mol	900 µg	1.4×10^{-5} mol
F	4 mg	2.1×10^{-4} mol	3 mg	1.5×10^{-4} mol
I	150 µg	1.2×10^{-6} mol	150 µg	1.2×10^{-6} mol
Fe	8 mg	1.4×10^{-4} mol	18 mg	3.2×10^{-4} mol
K	3,500 mg	9.0×10^{-2} mol	3,500 mg	9.0×10^{-2} mol
Mg	400 mg	1.6×10^{-2} mol	310 mg	1.3×10^{-2} mol
Mn	2.3 mg	4.2×10^{-5} mol	1.8 mg	3.3×10^{-5} mol
Mo	45 mg	4.7×10^{-7} mol	45 mg	4.7×10^{-7} mol
Na	2,400 mg	1.0×10^{-1} mol	2,400 mg	1.0×10^{-1} mol
P	700 mg	2.3×10^{-2} mol	700 mg	2.3×10^{-2} mol
Se	55 µg	7.0×10^{-7} mol	55 µg	7.0×10^{-7} mol
Zn	11 mg	1.7×10^{-4} mol	8 mg	1.2×10^{-4} mol

Table **2.7.1** illustrates several things. First, the needs of men and women for some minerals are different. The extreme case is for iron; women need over twice as much as men do. In all other cases where there is a different RDI, men need more than women.

Second, the amounts of the various minerals needed on a daily basis vary widely—both on a mass scale and a molar scale. The average person needs 0.1 mol of Na a day, which is about 2.5 g. On the other hand, a person needs only about 25–35 µg of Cr per day, which is under one millionth of a mole. As small as this amount is, a deficiency of chromium in the diet can lead to diabetes-like symptoms or neurological problems, especially in the extremities (hands and feet). For some minerals, the body does not require much to keep itself operating properly.

Although a properly balanced diet will provide all the necessary minerals, some people take dietary supplements. However, too much of a good thing, even minerals, is not good. Exposure to too much chromium, for example, causes a skin irritation, and certain forms of chromium are known to cause cancer (as presented in the movie *Erin Brockovich*).

Key Takeaway

- A mole is 6.022×10^{23} things.
- It is possible to convert between moles of material and mass of material.

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2.8: Physical and Chemical Properties and Changes

Learning Objectives

- To separate physical from chemical properties.
- Label a change as chemical or physical.
- List evidence that can indicate a chemical change occurred.

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

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

Physical Property

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

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 g/cm^3 at 25°C. The average densities of some common substances are in Table **2.8.1**. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table **2.8.1**: Densities of Common Substances

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

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



Figure **2.8.1**: Pencil and Diamond. Both are a form of carbon, but exhibit very different physical properties.

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

Physical Change

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

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

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



Figure 2.8.2: Ice Melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical

change. When salt is dissolved into water, the salt is said to have entered the aqueous solution. The salt may be regained by boiling off the water, leaving the salt behind.

Chemical Properties

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



Figure **2.8.3**: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure **2.8.2**). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table **2.8.2**: Contrasting Physical and Chemistry Properties

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

Chemical Change

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

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



Figure 2.8.4: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

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

Observations that help to indicate chemical change include:

- Release of Energy (light, heat, or sound)
    - Temperature changes (either the temperature increases or decreases)
    - Light is given off
- Permanent color changes (a substance with a different color is made, rather than just mixing the original colors together)
- Release of a gas
    - Bubbles are formed (but the substance is not boiling—you made a substance that is a gas at the temperature of the beginning materials, instead of a liquid)
    - The gas released has a different smell.
- A solid forms if two clear liquids are mixed (look for *floaties/haziness*—technically called a precipitate)

Example 2.8.1

Which of the following is a chemical property of iron?

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

Solution

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

Exercise 2.8.1

Which of the following is a physical property of matter?

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

Answer

c

Exercise 2.8.2

Which of the following is a chemical property?

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

Answer

a

Example 2.8.2

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

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

Solution:

1. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written: $H_2O(l) \rightarrow H_2O(g)$
2. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $Fe + O_2 \rightarrow Fe_2O_3$
3. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
4. Chemical: the formation of a precipitate and the color change from colorless to yellow indicate a chemical change.

Exercise 2.8.3

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

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

Answer a:

physical change

Answer b:

chemical change

Answer c:

physical change

Answer d:

chemical change

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Summary

A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points. Physical changes may happen during an observation of a physical property. A physical change does not alter the identity of a substance. To identify a chemical property, we look for a chemical change. Chemical changes are changes that occur when one substance is turned into another substance. Observations that indicate a chemical change has occurred include color change, energy change, evolution of a gas, formation of a precipitate, etc.

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2.9: End of Chapter Problems

Atoms

1. What are the charges and the relative masses of the three subatomic particles?

2. Describe the structure of an atom in terms of its protons, neutrons, and electrons.

Elements

1. What is an element? Why are chemical symbols so useful? What is the source of the letter(s) for a chemical symbol?

2. Write the chemical symbol for the following elements.

bromine

carbon

calcium

gold

manganese

magnesium

neon

nitrogen

silver

3. What element is represented by each chemical symbol?

Na

Hg

P

K

F

I

Pb

Sn

O

Atomic Number, Mass Number, Atomic Symbol, and Isotopes

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

aluminum

iron

carbon

sodium

oxygen

chlorine

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

sulfur

iron

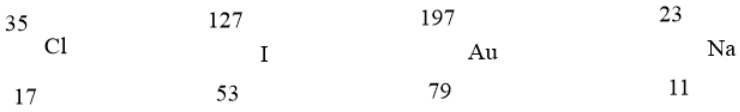
argon

magnesium

potassium

iodine

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



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

5. What is the relationship between the number of protons and the number of electrons in an atom?

6. How do isotopes of an element differ from each other?

7. What is the mass number of an element?

8. Which are isotopes? **Explain**.



9. Complete the following table.

Number of Protons	Number of Neutrons	Element Name	Isotope Symbol
			$^{3}_{2}He$
95	153		
	21	potassium	

Atomic Weight

1. In nature the element oxygen exists as three different isotopes, O-16, O-17, and O-18. Which isotope predominates? Explain.

2. Copper exists as a mixture of 2 isotopes. ^{63}Cu and ^{65}Cu. Which isotope predominates?

3. Define atomic mass. Why is it considered a weighted average?

4. a. What is an atomic mass unit?

b. What is the atomic weight of zinc in atomic mass units?

c. What is the atomic weight of barium in atomic mass units?

Periodic Table

5. How are the elements organized into the periodic table?

6. Looking at the periodic table, where do the following elements appear?

the metals

the nonmetals

the halogens

the transition metals

the noble gases

7. Describe the trends in atomic radii as related to an element's position on the periodic table.

8. Using the periodic table, which atom is larger? Explain.

N or Bi

Mg or Cl

Li or F

Na or K

9. a. A potassium atom has ____ core electrons and ____ valence electrons.

b. A silicon atom has ____ core electrons and ____ valence electrons.

c. How many total and valence electrons are in a neutral phosphorus atom?

d. How many total, valence, and core electrons are there in a neutral xenon atom?

Mole-Mass Conversions

1. How many moles are present in 100.0 g of Al?

2. How many moles are present in 25.5 g Na?

3. What is the mass of 8.603 mol of Fe metal?

4. What is the mass of 0.552 mol of Ag metal?

Physical & Chemical Properties

1. Which of the following is a chemical property of iron?

Iron corrodes in moist air

Density = 7.874 g/cm^3

Iron is soft

pure iron melts at 1808 K

2. Which of the following is a physical property of matter?

corrosiveness

pH (acidity)

density

flammability

3. Which of the following is a chemical property?

flammability

melting point

boiling point

density

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

a. Boiling water

b. A nail rusting

c. A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.

d. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

e. A mirror is broken.

f. An iron nail corroded in moist air

g. Copper metal is melted.

h. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

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CHAPTER OVERVIEW

3: Compounds

3.1: Ions - Losing and Gaining Electrons
3.2: Ions-Main Group Elements
3.3: Transition Elements and Ionic Compounds
3.4: Polyatomic Ions and Formulae for Ionic Compounds
3.5: Characteristics of Ionic Compounds-Electrolytes
3.6: Covalent Bonds
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3.8: Binary Molecules - Formulae and Names
3.9: Formula Weight
3.10: Molecular Weight
3.11: Mole-Mass Conversions
3.12: End of Chapter Problems

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3.1: Ions - Losing and Gaining Electrons

Learning Objectives

- Define the two types of ions.

Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge as a result because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called cations. Most metals become cations when they make ionic compounds.

Cations

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

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + \mathrm{e^-} \tag{3.1.1}$$

The cation produced in this way, Na^+, is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure **3.1.1** is a graphical depiction of this process.

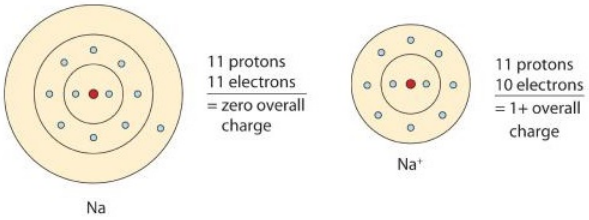


Figure **3.1.1**: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

Anions

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called anions. Most nonmetals become anions when they make ionic compounds.

A neutral chlorine atom has seven electrons in its outermost shell. Only one more electron is needed to achieve an octet in chlorine's valence shell. (In table salt, this electron comes from the sodium atom.)

$$\mathrm{e^-} + \mathrm{Cl} \longrightarrow \mathrm{Cl^-} \tag{3.1.2}$$

In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^-, is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure **3.1.2** is a graphical depiction of this process.

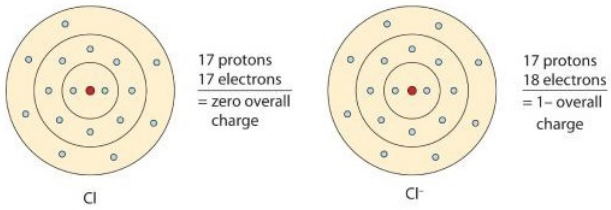


Figure **3.1.2**: The Formation of a Chloride Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge.

> *The names for positive and negative ions are pronounced CAT-eye-ons and ANN-eye-ons, respectively.*

Ionophores

Cell membranes surround each living cell. It is a barrier that prevents unwanted substances from entering and desired substances from leaving the cell. The health of cells depends on maintaining the proper levels of ions in intracellular fluids. Any change that affects the normal flow of ions across cell membranes could well cause an organism to die. Molecules that facilitate the transport of metal ions across membranes are called **ionophores** (ion plus phore from the Greek phorein, meaning "to carry"). Streptomyces bacteria make valinomycin an ionophore to defend themselves against competing bacteria. Valinomycin is a potent antibiotic that can kill or inhibit the growth of bacteria. It is a cyclic molecule with a central cavity lined with oxygen atoms to bind to metal cations. Valinomycin is highly selective: its affinity for K^+ is about 1000 times greater than that for Na^+. By increasing the flow of K^+ ions into cells, valinomycin disrupts the normal K^+ levels in the cell, thereby killing the cell.

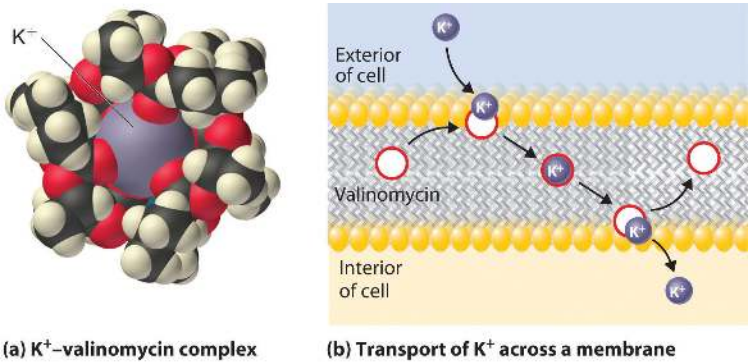


Figure **3.1.3**: Valinomycin Is an Antibiotic That Functions Like an Ionophore

Contributions & Attributions

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3.2: Ions-Main Group Elements

Learning Objectives

- Use electron dot diagrams to illustrate ion formation.

There is an electrostatic attraction between opposite charged ions such as sodium cation and chloride ion. The resulting combination is the compound sodium chloride. The number of electrons lost by the sodium atom (one) equals the number of electrons gained by the chlorine atom (one), so the compound is electrically neutral. In macroscopic samples of sodium chloride, there are billions and billions of sodium and chloride ions, although there is always the same number of cations and anions.

Ions of the Main Group Elements

Elements that belong to the same main group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions.

For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge.

Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge.

On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1− charge.

The figure 3.2.1 below shows how the charge on many ions can be predicted by the location of an element on the periodic table.

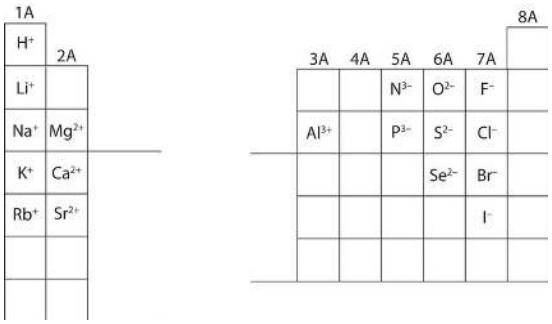


Figure **3.2.1**: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the position in the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

Example 3.2.2

Which of these ions is *not likely* to form?

a. Mg^+
b. K^+

Solution

(a) Mg is in Group 2A and has two valence electrons. It achieves octet by losing two electrons to form Mg^{2+} cation. Losing only one electron to form Mg^+ does not make an octet, hence, Mg^+ is not likely to form.

3.2.2 Exercise

Which of these ions is *not likely* to form?

a. S^{3-}
b. N^{3-}

Answer

(a) S is in Group 6A and has six valence electrons. It achieves octet by gaining two electrons to form S^{2-} anion. Gaining three electrons to form S^{3-} does not make it octet, hence, S^{3-} is not likely to form.

Cations of the Main Group Elements

The table 3.2.1 below lists some common ions of the main group metals. Monoatomic cations formed from the main group are given the same name as the original element. So Na^+ is sodium ion and Ca^{2+} is calcium ion. While most main group element form only one cation, elements of the group 4A form two different cations by losing 2 or 4 valence electrons. While naming these cations of lead and tin, a Roman numeral (in parentheses) is used to indicate the charge. Sn^{2+} is Tin(II) ion and Sn^{4+} is Tin(IV) ion.

Table **3.2.1**: Some Monatomic Cations of the Main Group Elements

Some Monatomic Cations of the Main Group Elements

Group #	Cation formula	Name of cation
1A	Li^{+}	Lithium ion
1A	Na^{+}	Sodium ion
1A	K^{+}	Potassium ion
2A	Mg^{+2}	Magnesium ion
2A	Ca^{+2}	Calcium ion
2A	Sr^{+2}	Strontium ion
2A	Ba^{+2}	Barium ion
3A	Al^{+3}	Aluminum ion
4A	Sn^{2+}	tin(II) ion or stannous ion
4A	Sn^{4+}	tin(IV) ion or stannic ion
4A	Pb^{2+}	lead(II) ion or plumbous ion
4A	Pb^{4+}	lead(IV) ion or plumbic ion

Anions of the Main Group Elements

The table 3.2.2 below lists some common anions of the main group nonmetals.The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, Cl^{-} is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, N^{3-} is the nitride ion, and so forth.

Table **3.2.2**: Some Monatomic Anions of the Main Group Elements

Some Monatomic Cations of the Main Group Elements

Group #	Anion formula	Name
7A	$\mathbf{F^{-}}$	fluoride ion
7A	$\mathbf{Cl^{-}}$	chloride ion
7A	$\mathbf{Br^{-}}$	bromide ion
7A	$\mathbf{I^{-}}$	iodide ion
6A	$\mathbf{O^{2-}}$	oxide ion
6A	$\mathbf{S^{2-}}$	sulfide ion
5A	$\mathbf{N^{3-}}$	nitride ion
5A	$\mathbf{P^{3-}}$	phosphide ion

The enamel of your teeth is composed of an ionic compound called calcium hydroxyapatite, $Ca_5(OH)(PO_4)_3$. Acids produced by bacteria present in dental plaque breaks down the enamel. As a result tooth decay happens. Fluoride ion (F^{-}) is used to prevent tooth decay since the anion converts hydroxy apatite to fluoroapatite, $Ca_5F(PO_4)_3$. Fluoroapatite is more durable and hence is not broken down easily by acids.

Electron Dot Diagrams (Lewis Diagrams)

Chemists use simple diagrams to show an atom's valence electrons and how they transfer. They show only valence electrons, each dot represents a valence electron. These dots are arranged to the right and left and above and below the symbol, **with no more than two dots on a side**. For example, the representation for sodium is as follows:



and the representation for chlorine is as follows:



> For the above diagrams, it does not matter what sides the dots are placed on in electron dot diagrams as long as each side has a maximum of two dots.

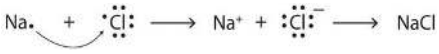


Figure **3.2.2**: The transfer of electrons can be illustrated with Lewis diagrams.

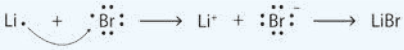
In representing the final formula, the dots are omitted.

Example 3.2.3

Starting with lithium and bromine atoms, use electron dot diagrams to show the formation of the ionic compound LiBr.

Solution

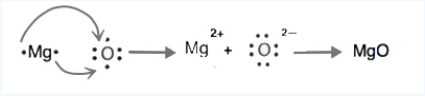
From the periodic table, we see that lithium is in the same column as sodium, so it will have the same valence shell electron configuration. That means that the neutral lithium atom will have the same Lewis diagram that the sodium atom has. Similarly, bromine is in the same column as chlorine, so it will have the same Lewis diagram that chlorine has. Therefore,



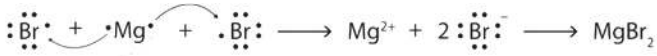
Exercise 3.2.3

Starting with magnesium and oxygen atoms, use electron dot diagrams to show the formation of the ionic compound MgO.

Answer



Some ionic compounds have different numbers of cations and anions. In those cases, electron transfer occurs between more than one atom. For example, here is the formation of $MgBr_2$:



Notice that in this example there are two bromide ions (1− charge) needed for every one magnesium ion (2+ charge) in order for the overall charge of the compound to equal zero. This is called **charge balance**. The number of each type of ion is indicated in the formula by the subscript.

Most of the elements that make ionic compounds form an ion that has a characteristic charge. For example, sodium makes ionic compounds in which the sodium ion always has a 1+ charge. Chlorine makes ionic compounds in which the chloride ion always has a 1− charge. Some elements, especially transition metals, can form ions of multiple charges.

Key Takeaways

- Ions can be positively charged or negatively charged.
- A Lewis diagram is used to show how electrons are transferred to make ions and ionic compounds.

Exercises

1. What are the two types of ions?
2. Use Lewis diagrams to illustrate the formation of an ionic compound from a potassium atom and an iodine atom.

3. When the following atoms become ions, what charges do they acquire?

a. Li
b. S
c. Ca
d. F

4. Identify each as a cation, an anion, or neither.

a. H^+
b. Cl^-
c. O_2
d. Ba^{2+}
e. CH_4
f. CS_2

5. Identify each as a cation, an anion, or neither.

a. NH_3
b. Br^-
c. H^-
d. Hg^{2+}

e. CCl_4
f. SO_3

6. Using Lewis diagrams, show the electron transfer for the formation of LiF.

7. Using Lewis diagrams, show the electron transfer for the formation of MgO.

8. Using Lewis diagrams, show the electron transfer for the formation of Li_2O.

9. Using Lewis diagrams, show the electron transfer for the formation of CaF_2.

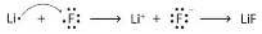
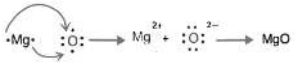
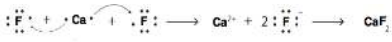
10. What characteristic charge do atoms in the first column of the periodic table have when they become ions?

11. What characteristic charge do atoms in the second column of the periodic table have when they become ions?

12. What characteristic charge do atoms in the third-to-last column of the periodic table have when they become ions?

13. What characteristic charge do atoms in the next-to-last column of the periodic table have when they become ions?

Answers

1. Cations (positive charged) and anions (negative charged).
2. 
3. a. 1+
   b. 2−
   c. 2+
   d. 1−
4. a. cation
   b. anion
   c. neither
   d. cation
   e. neither
   f. neither
5. a. neither
   b. anion
   c. anion
   d. cation
   e. neither
   f. neither
6. 
7. 
8. 
9. 
10. 1+
11. 2+
12. 2−
13. 1−

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3.3: Transition Elements and Ionic Compounds

Learning Objectives

- Name and write formula of the ions of the transition elements
- Name and write the formulas of simple ionic compounds from monoatomic ions

After learning a few more details about the names of individual ions, you will be a step away from knowing how to write the formula of ionic compounds and name them. This section begins the formal study of writing formula and nomenclature (the systematic naming of ionic compounds).

Naming Ions

The names of many monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

However, most transitional elements, tin, and lead lose different numbers of electrons, producing ions of different charges. Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}.

There are two ways to make this distinction. In the simpler, more modern approach, called the ***Stock system***, an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word ion. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name sodium ion.

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. This system does not work when there are more than two cations for a metal. We simply run out of names. The element manganese is an example. See table **3.3.1** for the list.

The element zinc in group 2B is an exception to the transitional elements. The metal loses two valence electrons to form the zinc ion (Zn^{2+}). The most common ion for the element silver in group 1B is the silver ion (Ag^+).

Table **3.3.1**: The Cation Names of Transitional Elements, Tin. and Lead

Element	Stem/Root	Charge	Name
iron	ferr-	2+	ferrous ion or iron(II) ion
		3+	ferric ion or iron(III) ion
copper	cupr-	1+	cuprous ion or copper(I) ion
		2+	cupric ion or copper(II) ion
tin	stann-	2+	stannous ion or tin(II) ion
		4+	stannic ion or tin(IV) ion
lead	plumb-	2+	plumbous ion or lead(II) ion
		4+	plumbic ion or lead(IV) ion
chromium	chrom-	2+	chromous ion or chromium(II) ion
		3+	chromic ion or chromium(III) ion
cobalt	colbalt-	2+	colbaltous ion or colbalt(II) ion
		3+	Cobaltic ion or cobalt(III) ion
manganese	-	2+	manganese(II) ion
		3+	manganese(III) ion
		4+	manganese(IV) ion
nickel	-	2+	nickel(II) ion
		3+	nickel(III) ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table **3.3.2** lists the names of some common monatomic ions.

Table **3.3.2**: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The polyatomic ions have their own characteristic names. They will be discussed in the next section.

Example 3.3.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. Br^{-}
d. K^{+}
e. Cu^{+}

Answer a

the calcium ion

Answer b

the sulfide ion (from Table 3.3.2)

Answer c

the bromide ion

Answer d

the potassium ion

Answer e

the copper(I) ion or the cuprous ion (copper can form cations with either a 1+ or 2+ charge, so we have to specify which charge this ion has

Exercise 3.3.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. N^{3-}
d. Ba^{2+}
e. F^{-}

Answer a

the iron (II) or ferrous ion

Answer b

the iron (III) or ferric ion

Answer c

the nitride ion

Answer d

the barium ion

Answer e

the fluoride ion

Example 3.3.2

Write the formula for each ion.

a. the bromide ion
b. the phosphide ion
c. the cupric ion
d. the magnesium ion

Answer a

Br^-

Answer b

P^{3-}

Answer c

Cu^{2+}

Answer d

Mg^{2+}

? Exercise 3.3.2

Write the formula for each ion.

a. the fluoride ion
b. the oxide ion
c. the ferrous ion
d. the potassium ion

Answer a

F^-

Answer b

O^{2-}

Answer c

Fe^{2+}

Answer d

K^+

Formulae and Names for Ionic Compounds from Monoatomic Ions

The formula for an ionic compound follows several conventions. First, the **cation** is written **before** the **anion**. Because most metals form cations and most nonmetals form anions, formulas typically list the metal first and then the nonmetal. Second, **charges** are **not written** in a formula. Remember that in an ionic compound, the components are ions, not neutral atoms, even though the formula does not contain charges. Finally, the proper formula for an ionic compound always has ***a net zero charge,*** meaning the total positive charge must equal the total negative charge. To determine the proper formula of any combination of ions, determine how many of each ion is needed to balance the total positive and negative charges in the compound.

> This rule is ultimately based on the fact that matter is, overall, electrically neutral.

If we look at the ionic compound consisting of lithium ions and bromide ions, we see that the lithium ion has a 1+ charge and the bromide ion has a 1− charge. Only one ion of each is needed to balance these charges. The formula for **lithium bromide** is $\mathbf{LiBr}$. By convention, assume that there is only one atom if a subscript is not present. We do not use 1 as a subscript.

When an ionic compound is formed from magnesium ion and oxide ion, the magnesium ion has a 2+ charge, and the oxide has a 2− charge. Although both of these ions have higher charges than the ions in lithium bromide, they still balance each other in a one-to-one ratio. Therefore, the proper formula for this **magnesium oxide** is $\mathbf{MgO}$.

Now consider the ionic compound formed by magnesium ion and chloride ion. A magnesium ion has a 2+ charge, while a chloride ion has a 1− charge:

$$\mathbf{Mg^{2+}Cl^-}$$

Combining one ion of each does not completely balance the positive and negative charges. The easiest way to balance these charges is to assume the presence of *two* chloride ions for each magnesium ion:

Mg^{2+} Cl^- Cl^-

Now the positive and negative charges are balanced, the convention is to use a numerical subscript when there is more than one ion of a given type: so $MgCl_2$. This chemical formula says that there are one magnesium ion and *two* chloride ions in the formula for **magnesium chloride**. (Do not read the "Cl_2" part of the formula as a molecule of the diatomic elemental chlorine. (Chlorine does not exist as a diatomic element in this compound. Rather, it exists as two individual chloride ions.) By convention, the **lowest whole number ratio** is used in the formulas of ionic compounds. The formula Mg_2Cl_4 has balanced charges with the ions in a 1:2 ratio, but it is not the lowest whole number ratio.

> By convention, the lowest whole-number ratio of the ions is used in ionic formulas. There are exceptions for certain ions, such as Hg_2^{2+} known as the mercurous ion.

For compounds in which the ratio of ions is not as obvious, the subscripts in the formula can be obtained by **crossing charges**: use the absolute value of the charge on one ion as the subscript for the other ion. This method is shown schematically in Figure **3.3.1**.

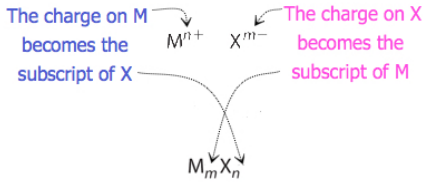


Figure **3.3.1*A***: Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges. The image shows M with a charge of n+ and X with a charge of m-. The charge on the M becomes the subscript of X, and the charge on X becomes the subscript of M, making the final product M subscript m X subscript n.

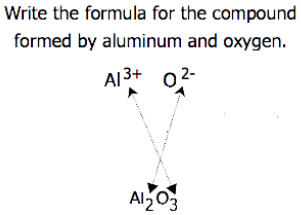


Figure **3.3.1*B***: Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges. The image asks to write the formula for the compound formed by aluminum ion and oxide ion. It shows Al with a charge of 3+ and O with a charge of 2-. By crossing charges, the final result is Al subscript 2, O subscript 3 for **aluminum oxide**.

When crossing charges, it is sometimes necessary to reduce the subscripts to their simplest ratio to write the formula. Consider, for example, the compound formed by Pb^{4+} and O^{2-}. Using the absolute values of the charges on the ions as subscripts gives the formula Pb_2O_4. This simplifies to its correct formula PbO_2 for **lead(IV) oxide.** The formula has one Pb^{4+} ion and two O^{2-} ions.

✔ Example 3.3.1

Write the chemical formula and name for an ionic compound composed of each pair of ions.

a. the sodium ion and the sulfide ion
b. the aluminum ion and the fluoride ion
c. the iron(III) ion and the oxide ion

Solution

a. To obtain an octet, sodium forms an ion with a 1+ charge, while the sulfide ion has a 2− charge. Two sodium 1+ ions are needed to balance the 2− charge on the sulfide ion. Rather than writing the formula as $NaNaS$, we shorten it by convention to Na_2S. The name of the compound is sodium sulfide.
b. The aluminum ion has a 3+ charge, while the fluoride ion formed by fluorine has a 1− charge. Three fluoride ions are needed to balance the 3+ charge on the aluminum ion. This combination is written as AlF_3. The name of the compound is aluminum fluoride.
c. Iron can form two possible ions, but the ion with a 3+ charge is specified here. The oxide has a 2− charge as an ion. To balance the positive and negative charges, we look to the least common multiple (6): two iron 3+ ions will give 6+, while three oxide ions will give 6−, thereby balancing the overall positive and negative charges. Thus, the formula for this ionic compound is Fe_2O_3. Alternatively, use the crossing charges method shown in Figure 3.3.1.

Exercise 3.3.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxide ion
b. the copper(II) ion and the sulfide ion
c. the copper(I) ion and the sulfide ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

Example 3.3.2

Name each ionic compound, using both Stock and common systems if necessary.

a. Ca_3P_2
b. Fe_3N_2
c. KCl
d. CuCl
e. SnF_2

Answer a

calcium phosphide

Answer b

iron(II) nitride or ferrous nitride

Answer c

potassium chloride

Answer d

copper(I) chloride or cuprous chloride

Answer e

tin(II) fluoride or stannous fluoride

Exercise 3.3.2

Name each ionic compound, using both Stock and common systems if necessary.

a. $ZnBr_2$
b. FeN
c. Al_2O_3
d. CuF_2
e. AgF

Answer a

zinc bromide

Answer b

iron (III) nitride or ferric nitride

Answer c

aluminum oxide

Answer d

copper (II) fluoride or cupric fluoride

Answer e

silver fluoride

KEY TAKEAWAY

- Each ionic compound has its own unique name and formula that comes from the names and formulae of the ions.

EXERCISES

1. Briefly describe the process for naming an ionic compound.

2. In what order do the names of ions appear in the names of ionic compounds?

3. Which ionic compounds can be named using two different systems? Give an example.

4. Name each ion.

   a. Ba^{2+}
   b. P^{3-}
   c. S^{2-}
   d. Sn^{4+}

5. Name each ion.

   a. Cs^{+}
   b. Al^{3+}
   c. I^{-}
   d. Sn^{2+}

6. Name the ionic compound formed by each pair of ions.

   a. Na^{+} and Br^{-}
   b. Mg^{2+} and Br^{-}
   c. Mg^{2+} and S^{2-}

7. Name the ionic compound formed by each pair of ions.

   a. K^{+} and Cl^{-}
   b. Mg^{2+} and Cl^{-}
   c. Mg^{2+} and S^{2-}

8. Name the ionic compound formed by each pair of ions.

   a. Na^{+} and N^{3-}
   b. Mg^{2+} and N^{3-}
   c. Al^{3+} and S^{2-}

9. Name the ionic compound formed by each pair of ions.

   a. Li^{+} and N^{3-}
   b. Mg^{2+} and P^{3-}
   c. Li^{+} and P^{3-}

10. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

    a. Fe^{3+} and Br^{-}
    b. Fe^{2+} and Br^{-}
    c. Au^{3+} and S^{2-}
    d. Au^{+} and S^{2-}

11. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

    a. Cr^{3+} and O^{2-}
    b. Cr^{2+} and O^{2-}

c. Pb^{2+} and Cl^-
d. Pb^{4+} and Cl^-

Answers

1. Name the cation and then the anion but don't use numerical prefixes.
2. the cation name followed by the anion name
3. Ionic compounds in which the cation can have more than one possible charge have two naming systems. $FeCl_3$ is either iron(III) chloride or ferric chloride (answers will vary).

4.
   a. the barium ion
   b. the phosphide ion
   c. the sulfide ion
   d. the tin(IV) ion or the stannic ion

5.
   a. the cesium ion
   b. the aluminum ion
   c. the iodide ion
   d. the tin(II) ion or the stannous ion

6.
   a. sodium bromide
   b. magnesium bromide
   c. magnesium sulfide

7.
   a. potassium chloride
   b. magnesium chloride
   c. magnesium sulfide

8.
   a. sodium nitride
   b. magnesium nitride
   c. aluminum sulfide

9.
   a. lithium nitride
   b. magnesium phosphide
   c. lithium phosphide

10.
   a. iron(III) bromide or ferric bromide
   b. iron(II) bromide or ferrous bromide
   c. gold(III) sulfide or auric sulfide
   d. gold(I) sulfide or aurous sulfide

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3.4: Polyatomic Ions and Formulae for Ionic Compounds

Learning Objectives

- Recognize polyatomic ions in chemical formulas.
- Write the chemical formula for an ionic compound and name them.

We have already encountered some chemical formulas for ionic compounds from monoatomic ions. A chemical formula is a concise list of the elements in a compound and the ratios of these elements. To better understand what a chemical formula means, we must consider how an ionic compound is constructed from its ions.

Ionic compounds exist as alternating positive and negative ions in regular, three-dimensional arrays called crystals (Figure **3.4.1**). As you can see, there are no individual **NaCl** "particles" in the array; instead, there is a continuous lattice of alternating sodium and chloride ions. However, we can use the ratio of sodium ions to chloride ions, expressed in the lowest possible whole numbers, as a way of describing the compound. In the case of sodium chloride, the ratio of sodium ions to chloride ions, expressed in lowest whole numbers, is 1:1, so we use **NaCl** (one **Na** symbol and one **Cl** symbol) to represent the compound. Thus, **NaCl** is the chemical formula for sodium chloride, which is a concise way of describing the relative number of different ions in the compound. A macroscopic sample is composed of myriads of NaCl pairs; each individual pair called a **formula unit**. Although it is convenient to think that **NaCl** crystals are composed of individual **NaCl** units, Figure **3.4.1** shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by ions of opposite charge.

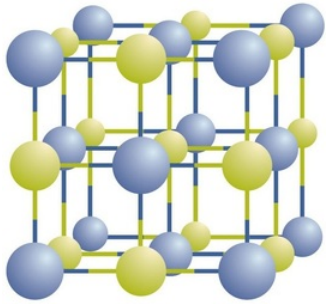


Figure **3.4.1**: A Sodium Chloride Crystal. A crystal contains a three-dimensional array of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions.

Polyatomic Ions

Some ions consist of groups of atoms ***covalently*** bonded together and have an **overall electric charge**. Because these ions contain more than one atom, they are called polyatomic ions. The Lewis structures, names and formulas of some polyatomic ions are found in Figure 3.4.1.

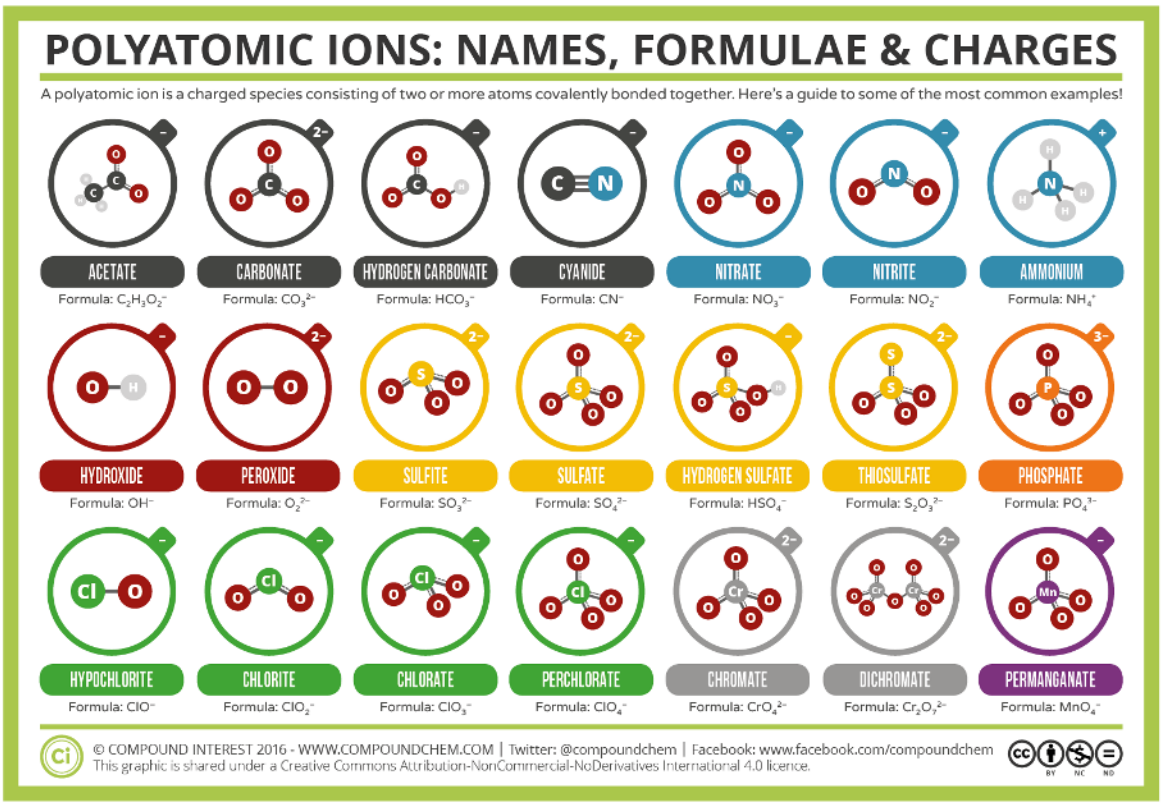


Figure 3.4.1: Names, formula, and charges of polyatomic ions

Polyatomic ions have defined formulas, names, and charges that cannot be modified in any way. Table **3.4.1** lists the ion names and ion formulas of the most common polyatomic ions. For example, $\mathbf{NO_3^-}$ is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge.

Table 3.4.1: Ion Names and Ion Formulas of Common Polyatomic Ions

Ion Name	Ion Formula
ammonium ion	NH_4^{+1}
hydronium ion	H_3O^{+1}
hydroxide ion	OH^{-1}
cyanide ion	CN^{-1}
carbonate ion	CO_3^{-2}
bicarbonate or hydrogen carbonate	HCO_3^-
acetate ion	$C_2H_3O_2^{-1}$ **or** $CH_3CO_2^{-1}$
nitrate ion	NO_3^{-1}
nitrite ion	NO_2^{-1}
sulfate ion	SO_4^{-2}
sulfite ion	SO_3^{-2}
phosphate ion	PO_4^{-3}
hydrogenphosphate ion	HPO_4^{-2}
dihydrogenphosphate ion	$H_2PO_4^{-1}$

Note that there are two polyatomic ions in this table, the ammonium ion and hydronium ion that are **cations**. The ammonium ion contains one nitrogen and four hydrogen's that collectively bear a +1 charge. The hydronium ion contains one oxygen and three hydrogen's that collectively bear a +1 charge.

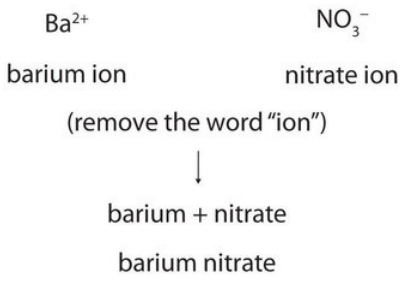
The remaining polyatomic ions are all negatively-charged and, therefore, are classified as anions. However, only two of these, the hydroxide ion and the cyanide ion, are named using the "-ide" suffix that is typically indicative of negatively-charged particles.

The remaining polyatomic anions, which all contain oxygen, in combination with another non-metal, exist as part of a series in which the number of oxygen's within the polyatomic unit can vary. No two chemical formulas should share a common chemical name. A single suffix, "-ide," is insufficient for distinguishing the names of the anions in a related polyatomic series. Therefore, "-ate" and "-ite" suffixes are employed, in order to denote that the corresponding polyatomic ions are part of a series. Additionally, these suffixes also indicate the relative number of oxygens that are contained within the polyatomic ions. Note that all of the polyatomic ions whose names end in "-ate" contain one more oxygen than those polyatomic anions whose names end in "-ite." Unfortunately, much like the common system for naming transition metals, these suffixes only indicate the *relative* number of oxygens that are contained within the polyatomic ions. For example, the nitrate ion, which is symbolized as NO_3^{-1}, has one more oxygen than the nitrite ion, which is symbolized as NO_2^{-1}. However, the sulfate ion is symbolized as SO_4^{-2}. While both the nitrate ion and the sulfate ion share an "-ate" suffix, the former contains three oxygens, but the latter contains four. Additionally, both the nitrate ion and the sulfite ion contain three oxygens, but these polyatomic ions do not share a common suffix. Unfortunately, the relative nature of these suffixes mandates that the ion formula/ion name combinations of the polyatomic ions must simply be memorized.

> Most of the polyatomic ions in table 3.4.1 are found in significant amounts in all living things and required for survival. For example, the hydrogen carbonate ion (HCO_3^-) is involved in the transport of carbon dioxide (CO_2) from the tissues to the lungs.

Naming Compounds

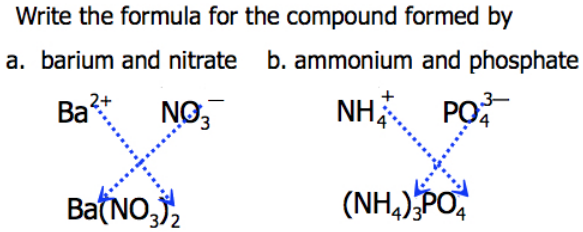
Now that we know how to name ions, we are ready to name ionic compounds. We do so by placing the name of the cation first, followed by the name of the anion, and dropping the word *ion* from both parts. For example, what is the name of the compound whose formula is $Ba(NO_3)_2$?



The compound's name does not indicate that there are two nitrate ions for every barium ion. You must determine the relative numbers of ions by balancing the positive and negative charges. If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride. If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively.

Formulae for Ionic Compounds from Polyatomic Ions

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. Two examples are shown below:



The image asks a two-part question: Write the formula for the compound formed by (a) barium and nitrate, and (b) ammonium and phosphate. For (a), Ba has a charge of 2+ and NO_3 has a charge of -1. By crossing charges, the formula is $Ba(NO_3)_2$. For (b), NH_4 has a charge of +1 and PO_4 has a charge of 3-, so by crossing charges the formula is $(NH_4)_3PO_4$.

Example 3.4.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the potassium ion and the sulfate ion
b. the calcium ion and the nitrate ion

Solution

a. Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.

b. Calcium ions have a charge of 2+, while nitrate ions have a charge of 1−. We will need two nitrate ions to balance the charge on each calcium ion. The formula for nitrate must be enclosed in parentheses. Thus, we write $Ca(NO_3)_2$ as the formula for this ionic compound.

? Exercise 3.4.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a:

Mg^{2+} and $CO_3^{2-} = MgCO_3$

Answer b:

Al^{3+} and $C_2H_3O_2^- = Al(C_2H_3O_2)_3$

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between **metal and nonmetal** elements are usually **ionic**. For example, $CaBr_2$ contains a metallic element (calcium, a group 2A metal) and a nonmetallic element (bromine, a group 7A nonmetal). Therefore, it is an ionic compound. In contrast, the compound NO_2 contains two elements that are **both nonmetals** (nitrogen, from group 5A, and oxygen, from group 6A). It is not an ionic compound; it belongs to the category of **covalent compounds** also know as molecular compounds. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a **polyatomic ion** in a compound, the compound is **ionic**. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, NO_3^-. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, **Ba**, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2− charge from the two nitrate ions. Thus, this compound is also ionic.

If the compound is not ionic then it is a molecular compound.

✓ Example 3.4.2

Identify each compound as ionic or molecular (not ionic).

a. Na_2O
b. PCl_3
c. NH_4Cl
d. OF_2

Solution

a. Sodium is a metal, and oxygen is a nonmetal; therefore, Na_2O is expected to be ionic.
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is a molecular compound.
c. The NH_4 in the formula represents the ammonium ion, NH_4^+, which indicates that this compound is ionic.
d. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.

? Exercise 3.4.2

Identify each compound as ionic or molecular (not ionic).

a. N_2O
b. $FeCl_3$
c. $(NH_4)_3PO_4$
d. $SOCl_2$

Answer a:

molecular

Answer b:

ionic

Answer c:

ionic

Answer d:

molecular

Looking Closer: Blood and Seawater

Science has long recognized that blood and seawater have similar compositions. After all, both liquids have ionic compounds dissolved in them. The similarity may be more than mere coincidence; many scientists think that the first forms of life on Earth arose in the oceans. A closer look, however, shows that blood and seawater are quite different. A 0.9% solution of sodium chloride approximates the salt concentration found in blood. In contrast, seawater is principally a 3% sodium chloride solution, over three times the concentration in blood. Here is a comparison of the amounts of ions in blood and seawater:

Table **3.4.2**: A comparison of the amounts of ions in blood and seawater.

Ion	Percent in Seawater	Percent in Blood
Na^+	2.36	0.322
Cl^-	1.94	0.366
Mg^{2+}	0.13	0.002
SO_4^{2-}	0.09	—
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
HCO_3^-	0.002	0.165
HPO_4^{2-}, $H_2PO_4^-$	—	0.01

Most ions are more abundant in seawater than they are in blood, with some important exceptions. There are far more hydrogen carbonate ions ($\mathbf{HCO_3^-}$) in blood than in seawater. This difference is significant because the hydrogen carbonate ion and some related ions have a crucial role in controlling the acid-base properties of blood. The amount of hydrogenphosphate ions—$\mathbf{HPO_4^{2-}}$ and dihydrogenphosphate ions—$\mathbf{H_2PO_4^-}$ in seawater is very low, but they are present in higher amounts in blood, where they also affect acid-base properties. Another notable difference is that blood does not have significant amounts of the sulfate ion ($\mathbf{SO_4^{2-}}$), but this ion is present in seawater.

A synopsis of how to name simple ionic compounds is shown in figure 3.4.2.

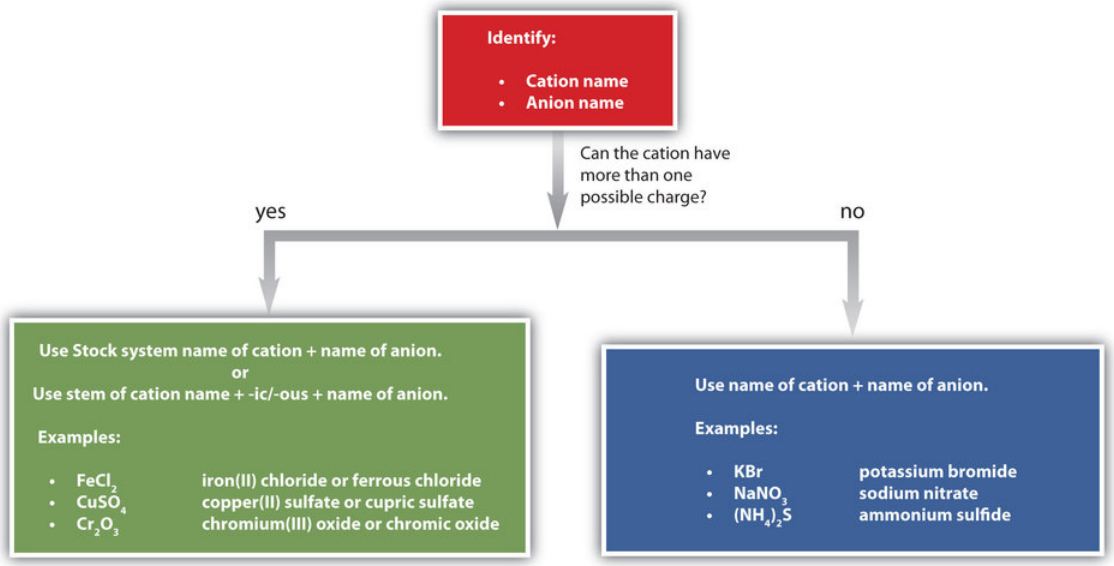


Figure 3.4.2: A Guide to Naming Simple Ionic Compounds. Follow these steps to name a simple ionic compound.

Identify the cation name and the anion name. If the cation can have more than one possible charge, either use the Stock system name of the cation and name of the anion, or use the stem of the cation name and -ic/-ous and the name of the anion. Examples of this would be $FeCl_2$, which is iron(II) chloride of ferrous chloride, $CuSO_4$, which is copper(II) sulfate or cupric sulfate, and Cr_2O_3, which is chromium(III) oxide or chromic oxide. If the cation can not have more than one possible charge, use the name of the cation and the name of the anion. Examples of this would be KBr, which is potassium bromide, $NaNO_3$, which is sodium nitrate, and $(NH_4)_2S$, which is ammonium sulfide.

Key Takeaways

- Proper chemical formulas for ionic compounds balance the total positive charge with the total negative charge.
- Groups of atoms with an overall charge, called polyatomic ions, also exist.

EXERCISES

1. What information is contained in the formula of an ionic compound?

2. Why do the chemical formulas for some ionic compounds contain subscripts, while others do not?

3. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Mg^{2+} and I^-
   b. Na^+ and O^{2-}

4. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Na^+ and Br^-
   b. Mg^{2+} and Br^-
   c. Mg^{2+} and S^{2-}

5. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. K^+ and Cl^-
   b. Mg^{2+} and Cl^-
   c. Mg^{2+} and Se^{2-}

6. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Na^+ and N^{3-}
   b. Mg^{2+} and N^{3-}
   c. Al^{3+} and S^{2-}

7. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Li^+ and N^{3-}
b. Mg^{2+} and P^{3-}
c. Li^+ and P^{3-}

8. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Fe^{3+} and Br^-
b. Fe^{2+} and Br^-
c. Au^{3+} and S^{2-}
d. Au^+ and S^{2-}

9. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Cr^{3+} and O^{2-}
b. Cr^{2+} and O^{2-}
c. Pb^{2+} and Cl^-
d. Pb^{4+} and Cl^-

10. Write the chemical formula for the ionic compound formed by each pair of ions.

a. Cr^{3+} and NO_3^-
b. Fe^{2+} and PO_4^{3-}
c. Ca^{2+} and CrO_4^{2-}
d. Al^{3+} and OH^-

11. Write the chemical formula for the ionic compound formed by each pair of ions.

a. NH_4^+ and NO_3^-
b. H^+ and $Cr_2O_7^{2-}$
c. Cu^+ and CO_3^{2-}
d. Na^+ and HCO_3^-

12. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

a. Ba and S
b. Cs and I

13. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

a. K and S
b. Sc and Br

14. Which compounds would you predict to be ionic?

a. Li_2O
b. $(NH_4)_2O$
c. CO_2
d. $FeSO_3$
e. C_6H_6
f. C_2H_6O

15. Which compounds would you predict to be ionic?

a. $Ba(OH)_2$
b. CH_2O
c. NH_2CONH_2
d. $(NH_4)_2CrO_4$
e. C_8H_{18}

f. NH_3

16. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

a. Cr^{3+} and NO_3^-
b. Fe^{2+} and PO_4^{3-}
c. Ca^{2+} and CrO_4^{2-}
d. Al^{3+} and OH^-

17. Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

a. NH_4^+ and NO_3^-
b. K^+ and $Cr_2O_7^{2-}$
c. Cu^+ and CO_3^{2-}
d. Na^+ and HCO_3^-

18. Give two names for each compound.

a. $Al(HSO_4)_3$
b. $Mg(HSO_4)_2$

19. Give two names for each compound.

a. $Co(HCO_3)_2$
b. $LiHCO_3$

20. Give the formula of each compound.

a. chromium(III) oxide or chromic oxide
b. chromium(II) oxide or chromous oxide
c. lead(II) chloride or plumbous chloride
d. lead(IV) chloride or plumbic chloride
e. chromium(III) nitrate or chromic nitrate
f. iron(II) phosphate or ferrous phosphate
g. calcium chromate
h. aluminum hydroxide
i. ammonium nitrate
j. copper(I) carbonate or cuprous carbonate
k. sodium hydrogen carbonate or sodium bicarbonate
l. aluminum hydrogen sulfate or aluminum bisulfate
m. magnesium hydrogen sulfate or magnesium bisulfate
n. cobalt hydrogen carbonate or cobalt bicarbonate
o. lithium hydrogen carbonate or lithium bicarbonate

Answers

1. the ratio of each kind of ion in the compound

2. Sometimes more than one ion is needed to balance the charge on the other ion in an ionic compound.

3.

a. MgI_2
b. Na_2O

4.

a. NaBr
b. $MgBr_2$
c. MgS

5.

a. KCl
b. $MgCl_2$
c. MgSe

6.

a. Na_3N
b. Mg_3N_2
c. Al_2S_3

7.

a. Li_3N
b. Mg_3P_2
c. Li_3P

8.

a. $FeBr_3$
b. $FeBr_2$
c. Au_2S_3
d. Au_2S

9.

a. Cr_2O_3
b. CrO
c. $PbCl_2$
d. $PbCl_4$

10.

a. $Cr(NO_3)_3$
b. $Fe_3(PO_4)_2$
c. $CaCrO_4$
d. $Al(OH)_3$

11.

a. NH_4NO_3
b. $H_2Cr_2O_7$
c. Cu_2CO_3
d. $NaHCO_3$

12.

a. Ba^{2+}, S^{2-}, BaS
b. Cs^{+}, I^{-}, CsI

13.

a. K^{+}, S^{2-}, K_2S
b. Sc^{3+}, Br^{-}, $ScBr_3$

14.

a. ionic
b. ionic
c. not ionic
d. ionic

e. not ionic
f. not ionic

15.

a. ionic
b. not ionic
c. not ionic
d. ionic
e. not ionic
f. not ionic

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3.5: Characteristics of Ionic Compounds-Electrolytes

Learning Objectives

- To describe the physical properties of ionic compounds.

The figure below shows just a few examples of the color and brilliance of naturally occurring ionic crystals. The regular and orderly arrangement of ions in the crystal lattice is responsible for the various shapes of these crystals, while transition metal ions give rise to the colors.

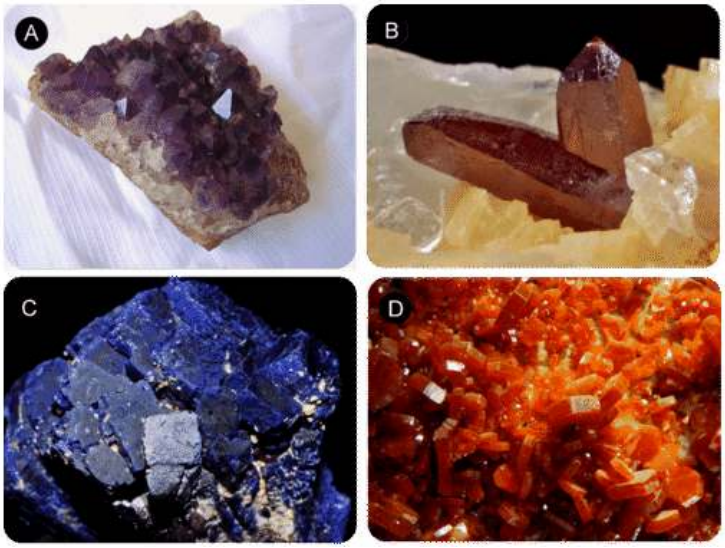


Figure 3.6.1: *In nature, the ordered arrangement of ionic solids gives rise to beautiful crystals. (A) Amethyst - a form of quartz, SiO_2, whose purple color comes from iron ions. (B) Cinnabar - the primary ore of mercury is mercury (II) sulfide, HgS (C) Azurite - a copper mineral, $Cu_3(CO_3)_2(OH)_2$. (D) Vanadinite - the primary ore of vanadium, $Pb_3(VO_4)_3Cl$.*

Physical Properties of Ionic Compounds

Melting Points

Because of the many simultaneous attractions between cations and anions that occur, ionic crystal lattices are very strong. The process of melting an ionic compound requires the addition of large amounts of energy in order to break all of the ionic bonds in the crystal. For example, sodium chloride has a melting temperature of about 800°C. As a comparison, the molecular compound water melts at 0 °C.

Shattering

Ionic compounds are generally hard, but **brittle.** Why? It takes a large amount of mechanical force, such as striking a crystal with a hammer, to force one layer of ions to shift relative to its neighbor. However, when that happens, it brings ions of the same charge next to each other (see below). The repulsive forces between like-charged ions cause the crystal to shatter. When an ionic crystal breaks, it tends to do so along smooth planes because of the regular arrangement of the ions.

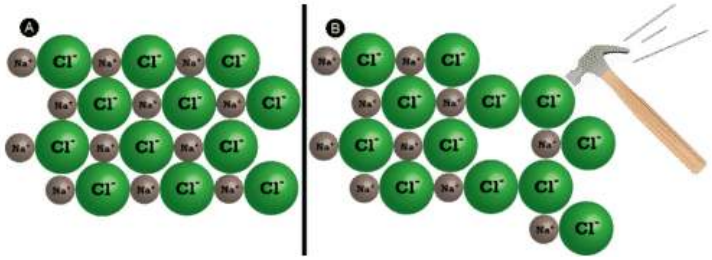


Figure 3.6.2: *(A) The sodium chloride crystal is shown in two dimensions. (B) When struck by a hammer, the negatively-charged chloride ions are forced near each other and the repulsive force causes the crystal to shatter.*

Conductivity

Another characteristic property of ionic compounds is their **electrical conductivity**. The figure below shows three experiments in which two electrodes that are connected to a light bulb are placed in beakers containing three different substances.

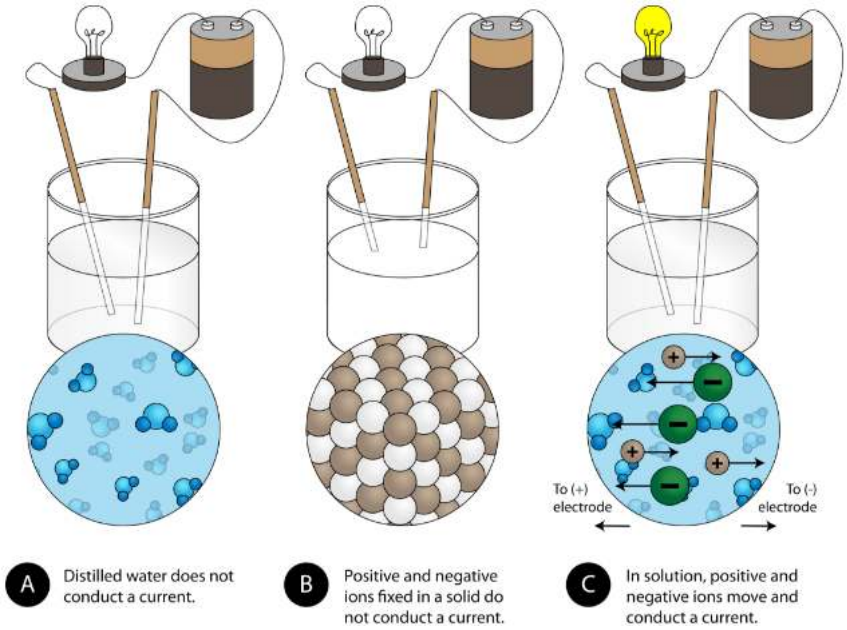


Figure 3.6.3: *(A) Distilled water does not conduct electricity. (B) A solid ionic compound also does not conduct. (C) A water solution of an ionic compound conducts electricity well.*

In the first beaker, distilled water does not conduct a current because water is a molecular compound. In the second beaker, solid sodium chloride also does not conduct a current. Despite being ionic and thus composed of charged particles, the solid crystal lattice does not allow the ions to move between the electrodes. Mobile charged particles are required for the circuit to be complete and the light bulb to light up. In the third beaker, the NaCl has been dissolved into the distilled water. Now the crystal lattice has been broken apart and the individual positive and negative ions can move. Cations move to one electrode, while anions move to the other, allowing electricity to flow (see figure below). Melting an ionic compound also frees the ions to conduct a current. Ionic compounds conduct an electric current when melted or dissolved in water.

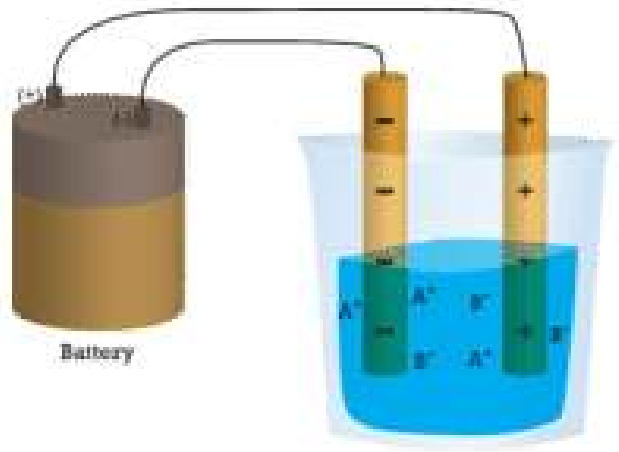


Figure 3.6.4: *In an ionic solution, the A^+ ions migrate toward the negative electrode, while the B^- ions migrate toward the positive electrode.*

Example 3.5.1

Write the dissociation equation of solid NaCl in water.

Solution

$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$

3.5.1 Exercise

Write the dissociation equation of solid NH_4NO_3 in water.

Answer

$NH_4NO_3(s) \rightarrow NH_4^{+}(aq) + NO_3^{-}(aq)$

Key Takeaways

- Ionic compounds have high melting points.
- Ionic compounds are hard and brittle.
- Ionic compounds dissociate into ions when dissolved in water.
- Solutions of ionic compounds and melted ionic compounds conduct electricity, but solid materials do not.
- An ionic compound can be identified by its chemical formula: cation+anion.

Contributors and Attributions

- CK-12 Foundation by Sharon Bewick, Richard Parsons, Therese Forsythe, Shonna Robinson, and Jean Dupon.

Exercises

1. Explain how the bonding in an ionic solid explains some of the properties of these solids.

2. Which type(s) of solid conduct(s) electricity in their liquid state but not in their solid state?

3. Based on chemical formula, identify which of the following is an ionic solid?

    a. Hg

    b. PH_3

    c. $Ba(NO_3)_2$

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3.6: Covalent Bonds

Learning Objectives

- To describe how a covalent bond forms.
- To apply the octet rule or duet rule to covalent compounds.

Water is not an ionic compound; it is not composed of a metal and a nonmetal. Consequently, its properties are different from those of ionic compounds.

Electron Sharing

Previously, we discussed ionic bonding where electrons can be transferred from one atom to another so that both atoms have an energy-stable outer electron shell. Because most filled electron shells have eight electrons in them, chemists called this tendency the octet rule. However, there is another way an atom can achieve a full valence shell: atoms can *share* electrons.

This concept can be illustrated by using two hydrogen atoms, each of which has a single electron in its valence shell. (For small atoms such as hydrogen atoms, the valence shell will be the first shell, which holds only two electrons.) We can represent the two individual hydrogen atoms as follows:

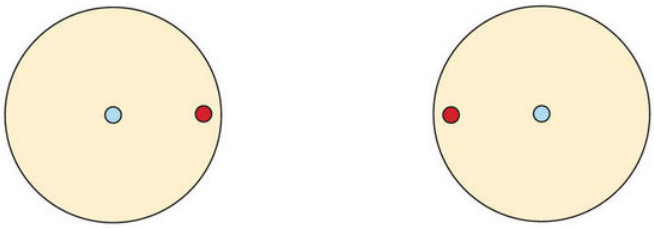


Figure **3.6.1**: Individual Hydrogen Atoms

In contrast, when two hydrogen atoms get close enough together to share their electrons, they can be represented as follows:

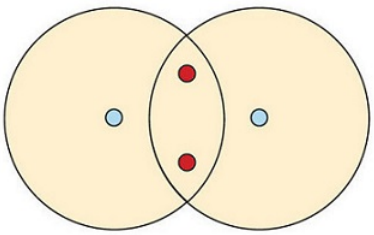


Figure **3.6.2**: H_2 Molecule

By sharing their valence electrons, both hydrogen atoms now have two electrons in their respective valence shells. Because each valence shell is now filled, this arrangement is more stable than when the two atoms are separate. The sharing of electrons between atoms is called a covalent bond, and the two electrons that join atoms in a covalent bond are called a bonding pair of electrons. A discrete group of atoms connected by covalent bonds is called a molecule.

Chemists frequently use electron dot diagrams to represent covalent bonding in molecular substances. For example, the electron dot diagram of two separate hydrogen atoms are as follows:

$$H\bullet \qquad\qquad \bullet H$$

The electron dot diagram of two hydrogen atoms sharing electrons looks like this:

$$H\!:\!H$$

This depiction of molecules is simplified further by using a dash to represent a covalent bond. The hydrogen molecule is then represented as follows:



Remember that the dash, also referred to as a single bond, represents a *pair* of electrons.

The bond in a hydrogen molecule, measured as the distance between the two nuclei, is about 7.4×10^{-11} m, or 74 picometers (pm; 1 pm = 1×10^{-12} m). This particular bond length represents a balance between several forces: the attractions between oppositely charged electrons and nuclei, the repulsion between two negatively charged electrons, and the repulsion between two positively charged nuclei. If the nuclei were closer together, they would repel each other more strongly; if the nuclei were farther apart, there would be less attraction between the positive and negative particles.

Fluorine is another element whose atoms bond together in pairs to form **diatomic** (two-atom) molecules. Two separate fluorine atoms have the following electron dot diagrams:

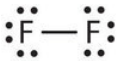


Each fluorine atom contributes one valence electron, making a single bond and giving each atom a complete valence shell, which fulfills the **octet rule**:



Two F's with two dots in between them, and two dots on the top, left/right side, and bottoms of each F.

The circles show that each fluorine atom has eight electrons around it. As with hydrogen, we can represent the fluorine molecule with a dash in place of the **bonding electrons**:



Two F's with a long dash in between them, and two dots on the top, left/right side, and bottoms of each F.

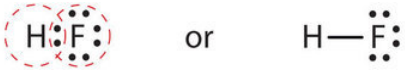
Each fluorine atom has six electrons, or three pairs of electrons, that are not participating in the covalent bond. Rather than being shared, they are considered to belong to a single atom. These are called **nonbonding pairs** (or **lone pairs**) of electrons.

Covalent Bonds between Different Atoms

Now that we have looked at electron sharing between atoms of the same element, let us look at covalent bond formation between atoms of different elements. Consider a molecule composed of one hydrogen atom and one fluorine atom:



Each atom needs one additional electron to complete its valence shell. By each contributing one electron, they make the following molecule:



An H and F connected by two dots, and the F has 2 dots on the top, right, and bottom. It also shows the same combination but with a long dash in between the H and F rather than two dots.

In this molecule, the hydrogen atom does not have nonbonding electrons, while the fluorine atom has six nonbonding electrons (three lone electron pairs). The circles show how the valence electron shells are filled for both atoms.

Example 3.6.1

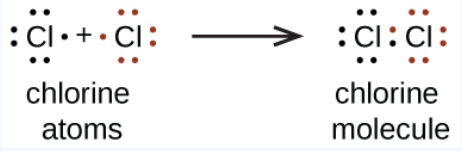
Use Lewis diagrams to indicate the formation of the following:

a. Cl_2
b. HBr

Solution

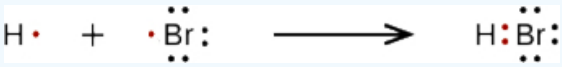
a. When two chlorine atoms form a chlorine molecule, they share one pair of electrons. In Cl_2 molecule, each chlorine atom is surrounded by an **octet** number of electrons.

The Lewis diagram for a Cl_2 molecule is similar to the one for F_2 (shown above).



Cl with 7 dots surrounding it plus Cl with 7 dots surrounding it means they are chlorine atoms. To turn into a chlorine molecule, have two dots in between each Cl and 6 dots surrounding each Cl on the sides that aren't connected.

b. When a hydrogen atom and a bromine atom form HBr, they share one pair of electrons. In the HBr molecule, H achieves a full valence of two electrons (**duet**) while Br achieves an **octet**. The Lewis diagram for HBr is similar to that for HF shown above.



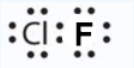
H with one dot plus Br with 7 dots turns into H connected to Br by two dots, with 6 other dots surrounding Br.

? Exercise 3.6.1

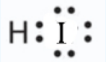
Draw the Lewis diagram for each compound.

a. a molecule composed of one chlorine atom and one fluorine atom
b. a molecule composed of one hydrogen atom and one iodine atom

Answer a:

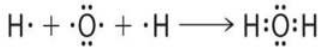


Answer b:



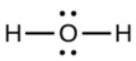
Covalent Bonds in Larger Molecules

The formation of a water molecule from two hydrogen atoms and an oxygen atom can be illustrated using Lewis dot symbols (shown below).



H with one dot + O with 6 dots + H with one dot turns into H connected to O with two dots connected to H with two dots. O also has two dots on top and top dots on bottom.

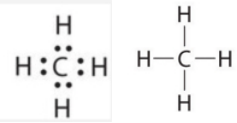
The structure shown below is the *electron dot structure* for H_2O. With two bonding pairs and two lone pairs, the oxygen atom has now completed its **octet**. Moreover, by sharing a bonding pair with oxygen, each hydrogen atom now has a full valence shell of two electrons. Chemists usually indicate a bonding pair by a single line, as shown (below).



Water

Other large molecules are constructed in a similar fashion, with some atoms participating in more than one covalent bond. For example, methane (CH_4), the central carbon atom bonded to four hydrogen atoms, can be represented using either of the Lewis

structures below. Again, sharing electrons between C and H atoms results in C achieving and octet while H achieving a duet number of electrons.



How Many Covalent Bonds Are Formed?

The number of bonds that an atom can form can often be predicted from the number of electrons needed to reach an octet (eight valence electrons). In the Lewis structure, the number of bonds formed by an element in a neutral compound is the same as the number of unpaired electrons it must share with other atoms to complete its octet of electrons. For example, each atom of a **group 4A** element has four electrons in its outermost shell and therefore requires four more electrons to reach an octet. These four electrons can be gained by forming **four covalent bonds**, as illustrated here for carbon in CH_4 (methane). **Group 5A** elements such as nitrogen have five valence electrons in the atomic Lewis symbol: one lone pair and three unpaired electrons. To obtain an octet, these atoms form **three covalent bonds**, as in NH_3 (ammonia). Oxygen and other atoms in **group 6A** obtain an octet by forming **two covalent bonds**. Fluorine and the other halogens in **group 7A** have seven valence electrons and can obtain an octet by forming **one covalent bond**. The number of electrons required to obtain an octet determines the number of covalent bonds an atom can form. This is summarized in the table below. In each case, the sum of the number of bonds and the number of lone pairs is 4, which is equivalent to eight (octet) electrons.

This table shows atoms and their group numbers, and how many bonds and lone pairs each has.

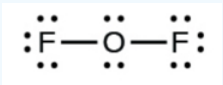
Atom (Group number)	Number of Bonds	Number of Lone Pairs
Carbon (Group 4A)	4	0
Nitrogen (Group 5A)	3	1
Oxygen (Group 6A)	2	2
Fluorine (Group 7A)	1	3

Because hydrogen only needs two electrons to fill its valence shell, it follows the **duet rule**. It is an **exception to the octet rule.** Hydrogen only needs to form one bond. This is the reason why H is always a terminal atom and never a central atom.

The transition elements and inner transition elements also do not follow the octet rule since they have d and f electrons involved in their valence shells.

Example 3.6.2

Examine the Lewis structure of OF_2 below. Count the number of bonds formed by each element. Based on the element's location in the periodic table, does it correspond to the expected number of bonds shown in Table 4.1? Does the Lewis structure below follow the octet rule?



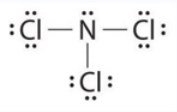
F surrounded by 6 dots, connected to O with a long dash. O surrounded by 4 dots and connects to another F with a long dash. F surrounded by 6 dots.

Solution

Yes. F (group 7A) forms one bond and O (group 6A) forms 2 bonds. Each atom is surrounded by 8 electrons. This structure satisfies the octet rule.

3.6.2 Exercise

Examine the Lewis structure of NCl_3 below. Count the number of bonds formed by each element. Based on the element's location in the periodic table, does it correspond to the expected number of bonds shown in Table 4.1? Does the Lewis structure below follow the octet rule?



Answer

Both Cl and N form the expected number of bonds. Cl (group 7A) has one bond and 3 lone pairs. The central atom N (group 5A) has 3 bonds and one lone pair. Yes, the Lewis structure of NCl_3 follows the octet rule.

Key Takeaways

- A covalent bond is formed between two atoms by sharing electrons.
- The number of bonds an element forms in a covalent compound is determined by the number of electrons it needs to reach octet.
- Hydrogen is an exception to the octet rule. H forms only one bond because it needs only two electrons.

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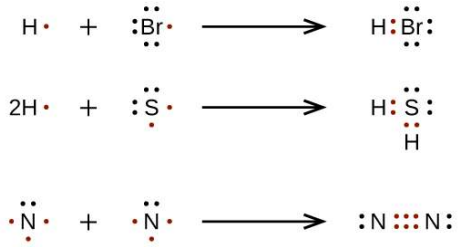
3.7: Multiple Bonds

Learning Objectives

- To draw electron dot structures.
- To recognize molecules that are likely to have multiple covalent bonds.

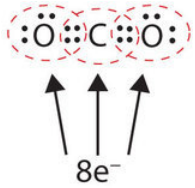
Electron Dot Structures of Molecules

Electron dot structures also called Lewis structures are drawn by pairing up the unpaired electrons on the constituent atoms so that each atom has an octet or duet complete. See these examples:

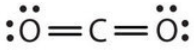


MULTIPLE BONDS

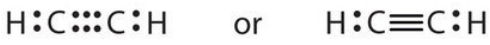
In many molecules, the octet rule would not be satisfied if each pair of bonded atoms shares only two electrons. In case of carbon dioxide, more than one pair of electrons must be shared between two atoms for both atoms to have an octet.



In this arrangement, the carbon atom shares four electrons (two pairs) with the oxygen atom on the left and four electrons with the oxygen atom on the right. There are now eight electrons around each atom. Two pairs of electrons shared between two atoms make a double bond between the atoms, which is represented by a double dash:

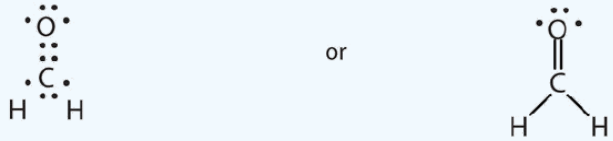


Triple bonds are covalent bonds in which *three* pairs of electrons are shared by two atoms. A compound that has a triple bond is acetylene (C_2H_2), whose Lewis diagram is as follows:



Example 3.7.1

The Lewis diagram for formaldehyde CH_2O is shown below. The carbon atom is the central atom. One application of formaldehyde, is the preservation of biological specimens. Aqueous solutions of CH_2O are called formalin and have a sharp, characteristic (pungent) odor.



How many double bonds, single bonds, and lone pairs of electrons are present in the molecule of formaldehyde? Are the octet/duet rules met for each atom?

Solution

There is one double bond between C and O atom, two single bonds between C and H atoms, and two lone pairs on the oxygen atom. The octet for the oxygen and carbon are complete. The two hydrogen atoms complete their duet.

? Exercise 3.7.1

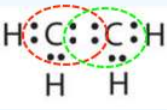
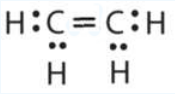
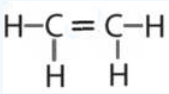
Draw the Lewis diagram for each molecule.

a. O_2
b. C_2H_4

Answer a:

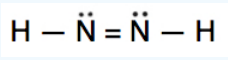
 or 

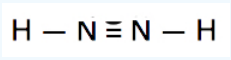
Answer b:

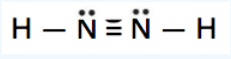
 or  or 

✓ Example 3.7.2

Which is the correct Lewis structure for N_2H_2?

A. 

B. 

C. 

Solution

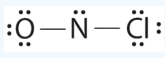
Lewis structure A is the correct answer. Each of the N atoms satisfy the octet requirement and the H atoms follow the duet rule.

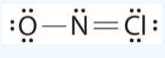
Structure B is electron deficient. It has only 10 valence electrons instead of 12.

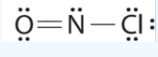
Structure C has 14 (2 extra) electrons. The N atoms do not satisfy the octet.

3.7.2 Exercise

Which is the correct Lewis structure for NOCl?

A. 

B. 

C. 

Answer

Structure A violates the octet rule; N is surrounded by only $6e^-$.

Structure B violates the octet rule; Cl has $10e^-$ around it.

Structure C is the correct structure. It has a total of $6e^- + 5e^- + 7e^- = 18e^-$. Each atom is surrounded by 8 electrons (octet rule).

Key Takeaways

- A Lewis structure shows the bonding and nonbonding electrons around individual atoms in a molecule.
- Some molecules must have multiple covalent bonds between atoms to satisfy the octet rule.
- A double bond contains four electrons and a triple bond contains six electrons.

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3.8: Binary Molecules - Formulae and Names

Learning Objectives

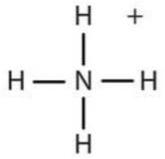
- Identify covalent and ionic compounds.
- Determine the chemical formula of a simple covalent compound from its name.
- Determine the name of a simple covalent compound from its chemical formula.

COVALENT AND IONIC COMPOUNDS

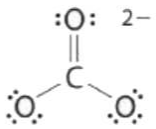
What elements make covalent bonds? Covalent bonds form when two or more nonmetals combine. For example, both hydrogen and oxygen are nonmetals, and when they combine to make water, they do so by forming covalent bonds. Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds.

As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals), silicon dioxide (SiO_2) will be a covalently bound molecule (a semi-metal and a non-metal) and $MgCl_2$ will be ionic (a metal and a non-metal).

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 NH_4^+. 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 CO_3^{2-}.



ammonium ion



carbonate ion

The atoms of a polyatomic ion are tightly bonded together and so the entire ion behaves as a single unit. Nonmetal atoms in polyatomic ions are joined by covalent bonds, but the ion as a whole participates in ionic bonding. For example, ammonium chloride (NH_4Cl) has ionic bonding between a polyatomic ion, $\mathbf{NH_4^+}$, and $\mathbf{Cl^-}$ ions, but within the ammonium ion (NH_4^+), the nitrogen and hydrogen atoms are connected by covalent bonds (shown above).

Both ionic and covalent bonding are also found in calcium carbonate. Calcium carbonate ($CaCO_3$) has ionic bonding between calcium ion $\mathbf{Ca^{2+}}$ and a polyatomic ion, $\mathbf{CO_3^{2-}}$, but within the carbonate ion (CO_3^{2-}), the carbon and oxygen atoms are connected by covalent bonds (shown above).

Characteristics of Covalent (Molecular) Compounds

Compounds that contain covalent bonds (also called molecular compounds) exhibit different physical properties than ionic compounds. Because the attraction between molecules, which are electrically neutral, is weaker than that between electrically charged ions, covalent compounds generally have much lower melting and boiling points than ionic compounds. For example, water (molecular compound) boils at 100 °C while sodium chloride (ionic compound) boils at 1413 °C. In fact, many covalent compounds are liquids or gases at room temperature, and, in their solid states, they are typically much softer than ionic solids. Furthermore, whereas ionic compounds are good conductors of electricity when dissolved in water, most covalent compounds, being electrically neutral, are poor conductors of electricity in any state.

Example 3.8.1

Is each compound formed from ionic bonds, covalent bonds, or both?

a. $\mathbf{Na_2O}$

b. Na_3PO_4
c. N_2O_4

Answer a

The elements in Na_2O are a metal and a nonmetal, which form ionic bonds.

Answer b

Because sodium is a metal and we recognize the formula for the phosphate ion, we know that this compound is ionic. However, within the polyatomic phosphate ion, the atoms are held together by covalent bonds, so this compound contains both ionic and covalent bonds.

Answer c

The elements in N_2O_4 are both nonmetals, rather than a metal and a nonmetal. Therefore, the atoms form covalent bonds.

? Exercise 3.8.1

Is each compound are formed from ionic bonds, covalent bonds, or both?

a. $Ba(OH)_2$
b. F_2
c. PCl_3

Answer a:

both

Answer b:

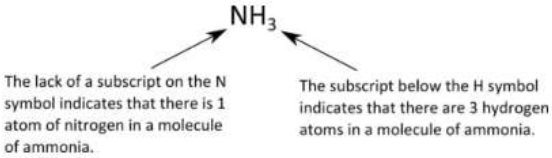
covalent

Answer c:

covalent

MOLECULAR FORMULAS

The chemical formulas for covalent compounds are referred to as **molecular formulas** because these compounds exist as separate, discrete molecules. Typically, a molecular formula begins with the nonmetal that is closest to the lower left corner of the periodic table, except that hydrogen is almost never written first (H_2O is the prominent exception). Then the other nonmetal symbols are listed. Numerical subscripts are used if there is more than one of a particular atom. For example, we have already seen CH_4, the molecular formula for methane. Below is the molecular formula of ammonia, NH_3.



NH_3. An arrow points to the N and says, "The lack of a subscript on the N symbol indicates that there is 1 atom of nitrogen in a molecule of ammonia". An arrow points to H and says, "The subscript below the H symbol indicates that there are 3 hydrogen atoms in a molecule of ammonia".

NAMING COVALENT COMPOUNDS

Naming *binary* (two-element) covalent compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table **3.8.1** lists these numerical prefixes. Normally, no prefix is added to the first element's name if there is only one atom of the first element in a molecule. If the second element is oxygen, the trailing vowel is usually omitted from the end of a polysyllabic prefix but not a monosyllabic one (that is, we would say "monoxide" rather than "monooxide" and "trioxide" rather than "troxide").

Table **3.8.1**: Numerical Prefixes for Naming Binary Covalent Compounds

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

*This prefix is not used for the first element's name.

Let us practice by naming the compound whose molecular formula is CCl_4. The name begins with the name of the first element—carbon. The second element, chlor*ine*, becomes chlor*ide*, and we attach the correct numerical prefix ("tetra-") to indicate that the molecule contains four chlorine atoms. Putting these pieces together gives the name *carbon tetrachloride* for this compound.

Example 3.8.2

Write the molecular formula for each compound.

a. chlorine trifluoride
b. phosphorus pentachloride
c. sulfur dioxide
d. dinitrogen pentoxide

Solution

If there is no numerical prefix on the first element's name, we can assume that there is only one atom of that element in a molecule.

a. ClF_3
b. PCl_5
c. SO_2
d. N_2O_5 (The *di-* prefix on nitrogen indicates that two nitrogen atoms are present.)

Exercise 3.8.2

Write the molecular formula for each compound.

a. nitrogen dioxide
b. dioxygen difluoride
c. sulfur hexafluoride
d. selenium monoxide

Answer a:

a. NO_2

Answer b:

O_2F_2

Answer c:

SF_6

Answer d:

SeO

Because it is so unreactive, sulfur hexafluoride is used as a spark suppressant in electrical devices such as transformers.

Example 3.8.3

Write the name for each compound.

a. BrF_5
b. S_2F_2
c. CO

Solution

a. bromine pentafluoride
b. disulfur difluoride
c. carbon monoxide

Exercise 3.8.3

Write the name for each compound.

a. CF_4
b. $SeCl_2$
c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O: water
- NH_3: ammonia
- CH_4: methane

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

The prefix of mono are not used for the two molecules below:

HCl: hydrogen chloride

HF: hydrogen fluoride

Oxides of Nitrogen-Applications in Medicine

Nitric Oxide

Nitrogen monoxide (NO) is known more commonly as nitric oxide. It is a colorless gas formed when nitrogen and oxygen in the air combine when gasoline is burned in the engines of cars and trucks. So this compound has a bad reputation regarding atmospheric pollution. However, NO has many biological functions. It regulates blood pressure, dilates arteries and blood vessels, and is used by the immune system to help fight infections.

These effects of NO on the human body were noticed in the mid 19 th century soon after the discovery of the explosive nitroglycerin that was used to make dynamite. It was noticed that people who worked with nitroglycerin experienced splitting headaches which was the result of vasodilation (relaxation) of the blood vessels in the brain. This led to the use of nitroglycerin being used as a medication for the treatment of angina which is caused by constriction of the arteries that carry blood to the heart. A century later scientists discovered that the body produces NO from nitroglycerin and it is NO that causes coronary arteries to dilate reducing the symptoms of angina. In 1998 Nobel Prize in Medicine was awarded to three scientists who worked on the nitric oxide research.

Nitrous oxide

Nitrous oxide (N_2O) is prepared by the careful thermal decomposition of ammonium nitrate (NH_4NO_3). Nitrous oxide is known as "laughing gas" due to the euphoric effects of inhaling it, a property that has led to its recreational use as a hallucinogen. However, it is as an anesthetic that it has a legitimate application. The first use of nitrous oxide as anesthetic drug was when dentist Horace Wells with assistance by Gardner Quincy Colton and John Mankey Riggs, demonstrated insensitivity to pain from a dental extraction in December 1844. Wells subsequently treated 12-15 patients, and according to his own record it only failed as an anesthetic in two cases.

The method did not come into general use until 1863, when Colton successfully used it for more than 25,000 patients. As such, the usage of nitrous oxide rapidly became the preferred anesthetic method in dentistry. The gas is mild enough to keep a patient in a conscious and conversational state, and yet in most cases strong enough to suppress the pain caused by dental work. It remains the preferred gas anesthetic in today's dentistry.

Key Takeaways

- The chemical formula of a simple covalent compound can be determined from its name.
- The name of a simple covalent compound can be determined from its chemical formula.

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3.9: Formula Weight

Learning Objectives

- To determine the formula mass of an ionic compound.

Formula Weight

One skill needed in future chapters is the ability to determine the mass of the formula of an ionic compound. This quantity is called the formula weight (formula mass). The formula weight is obtained by adding the atomic weight of each individual atom in the formula of the compound. Because a proper formula is electrically neutral (with no net electrons gained or lost), the ions can be considered atoms for the purpose of calculating the formula mass.

Let us start by calculating the formula weight of sodium chloride (NaCl). This formula weight is the sum of the atomic weights of one sodium atom and one chlorine atom, which we find from the periodic table; here, we use the atomic weights to two decimal places:

Na: 22.99 amu

Cl: +35.34 amu

Total: 58.44 amu

To two decimal places, the formula weight of NaCl is 58.44 amu.

When an ionic compound has more than one anion or cation, you must remember to use the proper multiple of the atomic weights for the element in question. For the formula mass of calcium fluoride (CaF_2), we must multiply the mass of the fluorine atom by 2 to account for the two fluorine atoms in the chemical formula:

Ca: 1 x 40.08 = 40.08 amu

F: 2 x 19.00 = +38.00 amu

Total = 78.08 amu

The formula weight of CaF_2 is 78.08 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. For example, potassium nitrate (KNO_3) has one potassium atom, one nitrogen atom, and three oxygen atoms:

K: 1 x 39.10 = 39.10 amu

N: 1 x 14.00 = +14.00 amu

O: 3 x 16.00 = +48.00 amu

Total = 101.10 amu

The formula weight of KNO_3 is 101.10 amu.

> Potassium nitrate is a key ingredient in gunpowder and has been used clinically as a diuretic.

When a formula contains more than one polyatomic unit in the chemical formula, as in $Ca(NO_3)_2$, do not forget to multiply the atomic weights of every atom inside of the parentheses by the subscript outside of the parentheses. This is necessary because the subscript refers to the *entire polyatomic ion*. Thus, for $Ca(NO_3)_2$, the subscript 2 implies two complete nitrate ions, so we must sum the masses of two (1 × 2) nitrogen atoms and six (3 × 2) oxygen atoms, along with the mass of a single calcium atom:

Ca: 1 x 40.08 = 40.08 amu

N: 2 x 14.00 = +28.00 amu

O: 6 x 16.00 = +96.00 amu

Total = 164.08 amu

The key to calculating the formula weight of an ionic compound is to correctly count each atom in the formula and multiply the atomic weights of its atoms accordingly.

Example 3.9.1

Use the atomic masses (rounded to two decimal places) to determine the formula weight for each ionic compound.

a. $FeCl_3$
b. $(NH_4)_3PO_4$

Solution

a.

Fe: 1 x 55.85 = 55.85 amu

Cl: 3 x 35.45 = +106.35 amu

Total = 162.20 amu

The formula weight of $FeCl_3$ is 162.2 amu.

b. When we distribute the subscript 3 through the parentheses containing the formula for the ammonium ion, we see that we have 3 nitrogen atoms and 12 hydrogen atoms. Thus, we set up the sum as follows:

N: 3 x 14.00 = 42.00 amu

H: 12 x 1.00 = +12.00 amu

P: 1 x 30.97 = +30.97 amu

O: 4 x 16.00 = +64.00 amu

Total = 148.97 amu

The formula weight for $(NH_4)_3PO_4$ is 149.0 amu.

Exercise 3.9.1

Use the atomic weights (rounded to two decimal places) to determine the formula weight for each ionic compound.

a. TiO_2
b. AgBr
c. $Au(NO_3)_3$
d. $Fe_3(PO_4)_2$

Answer

a. 79.87 amu

b. 187.77 amu

c. 383.0 amu

To Your Health: Hydrates

Some ionic compounds have water (H_2O) incorporated within their formula unit. These compounds, called *hydrates*, have a characteristic number of water units associated with each formula unit of the compound. Hydrates are solids, not liquids or solutions, despite the water they contain.

To write the chemical formula of a hydrate, write the number of water units per formula unit of compound after its chemical formula. The two chemical formulas are separated by a vertically centered dot. The hydrate of copper(II) sulfate has five water

units associated with each formula unit, so it is written as $CuSO_4 \cdot 5H_2O$. The name of this compound is copper(II) sulfate pentahydrate, with the *penta-* prefix indicating the presence of five water units per formula unit of copper(II) sulfate.



Figure used with permisson from Wikipedia.

Hydrates have various uses in the health industry. Calcium sulfate hemihydrate ($CaSO_4 \cdot \frac{1}{2}H_2O$), known as plaster of Paris, is used to make casts for broken bones. Epsom salt ($MgSO_4 \cdot 7H_2O$) is used as a bathing salt and a laxative. Aluminum chloride hexahydrate is an active ingredient in antiperspirants. The accompanying table lists some useful hydrates.

Table **3.9.1**: Names and Formulas of Some Widely Used Hydrates

Formula	Name	Uses
$AlCl_3 \cdot 6H_2O$	aluminum chloride hexahydrate	antiperspirant
$CaSO_4 \cdot \frac{1}{2}H_2O$	calcium sulfate hemihydrate (plaster of Paris)	casts (for broken bones and castings)
$CaSO_4 \cdot 2H_2O$	calcium sulfate dihydrate (gypsum)	drywall component
$CoCl_2 \cdot 6H_2O$	cobalt(II) chloride hexahydrate	drying agent, humidity indicator
$CuSO_4 \cdot 5H_2O$	copper(II) sulfate pentahydrate	fungicide, algicide, herbicide
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate (Epsom salts)	laxative, bathing salt
$Na_2CO_3 \cdot 10H_2O$	sodium carbonate decahydrate (washing soda)	laundry additive/cleaner

Key Takeaway

- Formula masses of ionic compounds can be determined from the masses of the atoms in their formulas.

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3.10: Molecular Weight

Learning Objectives

- Determine the molecular weight of a molecule.

Ions in ionic compounds, are arranged in lattices called crystals, molecules of covalent compounds exist as discrete units with a characteristic mass and a certain three-dimensional shape.

Molecular Weight

The mass of a molecule—the molecular weight (also called the molecular mass)—is simply the sum of the atomic weights of its atoms. As with formula weights, it is important that you keep track of the number of atoms of each element in the molecular formula to obtain the correct molecular weight.

Example 3.10.1

What is the molecular weight of each covalent compound?

a. H_2O
b. C_6H_6
c. NO_2
d. N_2O_5

Solution

Use the atomic weights from the Periodic Table.

The molecular formula H_2O indicates that there are two hydrogen atoms and one oxygen atom in each molecule. Summing the atomic weights of these atoms,

This periodic table finds the atomic weight of atoms.

2 H:	2 × 1.01 =	2.02 amu
1 O:		+ 16.00 amu
Total:		18.02 amu

The molecular weight of H_2O is **18.02** amu.

This periodic table finds the atomic weight of atoms.

6 C:	6 × 12.01 =	72.06 amu
6 H:	6 × 1.01 =	+ 6.06 amu
Total:		78.12 amu

The molecular weight of C_6H_6 is **78.12** amu.

This periodic table finds the atomic weight of atoms.

1 N:		14.01 amu
2 O:	2 × 16.00 =	+ 32.00 amu
Total:		46.01 amu

The molecular weight of NO_2 is **46.01** amu.

This periodic table finds the atomic weight of atoms.

2 N:	2 × 14.01 =	28.02 amu
5 O:	5 × 16.00 =	+ 80.00 amu
Total:		108.02 amu

The molecular weight of N_2O_5 is **108.02** amu.

Note that the two different nitrogen and oxygen compounds in these examples have different molecular masses.

Exercise 3.10.1

What is the molecular weight of each covalent compound?

a. C_2H_2
b. CO
c. CO_2
d. BF_3

Answer a:

26.04 amu

Answer b:

28.01 amu

Answer c:

44.01 amu

Answer d:

67.81 amu

Key Takeaways

- A molecule has a certain mass, called the molecular weight.

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3.11: Mole-Mass Conversions

Learning Objectives

- To convert quantities between mass units and mole units.

The mass of 2 mol of Uranium is twice the atomic weight of uranium in grams. Such a straightforward exercise does not require any formal mathematical treatment. Many questions concerning mass are not so straightforward, however, and require some mathematical manipulations.

Use the atomic weight in grams as a conversion factor in a mole-mass conversion (or its reverse, a mass-mole conversion).

1 mol of Al has a mass of 26.98 g (Example **3.11.1**). Stated mathematically,

$$1 \text{ mol Al} = 26.98 \text{ g Al}$$

We can divide both sides of this expression by either side to get one of two possible conversion factors:

$$\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \quad \text{and} \quad \frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$$

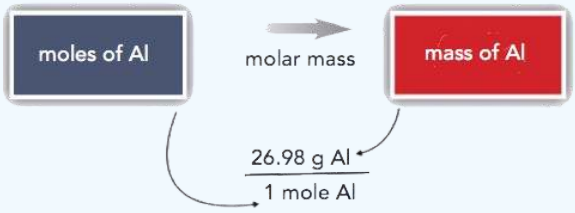
The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do "by eye."

Example 3.11.1

What is the mass of 3.987 mol of Al?

Solution

The first step in a conversion problem is to decide what conversion factor to use. Because we are starting with mole units, we want a conversion factor that will cancel the mole unit and introduce the unit for mass in the numerator. Therefore, we should use the $\frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$ conversion factor. We start with the given quantity and multiply by the conversion factor:



$$3.987 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$$

Note that the mol units cancel algebraically. (The quantity 3.987 mol is understood to be in the numerator of a fraction that has 1 in the unwritten denominator.) Canceling and solving gives

$$3.987 \text{ mol Al} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 107.6 \text{ g Al}$$

Our final answer is expressed to four significant figures.

Exercise 3.11.1

How many moles are present in 100.0 g of Al? (Hint: you will have to use the other conversion factor we obtained for aluminum.)

Answer

$$100.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 3.706 \text{ mol Al}$$

Conversions like this are possible for any substance, as long as the proper atomic weight, formula weight, or molecular weight is known (or can be determined) and expressed in grams per mole. Figure **3.11.1** is a chart for determining what conversion factor is needed, and Figure **3.11.2** is a flow diagram for the steps needed to perform a conversion.

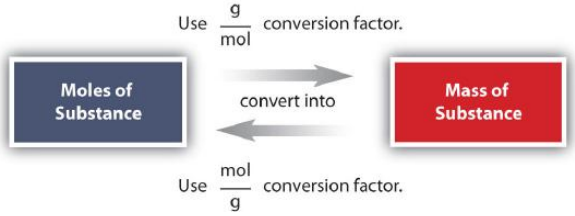


Figure **3.11.1** A Simple Flowchart for Converting between Mass and Moles of a Substance. One mathematical step to convert from moles to mass or from mass to moles.

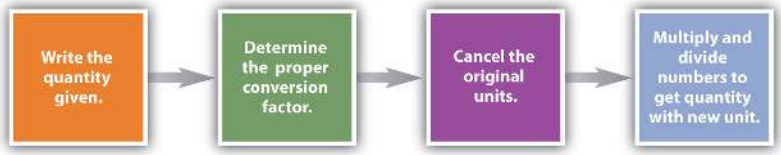


Figure **3.11.2** A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

Example 3.11.2

A biochemist needs 0.00655 mol of bilirubin ($C_{33}H_{36}N_4O_6$) for an experiment. How many grams of bilirubin will that be?

Solution

To convert from moles to mass, we need the molecular weight of bilirubin in grams, which we can determine from its chemical formula:

Solutions to Example 3.11.2

33 C	33 × 12.01 g =	396.33 g
36 H	36 × 1.01 g =	36.36 g
4 N	4 × 14.01 g =	56.04 g
6 O	6 × 16.00 g =	96.00 g
Total		584.73 g

The mass of 1 mole of bilirubin is 584.73 g. Using the relationship

$$1 \text{ mol bilirubin} = 584.73 \text{ g bilirubin}$$

we can construct the appropriate conversion factor for determining how many grams there are in 0.00655 mol. Following the steps from Figure **3.11.2**:

$$0.00655 \text{ mol bilirubin} \times \frac{584.73 \text{ g bilirubin}}{\text{mol bilirubin}} = 3.83 \text{ g bilirubin}$$

The mol bilirubin unit cancels. The biochemist needs 3.83 g of bilirubin.

Exercise 3.11.2

A chemist needs 457.8 g of $KMnO_4$ to make a solution. How many moles of $KMnO_4$ is that?

Answer

$$457.8 \text{ g } KMnO_4 \times \frac{1 \text{ mol } KMnO_4}{158.04 \text{ g } KMnO_4} = 2.897 \text{ mol } KMnO_4$$

To Your Health: Minerals

For our bodies to function properly, we need to ingest certain substances from our diets. Among our dietary needs are minerals, the noncarbon elements our body uses for a variety of functions, such developing bone or ensuring proper nerve transmission. The US Department of Agriculture has established some recommendations for the RDIs of various minerals. The accompanying table lists the RDIs for minerals, both in mass and moles, assuming a 2,000-calorie daily diet.

Table **3.11.1**: Essential Minerals and their Composition in Humans

Mineral	Male (age 19–30 y)		Female (age 19–30 y)	
Ca	1,000 mg	0.025 mol	1,000 mg	0.025 mol
Cr	35 µg	6.7×10^{-7} mol	25 µg	4.8×10^{-7} mol
Cu	900 µg	1.4×10^{-5} mol	900 µg	1.4×10^{-5} mol
F	4 mg	2.1×10^{-4} mol	3 mg	1.5×10^{-4} mol
I	150 µg	1.2×10^{-6} mol	150 µg	1.2×10^{-6} mol
Fe	8 mg	1.4×10^{-4} mol	18 mg	3.2×10^{-4} mol
K	3,500 mg	9.0×10^{-2} mol	3,500 mg	9.0×10^{-2} mol
Mg	400 mg	1.6×10^{-2} mol	310 mg	1.3×10^{-2} mol
Mn	2.3 mg	4.2×10^{-5} mol	1.8 mg	3.3×10^{-5} mol
Mo	45 mg	4.7×10^{-7} mol	45 mg	4.7×10^{-7} mol
Na	2,400 mg	1.0×10^{-1} mol	2,400 mg	1.0×10^{-1} mol
P	700 mg	2.3×10^{-2} mol	700 mg	2.3×10^{-2} mol
Se	55 µg	7.0×10^{-7} mol	55 µg	7.0×10^{-7} mol
Zn	11 mg	1.7×10^{-4} mol	8 mg	1.2×10^{-4} mol

Table **3.11.1** illustrates several things. First, the needs of men and women for some minerals are different. The extreme case is for iron; women need over twice as much as men do. In all other cases where there is a different RDI, men need more than women.

Second, the amounts of the various minerals needed on a daily basis vary widely—both on a mass scale and a molar scale. The average person needs 0.1 mol of Na a day, which is about 2.5 g. On the other hand, a person needs only about 25–35 µg of Cr per day, which is under one millionth of a mole. As small as this amount is, a deficiency of chromium in the diet can lead to diabetes-like symptoms or neurological problems, especially in the extremities (hands and feet). For some minerals, the body does not require much to keep itself operating properly.

Although a properly balanced diet will provide all the necessary minerals, some people take dietary supplements. However, too much of a good thing, even minerals, is not good. Exposure to too much chromium, for example, causes a skin irritation, and certain forms of chromium are known to cause cancer (as presented in the movie *Erin Brockovich*).

Key Takeaway

Conversions between moles and mass of a compound can be achieved in one mathematical step using the atomic weight of the compound in grams.

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3.12: End of Chapter Problems

Ions

1. Identify each as a cation, an anion, or molecule.

a. H^+

b. Cl^-

c. O_2

d. Ba^{2+}

e. CH_4

f. CS_2

g. NH_3

h. Br^-

i. H^-

j. Hg^{2+}

k. CCl_4

l. SO_3

2. What are ionophores? Explain how ionophores act as antibiotics.

Main Group Elements

1. Which of these ions is *not likely* to form?

a. Mg^+

b. K^+

c. S^{3-}

d. N^{3-}

2. When the following atoms become ions, what charges do they acquire?

a. Li

b. S

c. Ca

d. F

3. What characteristic charge do atoms in the first column of the periodic table have when they become ions?

4. What characteristic charge do atoms in the second column of the periodic table have when they become ions?

5. What characteristic charge do atoms in the third-to-last column of the periodic table have when they become ions?

6. What characteristic charge do atoms in the next-to-last column of the periodic table have when they become ions?

7. Starting with lithium and bromine atoms, use Lewis diagrams to show the formation of the ionic compound.

a. LiBr

b. MgO

Nomenclature of Ions

1. Name each ion.

a. Ca^{2+}

b. S^{2-}

c. SO_4^{2-}

d. NH_4^+

e. Ba^{2+}

f. HCO_3^-

g. Ca^{2+}

h. P^{3-}

i. $H_2PO_4^-$

j. Sn^{4+}

k. K^+

l. Fe^{3+}

m. HSO_4^-

n. Sn^{2+}

o. Ca^{2+}

p. S^{2-}

q. SO_3^{2-}

r. NH_4^+

s. Cu^+

t. Fe^{2+}

2. Write the formula for each ion.

a. the bromide ion

b. the phosphate ion

c. the cupric ion

d. the magnesium ion

e. the fluoride ion

f. the carbonate ion

g. the ferrous ion

h. the potassium ion

i. the bromide ion

j. the phosphate ion

k. the cupric ion

l. the magnesium ion

m. the fluoride ion

n. the carbonate ion

o. the potassium ion

Formula of Ionic Compounds

1. Starting with lithium and bromine atoms, use Lewis diagrams to show the formation of the ionic compounds.

LiBr

MgO

2. Write the chemical formula for an ionic compound composed of each pair of ions.

a. the 3+ iron ion and the oxide ion

b. the calcium ion and the oxide ion

c. the 2+ copper ion and the sulfide ion

d. the 1+ copper ion and the sulfide ion

e. the potassium ion and the sulfate ion

f. the calcium ion and the nitrate ion

g. the magnesium ion and the carbonate ion

h. the aluminum ion and the acetate ion

i. the sodium ion and the sulfide ion

j. the aluminum ion and the fluoride ion

Nomenclature of Ionic Compounds

1. Name each ionic compound using both Stock and common systems if necessary.

$Ca_3(PO_4)_2$

$(NH_4)_2Cr_2O_7$

KCl

CuCl

SnF_2

$ZnBr_2$

$Fe(NO_3)_3$

Al_2O_3

CuF_2

AgF

$Ca_3(PO_4)_2$

$(NH_4)_2Cr_2O_7$

KCl

CuCl

SnF_2

$ZnBr_2$

$Fe(NO_3)_3$

Al_2O_3

CuF_2

AgF

2. Give two names for each compound

a. $Al(HSO_4)_3$

b. $Mg(HSO_4)_2$

c. $Co(HCO_3)_2$

d. $LiHCO_3$

Formula and Nomenclature of Ionic Compounds

Give the formula and the name of the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.

a. Na^+ and Br^-

b. Mg^{2+} and Br^-

c. Mg^{2+} and S^{2-}

d. K^+ and Cl^-

e. Mg^{2+} and Cl^-

f. Mg^{2+} and F^-

g. Na^+ and N^{3-}

h. Mg^{2+} and N^{3-}

i. Al^{3+} and S^{2-}

j. Li^+ and N^{3-}

k. Mg^{2+} and P^{3-}

l. Li^+ and P^{3-}

m. Fe^{3+} and Br^-

n. Fe^{2+} and Br^-

o. Fe^{3+} and S^{2-}

p. Cu^+ and S^{2-}

q. Cr^{3+} and O^{2-}

r. Cr^{2+} and O^{2-}

s. Pb^{2+} and Cl^-

t. Pb^{4+} and Cl^-

u. Cr^{3+} and NO_3^-

v. Fe^{2+} and PO_4^{3-}

w. Ca^{2+} and SO_4^{2-}

x. Al^{3+} and OH^-

y. NH_4^+ and NO_3^-

z. K^+ and SO_4^{2-}

Covalent Bonds

Predict the number of covalent bonds formed by each nonmetal atom.

N

Cl

C

O

Br

H

Ionic vs. Molecular Compounds

1. Which compounds would you predict to be ionic? Circle.

a. Li_2O

b. $(NH_4)_2S$

c. CO_2

d. $FeSO_3$

e. C_6H_6

f. C_2H_6O

g. $Ba(OH)_2$

h. CH_2O

i. NH_2CONH_2

j. $(NH_4)_2CrO_4$

k. C_8H_{18}

l. NH_3

2. Identify each compound as ionic or molecular.

a. Na_2O

b. PCl_3

c. NH_4Cl

d. OF_2

e. N_2O

f. $FeCl_3$

g. $(NH_4)_3PO_4$

h. $SOCl_2$

3. Is each compound formed from ionic bonds, covalent bonds, or both?

a. Na_2O

b. Na_3PO_4

c. N_2O_4

d. $Ba(OH)_2$

e. F_2

f. PCl_3

4. Write the molecular formula for each compound.

a. chlorine trifluoride
b. phosphorus pentachloride
c. sulfur dioxide

d. nitrogen dioxide

e. dioxygen difluoride

f. dinitrogen pentoxide

g. sulfur hexafluoride

h. nitrogen monoxide

5. Write the name for each compound.

a. BrF_5
b. S_2F_2
c. CO

d. CF_4

e. $SeCl_2$

f. SO_3

Formula Weight and Molecular Weight

1. Use the atomic weights rounded to two decimal places to determine the formula weight for each ionic compound.

a. TiO_2
b. AgBr
c. $Au(NO_3)_3$
d. $Fe_3(PO_4)_2$

2. Calculate the molecular weight of each covalent compound? Use the atomic weights rounded to two decimal places.

a. C_2H_2

b. CO

c. CO_2

d. BF_3

Mass-Mole Conversions

1. A biochemist needs 0.00655 mol of bilirubin ($C_{33}H_{36}N_4O_6$) for an experiment. How many grams of bilirubin will that be?

2. A chemist needs 457.8 g of $KMnO_4$ to make a solution. How many moles of $KMnO_4$ is that?

3. What is the mass of 62.5 moles of carbon tetrachloride?

4. How many moles of ammonium hydroxide are present in 0.475 g?

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CHAPTER OVERVIEW

4: Organic Chemistry-An Introduction

4.1: Molecular Formula

4.2: Isomers

4.3: Condensed Formula and Skeletal Structures

4.4: Shapes of Molecules

4.5: Electronegativity, Bond Polarity, and Overall Polarity

4.6: Intermolecular Interactions

4.7: Functional Groups or Families of Compounds

4.8: Cis-Trans Isomers (Geometric Isomers)

4.9: End of Chapter Problems

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4.1: Molecular Formula

Learning Objectives

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

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

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

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

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

$$Ca_3(PO_4)_2$$

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

Molecular Formula

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

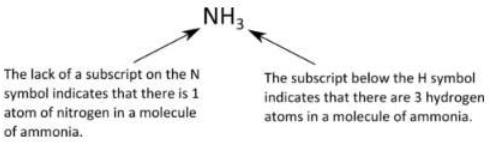


Figure 4.1.1: The molecular formula for ammonia.

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

Empirical Formula

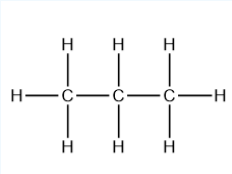
An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical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

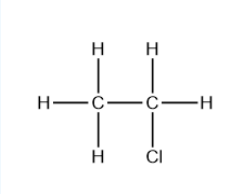
- molecular formula: $C_6H_{12}O_6$
- empirical formula: CH_2O

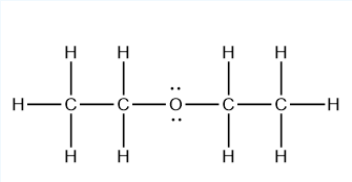
There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.

Exercise 4.1.1

1. Write the molecular formula for the following compounds shown in their line bond structure.

a. 

b. 

c. 

Answer

List the atoms and the number of the atoms in the compound.

a. C_3H_8

b. C_2H_5Cl

c. $C_4H_{10}O$

Summary

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

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4.2: Isomers

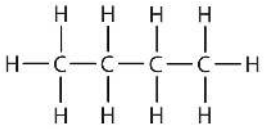
Learning Objectives

- Recognize compounds that are isomers.

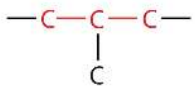
We can write the structure of butane (C_4H_{10}) by stringing four carbon atoms in a row,

–C–C–C–C–

and then adding enough hydrogen atoms to give each carbon atom four bonds:



The compound butane has this structure, but there is another way to put 4 carbon atoms and 10 hydrogen atoms together. Place 3 of the carbon atoms in a row and then branch the fourth one off the middle carbon atom to make an isomer of butane called isobutane.



Now we add enough hydrogen atoms to give each carbon four bonds.

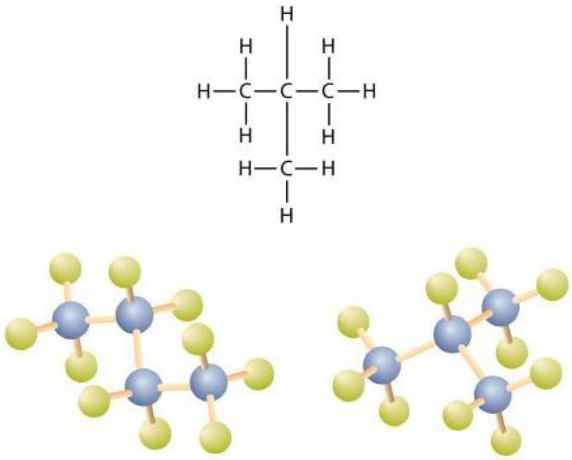


Figure **4.2.1**: Butane and Isobutane. The ball-and-stick models of these two compounds show them to be isomers; both have the molecular formula C_4H_{10}.

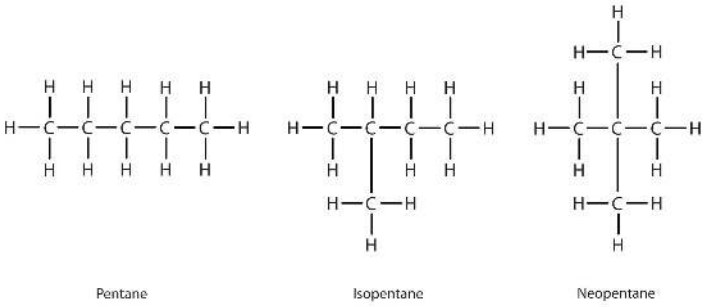
The structure of isobutane shows a continuous chain of three carbon atoms only, with the fourth attached as a branch off the middle carbon atom of the continuous chain, which is different from the structures of butane (compare the two structures in Figure **4.2.1**.

Butane and Isobutane are isomers. Isomers are compounds that have the same molecular formula, but different structural formula. Both butane and isobutane have the same molecular formula C_4H_{10}. However they differ in the connectivity of the atoms. The line bond formula of the two isomers are different and hence they are different compounds.

Unlike C_4H_{10}, the compounds methane (CH_4), ethane (C_2H_6), and propane (C_3H_8) do not exist in isomeric forms because there is only one way to arrange the atoms in each formula so that each carbon atom has four bonds.

Pentane has three isomers. Each compound has the same molecular formula: C_5H_{12}. The compound on the left is pentane. It has all five carbon atoms in a continuous chain. The compound in the middle is isopentane; like isobutane, it has a one CH_3 branch off the second carbon atom of the continuous chain. The compound on the right, discovered after the other two, was named neopentane

(from the Greek *neos*, meaning "new"). Although all three have the same molecular formula, they have different properties, including boiling points: pentane, 36.1°C; isopentane, 27.7°C; and neopentane, 9.5°C.

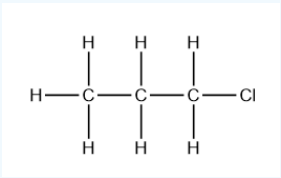


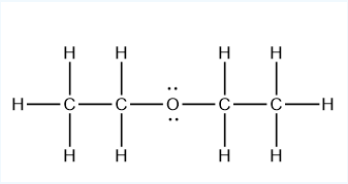
Pentane Isopentane Neopentane

A continuous (unbranched) chain of carbon atoms is often called a *straight chain* even though the tetrahedral arrangement about each carbon gives it a zigzag shape.

Exercise 4.2.1

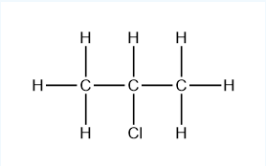
Draw an isomer of the following compounds.

a. 

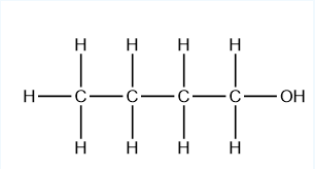
b. 

Answer

a. Move the chlorine to the second carbon atom to obtain an isomer. This is the only isomer possible for this compound.



b. Move the oxygen to a different carbon atom to obtain an isomer. There are 2 other isomers possible.



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4.3: Condensed Formula and Skeletal Structures

Learning Objectives

- Write condensed structural formulas for compounds given line bond formulas.
- Draw skeletal structures given line bond or condensed formulas.

We use several kinds of formulas to describe organic compounds. A *molecular formula* shows only the kinds and numbers of atoms in a molecule. For example, the molecular formula C_4H_{10} tells us there are 4 carbon atoms and 10 hydrogen atoms in a molecule, but it doesn't distinguish between butane and isobutane. A line bond formula shows all the carbon and hydrogen atoms and the bonds attaching them. Thus, line bond formulas show the order of attachment of the various atoms.

Unfortunately, line bond formulas are difficult to type/write and take up a lot of space. Chemists often use condensed structural formulas to alleviate these problems. The condensed formulas show hydrogen atoms right next to the carbon atoms to which they are attached, as illustrated for butane:



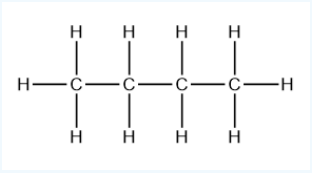
The ultimate condensed formula is a skeletal structure, in which carbon atoms are implied at the corners and ends of lines, and each carbon atom is understood to be attached to enough hydrogen atoms to give each carbon atom four bonds. For example, we can represent pentane ($CH_3CH_2CH_2CH_2CH_3$) and isopentane [$(CH_3)_2CHCH_2CH_3$] as follows:

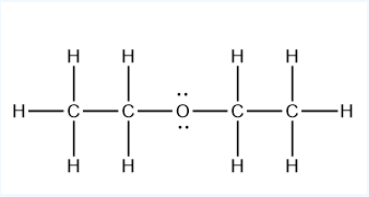


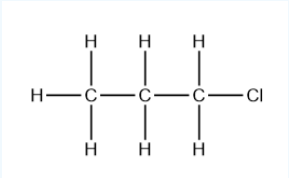
Pentane Isopentane

Exercise 4.3.1

Draw condensed formula for the following compounds shown in their line bond structure.

a. 

b. 

c. 

Answer

a. $CH_3CH_2CH_2CH_3$

b. $CH_3CH_2OCH_2CH_3$

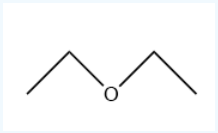
c. $CH_3CH_2CH_2Cl$

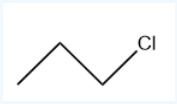
? Exercise 4.3.2

Draw skeletal formula for the compounds in Exercise 4.3.1 shown in their line bond structure.

Answer

a. 

b. 

c. 

In the skeletal formula, the hydrogen atoms are not shown, the carbons are the corners, atoms such as oxygen, nitrogen or chlorine need to be shown.

Key Takeaways

- Condensed chemical formulas show the hydrogen atoms (or other atoms or groups) right next to the carbon atoms to which they are attached.
- Skeletal formulas imply a carbon atom at the corners and ends of lines. Each carbon atom is understood to be attached to enough hydrogen atoms to give each carbon atom four bonds.

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4.4: Shapes of Molecules

Learning Objectives

- Predict the general shape of a simple covalent compounds.

Molecular Shape

Ionic compounds have extended crystal lattices. Atoms in a molecule or polyatomic ion are arranged into geometric patterns that allow their electron pairs to get as far away from each other as possible. This minimizes the repulsive forces between them. Covalent bonds, are composed of negatively charged electrons, that tend to repel one another. This concept is called the **valence shell electron pair repulsion (VSEPR) theory**.

Most simple geometric structures fall into the following patterns:

Tetrahedral

When four pairs of shared electrons for a total of 4 electron groups are around a central atom the result is a tetrahedral shape. CH_4, CCl_4, $CHCl_3$, CH_2Cl_2 etc., all have tetrahedral geometry around the central atom.

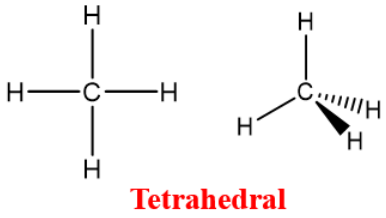


Figure **4.4.1**: Tetrahedral shape of CH_4.

Pyramidal

When there are three pairs of shared electrons and one lone pair of unshared electrons for a total of 4 electron groups around a central atom the result is a pyramidal shape. NH_3 has **one lone electron pair** and **three bonded electron pairs**. These four electron pairs repel each other and adopt a tetrahedral arrangement. However, the shape of the molecule is described in terms of the positions of the atoms, not the lone electron pairs. Thus, NH_3 is said to have a ***pyramidal*** shape, not a tetrahedral one.

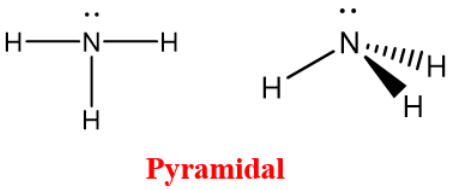


Figure **4.4.2**: Pyramidal shape of NH_3.

Planar

When there are three groups of shared electrons around a central atom, two of these groups are single bonds and one group is a double bond made up of two pairs of shared electrons. There are no lone pairs of electrons around the central atom.

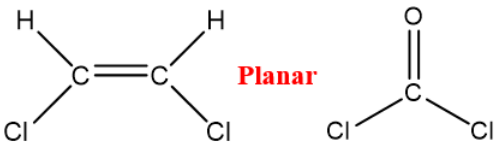


Figure **4.4.3**: Planar shape around the central atom carbon.

The shapes of molecules with multiple bonds are determined by treating the double bond as one group. Thus, formaldehyde (CH_2O) has a ***planar*** shape.

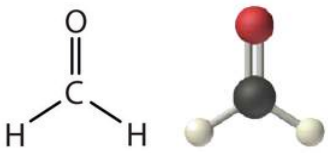


Figure **4.4.4**: Planar shape of CH_2O.

Bent

When two groups of shared electrons in single bonds and two pairs of unshared electrons are around a central atom the result is the bent shape around the central atom. An important example is the bent shape around the oxygen central atom in water (H_2O).

Or

When two groups of shared electrons in a single bond and a double bond and one pair of unshared electrons are around a central atom the result is also a bent geometry around the central atom.

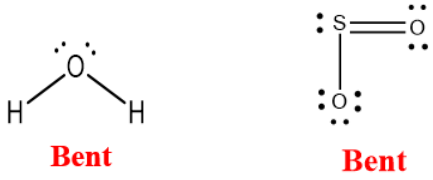


Figure **4.4.5**: Bent shape of H_2O and SO_2.

Linear

When there are two groups of shared electrons, usually double bonds with two shared electron pairs between two atoms, and no unshared electrons around a central atom the geometry is linear.

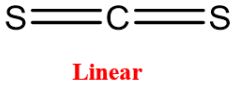


Figure **4.4.6**: Linear shape of CS_2.

When there are two atoms in a molecule or ion, and there is no central atom (HBr for example), the geometry is also linear.

In summary, to determine the molecular geometry:

Step 1: Draw the line bond structure.

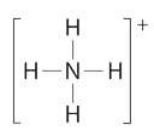
Step 2: Count the number of bonds (a double/triple bond counts as one) and lone pairs around the **central atom**.

Step 3: Determine the molecular geometry. The molecular geometry depends on the number of bonds and/or lone pairs around the **central atom.**

If a molecule has 2 bonds and 0 lone pairs, it is linear. If a molecule has 2 bonds and 1 lone pair, it is bent or angular. If a molecule has 4 bonds and 0 lone pairs, it is tetrahedral. If a molecule has 3 bonds and 1 lone pair, it is a pyramid. If a molecule has 2 bonds and 2 lone pairs, it is bent or angular.

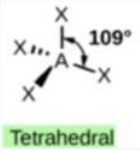
Example 4.4.1

What is the geometry of the ammonium ion, NH_4^+? Its Lewis structure is shown below. How is this different from ammonia, NH_3?

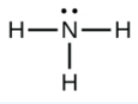


Solution

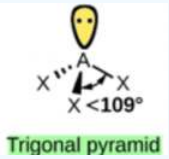
In ammonium ion, the central atom N has 4 bonds and no lone pair. Hence, this is *tetrahedral*.



In ammonia (NH_3), shown below, N has 3 bonds and one lone pair.

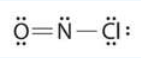


Hence, the shape of this molecule is *pyramidal*.



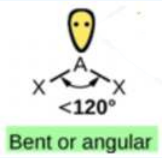
4.4.1 Exercise

What is the molecular shape of nitrosyl chloride, a highly corrosive, reddish-orange gas? Its Lewis structure is shown below.



Answer

Focus on the central atom, N. It has a double bond to O, count this as one bond. It also has a single bond to Cl. Thus, N has 2 bonds and one lone pair. These 3 electron pairs will spread out 120 degrees from each other. But, since the shape is defined by the arrangement of the atoms only, the shape is bent or angular.



Biochemistry Link: Prions

The shape that a molecule takes can influence its biological function. Prions (proteinaceous infection particles) are a class of proteins that have been identified as the cause of mad cow disease in cattle and scrapie in sheep. Protein are large macromolecules made of amino acids. They have a favored shape and fold into a compact structure. In prion diseases the normal prion PrP^{C} is twisted into an infectious abnormal shape PrP^{SC}. PrP^{SC} induces PrP^{C} to change shape and become PrP^{SC}. As PrP^{SC} accumulates in animals, sponge like holes form in the brain causing dizziness, seizures, and death. Prion disease is infectious and can be transferred from one species to another. Infections in humans are linked to eating mad cow infected beef.



Figure **4.4.7**: Models of PrP^{C} (left) and PrP^{SC}(right)

Key Takeaways

- Simple molecules have geometries around a central atom.

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4.5: Electronegativity, Bond Polarity, and Overall Polarity

Learning Objectives

- Use electronegativity values to predict bond polarity.

Electronegativity and Bond Polarity

Although we defined covalent bonding as electron sharing, the electrons in a covalent bond are not always shared equally by the two bonded atoms. Unless the bond connects two atoms of the same element, as in H_2, there will always be one atom that attracts the electrons in the bond more strongly than the other atom does, as in HCl, shown in Figure **4.5.1**. A covalent bond that has an equal sharing of electrons (Figure **4.5.1***a*) is called a **nonpolar covalent bond**. A covalent bond that has an unequal sharing of electrons, as in Figure **4.5.1***b*, is called a **polar covalent bond**.

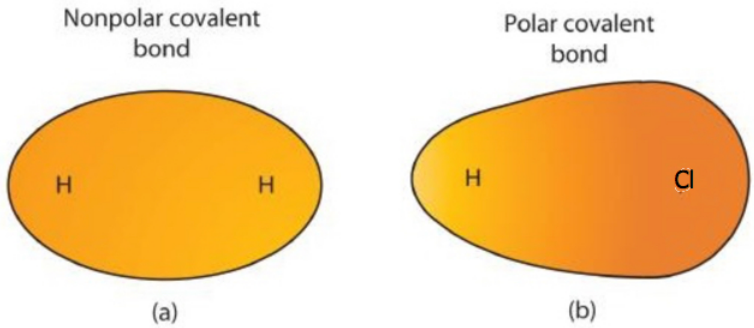


Figure **4.5.1** Polar versus Nonpolar Covalent Bonds. (a) The electrons in the covalent bond are equally shared by both hydrogen atoms. This is a nonpolar covalent bond. (b) The chlorine atom attracts the electrons in the bond more than the hydrogen atom does, leading to an imbalance in the electron distribution. This is a polar covalent bond.

The distribution of electron density in a polar bond is uneven. It is greater around the atom that attracts the electrons more than the other. For example, the electrons in the H–Cl bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom. Note that the shaded area around Cl in Figure **4.5.1***b* is much larger than it is around H.

This imbalance in electron density results in a buildup of partial negative charge (designated as δ−) on one side of the bond (Cl) and a partial positive charge (designated δ+) on the other side of the bond (H). This is seen in Figure **4.5.2***a*. The separation of charge in a polar covalent bond results in an electric dipole (two poles), represented by the arrow in Figure **4.5.2***b*. The direction of the arrow is pointed toward the δ− end while the + tail of the arrow indicates the δ+ end of the bond.

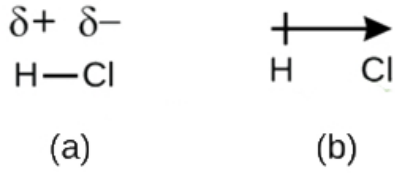


Figure **4.5.2**: (a) Unequal sharing of the bonding pair of electrons between H and Cl leads to partial positive charge on the H atom and partial negative charge on the Cl. Symbols δ+ and δ– indicate the polarity of the H–Cl bond. (b) The dipole is represented by an arrow with a cross at the tail. The cross is near the δ+ end and the arrowhead coincides with the δ–.

Any covalent bond between atoms of different elements is a polar bond, but the degree of polarity varies widely. Some bonds between different elements are only minimally polar, while others are strongly polar. Ionic bonds can be considered the ultimate in polarity, with electrons being transferred rather than shared. To judge the relative polarity of a covalent bond, chemists use **electronegativity**, which is a relative measure of how strongly an atom attracts electrons when it forms a covalent bond. There are various numerical scales for rating electronegativity. Figure **4.5.3** shows one of the most popular—the Pauling scale.

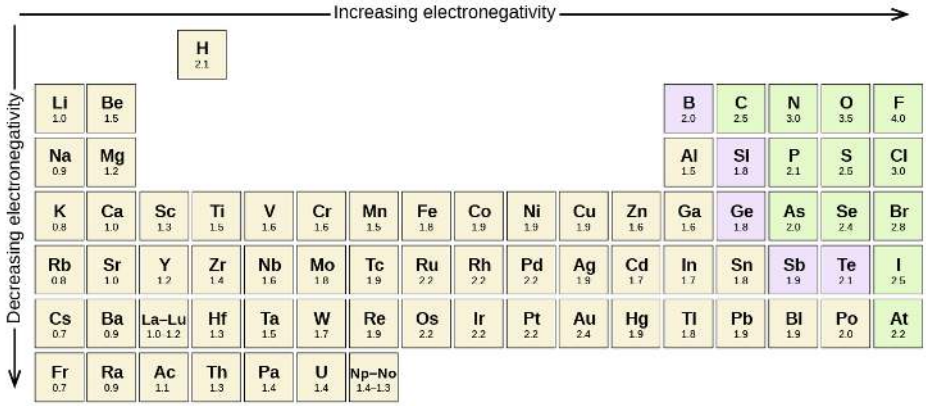
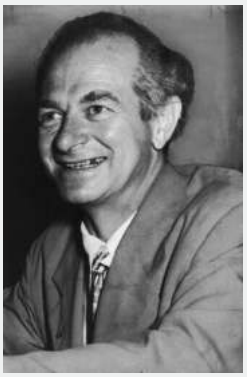


Figure **4.5.3** The electronegativity values derived by Pauling follow predictable periodic trends with the higher electronegativities toward the upper right of the periodic table. Fluorine has the highest value (4.0).

Looking Closer: Linus Pauling

Arguably the most influential chemist of the 20th century, Linus Pauling (1901–94) is the only person to have won two individual (that is, unshared) Nobel Prizes. In the 1930s, Pauling used new mathematical theories to enunciate some fundamental principles of the chemical bond. His 1939 book *The Nature of the Chemical Bond* is one of the most significant books ever published in chemistry.

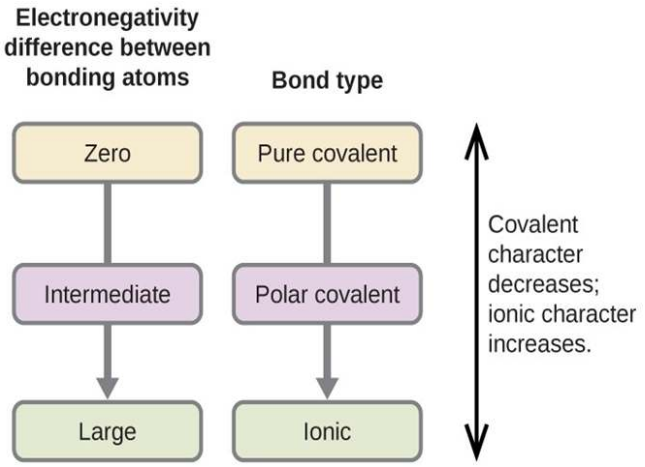
By 1935, Pauling's interest turned to biological molecules, and he was awarded the 1954 Nobel Prize in Chemistry for his work on protein structure. (He was very close to discovering the double helix structure of DNA when James Watson and James Crick announced their own discovery of its structure in 1953.) He was later awarded the 1962 Nobel Peace Prize for his efforts to ban the testing of nuclear weapons.



Linus Pauling was one of the most influential chemists of the 20th century.

In his later years, Pauling became convinced that large doses of vitamin C would prevent disease, including the common cold. Most clinical research failed to show a connection, but Pauling continued to take large doses daily. He died in 1994, having spent a lifetime establishing a scientific legacy that few will ever equal.

The polarity of a covalent bond can be judged by determining the difference in the electronegativities of the two atoms making the bond. The greater the difference in electronegativities, the greater the imbalance of electron sharing in the bond. Although there are no hard and fast rules, the general rule is if the difference in electronegativities is **less than about 0.5,** the bond is considered **nonpolar covalent bond;** if the difference is **0.5 or greater,** the bond is considered **polar covalent bond.** If the difference in electronegativities is large enough **(**generally **greater than about 1.9),** the resulting compound is considered **ionic** rather than covalent**.** An electronegativity difference of **zero,** of course, indicates a **nonpolar covalent bond.**



Bond Type	Electronegativity Difference
pure covalent	less than 0.5
polar covalent	between 0.5 and 1.9
ionic	greater than 1.9

Figure **4.5.4**: As the electronegativity difference increases between two atoms, the bond becomes more ionic.

✔ Example 4.5.1

Describe the electronegativity difference between each pair of atoms and the resulting polarity (or bond type).

a. C and H
b. H and H
c. Na and Cl
d. O and H

Solution

a. Carbon has an electronegativity of 2.5, while the value for hydrogen is 2.1. The difference is 0.4, which is rather small. The C–H bond is therefore considered nonpolar.
b. Both hydrogen atoms have the same electronegativity value—2.1. The difference is zero, so the bond is nonpolar.
c. Sodium's electronegativity is 0.9, while chlorine's is 3.0. The difference is 2.1, which is rather high, and so sodium and chlorine form an ionic compound.
d. With 2.1 for hydrogen and 3.5 for oxygen, the electronegativity difference is 1.4. We would expect a very polar bond. The sharing of electrons between O and H is unequal with the electrons more strongly drawn towards O.

? Exercise 4.5.1

Describe the electronegativity (EN) difference between each pair of atoms and the resulting polarity (or bond type).

a. C and O
b. K and Br
c. N and N
d. Cs and F

Answer a:

The EN difference is 1.0 , hence polar. The sharing of electrons between C and O is unequal with the electrons more strongly drawn towards O.

Answer b:

The EN difference is greater than 1.8, hence ionic.

Answer c:

Identical atoms have zero EN difference, hence nonpolar.

Answer d:

The EN difference is greater than 1.8, hence ionic.

Bond Polarity and Molecular Polarity

If there is only one bond in the molecule, the bond polarity determines the molecular polarity. Any diatomic molecule in which the two atoms are the same element must be a **nonpolar** molecule. Examples are I_2, O_2, and H_2.

A diatomic molecule that consists of a polar covalent bond, such as HF, is a **polar molecule**. A **polar molecule** *is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative*. 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. Hence, a *molecule with two poles is called a* **dipole**. A simplified way to depict polar molecules like HF is pictured below (see figure below).

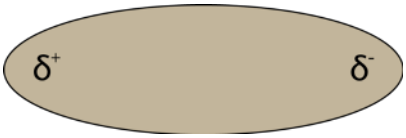


Figure 4.5.5: A molecular dipole results from the unequal distribution of electron density throughout a molecule.

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 4.4.6 below).

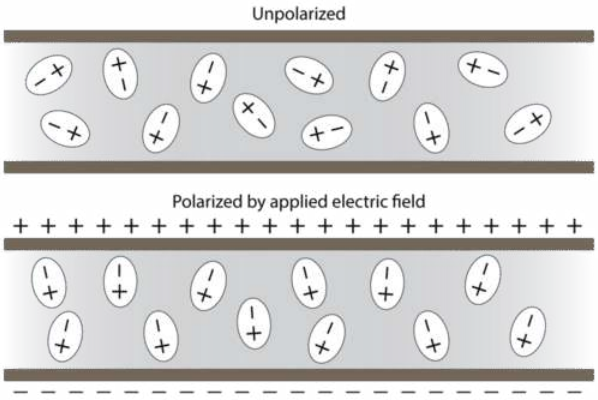


Figure 4.5.6: Polar molecules are randomly oriented in the absence of an applied electric field (top). In an electric field, the molecules orient themselves to maximize the attraction between opposite charges (bottom).

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. (Figure 4.5.7) 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 directly opposite each other, they cancel each other out, and the overall molecular polarity of CO_2 is zero. CO_2 is a **nonpolar** molecule. In H_2O, the orientation of the two O–H bonds is **bent**: one end of the molecule has a partial positive charge, and the other end has a partial negative charge. In short, the molecule itself is **polar**. The polarity of water has an enormous impact on its physical and chemical properties. For example, the boiling point of water (100°C) is high for such a small molecule due to the fact that polar molecules attract each other strongly. On the other hand, the nonpolar carbon dioxide becomes a gas at −77°C, almost 200° lower than the temperature at which water boils.



Figure 4.5.7 Physical Properties and Polarity. The physical properties of water and carbon dioxide are affected by their polarities.

Similarly, in BF_3 (planar), the effect of a B-F bond is cancelled by the sum of the other two B-F bonds. Hence, a planar molecule (BF_3) is nonpolar because the bond polarities cancel each other, but a pyramidal molecule(NH_3) is polar.

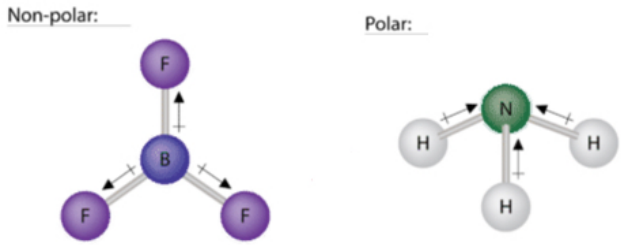


Figure 3.5.8: The molecular geometry of a molecule (planar vs. pyramid) affects its polarity.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as CCl_4 is ***nonpolar***. However, if the peripheral atoms are not of the same electronegativity, the bond polarities don't cancel and the molecule becomes ***polar***, as in CH_3Cl.

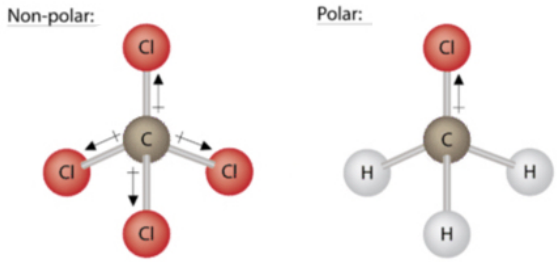


Figure 3.5.9: *The same molecular geometry but peripheral bonds are of different electronegativity. CCl_4 is nonpolar but CH_3Cl is polar*.

Key Takeaways

- Covalent bonds between different atoms have different bond lengths.
- Covalent bonds can be polar or nonpolar, depending on the electronegativity difference between the atoms involved.

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4.6: Intermolecular Interactions

Learning Objectives

- Define *phase*.
- Identify the types of interactions between molecules.

A phase is a certain form of matter that includes a specific set of physical properties. That is, the atoms, the molecules, or the ions that make up the phase do so in a consistent manner throughout the phase. Science recognizes three stable phases: the ***solid** phase*, in which individual particles can be thought of as in contact and held in place; the ***liquid** phase*, in which individual particles are in contact but moving with respect to each other; and the ***gas** phase*, in which individual particles are separated from each other by relatively large distances. Not all substances will readily exhibit all phases. For example, carbon dioxide does not exhibit a liquid phase unless the pressure is greater than about six times normal atmospheric pressure. Other substances, especially complex organic molecules, may decompose at higher temperatures, rather than becoming a liquid or a gas.

Which phase a substance adopts depends on the pressure and the temperature it experiences. Of these two conditions, temperature variations are more obviously related to the phase of a substance. When it is very cold, H_2O exists in the solid form as ice. When it is warmer, the liquid phase of H_2O is present. At even higher temperatures, H_2O boils and becomes steam. Pressure changes can also affect the presence of a particular phase (as we indicated for carbon dioxide), but its effects are less obvious most of the time. We will mostly focus on the temperature effects on phases. Most chemical substances follow the same pattern of phases when going from a low temperature to a high temperature: the solid phase, then the liquid phase, and then the gas phase. However, the temperatures at which these phases are present differ for all substances and can be rather extreme. Table **4.6.1** shows the temperature ranges for solid, liquid, and gas phases for three substances. As you can see, there is extreme variability in the temperature ranges.

Table **4.6.1**: Temperature Ranges for the Three Phases of Various Substances

Substance	Solid Phase Below	Liquid Phase Above	Gas Phase Above
hydrogen (H_2)	−259°C	−259°C	−253°C
water (H_2O)	0°C	0°C	100°C
sodium chloride (NaCl)	801°C	801°C	1413°C

The *melting point* of a substance is the temperature that separates a solid and a liquid. The *boiling point* of a substance is the temperature that separates a liquid and a gas.

What accounts for this variability? Why do some substances become liquids at very low temperatures, while others require very high temperatures before they become liquids? It all depends on the strength of the **intermolecular forces (IMF)** between the **particles** of substances and the **kinetic energies** (KE) of its molecules. Substances that experience **strong intermolecular interactions** require **higher temperatures to become liquids and, finally, gases.** Substances that experience **weak intermolecular interactions** do not need much energy (as measured by temperature) to become liquids and gases and will exhibit these phases at **lower temperatures.**

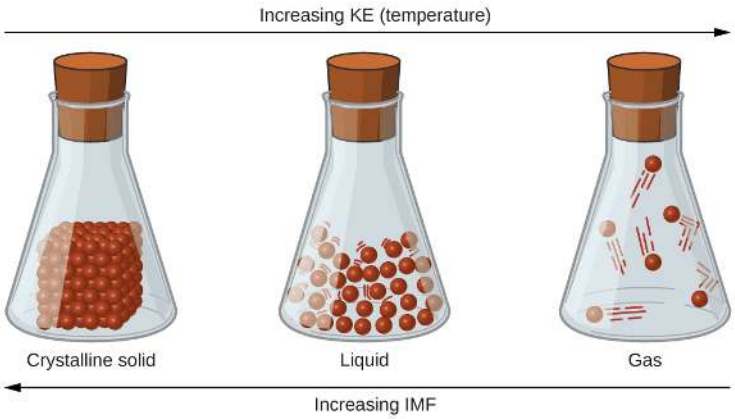


Figure 4.6.1: Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

Three sealed flasks are labeled, “Crystalline solid,” “Liquid,” and “Gas,” from left to right. The first flask holds a cube composed of small spheres sitting on the bottom while the second flask shows a lot of small spheres in the bottom that are spaced a small distance apart from one another and have lines around them to indicate motion. The third flask shows a few spheres spread far from one another with larger lines to indicate motion. There is a right-facing arrow that spans the top of all three flasks. The arrow is labeled, “Increasing K E (temperature).” There is a left-facing arrow that spans the bottom of all three flasks. The arrow is labeled, “Increasing IMF.”

Covalent Molecular Compounds

Molecules or molecular compounds are made of nonmetals connected via a covalent bond. The bonding between atoms in the individual molecule is a covalent bond. The attractive forces between the molecules are called ***intermolecular forces (IMF)***.

The covalent bonds hold atoms together in molecules and polyatomic ion. Intermolecular forces hold molecules together in a **liquid or solid**. Intermolecular forces are generally much weaker than covalent bonds. For example, it requires 927 kJ to break both O–H bonds in 1 mol of water, but it takes only about 41 kJ to overcome the intermolecular attractions and convert 1 mol of liquid water to water vapor at 100°C. Despite this seemingly low value, the intermolecular forces in liquid water are among the strongest such forces between molecules known. Changes between the solid, liquid, and gaseous states almost invariably occur for molecular substances ***without breaking covalent bonds.***

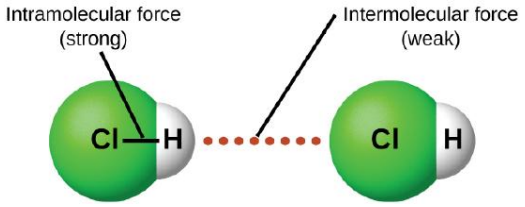


Figure 4.6.2: **Intramolecular** forces keep a molecule intact. **Intermolecular** forces hold multiple molecules together and determine many of a substance’s properties.

In this section, we will discuss the three types of IMF in molecular compounds:

- dipole-dipole interactions
- hydrogen bonding
- London dispersion forces

Dipole-Dipole Intermolecular Forces

A covalent bond that has an unequal sharing of electrons is called a **polar covalent** bond. A covalent bond that has an equal sharing of electrons is called a **nonpolar covalent** bond. A molecule with a net unequal distribution of electrons in its covalent bonds is a polar molecule. HF is an example of a polar molecule or called a dipole.



Figure **4.6.3**: Polar Covalent Bonds. The electrons in the HF molecule are not equally shared by the two atoms in the bond. The fluorine atom has nine protons in its nucleus, it attracts the negatively charged electrons in the bond more than the hydrogen atom does with its one proton in its nucleus. Thus, electrons are more strongly attracted to the fluorine atom, leading to an imbalance in the electron distribution between the atoms. The fluorine side of the bond picks up a partial overall negative charge (represented by the δ− in the diagram), while the hydrogen side of the bond has an overall partial positive charge (represented by the δ+ in the diagram). Such a bond is called a *polar covalent bond*.

The charge separation in a **polar covalent bond** is not as extreme as is found in ionic compounds, but there is a related result: oppositely charged ends of different molecules will attract each other. This type of intermolecular interaction is called a **dipole-dipole interaction**. Most molecules with polar covalent bonds experience dipole-dipole interactions.

Polar molecules have a partial positive charge on one side and a partial negative charge on the other side of the molecule—a separation of charge called a ***dipole*.** Consider a polar molecule such as hydrogen chloride, HCl. In the HCl molecule, the more electronegative Cl atom bears the partial negative charge, whereas the less electronegative H atom bears the partial positive charge. An attractive force between HCl molecules results from the attraction between the positive end of one HCl molecule and the negative end of another. This attractive force is called a dipole-dipole attraction—the electrostatic force between the partially positive end of one polar molecule and the partially negative end of another, as illustrated in Figure **4.6.4**.

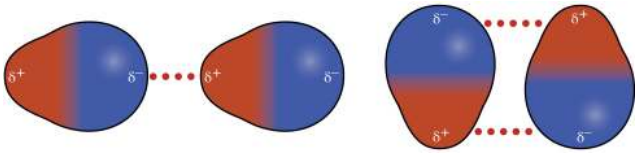


Figure **4.6.4**: This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

Two pairs of molecules are shown where each molecule has one larger blue side labeled “delta sign, negative sign” and a smaller red side labeled “delta sign, positive sign. In the first pair, the red sides of the two molecules both face to the left and the blue side to the right. A horizontal dotted line lies in between the two. In the second pair, the molecules face up and down, with the red and blue ends aligning. A horizontal dotted line lies between the red and blue ends facing upward and another lies between the red and blue ends facing downward.

The effect of a dipole-dipole attraction is apparent when we compare the properties of HCl molecules to nonpolar F_2 molecules. Both HCl and F_2 consist of the same number of atoms and have similar molecular weight. At a temperature of 150 K, molecules of both substances would have the same average kinetic energy. However, the dipole-dipole attractions between HCl molecules are sufficient to cause them to “stick to each other” to form a liquid, whereas the nonpolar F_2 molecules do not experience this effect. So F_2 is a gas at this temperature. The higher normal boiling point of HCl (188 K) compared to F_2 (85 K) is a reflection of the greater strength of dipole-dipole attractions between HCl molecules, compared to the attractions between nonpolar F_2 molecules. We will often use values such as boiling or freezing points as indicators of the relative strengths of IMF's of attraction present within different substances.

Hydrogen Bonding Intermolecular Forces

Molecules with **hydrogen atoms** bonded to electronegative atoms such as **O, N, and F** tend to exhibit unusually strong intermolecular interactions due to a particularly strong type of dipole-dipole attraction called **hydrogen bonding**. The large difference in electronegativity between the H atom (2.1) and the atom to which it is bonded (4.0 for an F atom, 3.5 for an O atom, or 3.0 for a N atom), combined with the very small size of a H atom and the relatively small sizes of F, O, or N atoms, leads to ***highly concentrated partial charges*** with these atoms. 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.

Examples of hydrogen bonds include HF···HF, H_2O···HOH, and H_3N···HNH_2, in which the **hydrogen bonds** are denoted by **dots**. Figure **4.6.5** illustrates hydrogen bonding between water molecules. The physical properties of water, which has two O–H bonds, are strongly affected by the presence of hydrogen bonding between water molecules. 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.

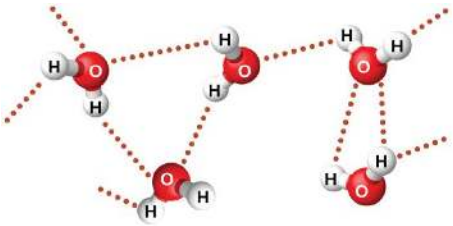


Figure **4.6.5**: Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.
Five water molecules are shown near one another, but not touching. A dotted line lies between many of the hydrogen atoms on one molecule and the oxygen atom on another molecule.

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. Figure **4.6.6** shows how methanol (CH_3OH) molecules experience hydrogen bonding. Methanol contains both a hydrogen atom attached to O; methanol can thus form hydrogen bonds by acting as either a hydrogen bond donor or a hydrogen bond acceptor (lone pair). The hydrogen-bonded structure of methanol is as follows:

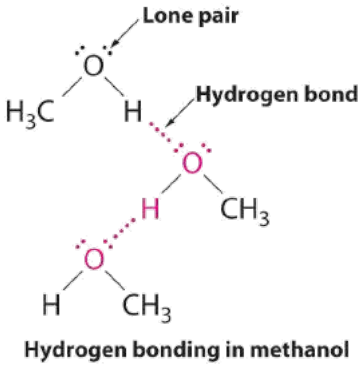


Figure **4.6.6**: Methanol molecule is capable of hydrogen-bonding interactions to other methanol molecules like itself.

Despite use of the word "bond," keep in mind that hydrogen bonds are *intermolecular* attractive forces, not *intramolecular* attractive forces (covalent bonds). Hydrogen bonds are much weaker than covalent bonds, but are generally much stronger than other dipole-dipole attractions.

London Dispersion Forces

Finally, there are forces between all molecules that are caused by electrons being in different places in a molecule at any one time, which sets up a temporary separation of charge that disappears almost as soon as it appears. These are very weak intermolecular interactions and are called **dispersion forces or London forces**. An alternate name is London dispersion forces. Molecules that experience no other type of intermolecular interaction will at least experience dispersion forces. Substances that experience only dispersion forces are typically soft in the solid phase and have relatively low melting points. Examples include *waxes*, which are long hydrocarbon chains that are solids at room temperature because the molecules have so many electrons. The resulting dispersion forces between these molecules make them assume the solid phase at normal temperatures.

Dispersion forces that develop between atoms in different molecules can attract the two molecules to each other. The forces are relatively weak, however, and become significant only when the molecules are very close. **Larger and heavier atoms and molecules** exhibit **stronger dispersion forces** than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in Table 4.6.2.

Table **4.6.2**: **Melting and Boiling Points of the Halogens**

Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K
chlorine, Cl_2	71 g/mol	99 pm	172 K	238 K
bromine, Br_2	160 g/mol	114 pm	266 K	332 K
iodine, I_2	254 g/mol	133 pm	387 K	457 K
astatine, At_2	420 g/mol	150 pm	575 K	610 K

The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther from the nuclei than in a smaller atom. Thus, they are less tightly held and can more easily form the temporary dipoles that produce the attraction. The measure of how easy or difficult it is for another electrostatic charge (for example, a nearby ion or polar molecule) to distort a molecule's charge distribution (its electron cloud) is known as **polarizability**. A molecule that has a charge cloud that is easily distorted is said to be very polarizable and will have large dispersion forces; one with a charge cloud that is difficult to distort is not very polarizable and will have small dispersion forces.

The covalent bonds in some molecules are oriented in space in such a way that the effect of the polar bonds in the molecules cancel each other out. The individual bonds are polar, but due to molecular symmetry, the overall molecule is ***nonpolar*.** Such molecules experience little or no dipole-dipole interactions. Carbon dioxide (CO_2) and carbon tetrachloride (CCl_4) are examples of such molecules (Figure **4.6.7**).

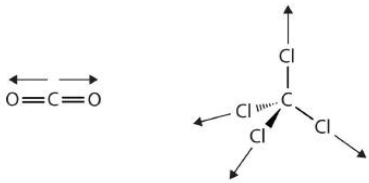


Figure **4.6.7**: Nonpolar Molecules. Although the individual bonds in both CO_2 and CCl_4 are polar, their effects cancel out because of the spatial orientation of the bonds in each molecule. As a result, such molecules experience little or no dipole-dipole interaction.

Ionic Compounds

The strongest force between any two particles is the ionic bond, in which two ions of opposing charge are attracted to each other. Thus, ionic interactions between particles are another type of intermolecular interaction. Substances that contain ionic interactions are relatively strongly held together, so these substances typically have high melting and boiling points. Sodium chloride (Figure **4.6.8**) is an example of a substance whose particles experience ionic interactions (Table **4.6.1**). These attractive forces are sometimes referred to as ion-ion interactions.

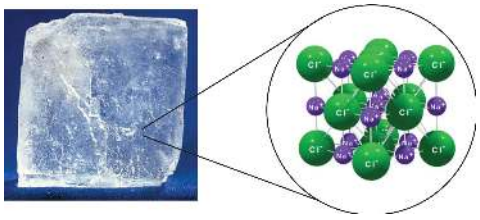


Figure **4.6.8**: Solid NaCl (sodium chloride) is held together by ionic interactions. Source: Photo © Thinkstock

These interactions are common in proteins and result in folding of the protein. Ion-ion interactions are called salt bridges.

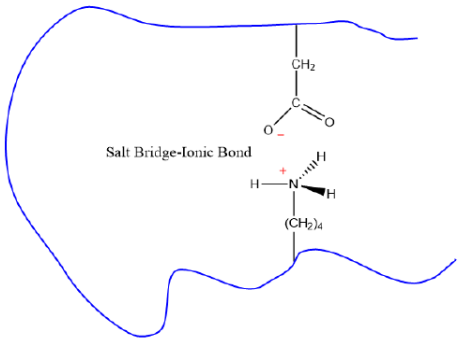


Figure 4.6.9: A protein strand in blue folds into a compact structure via ionic interactions called salt bridges.

Covalent Network Materials

Substances with the highest melting and boiling points have **covalent network** bonding. This type of intermolecular interaction is actually a covalent bond. In these substances, all the atoms in a sample are covalently bonded to one another; in effect, the entire sample is essentially **one giant molecule**. Many of these substances are solid over a large temperature range because it takes a lot of energy to disrupt all the covalent bonds at once. One example of a substance that shows covalent network bonding is diamond (Figure 4.6.9). Diamond is composed entirely of carbon atoms, each bonded to four other carbon atoms in a tetrahedral geometry. Melting a covalent network solid is not accomplished by overcoming the relatively weak intermolecular forces. Rather, all of the covalent bonds must be broken, a process that requires extremely high temperatures. Diamond, in fact, does not melt at all. Instead, it vaporizes to a gas at temperatures above 3,500°C. Diamond is extremely hard and is one of the few materials that can cut glass.

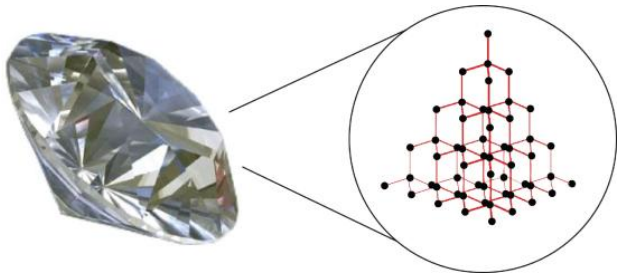


Figure 4.6.10: Diamond is a form of pure carbon, has covalent network bonding. It takes a very high temperature—over 3,500°C—for diamond to leave the solid state. Source: Photo © Thinkstock

Example 4.6.1

Predict which will have the higher boiling point: N_2 or CO. Explain your reasoning.

Solution

CO and N_2 are both diatomic molecules with molecular weights of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N_2 is nonpolar, its molecules cannot exhibit dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N_2 molecules, so CO is expected to have the higher boiling point.

4.6.1 Exercise

Predict which will have the higher boiling point: **ICl** or **Br_2**. Explain your reasoning.

Answer

ICl will have a higher boiling point.

ICl and Br_2 have similar molecular weights (~160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br_2 is nonpolar and does not. The relatively stronger dipole-dipole

attractions require more energy to overcome, so ICl will have the higher boiling point.

Effect of Hydrogen Bonding on Boiling Points

Consider the compounds dimethylether (CH_3OCH_3), ethanol (CH_3CH_2OH), and propane ($CH_3CH_2CH_3$). Their boiling points, not necessarily in order, are −42.1 °C, −24.8 °C, and 78.4 °C. Match each compound with its boiling point. Explain your reasoning.

Solution

The shapes of CH_3OCH_3, CH_3CH_2OH, and $CH_3CH_2CH_3$ are similar, as are their molar masses (46 g/mol, 46 g/mol, and 44 g/mol, respectively), so they will exhibit similar dispersion forces. Since $CH_3CH_2CH_3$ is nonpolar, it may exhibit *only* dispersion forces. Because CH_3OCH_3 is polar, it will also experience dipole-dipole attractions. Finally, CH_3CH_2OH has an −OH group, and so it will experience the uniquely strong dipole-dipole attraction known as hydrogen bonding. So the ordering in terms of strength of IMFs, and thus boiling points, is $CH_3CH_2CH_3 < CH_3OCH_3 < CH_3CH_2OH$. The boiling point of propane is −42.1 °C, the boiling point of dimethylether is −24.8 °C, and the boiling point of ethanol is 78.5 °C.

? Exercise 4.6.2

Ethane (CH_3CH_3) has a melting point of −183 °C and a boiling point of −89 °C. Predict the melting and boiling points for methylamine (CH_3NH_2). Explain your reasoning.

Answer

The melting point and boiling point for methylamine are predicted to be significantly greater than those of ethane. CH_3CH_3 and CH_3NH_2 are similar in size and mass, but methylamine possesses an −NH group and therefore may exhibit hydrogen bonding. This greatly increases its IMFs, and therefore its melting and boiling points. It is difficult to predict values, but the known values are a melting point of −93 °C and a boiling point of −6 °C.

Hydrogen Bonding and DNA

Deoxyribonucleic acid (DNA) is found in every living organism and contains the genetic information that determines the organism's characteristics, provides the blueprint for making the proteins necessary for life, and serves as a template to pass this information on to the organism's offspring. A DNA molecule consists of two anti-parallel chains of repeating nucleotides, which form its well-known double helical structure, as shown in Figure **4.6.10**.

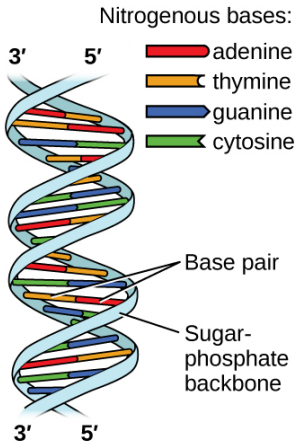


Figure **4.6.11**: Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (credit: modification of work by Jerome Walker, Dennis Myts)

Each nucleotide contains a (deoxyribose) sugar bound to a phosphate group on one side, and one of four nitrogenous bases on the other. Two of the bases, cytosine (C) and thymine (T), are single-ringed structures known as pyrimidines. The other two, adenine (A) and guanine (G), are double-ringed structures called purines. These bases form complementary base pairs consisting of one

purine and one pyrimidine, with adenine pairing with thymine, and cytosine with guanine. Each base pair is held together by hydrogen bonding. A and T share two hydrogen bonds, C and G share three, and both pairings have a similar shape and structure Figure **4.6.11**

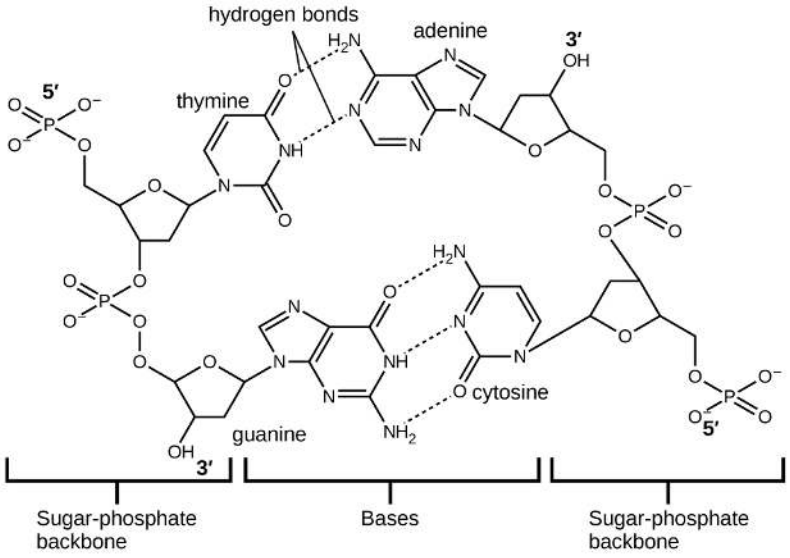


Figure **4.6.12**: The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called "complementary base pairs."

The cumulative effect of millions of hydrogen bonds effectively holds the two strands of DNA together. Importantly, the two strands of DNA can relatively easily "unzip" down the middle since hydrogen bonds are relatively weak compared to the covalent bonds that hold the atoms of the individual DNA molecules together. This allows both strands to function as a template for replication.

Example 4.6.3

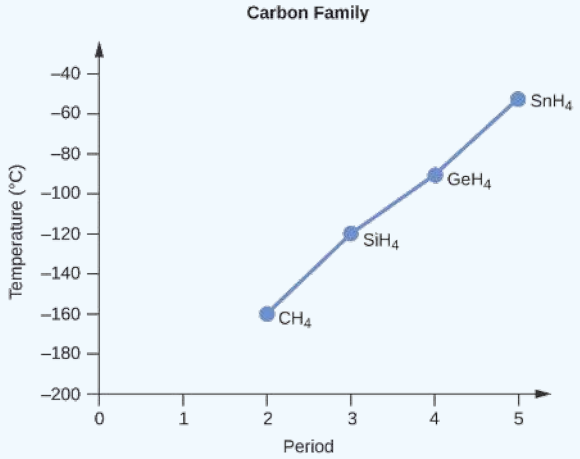
Order the following compounds of group 4A elements and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

Solution

Applying the skills acquired in the chapter on chemical bonding and molecular geometry, all of these compounds are predicted to be nonpolar, so they may experience only dispersion forces: the smaller the molecule, the less polarizable and the weaker the dispersion forces; the larger the molecule, the larger the dispersion forces. The molar masses of CH_4, SiH_4, GeH_4, and SnH_4 are approximately 16 g/mol, 32 g/mol, 77 g/mol, and 123 g/mol, respectively. Therefore, CH_4 is expected to have the lowest boiling point and SnH_4 the highest boiling point. The ordering from lowest to highest boiling point is expected to be

$CH_4 < SiH_4 < GeH_4 < SnH_4$

A graph of the actual boiling points of these compounds versus the period of the group 14 elementsshows this prediction to be correct:



A line graph, titled “Carbon Family,” is shown where the y-axis is labeled “Temperature, (degree sign C)” and has values of “negative 200” to “negative 40” from bottom to top in increments of 20. The x-axis is labeled “Period” and has values of “0” to “5” in increments of 1. The first point on the graph is labeled “C H subscript 4” and is at point “2, negative 160.” The second point on the graph is labeled “S i H subscript 4” and is at point “3, negative 120” while the third point on the graph is labeled “G e H subscript 4” and is at point “4, negative 100.” The fourth point on the graph is labeled “S n H subscript 4” and is at point “5, negative 60.”

4.6.3 Exercise

Order the following hydrocarbons from lowest to highest boiling point: C_2H_6, C_3H_8, and C_4H_{10}.

Answer

All of these compounds are nonpolar and only have London dispersion forces: the larger the molecule, the larger the dispersion forces and the higher the boiling point. The ordering from lowest to highest boiling point is therefore

$$C_2H_6 < C_3H_8 < C_4H_{10}.$$

Applications: Geckos and Intermolecular Forces

Geckos have an amazing ability to adhere to most surfaces. They can quickly run up smooth walls and across ceilings that have no toe-holds, and they do this without having suction cups or a sticky substance on their toes. And while a gecko can lift its feet easily as it walks along a surface, if you attempt to pick it up, it sticks to the surface. How are geckos (as well as spiders and some other insects) able to do this? Although this phenomenon has been investigated for hundreds of years, scientists only recently uncovered the details of the process that allows geckos’ feet to behave this way.

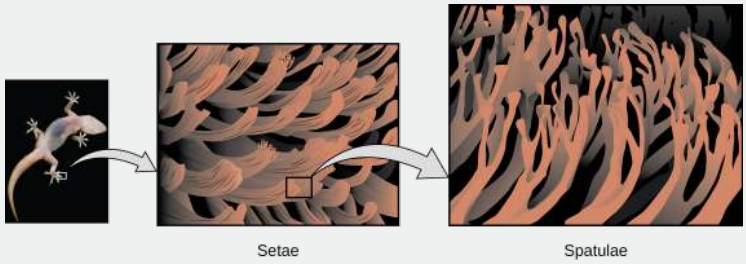


Figure **4.6.12**: *Geckos’ toes contain large numbers of tiny hairs (setae), which branch into many triangular tips (spatulae). Geckos adhere to surfaces because of van der Waals attractions between the surface and a gecko’s millions of spatulae. By changing how the spatulae contact the surface, geckos can turn their stickiness “on” and “off.” (credit photo: modification of work by “JC*+A!”/Flickr)*

Three figures are shown. The first is a photo of the bottom of a gecko’s foot. The second is bigger version which shows the setae. The third is a bigger version of the setae and shows the spatulae.

Geckos’ toes are covered with hundreds of thousands of tiny hairs known as *setae*, with each seta, in turn, branching into hundreds of tiny, flat, triangular tips called *spatulae*. The huge numbers of spatulae on its setae provide a gecko, shown in Figure 8.1.12, with a large total surface area for sticking to a surface. In 2000, Kellar Autumn, who leads a multi-institutional

gecko research team, found that geckos adhered equally well to both polar silicon dioxide and nonpolar gallium arsenide. This proved that geckos stick to surfaces because of dispersion forces—weak intermolecular attractions arising from temporary, synchronized charge distributions between adjacent molecules. Although dispersion forces are very weak, the total attraction over millions of spatulae is large enough to support many times the gecko's weight.

In 2014, two scientists developed a model to explain how geckos can rapidly transition from "sticky" to "non-sticky." Alex Greaney and Congcong Hu at Oregon State University described how geckos can achieve this by changing the angle between their spatulae and the surface. Geckos' feet, which are normally nonsticky, become sticky when a small shear force is applied. By curling and uncurling their toes, geckos can alternate between sticking and unsticking from a surface, and thus easily move across it. Further investigations may eventually lead to the development of better adhesives and other applications.

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 **4.6.3**: Intermolecular Forces and Boiling Points

Substance	Strongest Intermolecular Force	Boiling Point (°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

Example 4.6.4: Intermolecular Forces

What intermolecular forces besides dispersion forces, if any, exist in each substance? Are any of these substances solids at room temperature?

a. potassium chloride (KCl)
b. ethanol (C_2H_5OH)
c. bromine (Br_2)

Solution

a. Potassium chloride is composed of ions, so the intermolecular interaction in potassium chloride is ionic forces. Because ionic interactions are strong, it might be expected that potassium chloride is a solid at room temperature.
b. Ethanol has a hydrogen atom attached to an oxygen atom, so it would experience hydrogen bonding. If the hydrogen bonding is strong enough, ethanol might be a solid at room temperature, but it is difficult to know for certain. (Ethanol is actually a liquid at room temperature.)

c. Elemental bromine has two bromine atoms covalently bonded to each other. Because the atoms on either side of the covalent bond are the same, the electrons in the covalent bond are shared equally, and the bond is a nonpolar covalent bond. Thus, diatomic bromine does not have any intermolecular forces other than dispersion forces. It is unlikely to be a solid at room temperature unless the dispersion forces are strong enough. Bromine is a liquid at room temperature.

? Exercise 4.6.4

What intermolecular forces besides dispersion forces, if any, exist in each substance? Are any of these substances solids at room temperature?

a. methylamine (CH_3NH_2)
b. calcium sulfate ($CaSO_4$)
c. carbon monoxide (CO)

Answer

a. dipole-dipole, hydrogen bonding

b. ionic forces (solid at room temperature)

c. dipole-dipole

Key Takeaways

- A phase is a form of matter that has the same physical properties throughout.
- Molecules interact with each other through various forces: dipole-dipole interactions, hydrogen bonding, and dispersion forces.
- Dipole-dipole attractions result from the electrostatic attraction of the partial negative end of one dipolar molecule for the partial positive end of another.
- Hydrogen bonds are a special type of dipole-dipole attraction that results when hydrogen is covalently bonded to one of the three most electronegative elements: F, O, or N.
- The temporary dipole that results from the motion of the electrons in an atom can induce a dipole in an adjacent atom and give rise to the London dispersion force. London forces increase with increasing molecular size.

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4.7: Functional Groups or Families of Compounds

Learning Outcomes

- Identify and name families of organic compounds or 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 that belong to the same family of compounds or 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.

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, alkenes, alkynes and aromatic compounds are hydrocarbons. Hydrocarbons interact with each other through London forces.

Alkanes

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

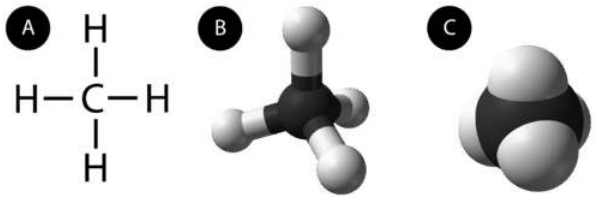


Figure **4.7.1**: Methane is the simplest hydrocarbon and is shown with a line bond formula, a ball-and-stick model, and a space-filling model.

Alkanes are generally unreactive. We primarily use alkanes as a source of energy when they are combusted. Alkanes interact with each other through London forces

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 alkanes. Alkenes and alkynes interact with each other through London forces.



Figure **4.7.2**: 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 **4.7.3**: Aromatic rings contain alternating double bonds.

Alcohols

The **alcohol** functional group involves an oxygen atom that is bonded to one hydrogen atom and one alkane type carbon atom. The carbon atom can be part of a larger organic structure. The primary IMF is hydrogen bonding between alcohol molecules. Alcohols can hydrogen bond to water.

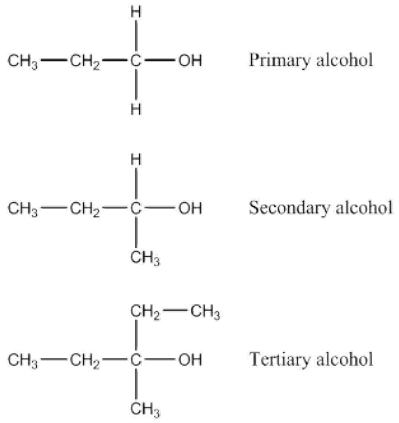


Figure **4.7.4**: Primary, secondary, and tertiary alcohols.

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.

We are familiar with several common alcohols. For example, ethanol ($\mathbf{CH_3CH_2OH}$) is the alcohol present in alcoholic beverages. It is also widely used in the industrial manufacture of other chemicals. Methanol ($\mathbf{CH_3OH}$) 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 $CH_3CH(OH)CH_3$ 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).

Phenols

Compounds in which an OH group is attached directly to an aromatic ring are called phenols. The primary IMF is hydrogen bonding between phenol molecules. Phenols can hydrogen bond to water.

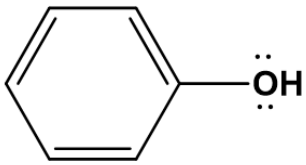
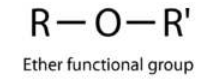


Figure **4.7.5**: A hydroxy attached to an aromatic ring is a phenol.

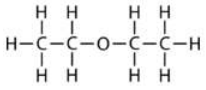
Phenols differ from alcohols in that they are slightly acidic in water. The parent compound, C_6H_5OH, is itself called phenol. Phenol is a white crystalline compound that has a distinctive (“hospital smell”) odor. Phenols are widely used as antiseptics (substances that kill microorganisms on living tissue) and as disinfectants (substances intended to kill microorganisms on inanimate objects such as furniture or floors). The first widely used antiseptic was phenol. Joseph Lister used it for antiseptic surgery in 1867.

Ethers

The **ether** functional group consists of an oxygen atom that forms single bonds with two carbon atoms. Ethers interact with each other via dipole-dipole interactions. They are polar molecules and hence they can hydrogen bond to water molecules.



Ether functional group



diethyl ether

Figure 4.7.6: 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, Sulfides, and Disulfides

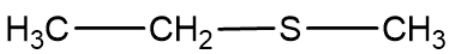
The **thiol** functional group contains a sulfur atom bonded to a hydrogen atom. It is very similar to an alcohol functional group with the sulfur replacing the O.



Figure 4.7.7: 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 present in the amino acid cysteine.

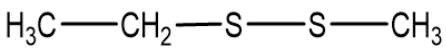
Sulfides also called thioethers, are sulfur analogs of ethers, have the form general formula RSR′. An example is dimethylsulfide (CH_3SCH_3), which is responsible for the sometimes unpleasant odor of cooking cabbage and related vegetables. The amino acid methionine has a sulfide functional group.



A sulfide

Figure 4.7.8: Sulfide

The mild oxidation of thiols gives compounds called **disulfides**. Disulfide linkages(–S–S–) between protein chains are extremely important in protein structure.



A disulfide

Figure 4.7.9: Disulfide

Thiols, sulfides and disulfides have only nonpolar covalent bonds. They interact with each other via London forces.

Amines

An **amine** consists of a nitrogen atom bonded to some combination of carbons and hydrogens. The amine functional group is found in amino acids which are building blocks of proteins.

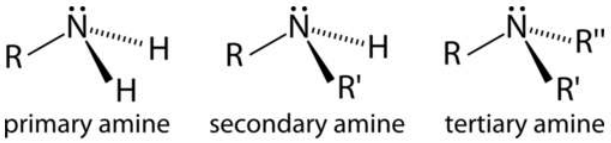


Figure 4.7.10: Amines

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.

In 1° amine and 2° amine, a hydrogen atom is covalently bonded to a nitrogen atom. So hydrogen bonding is possible between molecules of 1°-amine and 2°-amine. A 3°-amine can do dipole-dipole interactions with each other. All amines are polar and can hydrogen bond to polar water molecules.

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.

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 polarity of the C=O bond. While carbonyl is a component of many functional groups, it is not itself a functional group.

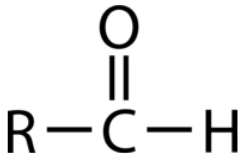


Figure 4.7.11: Aldehyde

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.

Aldehydes are polar compounds and so the primary mode of interaction is dipole-dipole interactions with each other. Aldehydes can interact with water via hydrogen bonding.

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.

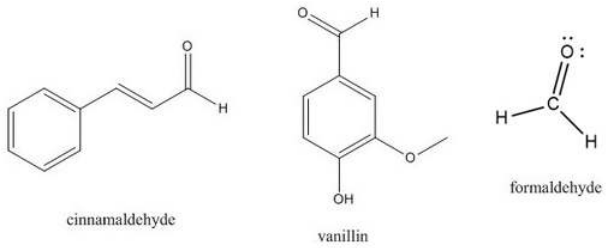


Figure **4.7.12**: Examples of molecules with aldehyde functional groups.

Ketones

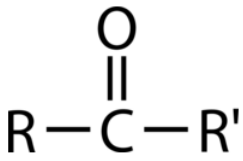


Figure **4.7.13**: Ketone

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 $\mathbf{CH_3}$ 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.



Figure **4.7.14**: Two examples of ketones.

Ketones like aldehydes are polar compounds and so the primary mode of interaction is dipole-dipole interactions with each other. Ketones can interact with water via hydrogen bonding.

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. Carboxylic acids can interact with each other via hydrogen bonds and they can interact with water via hydrogen bonding.

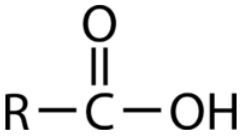


Figure **4.7.15***: Carboxylic acid*

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 resulting anion is a carboxylate ion (see figure below). Carboxylic acids are present as carboxylate ions in amino acids.

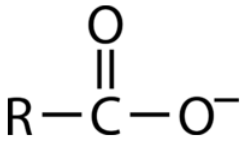


Figure **4.7.16**: Carboxylate ion

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.

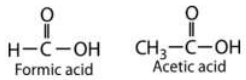


Figure 4.7.17: Carboxylic acids

Esters

An **ester** is similar to a carboxylic acid, in that it contains a carbonyl. Esters contain a C-O-C linkage in which one of the carbon atoms belongs to a carbonyl group. Structurally, carboxylic acids and esters are related to one another in the same way as alcohols and ethers. Esters are polar compounds and so the primary mode of interaction is dipole-dipole interactions with each other. Esters can interact with water via hydrogen bonding.

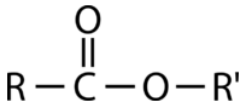


Figure 4.7.18: Ester

Esters can be formed by heating carboxylic acids and alcohols in the presence of an acid catalyst.

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 think of an amide as an amine in which one of the carbon atoms attached to the nitrogen is part of a carbonyl.

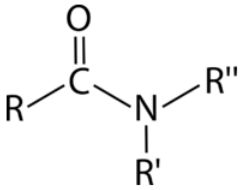


Figure 4.7.19: Amide

An amide can be formed by combining a carboxylic acid and an amine. Only primary and secondary amines can 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 the treatment of drinking water and sewage, and in plastics manufacture. The amide Kevlar is widely employed for the production of body armor. 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 7A atom). There is a fairly large distinction between the structural and physical properties of haloalkanes and the structural and physical properties of alkanes.

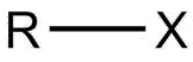


Figure 4.7.20: Alkyl Halide

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.

Application in Biochemistry

Most functional groups you have encountered in this section are important in molecules found in living things. For example the functional groups carboxylic acid and alkene (highlighted with a red arrow in figure 4.7.21) is found in fatty acids which are building blocks of triglycerides such as butter, vegetable oil, and corn oil.

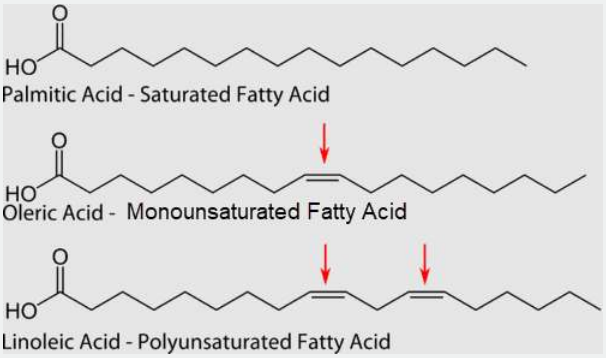


Figure **4.7.21**: Fatty acids contain a carboxylic acid functional group. Unsaturated fatty acids contain one or more alkene functional group.

The functional groups aldehyde and alcohol are found in the monosaccharide glucose (figure 4.7.22) which is a building block for starch found in rice, potato, or bread.

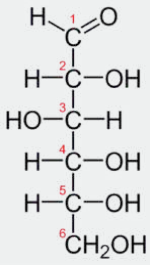


Figure **4.7.22**: Glucose contains an aldehyde functional group and five alcohol functional groups.

Contributors and Attributions

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4.8: Cis-Trans Isomers (Geometric Isomers)

Learning Objectives

- Recognize that alkenes that can exist as cis-trans isomers.
- Classify isomers as cis or trans.
- Draw structures for cis-trans isomers given their names.

There is free rotation about the carbon-to-carbon single bonds (C–C) in alkanes. In contrast, the structure of alkenes requires that the carbon atoms of a double bond and the two atoms bonded to each carbon atom all lie in a single plane, and that each doubly bonded carbon atom lies in the center of a triangle. This part of the molecule's structure is rigid; rotation about the double bonded carbon atoms is *not* possible without rupturing the bond. Look at the two chlorinated hydrocarbons in Figure **4.8.1**.

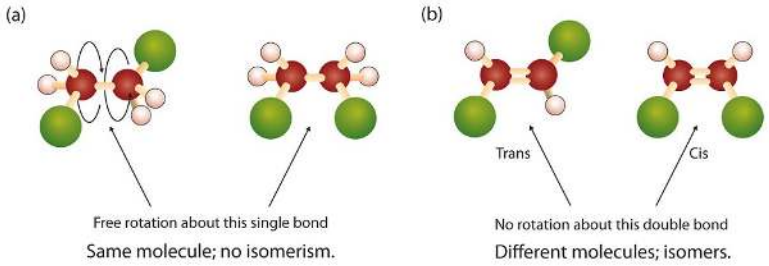


Figure **4.8.1**: Rotation about Bonds. In 1,2-dichloroethane (a), free rotation about the C–C bond allows the two structures to be interconverted by a twist of one end relative to the other. In 1,2-dichloroethene (b), restricted rotation about the double bond means that the relative positions of substituent groups above or below the double bond are significant.

In 1,2-dichloroethane Figure **4.8.2**, there is free rotation about the C–C bond. The two models shown represent exactly the same molecule; they are *not* isomers. You can draw structural formulas that look different, but bear in mind the possibility of free rotation about single bonds. These two structures represent the same molecule:



Figure **4.8.2**: 1,2-dichloroethane: Both drawings represent the same molecule.

In 1,2-dichloroethene (Figure **4.8.2**), however, restricted rotation about the double bond means that the relative positions of substituent groups above or below the double bond become significant. This leads to a special kind of isomerism. The isomer in which the two chlorine (Cl) atoms lie on the same side of the molecule is called the cis isomer (Latin *cis*, meaning "on this side") and is named cis-1,2-dichloroethene. The isomer with the two Cl atoms on opposite sides of the molecule is the trans isomer (Latin *trans*, meaning "across") and is named *trans*-1,2-dichloroethene. These two compounds are cis-trans isomers (or geometric isomers), compounds that have different configurations (groups permanently in different places in space) because of the presence of restricted rotation around the double bond.

Consider the alkene with the condensed structural formula $CH_3CH{=}CHCH_3$. We could name it 2-butene, but there are actually two such compounds; the double bond results in cis-trans isomerism (Figure **4.8.3**).

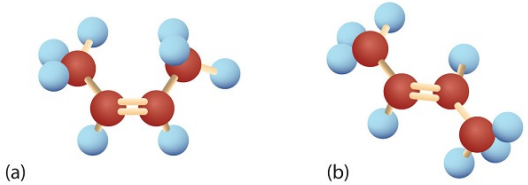


Figure **4.8.3**: Ball-and-Stick Models of (a) Cis-2-Butene and (b) Trans-2-Butene. Cis-trans isomers have different physical, chemical, and physiological properties.

Cis-2-butene has both methyl groups on the same side of the molecule. *Trans*-2-butene has the methyl groups on opposite sides of the molecule. Their structural formulas are as follows:



***Figure* 4.8.4**: *Models of (left) Cis-2-Butene and (right) Trans-2-Butene.*

Note, however, that the presence of a double bond does **not** necessarily lead to cis-trans isomerism (Figure **4.8.4**). We can draw two *seemingly* different propenes:



***Figure* 4.8.5**: *Different views of the propene molecule (flip vertically). These are not isomers.*

However, these two structures are not really different from each other. If you could pick up either molecule from the page and flip it over top to bottom, you would see that the two formulas are identical.

Thus there are two requirements for cis-trans isomerism:

1. Rotation must be restricted in the molecule.
2. There must be **two nonidentical groups** on **each** double bonded carbon atom.

In addition, cis-trans isomers have the same molecular formula, have the same atomic connections, have different 3-dimensional shape, and are interchanged only by breaking bonds.

In the propene structures, the second requirement for cis-trans isomerism is not fulfilled. One of the double bonded carbon atoms does have two different groups attached, but the rules require that *both* carbon atoms have two different groups. In general, the following statements hold true in cis-trans isomerism:

- Alkenes with a $C{=}CH_2$ unit do not exist as cis-trans isomers.
- Alkenes with a $C{=}CR_2$ unit, where the two R groups are the same, do not exist as cis-trans isomers.
- Alkenes of the type R–CH=CH–R can exist as cis and trans isomers; cis if the two R groups are on the same side of the carbon-to-carbon double bond, and trans if the two R groups are on opposite sides of the carbon-to-carbon double bond.

Advanced Note: E/Z Isomerization

If a molecule has a C=C bond with one non-hydrogen group attached to each of the carbons, cis/trans nomenclature descried above is enough to describe it. However, if you have three different groups (or four), then the cis/trans approach is insufficient to describe the different isomers, since we do not know which two of the three groups are being described. For example, if you have a C=C bond, with a methyl group and a bromine on one carbon , and an ethyl group on the other, it is neither trans nor cis, since it is not clear whether the ethyl group is trans to the bromine or the methyl. This is addressed with a more advanced E,Z Convention.

Cis-trans isomerism also occurs in cyclic compounds. In ring structures, groups are unable to rotate about any of the ring carbon–carbon bonds. Therefore, groups can be either on the same side of the ring (cis) or on opposite sides of the ring (trans). For our purposes here, we represent all cycloalkanes as planar structures, and we indicate the positions of the groups, either above or below the plane of the ring.



Cis-1,2-dimethylcyclopropane *Trans*-1,2-dimethylcyclopropane

***Figure* 4.8.6**:- Cis-trans isomerism in cyclic compounds

Application in Biochemistry

Esters and alkenes are functional groups common to biological molecules found in living things. An example is a triglyceride (vegetable oil or corn oil). Note the presence of the cis-double bond in the triglyceride. Triglycerides with a high percentage of alkenes in the cis configuration is considered heart healthy. Trans-fats increase the risk of heart disease.

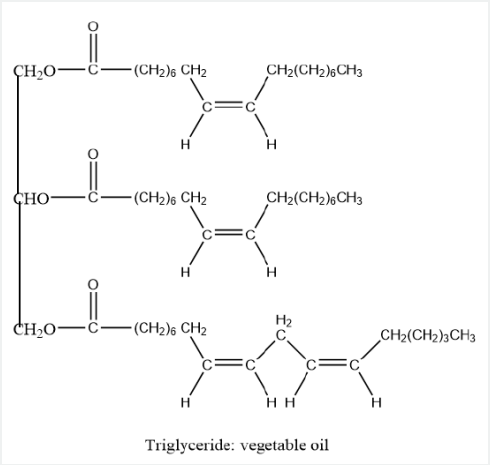


***Figure* 4.8.7**: The alkenes present in vegetable oil are cis.

Example 4.8.1

Which compounds can exist as cis-trans (geometric) isomers? Draw them.

1. CHCl=CHBr
2. CH_2=$CBrCH_3$
3. $(CH_3)_2C$=$CHCH_2CH_3$
4. CH_3CH=$CHCH_2CH_3$

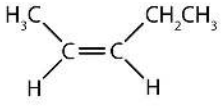
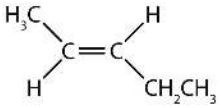
Solution

All four structures have a double bond and thus meet rule 1 for cis-trans isomerism.

1. This compound meets rule 2; it has two nonidentical groups on *each* carbon atom (H and Cl on one and H and Br on the other). It exists as both cis and trans isomers:

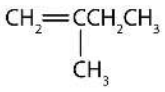
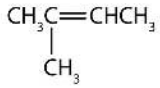


2. This compound has two hydrogen atoms on one of its doubly bonded carbon atoms; it fails rule 2 and does not exist as cis and trans isomers.
3. This compound has two methyl (CH_3) groups on one of its doubly bonded carbon atoms. It fails rule 2 and does not exist as cis and trans isomers.
4. This compound meets rule 2; it has two nonidentical groups on *each* carbon atom and exists as both cis and trans isomers:

? Exercise 4.8.1

Which compounds can exist as cis-trans isomers? Draw them.

a. $CH_2=CHCH_2CH_2CH_3$
b. $CH_3CH=CHCH_2CH_3$
c. $CH_3CH_2CH=CHCH_2CH_3$
d. 
e. 

Key Takeaway

- Cis-trans (geometric) isomerism exists when there is restricted rotation in a molecule and there are two nonidentical groups on *each* doubly bonded carbon atom.

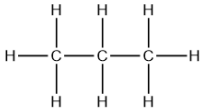
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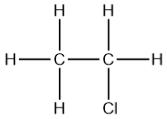
4.9: End of Chapter Problems

Formulae of Compounds

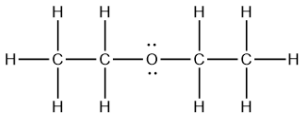
1. Write the molecular formula for the following compounds shown in their line bond structure.



a.

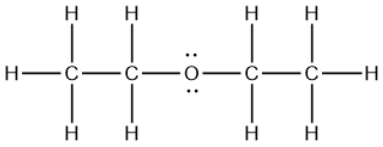


b.

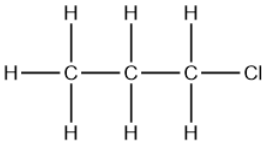


c.

2. Draw an isomer of the following compounds.

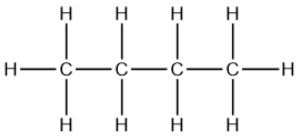


a.

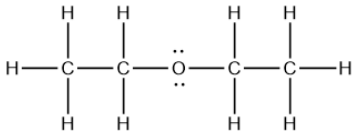


b.

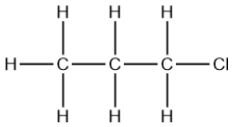
3. Draw condensed formula for the following compounds shown in their line bond structure.



a.



b.



c.

4. Draw skeletal formula for the compounds in question 3 shown in their line bond structure.

5. The molecules shown below are in their condensed structural formula. Draw line bond formula, molecular formula, and skeletal formula of the compounds.

a. $CH_3CH_2NH_2$

b. $CH_3OCH_2CH_2CH_3$

c. $CH_3CH_2CH_2Cl$

d. $CH_3CH_2CH_2CH_2CH_3$

Shapes of Molecules

What is the geometry around the central atom?

a. N of the ammonium ion, NH_4^+

b. N of ammonia, NH_3

c. C of methane, CH_4

d. O of water, H_2O

Bond Polarity, and Overall Polarity

1. Describe the electronegativity difference between each pair of atoms and the resulting polarity (or bond type).

a. C and H

b. H and H

c. Na and Cl

d. O and H

2. Describe the electronegativity (EN) difference between each pair of atoms and the resulting polarity (or bond type).

a. C and O

b. K and Br

c. N and N

d. Cs and F

3. Which of the following molecules are polar? (Hint: Draw the molecules in their line bond formula around the central atom carbon)

a. CH_3F

b. CH_2F_2

c. CF_4

4. Which of the following molecules are polar? (Hint: Draw the molecules in their line bond formula and label the polar bonds.)

a. CH_3OH

b. CH_3COCH_3

c. $CH_3CH_2CH_3$

d. CH_3NH_2

Intermolecular Forces

1. Predict which will have the higher boiling point: N_2 or CO. Explain your reasoning.

2. Predict which will have the higher boiling point: ICl or Br_2. Explain your reasoning.

3. Consider the compounds dimethylether (CH_3OCH_3), ethanol (CH_3CH_2OH), and propane ($CH_3CH_2CH_3$). Their boiling points are −42.1 °C, −24.8 °C, and 78.4 °C. Match each compound with its boiling point. Explain your reasoning.

4. Ethane (CH_3CH_3) has a melting point of −183 °C and a boiling point of −89 °C. Will the melting and boiling points for methylamine (CH_3NH_2) be higher or lower than ethane. Explain your reasoning.

5. Order the following compounds of group 4A elements and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

6. Order the following hydrocarbons from lowest to highest boiling point: C_2H_6, C_3H_8, and C_4H_{10}. Explain.

7. What intermolecular forces besides dispersion forces, if any, exist in each substance?

a. potassium chloride (KCl)

b. ethanol (C_2H_5OH)

c. bromine (Br_2)

8. What intermolecular forces besides dispersion forces, if any, exist in each substance?

methylamine (CH_3NH_2)

calcium sulfate ($CaSO_4$)

carbon monoxide (CO)

9. Which pair of molecules can form a hydrogen bond with one another?

a. CH_3CH_3 and CH_3CH_3

b. CH_3SH and CH_3SH

c. CH_3OH and CH_3OH

d. CH_3CHO and CH_3CHO

10. Which of the molecules in question 9 can form hydrogen bonds with water molecules?

Functional Groups or Families of Compounds

Draw the structures of 10 compounds. Each of the 10 compounds should be an example of the functional group below with the correct # of carbon atoms specified.

a. An alkene with 4 carbons

b. An alcohol with 3 carbons

c. A thiol with 2 carbons.

d. A disulfide with 4 carbons.

e. An amine with 3 carbons. Label as primary, secondary, or tertiary.

f. A ketone with 3 carbons.

g. An aldehyde with 3 carbons.

h. A carboxylic acid with 3 carbons.

i. An ester with 3 carbons.

j. An amide with 2 carbons.

Cis-Trans Isomers

Which compounds can exist as cis-trans (geometric) isomers? Draw the cis and trans isomers.

a. ClCH=CHBr

b. CH_2=$CBrCH_3$

c. $(CH_3)_2C$=$CHCH_2CH_3$

d. CH_3CH=$CHCH_2CH_3$

e. CH_2=$CHCH_2CH_2CH_3$

f. CH_3CH=$CHCH_2CH_3$

g. CH_3CH_2CH=$CHCH_2CH_3$

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CHAPTER OVERVIEW

5: Reactions

5.1: Chemical Equations
5.2: Evidence of a Chemical Reaction
5.3: Balancing Chemical Equations
5.4: Classifying Chemical Reactions
5.5: Reactions Involving Water
5.6: Classification of alcohols
5.7: Oxidations and Reductions
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5.11: Citric Acid Cycle
5.12: End of Chapter Problems

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5.1: Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

When reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that—what is produced—or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl and baking sheet.

$$\underbrace{\text{Flour} + \text{Butter} + \text{Sugar} + \text{Vanilla} + \text{Baking Soda} + \text{Eggs} + \text{Chocolate Chips}}_{\text{Ingredients = Reactants}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}}$$

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $SO_2 + O_2$, are reactants and sulfur trioxide, SO_3, is the product.

$$\underbrace{2\,SO_2\,(g) + O_2\,(g)}_{\text{Reactants}} \rightarrow \underbrace{2\,SO_3\,(g)}_{\text{Products}}$$

In chemical reactions, the reactants are found before the symbol "$\rightarrow$" and the products are found after the symbol "$\rightarrow$". The general equation for a reaction is:

$$\text{Reactants} \rightarrow \text{Products}$$

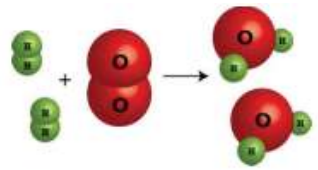
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table 5.1.1 shows a listing of symbols used in chemical equations.

Table 5.1.1: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Description
$+$	used to separate multiple reactants or products	(s)	reactant or product in the solid state
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

Chemists have a choice of methods for describing a chemical reaction.

1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. They could write the equation in chemical shorthand.

$$2\text{H}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{H}_2\text{O}\,(g) \tag{5.1.1}$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:

$$\text{Ca(NO}_3)_2\,(aq) + 2\text{NaOH}\,(aq) \rightarrow \text{Ca(OH)}_2\,(s) + 2\text{NaNO}_3\,(aq) \tag{5.1.2}$$

How much easier is that to read? Let's try it in reverse. Look at the following reaction in shorthand and write the word equation for the reaction:

$$\text{Cu}\,(s) + \text{AgNO}_3\,(aq) \rightarrow \text{Cu(NO}_3)_2\,(aq) + \text{Ag}\,(s) \tag{5.1.3}$$

The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver."

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned with ionic compounds (charge balanced).

3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table **5.1.1**: *Diatomic Elements*

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

Example 5.1.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. $\mathbf{HCl}\,(aq) + \mathbf{NaOH}\,(aq) \rightarrow \mathbf{NaCl}\,(aq) + \mathbf{H_2O}\,(l)$
b. Gaseous propane, $\mathbf{C_3H_8}$, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
c. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane ($\mathbf{C_3H_8}$) and oxygen ($\mathbf{O_2}$)

Product: carbon dioxide ($\mathbf{CO_2}$) and water ($\mathbf{H_2O}$)

$$\mathbf{C_3H_8}\,(g) + \mathbf{O_2}\,(g) \rightarrow \mathbf{CO_2}\,(g) + \mathbf{H_2O}\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathbf{HF}\,(g) + \mathbf{K_2CO_3}\,(aq) \rightarrow \mathbf{KF}\,(aq) + \mathbf{H_2O}\,(l) + \mathbf{CO_2}\,(g)$$

Exercise 5.1.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

a. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.
b. $\mathbf{HCl}\,(aq) + \mathbf{LiOH}\,(aq) \rightarrow \mathbf{NaCl}\,(aq) + \mathbf{H_2O}\,(l)$
c. Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer a

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

Answer b

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

Answer c

$Cu(s) + O_2(g) \rightarrow CuO(s)$

Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

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5.2: Evidence of a Chemical Reaction

Learning Objectives

- Identify the evidence for chemical reactions.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

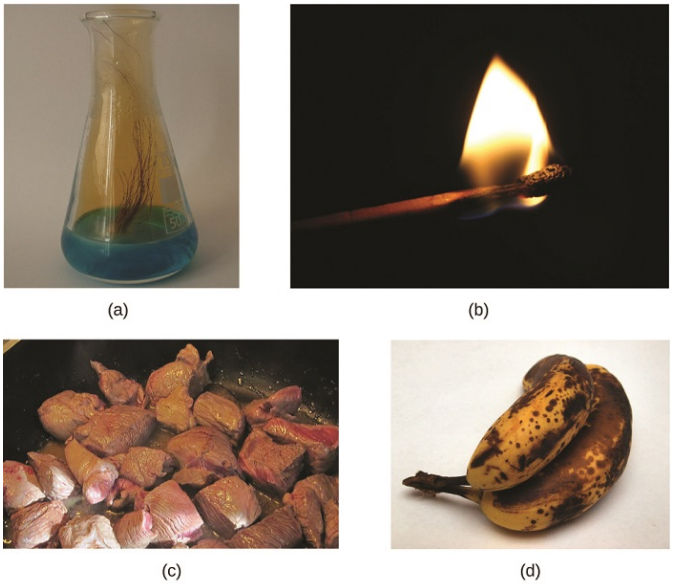


Figure **5.2.1**: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (Credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo.)

To identify a chemical reaction, we look for a **chemical change**. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include: reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure **5.2.1**).



***Video*5.2.1*:** *Evidence of a Chemical Reaction*

Example 5.2.1: Evidence of a Chemical Reaction

Which of the following is a chemical reaction?

a. Freezing liquid Mercury.
b. Adding yellow to blue to make green.
c. Cutting a piece of paper into two pieces.
d. Dropping a sliced orange into a vat of sodium hydroxide.
e. Filling a balloon with natural air.

Solution

A, b ,c , & e involve only physical changes. A sliced orange has acid (citric acid) that can react with sodium hydroxide, so the answer is d.

Exercise 5.2.1

Which of the following is a chemical reaction?

a. Painting a wall blue
b. A bicycle rusting
c. Ice cream melting
d. Scratching a key across a desk
e. Making a sand castle

Answer

b

Example 5.2.2: Evidence of a Chemical Reaction

Which of the following is not a chemical reaction?

a. Shattering glass with a baseball
b. Corroding Metal
c. Fireworks exploding
d. Lighting a match
e. Baking a cake

Solution

Shattering glass with a baseball results in glass broken into many pieces but no chemical change happens, so the answer is a.

Exercise 5.2.2

Which of the following is NOT a chemical reaction?

a. Frying an egg
b. Slicing carrots
c. A Macbook falling out of a window
d. Creating ATP in the human body
e. Dropping a fizzy tablet into a glass of water

Answer

b and c

Summary

Chemical reactions can be identified via a wide range of different observable factors including change in color, energy change (temperature change or light produced), gas production, formation of precipitate and change in properties.

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5.3: Balancing Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products; and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

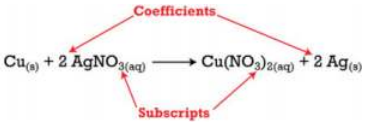
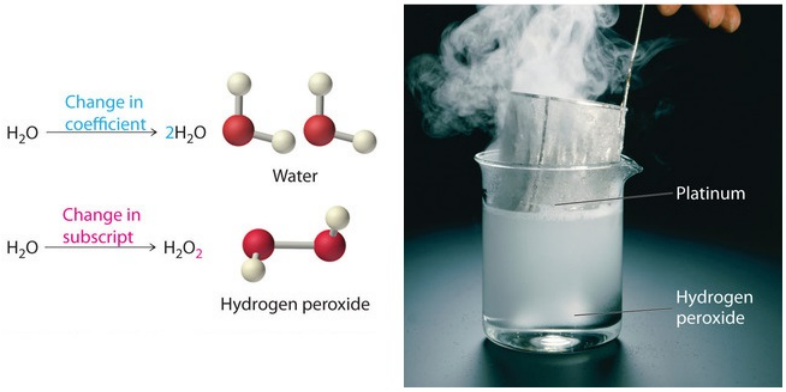


Figure 5.3.1: Balancing Equations. *You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.*

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 5.3.1.



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

Steps in Balancing a Chemical Equation

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

Example 5.3.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane (C_7H_{16}).

$$C_7H_{16}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Solution

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $$C_7H_{16}(l) + O_2(g) \longrightarrow \underline{7}CO_2(g) + H_2O(g)$$ • 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $$C_7H_{16}(l) + O_2(g) \longrightarrow 7CO_2(g) + \underline{8}H_2O(g)$$ • 16 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $$C_7H_{16}(l) + \underline{11}O_2(g) \longrightarrow 7CO_2(g) + 8H_2O(g)$$ • 22 oxygen atoms on both reactant and product sides
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

Example 5.3.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}\,(l) + O_2\,(g) \longrightarrow CO_2\,(g) + H_2O\,(g)$$

Solution

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products: $$C_8H_{18}\,(l) + O_2\,(g) \longrightarrow \underline{8}CO_2\,(g) + H_2O\,(g)$$ • 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $$C_8H_{18}\,(l) + O_2\,(g) \longrightarrow 8\,CO_2\,(g) + \underline{9}H_2O\,(g)$$ • 18 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $$C_8H_{18}\,(l) + \underline{\frac{25}{2}}O_2\,(g) \longrightarrow 8\,CO_2\,(g) + 9\,H_2O\,(g)$$ • 25 oxygen atoms on both reactant and product sides The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by 2: $$\underline{2}C_8H_{18}\,(l) + \underline{25}O_2\,(g) \longrightarrow \underline{16}CO_2\,(g) + \underline{18}H_2O\,(g)$$
5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

Example 5.3.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $\mathrm{Pb(NO_3)_2\,(aq) + NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$
2. **Adjust the coefficients.**	There are twice as many chloride ions in the reactants than in the products. Place a 2 in front of the NaCl in order to balance the chloride ions. $\mathrm{Pb(NO_3)_2\,(aq) + \underline{2}\,NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on reactant side, 1 Na atom on product side • 2 Cl atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow \underline{2}\,NaNO_3\,(aq) + PbCl_2\,(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides
4. **Balance the remaining atoms.**	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow 2\,NaNO_3\,(aq) + PbCl_2\,(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides

Exercise 5.3.1

Is each chemical equation balanced?

a. $\mathrm{2\,Hg(\ell) + O_2\,(g) \rightarrow 2\,Hg_O\,(s)}$
b. $\mathrm{C_2H_4\,(g) + 2\,O_2(g) \longrightarrow 2\,CO_2\,(g) + 2\,H_2O\,(g)}$
c. $\mathrm{Mg(NO_3)_2\,(s) + 2\,Li\,(s) \rightarrow Mg\,(s) + 2\,LiNO_3\,(s)}$

Answer a

yes

Answer b

no

Answer c

yes

? Exercise 5.3.2

Balance the following chemical equations.

a. $\mathrm{N_2\,(g) + O_2\,(g) \longrightarrow NO_2\,(g)}$
b. $\mathrm{Pb(NO_3)_2\,(aq) + FeCl_3\,(aq) \longrightarrow Fe(NO_3)_3\,(aq) + PbCl_2\,(s)}$
c. $\mathrm{C_6H_{14}\,(l) + O_2(g) \longrightarrow CO_2\,(g) + H_2O\,(g)}$

Answer a

$\mathrm{N_2\,(g) + 2O_2\,(g) \rightarrow 2NO_2\,(g)}$

Answer b

$\mathrm{3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)}$

Answer c

$\mathrm{2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)}$

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- Chemical reaction: The process in which one or more substances are changed into one or more new substances.
- Reactants: The starting materials in a reaction.
- Products: Materials present at the end of a reaction.
- Balanced chemical equation: A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- Subscripts: Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- Coefficient: A small whole number that appears in front of a formula in a balanced chemical equation.

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5.4: Classifying Chemical Reactions

Learning Objectives

- Classify a chemical reaction as a synthesis, decomposition, single replacement, double replacement, or a combustion reaction.
- Predict the products of simple reactions.

The chemical reactions we have described are only a tiny sampling of the infinite number of chemical reactions possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Types of Chemical Reactions

Most chemical reactions can be classified into one or more of seven basic types: combination reactions, decomposition reactions, the two exchange reactions, acid–base reactions, combustion reactions, and oxidation–reduction reactions. The general forms of these seven kinds of reactions are summarized in Table **5.4.1**, along with examples of each. It is important to note, however, that many reactions can be assigned to more than one classification, as you will see in our discussion.

Table **5.4.1**: *Basic Types of Chemical Reactions*

Name of Reaction	General Form	Examples
Combination (Synthesis or Condensation)	A + B → AB	$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$ $N_2(g) + 2O_2(g) \rightarrow 2NO_2(g)$
Decomposition	AB → A + B	$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
Exchange: Single Replacement	AB + C → AC + B	$ZnCl_2(aq) + Mg(s) \rightarrow MgCl_2(aq) + Zn(s)$
Exchange: Double Replacement	AB + CD → AD + CB	$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$
Acid–Base	acid + base → salt	$NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
Combustion	hydrocarbon + oxygen → carbondioxide + water	$C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(g)$
Oxidation–Reduction (redox)	oxidant + reductant → reduced oxidant + oxidized reductant	

The classification scheme is only for convenience; the same reaction can be classified in different ways. Acid–base reactions may be classified as double replacement reaction. A combustion reaction may be classified as an Oxidation–Reduction reaction. The combination reaction of nitrogen and oxygen to give nitrogen dioxide may also be classified as a combustion reaction. The first four types of reactions will be discussed in this section.

Combination Reactions

A **combination 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 or condensation or addition reactions. The general form of a combination reaction is:

$$\mathbf{A + B \rightarrow AB} \tag{5.4.1}$$

One combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to produce solid sodium chloride.

$$2Na\,(s) + Cl_2\,(g) \rightarrow 2NaCl\,(s)$$

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 type 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.

$$2Mg\,(s) + O_2\,(g) \rightarrow 2MgO\,(s)$$

Sulfur reacts with oxygen to form sulfur dioxide.

$$S\,(s) + O_2\,(g) \rightarrow SO_2\,(g)$$

When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form sulfur trioxide.

$$2S\,(s) + 3O_2\,(g) \rightarrow 2SO_3\,(g)$$

Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron (II) oxide and iron (III) oxide.

$$2Fe\,(s) + O_2\,(g) \rightarrow 2FeO\,(s)$$

$$4Fe\,(s) + 3O_2\,(g) \rightarrow 2Fe_2O_3\,(s)$$

The reaction of metals and nonmetals with oxygen may also be called a combustion reaction.

Example 5.4.1: Combustion of Solid Potassium

Potassium is a very reactive alkali metal that must be stored under oil in order to prevent it from reacting with air. Write the balanced chemical equation for the combination reaction of potassium with oxygen.

Solution

Steps	Example Solution
Plan the problem.	Make sure formulas of all reactants and products are correct before balancing the equation. Oxygen gas is a diatomic molecule. Potassium oxide is an ionic compound and so its formula is constructed by the crisscross method. Potassium as an ion becomes K^+, while the oxide ion is O^{2-}.
Solve.	The skeleton (unbalanced) equation: $$K\,(s) + O_2\,(g) \rightarrow K_2O\,(s)$$ The equation is then easily balanced with coefficients. $$4K\,(s) + O_2\,(g) \rightarrow 2K_2O\,(s)$$
Think about your result.	Formulas are correct and the resulting combination reaction is balanced.

Combination reactions can also take place when an element reacts with a compound to form a new compound composed of a larger number of atoms. Carbon monoxide reacts with oxygen to form carbon dioxide according to the equation:

$$2CO\,(g) + O_2\,(g) \rightarrow 2CO_2\,(g) \qquad (5.4.2)$$

Two compounds may also react to form a more complex compound. A very common example is the reactions of oxides with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide.

$$\mathrm{CaO}\,(s) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{Ca(OH)_2}\,(aq) \tag{5.4.3}$$

Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$\mathrm{SO_3}\,(g) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2SO_4}\,(aq) \tag{5.4.4}$$

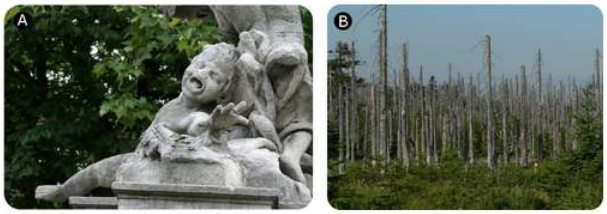


Figure 5.4.1: Acid rain has severe consequences on both natural and manmade objects. Acid rain degrades marble statues like the one on the left (A). The trees in the forest on the right (B) have been killed by acid rain.

? Exercise 5.4.1

a. Write the chemical equation for the synthesis of silver bromide, AgBr.
b. Predict the products for the following reaction: $\mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$

Answer a:

$$2\mathrm{Ag} + \mathrm{Br_2} \rightarrow 2\mathrm{AgBr}$$

Answer b:

$$\mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2CO_3}$$

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:

$$\mathrm{AB} \rightarrow \mathrm{A} + \mathrm{B} \tag{5.4.5}$$

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.

$$2\mathrm{HgO}\,(s) \rightarrow 2\mathrm{Hg}\,(l) + \mathrm{O_2}\,(g) \tag{5.4.6}$$



***Video* 5.4.2:** *Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.*

A reaction is also considered to be a decomposition reaction even when one or more of the products are still compounds. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide.

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \tag{5.4.7}$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water.

$$\mathrm{2NaOH}\,(s) \rightarrow \mathrm{Na_2O}\,(s) + \mathrm{H_2O}\,(g) \tag{5.4.8}$$

Some unstable acids decompose to produce nonmetal oxides and water. Carbonic acid decomposes easily at room temperature into carbon dioxide and water.

$$\mathrm{H_2CO_3}\,(aq) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l) \tag{5.4.9}$$

Example 5.4.2: Electrolysis of Water

When an electric current is passed through pure water, it decomposes into its elements. Write a balanced equation for the decomposition of water.

Solution

Steps	Example Solution
Plan the problem.	Water is a binary compound composed of hydrogen and oxygen. The hydrogen and oxygen gases produced in the reaction are both diatomic molecules.
Solve.	The skeleton (unbalanced) equation: $\mathrm{H_2O}\,(l) \overset{\text{elec}}{\rightarrow} \mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$ Note the abbreviation "**elec**" above the arrow to indicate the passage of an electric current to initiate the reaction. Balance the equation. $\mathrm{2H_2O}\,(l) \overset{\text{elec}}{\rightarrow} \mathrm{2H_2}\,(g) + \mathrm{O_2}\,(g)$
Think about your result.	The products are elements and the equation is balanced.

? Exercise 5.4.2

Write the chemical equation for the decomposition of:

a. Al_2O_3
b. Ag_2S

Answer a

$$2Al_2O_3 \rightarrow 4Al + 3O_2$$

Answer b

$$Ag_2S \rightarrow 2Ag + S$$

Single Replacement Reactions

A third type of reaction is the single replacement reaction, in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$A + BC \rightarrow AC + B \tag{5.4.10}$$

In this general reaction, element **A** is a metal and replaces element **B**, also a metal, in the compound BC. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$Y + XZ \rightarrow XY + Z \tag{5.4.11}$$

where **Y** is a nonmetal and replaces the nonmetal **Z** in the compound with **X**.

Metal Replacement

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.

$$Mg\,(s) + Cu(NO_3)_2\,(aq) \rightarrow Mg(NO_3)_2\,(aq) + Cu\,(s) \tag{5.4.12}$$

This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

Hydrogen Replacement

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 (figure below).

$$Zn\,(s) + 2HCl\,(aq) \rightarrow ZnCl_2\,(aq) + H_2\,(g) \tag{5.4.13}$$

In a hydrogen replacement reaction, the hydrogen in the acid is replaced by an active metal. Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All group 1A metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$2Na\,(s) + 2H_2O\,(l) \rightarrow 2NaOH\,(aq) + H_2\,(g) \tag{5.4.14}$$



Figure 5.4.2: (First image) Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Second image) Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite.

Halogen Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine.

$$\mathrm{Cl_2}\,(g) + 2\mathrm{NaBr}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{Br_2}\,(l) \tag{5.4.15}$$

The reactivity of the halogen group 7A decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

Example 5.4.3

What are the products of the reaction between solid aluminum (Al) and iron (III) oxide ($\mathrm{Fe_2O_3}$)?

Solution

Steps	Example Solution
Plan the problem.	To predict the products, we need to know that aluminum will replace iron and form aluminum oxide (the metal will replace the metal ion in the compound). Aluminum has a charge of $+3$ and oxygen has a charge of -2. The compound formed between aluminum and oxygen, therefore, will be $\mathrm{Al_2O_3}$. Since iron is replaced in the compound by aluminum, the iron will now be the single element in the products.
Solve.	The unbalanced equation will be: $$\mathrm{Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe}$$ and the balanced equation will be: $$\mathrm{2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe}$$
Think about your result.	This is a single replacement reaction, and when balanced the coefficients accurately reflect that the iron and aluminum have the same charge in this reaction.

Exercise 5.4.3

a. Write the chemical equation for the single replacement reaction between zinc solid and lead (II) nitrate solution to produce zinc nitrate solution and solid lead. (*Note: zinc forms ions with a $+2$ charge)

b. Predict the products for the following reaction: $\mathrm{Fe + CuSO_4}$ (in this reaction, assume iron forms ions with a $+2$ charge)

Answer a

$$Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2$$

Answer b

$$Fe + CuSO_4 \rightarrow Cu + FeSO_4$$

Double Replacement Reactions

A **double-replacement reaction** 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 (also called double-displacement) reaction is:

$$AB + CD \rightarrow AD + BC \tag{5.4.16}$$

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.

Formation of a Precipitate

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.

$$2KI\,(aq) + Pb(NO_3)_2\,(aq) \rightarrow 2KNO_3\,(aq) + PbI_2\,(s) \tag{5.4.17}$$

There are very strong attractive forces that occur between Pb^{2+} and I^- ions and the result is a brilliant yellow precipitate (Figure **5.4.3**). The other product of the reaction, potassium nitrate, remains soluble.



Figure **5.4.3**: Lead (II) iodide precipitates when potassium iodide is mixed with lead (II) nitrate (Equation 5.4.17). (CC BY-SA 3.0; PRHaney).

Formation of a Gas

Some double-replacement reactions produce a gaseous product which then bubbles out of the solution and escapes into the air. When solutions of sodium sulfide and hydrochloric acid are mixed, the products of the reaction are aqueous sodium chloride and hydrogen sulfide gas.

$$Na_2S\,(aq) + 2HCl\,(aq) \rightarrow 2NaCl\,(aq) + H_2S\,(g) \tag{5.4.18}$$

Formation of a Molecular Compound

Another kind of double-replacement reaction is one that produces a molecular compound as one of its products. Many examples in this category are reactions that produce water. When aqueous hydrochloric acid is reacted with aqueous sodium hydroxide, the products are aqueous sodium chloride and water.

$$HCl\,(aq) + NaOH\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l) \tag{5.4.19}$$

Example 5.4.4

Write a complete and balanced chemical equation for the double-replacement reaction $NaCN\,(aq) + HBr\,(aq) \rightarrow$ (hydrogen cyanide gas is formed).

Solution

Steps	Example Solution
Plan the problem.	The production of a gas drives the reaction.
Solve.	The cations of both reactants are $+1$ charged ions, while the anions are -1 charged ions. After exchanging partners, the balanced equation is: $$\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow \mathrm{NaBr}\,(aq) + \mathrm{HCN}\,(g) \quad (5.4.20)$$
Think about your result.	This is a double replacement reaction. All formulas are correct and the equation is balanced.

? Exercise 5.4.4

Write a complete and balanced chemical equation for the double-replacement reaction $(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow$ (a precipitate of barium sulfate forms).

Answer a:

$$(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow 2\mathrm{NH_4NO_3}\,(aq) + \mathrm{BaSO_4}\,(s) \quad (5.4.21)$$

Occasionally, a reaction will produce both a gas and a molecular compound. The reaction of a sodium carbonate solution with hydrochloric acid produces aqueous sodium chloride, carbon dioxide gas, and water.

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\ \mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

Contributions & Attributions

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5.5: Reactions Involving Water

Learning Objectives

- Identify the products of hydrolysis of an ester under acidic conditions.
- Describe how to prepare alcohols from alkenes via hydration reaction.
- Describe how to dehydrate an alcohol to an alkene by elimination of a water molecule.

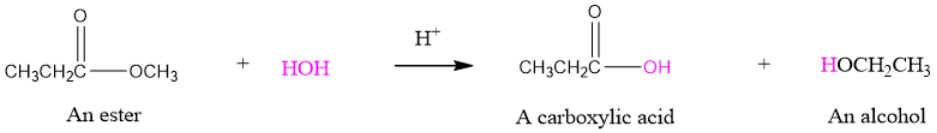
Water is a reactant or product in a number of reactions important to biochemistry. In this section we will take a look at three of them with examples from organic chemistry. They are

- Hydrolysis of esters
- Hydration of alkenes
- Dehydration of alcohols

Hydrolysis

Hydrolysis is, literally "splitting with water." Hydrolysis is the most important reaction of esters. The hydrolysis of esters is catalyzed by either an acid or a base. The ester is heated with a large excess of water containing a strong-acid catalyst such as sulfuric acid. The sulfuric acid is a source of the hydrogen ion (H^+), the catalyst in the reaction.

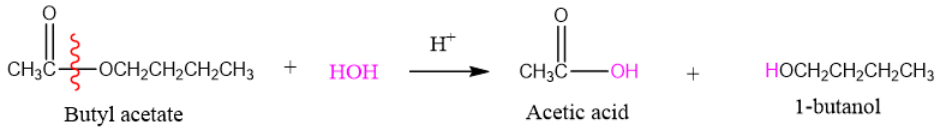
Acidic hydrolysis of an ester gives a carboxylic acid and an alcohol. Note that the -OH of the carboxylic acid and the H of the alcohol comes from water (shown in purple).



In this next example, butyl acetate and water react to form acetic acid and 1-butanol.

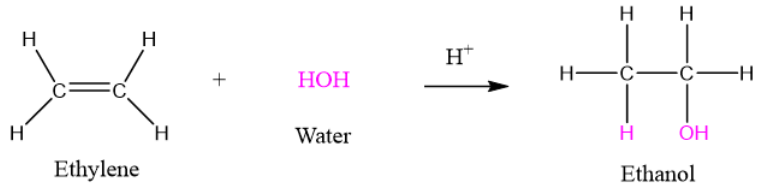
The lysis (splitting) happens at the C-O single bond of the ester indicated with a ~~~.

The role of the hydrogen ion (H^+) is to increase the speed of the hydrolysis reaction. Therefore H^+ is a catalyst in the reaction.



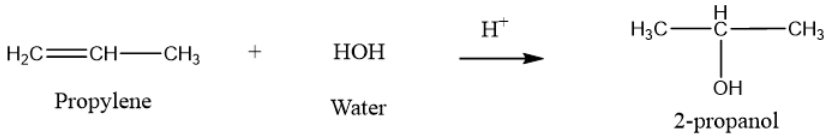
Hydration

Many simple alcohols are made by the hydration of alkenes. Ethanol is made by the hydration of ethylene in the presence of a catalyst such as hydrogen ion (H^+). Sulfuric acid (H_2SO_4) is the source of the hydrogen ion. The alcohol functional group and the hydrogen come from the water molecule (shown in purple). The double bond in the alkene is replaced by a single bond in the alcohol.



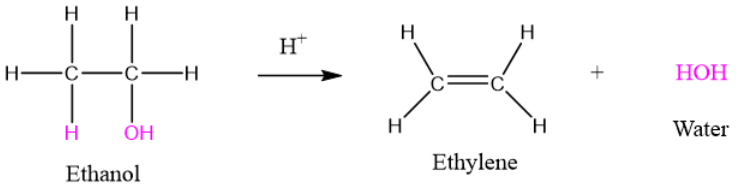
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In a similar manner, isopropyl alcohol (rubbing alcohol) is produced by the addition of water to propylene. Hydration of an alkene is an example of an addition reaction.



Dehydration

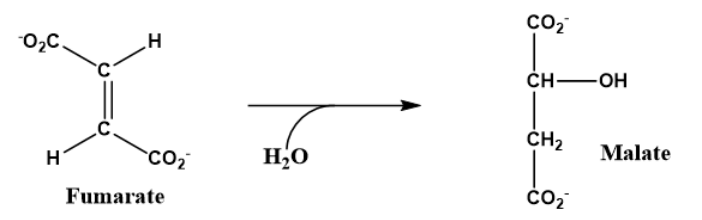
An alcohol undergoes dehydration in the presence of a catalyst to form an alkene and water. The reaction removes the OH group from the alcohol carbon atom and a hydrogen atom from an adjacent carbon atom in the same molecule. Dehydration is the reverse of hydration and is classified as a decomposition reaction.



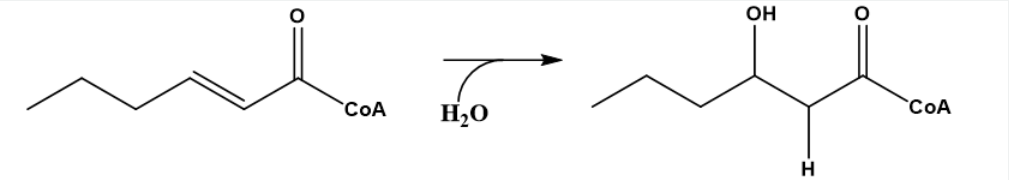
Reactions With Water Important in Biochemistry

The reaction of organic compounds with water are found in many metabolic processes and happen within the living cells. Here are a few examples from the citric acid cycle, fatty acid catabolism, and glucose metabolism.

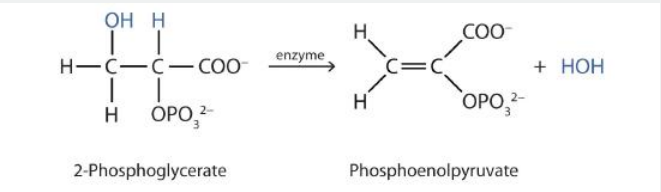
a. The **citric acid cycle** is a series of reactions important to metabolism. In step 7 of the cycle the alkene fumarate undergoes hydration to malate an alcohol.



b. In step 2 of the spiral of **fatty acid catabolism** is a hydration of an alkene to a 2°-alcohol.

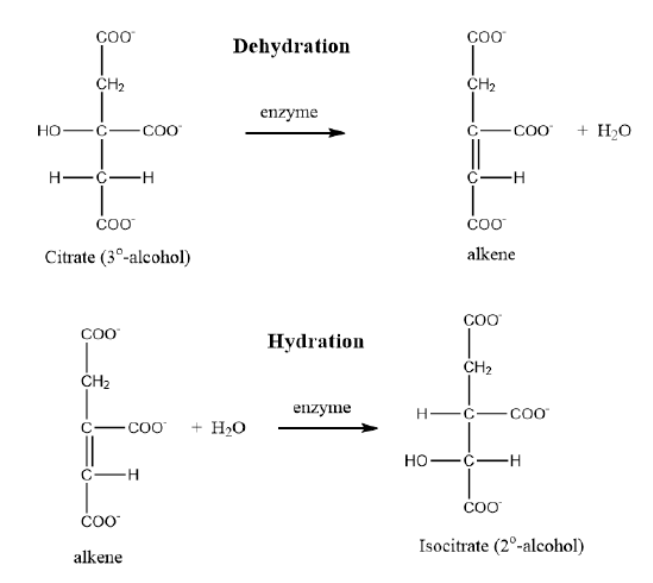


c. In step 9 of **glycolysis** (glucose metabolism) the alcohol 2-phosphoglycerate undergoes dehydration to an alkene.



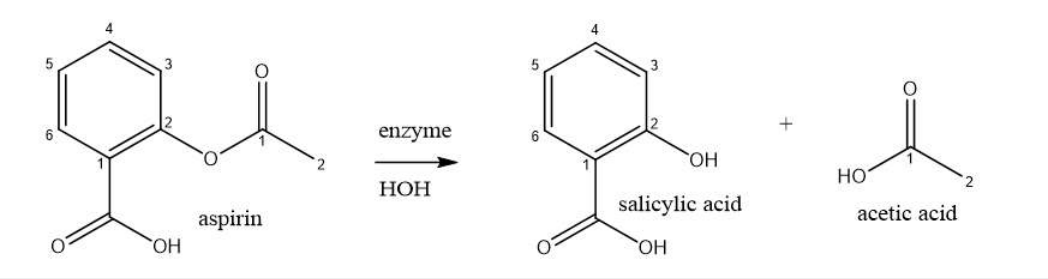
d. Another step in the **citric acid cycle** involves the dehydration of the citrate ion (3°-alcohol) to an alkene. The alkene then undergoes hydration to form a 2°-alcohol (isocitrate). Citrate and isocitrate are isomers. The subsequent step allows for the

oxidation of the 2°-alcohol in isocitrate.



Biochemistry Link

Hydrolysis determines the length of time some drugs remain active. Dental anesthetics like novocain and lidocaine are short acting anesthetics that remain effective for a short time as a pain killer due to hydrolysis reaction. Shown below is the hydrolysis of aspirin which is a pain killer. The ester functional group of aspirin undergoes hydrolysis and forms products salicylic acid and acetic acid. Enzymes are biological catalysts that speed up the hydrolysis reaction.



Although the participating compounds in living systems are complex, the reactions are the same. There is either elimination of water from the starting material, or addition of a water molecule to a starting material, or hydrolysis with water. The idea is that if you know the chemistry of a particular functional group, you know the chemistry of hundreds of different compounds that contain the particular functional group.

This page was constructed from the following sources by Deboleena Roy (American River College):

- Hydrolysis 15.8: Hydrolysis of Esters - Chemistry LibreTexts
- Hydration 14.4: Reactions that Form Alcohols - Chemistry LibreTexts
- Dehydration 14.5: Reactions of Alcohols - Chemistry LibreTexts

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5.6: Classification of alcohols

Learning Objectives

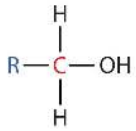
- Identify the general structure for an alcohol.
- Identify the structural feature that classifies alcohols as primary, secondary, or tertiary.

An alcohol is an organic compound with a hydroxyl (OH) functional group on an alkane type of carbon atom or a tetrahedral carbon atom. Because OH is the functional group of all alcohols, we often represent alcohols by the general formula ROH, where R is an alkyl group (R stands for Rest of the Molecule). Alcohols are common in nature. Most people are familiar with ethyl alcohol (ethanol), the active ingredient in alcoholic beverages, but this compound is only one of a family of organic compounds known as alcohols. The family also includes such familiar substances as cholesterol and the carbohydrates. Methanol (CH_3OH) and ethanol (CH_3CH_2OH) are a couple of examples of alcohols.

Classification of Alcohols

Some of the properties of alcohols depend on the number of carbon atoms attached to the specific carbon atom that is attached to the OH group. Alcohols can be grouped into three classes on this basis.

- A primary (1°) alcohol is one in which the carbon atom (in red) with the OH group is attached to *one* other carbon atom (in blue). Its general formula is RCH_2OH.



- A secondary (2°) alcohol is one in which the carbon atom (in red) with the OH group is attached to *two* other carbon atoms (in blue). Its general formula is R_2CHOH.



- A tertiary (3°) alcohol is one in which the carbon atom (in red) with the OH group is attached to *three* other carbon atoms (in blue). Its general formula is R_3COH.

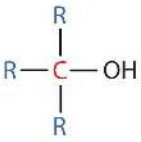
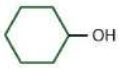


Table **5.6.1** names and classifies some of the simpler alcohols. Some of the common names reflect a compound's classification as secondary (*sec*-) or tertiary (*tert*-). These designations are not used in the IUPAC nomenclature system for alcohols. IUPAC stands for International Union of Pure and Applied Chemistry. Note that there are four butyl alcohols in the table, corresponding to the four butyl groups. One butyl group is ($CH_3CH_2CH_2CH_2$-), and three others are:

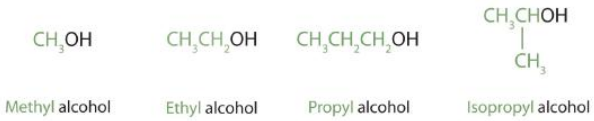


Table **5.6.1**: Classification and Nomenclature of Some Alcohols

Condensed Structural Formula	Class of Alcohol	Common Name	IUPAC Name
CH_3OH	—	methyl alcohol	methanol
CH_3CH_2OH	primary	ethyl alcohol	ethanol
$CH_3CH_2CH_2OH$	primary	propyl alcohol	1-propanol
$(CH_3)_2CHOH$	secondary	isopropyl alcohol	2-propanol
$CH_3CH_2CH_2CH_2OH$	primary	butyl alcohol	1-butanol
$CH_3CH_2CHOHCH_3$	secondary	*sec*-butyl alcohol	2-butanol
$(CH_3)_2CHCH_2OH$	primary	isobutyl alcohol	2-methyl-1-propanol
$(CH_3)_3COH$	tertiary	*tert*-butyl alcohol	2-methyl-2-propanol
	secondary	cyclohexyl alcohol	cyclohexanol

Nomenclature of Alcohols

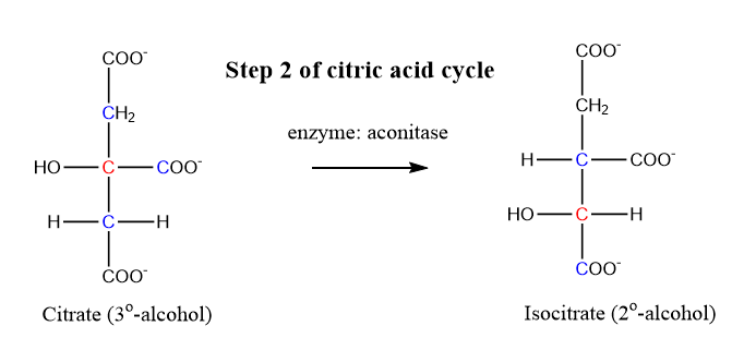
Alcohols with one to four carbon atoms are frequently called by common names, in which the name of the alkyl group is followed by the word *alcohol*:



Structural formula of methyl alcohol, ethyl alcohol, propyl alcohol, and isopropyl alcohol with the methyl, ethyl propyl, and isopropyl groups highlighted in green.

Biochemistry Link

The **citric acid cycle** is a series of reactions important to metabolism. One step in the process involves the conversion of citrate ion (3°-alcohol) to isocitrate ion (2°-alcohol). Citrate and isocitrate are isomers of each other.



Citrate is a 3°-alcohol, the alcohol (-OH) functional group is attached to a carbon atom (red) to three other carbon atoms (blue).

Isocitrate is a 2°-alcohol, the alcohol (-OH) functional group is attached to a carbon atom (red) to two other carbon atoms (blue).

Summary

Alcohols are classified according to the number of carbon atoms attached to the carbon atom bears the OH group.

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5.7: Oxidations and Reductions

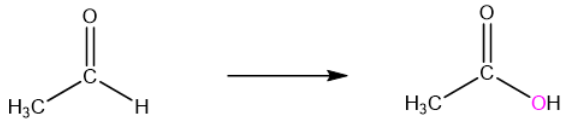
Learning Objectives

- Define oxidation and reductions
- Identify a chemical reaction as an oxidation-reduction reaction based on their definition.

Oxidation-reduction reactions are of central importance in organic chemistry and biochemistry. The burning of fuels that provides the energy to maintain our civilization and the metabolism of foods that furnish the energy that keeps us alive both involve oxidation-reduction reactions.

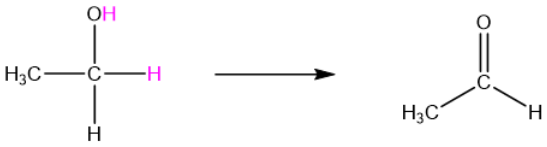
Oxidation

The origin of the term **oxidation** is "**adding oxygen**," so when oxygen is added to a molecule, the molecule is being oxidized. For example, the aldehyde molecule takes on an oxygen atom (marked purple) to become a carboxylic acid. Thus, the aldehyde molecule is being oxidized.



aldehyde carboxylic acid

Similarly, **oxidation can be defined in terms of the loss of hydrogen atoms**. If a molecule loses hydrogen atoms, the molecule is being oxidized. For example, in the conversion of the 1° alcohol to aldehyde, the alcohol loses two hydrogen atoms (marked purple).



1° alcohol aldehyde

Oxidation may also be **defined as the loss of electrons by an atom**. An example, is the single displacement reaction of sodium metal with water

$2Na(s) + 2HOH(l) \rightarrow H_2(g) + 2NaOH(aq)$

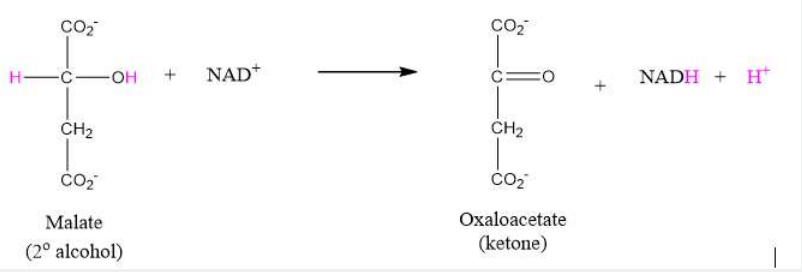
In this reaction sodium metal Na(s) loses and an electron to become sodium ion (Na^+) in the ionic compound NaOH. So, sodium atom is oxidized.

$Na \rightarrow Na^+ + e^-$

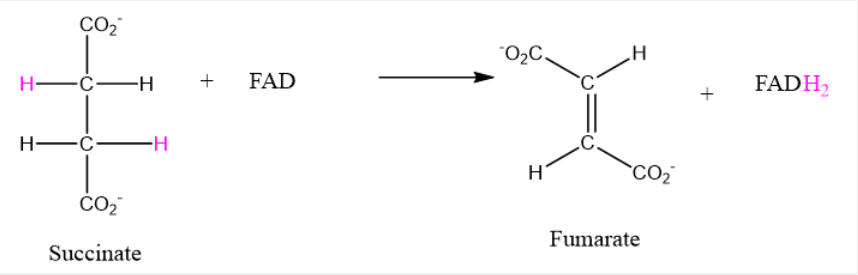
Applications of Oxidation in Biochemistry

Oxidation reactions are found in many metabolic process and happen within the living cells. Here are a couple of examples from the citric acid cycle.

a. The citric acid cycle is a series of reactions important to metabolism. One of the steps (step 8) in the cycle is the conversion of the malate (2°-alcohol) to oxaloacetate (ketone). The alcohol loses two hydrogen atoms (marked purple). Malate is **oxidized** because it has undergone oxidation. NAD^+ is the oxidizing agent. NAD^+ is reduced to NADH and H^+ ion.



b. Another step in the cycle (step 6) involves the conversion of the succinate to fumarate, succinate loses two hydrogen atoms (marked purple) to an alkene fumarate. FAD is the oxidizing agent. FAD is reduced to $FADH_2$.



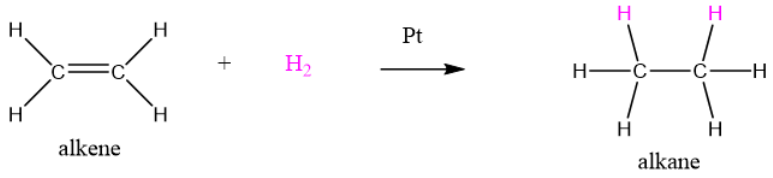
NAD^+ and FAD are biological oxidizing agents called oxidized coenzymes.

Reduction

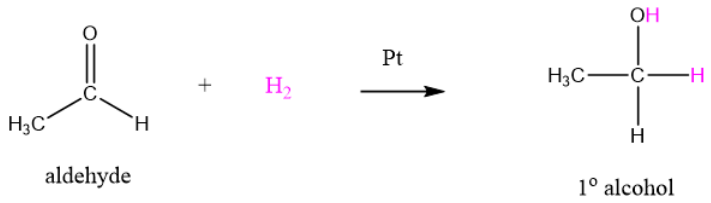
The reverse is true for **reduction**: if a molecule **loses oxygen** atoms, the molecule is being reduced.

If a molecule **adds hydrogen atoms**, it is being reduced.

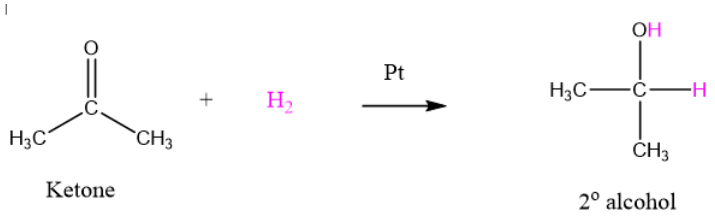
For example, in the conversion of alkene into alkane, hydrogen atoms are added to alkene (in purple), so the alkene is being reduced.



In the conversion of aldehyde into 1° alcohol below, hydrogen atoms are added (in purple) to the aldehyde, so the aldehyde is being reduced.



In the conversion of ketone into 2° alcohol below, hydrogen atoms are added (in purple) to the ketone, so the ketone is being reduced.

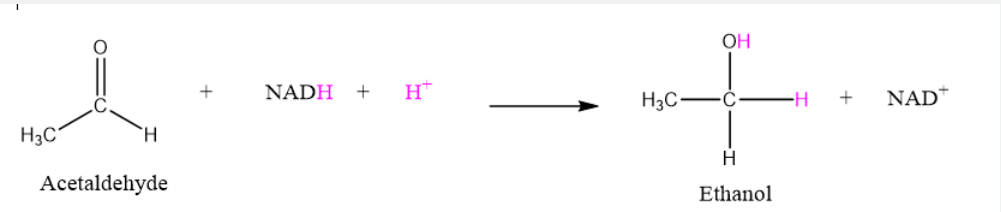


Reduction may also be **defined as the gain of electrons by an atom**.

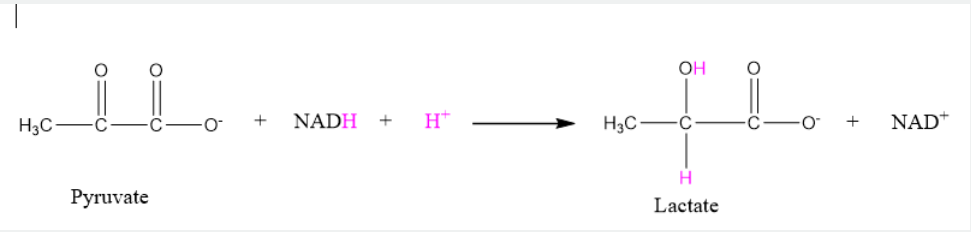
Applications of Reduction in Biochemistry

Reduction reactions are found in many metabolic process and happen within the living cells. Here are a couple of examples from glucose metabolism.

a. In the conversion of acetaldehyde to ethanol, the aldehyde gains two hydrogen atoms (marked purple). Acetaldehyde is **reduced** because it has undergone reduction. NADH is the reducing agent. NADH is reduced to NAD^+. This step is important in yeast during the **fermentation of glucose** from fruits and grains to produce ethanol (drinking alcohol).



b. In the conversion of pyruvate to lactate, the ketone functional group gains two hydrogen atoms (marked purple). Pyruvate is **reduced** because it has undergone reduction. NADH is the reducing agent. NADH is reduced to NAD^+. Lactate is produced during anaerobic (O_2 deficient) conditions of the cell such as when a person is exercising vigorously.

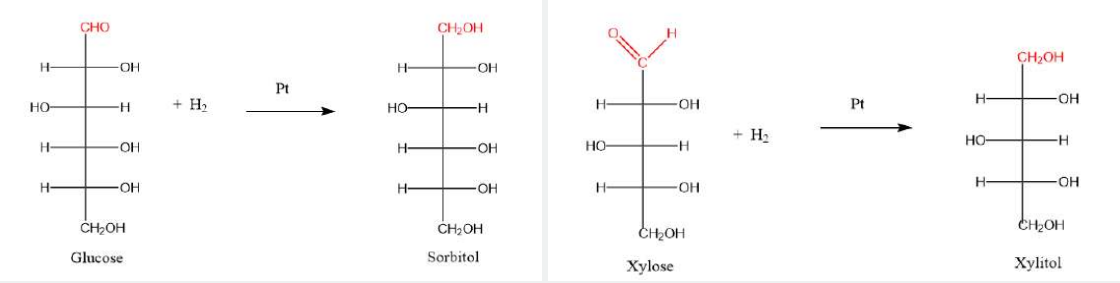


NADH and $FADH_2$ are biological reducing agents called **reduced coenzymes.**

Other Applications of Reduction

a. Artificial Sweeteners:

Alcohol sugars such as sorbitol and xylitol are made by the reduction reaction of aldehydes.

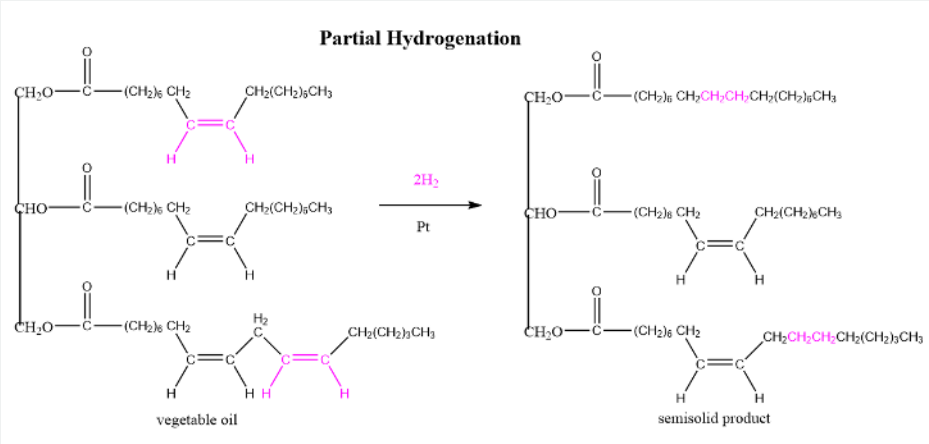


Sorbitol is used to sweeten food for diabetics. The aldehyde functional group in glucose gains two hydrogen atoms from H_2 to form sorbitol (a 1°-alcohol).

Xylitol is used for sugarless gums. The aldehyde functional group in xylose gains two hydrogen atoms from H_2 to form xylitol (a 1°-alcohol).

b. Partially Hydrogenated Vegetable Oils

Partial hydrogenation of a triglyceride converts a vegetable oil into a semisolid product called margarine which has a good consistency for cooking but does not contain cholesterol found in animal fats. The vegetable oil shown below has 4 alkene functional groups. Two of the alkenes (shown in purple) undergo hydrogenation to alkane.



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5.6: Redox Reactions in Organic Chemistry and Biochemistry - Chemistry LibreTexts

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5.8: Exothermic and Endothermic Reactions

Learning Objectives

- Determine if a chemical process is exothermic or endothermic.

What happens when you take a basketball, place it halfway up a playground slide, and then let it go? The basketball rolls down the slide. What happens if you do it again? Does the basketball roll down the slide? It should.

If you were to perform this experiment over and over again, do you think the basketball would ever roll *up* the slide? Probably not. Why not? Well, for starters, in all of our experience, the basketball has always moved to a lower position when given the opportunity. The gravitational attraction of Earth exerts a force on the basketball, and given the chance, the basketball will move down. We say that the basketball is going to a lower *gravitational potential energy*. The basketball can move up the slide, but only if someone exerts some effort (that is, work) on the basketball. A general statement, based on countless observations over centuries of study, is that all objects tend to move spontaneously to a position of minimum energy unless acted on by some other force or object.

A similar statement can be made about atoms in compounds. Atoms bond together to form compounds because in doing so they attain lower energies than they possess as individual atoms. A quantity of energy, equal to the difference between the energies of the bonded atoms and the energies of the separated atoms, is released, usually as heat. That is, the bonded atoms have a lower energy than the individual atoms do. *When atoms combine to make a compound, energy is always given off, and the compound has a lower overall energy.* In making compounds, atoms act like a basketball on a playground slide; they move in the direction of decreasing energy. We can reverse the process, just as with the basketball. If we put energy into a molecule, we can cause its bonds to break, separating a molecule into individual atoms.

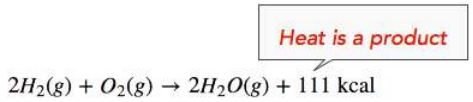
Heat of Reaction, ΔH

When a chemical reaction occurs, the atoms in the reactants rearrange their chemical bonds to make products. The new arrangement of bonds does not have the same total energy as the bonds in the reactants. Therefore, when chemical reactions occur, *there will always be an accompanying energy change called the* ***Heat of Reaction***.

Let's consider the reaction of 2 mols of hydrogen gas (H_2) with 1 mol of oxygen gas (O_2) to give 2 mol water:

$$\mathbf{2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g)}$$

The enthalpy change (ΔH) of the reaction is approximately −111 kcal/mol. This means that bonds in the products are stronger than the bonds in the reactants by about 111 kcal/mol. Because the bonds in the products are stronger than those in the reactants, the reaction releases more energy than it consumes. This excess energy is **released** as heat, so the reaction is **exothermic**. Hence, we can re-write the reaction with the heat released (111 kcal) on the **product side** of the equation, as follows:



We can also re-write the reaction equation with the ΔH information (see below). Note that an **exothermic** reaction has a **negative ΔH** value.

$$\mathbf{2\,H_2(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g)} \qquad \Delta H = -111\ \text{kcal}$$

Exothermic and Endothermic Reactions

In endothermic and exothermic reactions, energy can be thought of as either a reactant of the reaction or a product.

In an **exothermic** reaction, heat is released (considered a product) and the energy of the system decreases (ΔH is negative). A chemical reaction is **exothermic** if heat is released by the system into the surroundings. Because the surroundings is gaining heat from the system, **the temperature of the surroundings increases.** See Figure 5.8.1.

When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of **1 mol** of methane releases 890.4 kilojoules of heat energy. This information can be shown as part of the balanced equation in two ways. First, the amount of heat released can be written in the **product** side of the reaction. Another way is to write the **heat of reaction** (ΔH) information with a **negative** sign, **−890.4 kJ**.

Heat is a product

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(l) + 890.4\ \text{kJ}$$

$$\mathbf{CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(l) \quad \Delta H = -890.4\ kJ}$$

Endothermic reactions require energy, so energy is a reactant. Heat flows from the surroundings to the system (reaction mixture) and the energy of the system increases (ΔH is positive). In the course of an **endothermic** process, the system gains heat from the surroundings and so the **temperature of the surroundings decreases (gets cold).** See Figure 5.8.1.

When **1 mol** of calcium carbonate decomposes into **1 mol** of calcium oxide and **1 mol** of carbon dioxide, **177.8 kJ** of heat is absorbed. Because the heat is absorbed by the system, the **177.8 kJ** is written as a **reactant**. The **ΔH is positive** for an endothermic reaction.

Heat is a reactant

$$CaCO_3\,(s) + 177.8\ \text{kJ} \rightarrow CaO\,(s) + CO_2\,(g)$$

$$\mathbf{CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g) \quad \Delta H = +177.8\ kJ}$$

Green plants are capable of synthesizing glucose ($C_6H_{12}O_6$) from carbon dioxide (CO_2) and water (H_2O) by using solar energy in the process known as photosynthesis:

$$6CO_2 + 6H_2O + 686\ \text{kcal} \rightarrow C_6H_{12}O_6 + 6O_2$$

(The 686 kcal came from solar energy and this is an example of an endothermic reaction.)

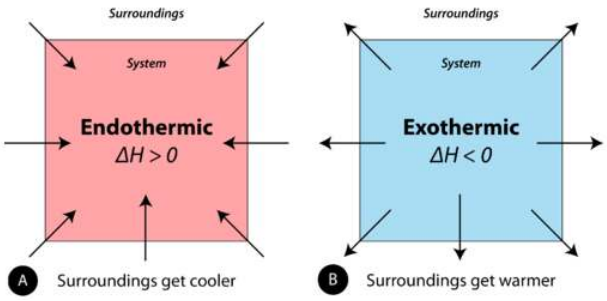


Figure 5.8.1: (A) Endothermic reaction. (B) Exothermic reaction.
Endothermic reaction: surroundings get cooler and ΔH is greater than 0,
Exothermic reaction: surroundings get warmer and ΔH is less than 0

Example 5.8.2

Is each chemical reaction exothermic or endothermic? What is the ΔH of the reaction?

a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 213$ kcal
b. $N_2(g) + O_2(g) + 45$ kcal $\rightarrow 2NO(g)$

Solution

a. Because energy (213 kcal) is a product, energy is given off by the reaction. Therefore, this reaction is exothermic. ΔH = -213 kcal.
b. Because energy (45 kcal) is a reactant, energy is absorbed by the reaction. Therefore, this reaction is endothermic. ΔH = +45 kcal

? Exercise 5.8.2

Is each chemical reaction exothermic or endothermic? What is the ΔH of the reaction?

a. $H_2(g) + F_2(g) \rightarrow 2HF\ (g) + 130\ kcal$
b. $2C(s) + H_2(g) + 5.3\ kcal \rightarrow C_2H_2(g)$

Answer

a. The energy (130 kcal) is produced, hence the reaction is exothermic. ΔH = -130 kcal

b. The energy (5.3 kcal) is supplied or absorbed to react, hence, the reaction is endothermic. ΔH = +5.3 kcal.

Energy Diagrams

Endothermic and exothermic reactions can be visually represented by *energy-level diagrams* like the ones in **Figure 5.8.2**. In endothermic reactions, the reactants have stronger bonds than the products. **Strong bonds have lower potential energy than weak bonds**. Hence, the **energy of the reactants is lower than that of the products**. This type of reaction is represented by an "uphill" energy-level diagram shown in Figure **5.8.2*A***. For an endothermic chemical reaction to proceed, the reactants must absorb energy from their environment to be converted to products.

In an exothermic reaction, the bonds in the product have stronger bonds than the reactants. In other words, the **energy of the products is lower than the energy of the reactants,** hence is energetically downhill, shown in Figure **5.8.2*B***. Energy is given off as reactants are converted to products. The energy given off is usually in the form of heat (although a few reactions give off energy as light). In the course of an exothermic reaction, heat flows from the system to its surroundings, and thus, gets warm.

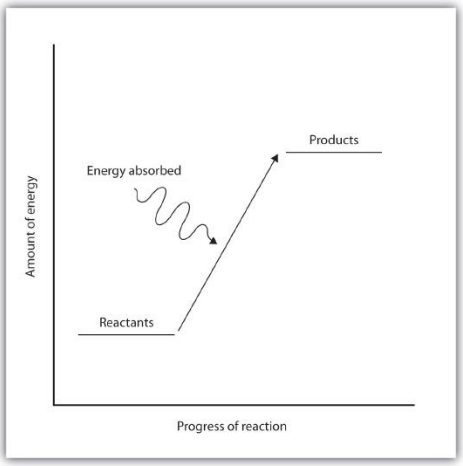


Figure **5.8.2*A***: Endothermic Reactions

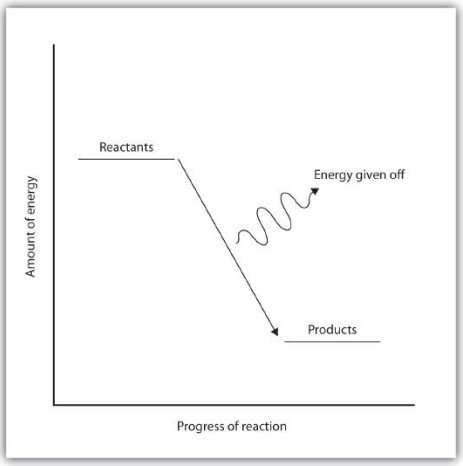


Figure **5.8.2*B***: Exothermic Reactions

Table **5.8.2**: Endothermic and Exothermic Reactions

Endothermic Reactions	Exothermic Reactions
Heat is absorbed by reactants to form products.	Heat is released.
Heat is absorbed from the surroundings; as a result, the surroundings get cold.	Heat is released by the reaction to surroundings; surroundings feel hot.
ΔH_{rxn} is positive	ΔH_{rxn} is negative
The bonds broken in the reactants are stronger than the bonds formed in the products	The bonds formed in the products are stronger than the bonds broken in the reactants
The reactants are lower in energy than the products	The products are lower in energy than the reactants
Represented by an "uphill" energy diagram	Represented by an "downhill" energy diagram

Key Takeaway

- Chemical processes are labeled as exothermic or endothermic based on whether they give off or absorb energy, respectively.

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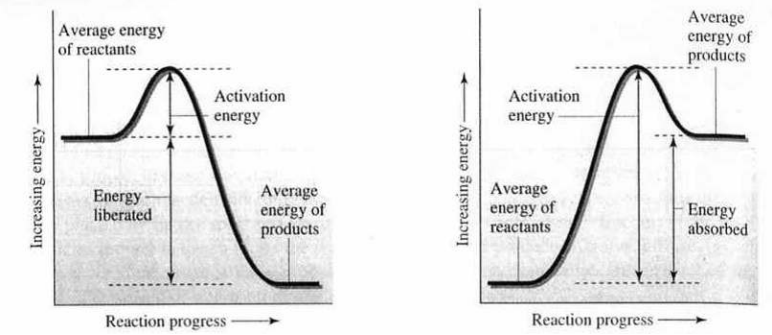
5.9: Reaction Rate

Reaction Rates (Kinetics)

When a reaction releases energy, the chemical change is exothermic. Release of energy does not guarantee that the reaction will take place within a reasonable amount of time. A measure of how quickly products form, is called the reaction rate.

Reaction Energy Diagrams

Reaction energy diagrams summarize the important information about a reaction in a single graph. The y-axis represents the energy and the x-axis represents the reaction progress (time). The relative energies of the reactants and products are determined by the potential energy stored within the molecules as chemical bonds. For an exothermic reaction A, the products are lower in energy than the reactants resulting in a release of heat. For an endothermic reaction B, the products are higher in energy than the reactants requiring the constant addition of energy to keep the reaction going. The reactants must possess a minimum amount of energy for a reaction to occur. In a reaction energy diagram, the minimum energy is described as the "activation energy" (E_a'). Because the activation energy can be viewed as a barrier to a reaction, a larger activation energy indicates a slower the reaction rate and a smaller activation energy indicates a faster the reaction rate.

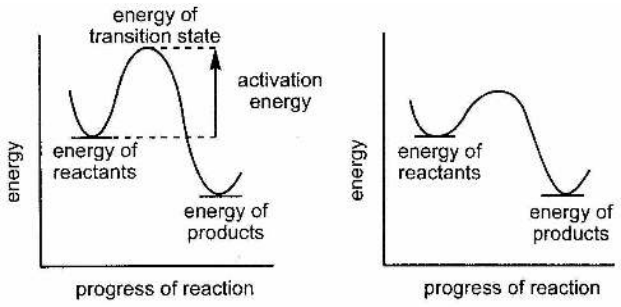


Reaction A Reaction B

Figure 5.9.1: Reaction A is exothermic, Reaction B is endothermic.

Catalysts are substances that speed up a reaction without being consumed or changed by the reaction. Because of these properties, catalysts can be added in small quantities to increase reaction rates. Catalysts increase reaction rates by decreasing the activation energy of a reaction. The transition state represents the reactants in the correct orientation with the minimum energy for a reaction to occur. The energy difference between the reactants and the transition state is the activation energy. The addition of a catalyst to Reaction C increases the reaction rate by decreasing the activation energy as shown below. Catalysts do not affect the ΔH of the reaction. Most chemical reactions that take place in living things are catalyzed by proteins called **enzymes**.

Reaction C



No Catalyst Reaction With Catalyst Reaction

Figure 5.9.2: Addition of a catalyst lowers the activation energy of the reaction and speeds up a chemical reaction.

Another way to alter the rate of a reaction is to change the **temperature**. The higher the temperature the faster the reaction. At higher temperatures the kinetic energy of the of the reactants are greater, collisions between reactants take place more frequently and with greater energy, so reactants are able to find their way over the activation barrier to form products.

The rate of the reaction can also be varied by changing the **concentration** of one or more reactants. Concentration refers to the amount of substance that is present in a mixture. A higher concentration of reactants can increase the rate of a reaction. The more crowded the reaction mixture is with reactants, the more likely is the collisions between reactant molecules.

Example 5.9.1

1. Use the reaction diagrams from Figure 5.9.1 to answer the following questions.
    a. Label the ΔH and indicate whether it is positive or negative.
    b. What is activation energy?
    c. Overlay a second reaction energy curve for Reaction B when a catalyst is added.
2. Predict which reaction occurs at a faster rate in figure 5.9.2.

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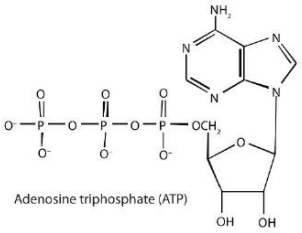
5.10: Catabolism

Learning Objectives

- Describe the importance of ATP as a source of energy in living organisms.
- To understand the stages of catabolism.

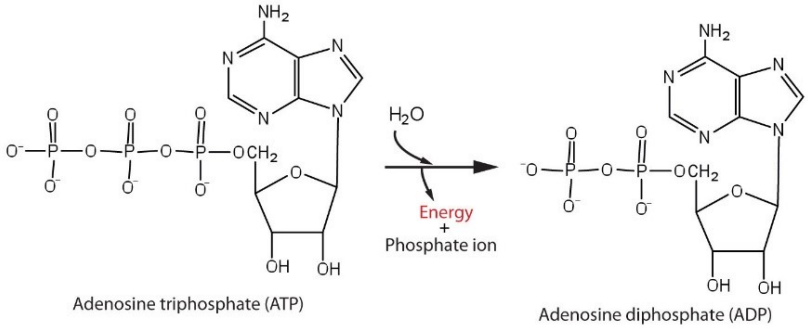
ATP

Adenosine triphosphate (ATP), is a nucleotide composed of adenine, ribose, and three phosphate groups. It is the most important of the energy-rich compounds in a cell.



Energy-rich compounds are substances having particular structural features that lead to a release of energy after hydrolysis. As a result, these compounds are able to supply energy for biochemical processes that require energy. The structural feature important in ATP is the phosphoric acid anhydride linkage. This linkage is hydrolyzed when ATP is converted to adenosine diphosphate (ADP). In this hydrolysis reaction, the products contain less energy than the reactants; there is a release of energy (ΔH = - 7.4 kcal). The general equation for ATP hydrolysis is as follows:

$$ATP + H_2O \rightarrow ADP + P_i + 7.4 \text{ kcal/mol}$$



Guanosine triphosphate (GTP), is another nucleotide composed of guanine, ribose, and three phosphate groups. It is addressed as ATP equivalent (ATP_{eq}).

Overview of Metabolism

Animals obtain chemical energy from the food they eat. The food they eat are carbohydrates, fats, and proteins. The energy from food is derived through reactions defined collectively as *catabolism*. We can think of catabolism as occurring in three stages (Figure **5.10.1**). In stage I, carbohydrates, fats, and proteins are broken down into their individual monomer units: carbohydrates into simple sugars (monosaccharides), fats into fatty acids and glycerol, and proteins into amino acids. Stage I of catabolism is the breakdown of food molecules by hydrolysis reactions into the individual monomer units. This occurs in the mouth, stomach, and small intestine. It is referred to as digestion.

In stage II, these monomer units (or building blocks) are further broken down through different reaction pathways. They produces energy-rich molecules such as NADH, $FADH_2$, and ATP or GTP as common end products. NADH and $FADH_2$ can then be used in stage III to produce even more ATP. Fatty acids undergo fatty acid catabolism and is converted to Acetyl-CoA. Monosaccharides

such as glucose undergoes glycolysis to pyruvate, which then may be converted to Acetyl-CoA. There a 20 amino acids obtained from the hydrolysis of proteins. Some of them are converted to pyruvate, others to Acetyl-CoA, which the rest are converted to one of the intermediates in the citric acid cycle or the glycolysis pathway.

Acetyl-CoA is a two-carbon molecule that enters in the citric acid cycle in stage III. This cycle is understood in 8 steps and it happens in the mitochondria of the cell. Energy rich molecules such as NADH, $FADH_2$, and GTP are produced during this cycle. The reduced coenzymes such as NADH and $FADH_2$ enter the electron transport chain (ETC) in stage III and are converted to ATP.

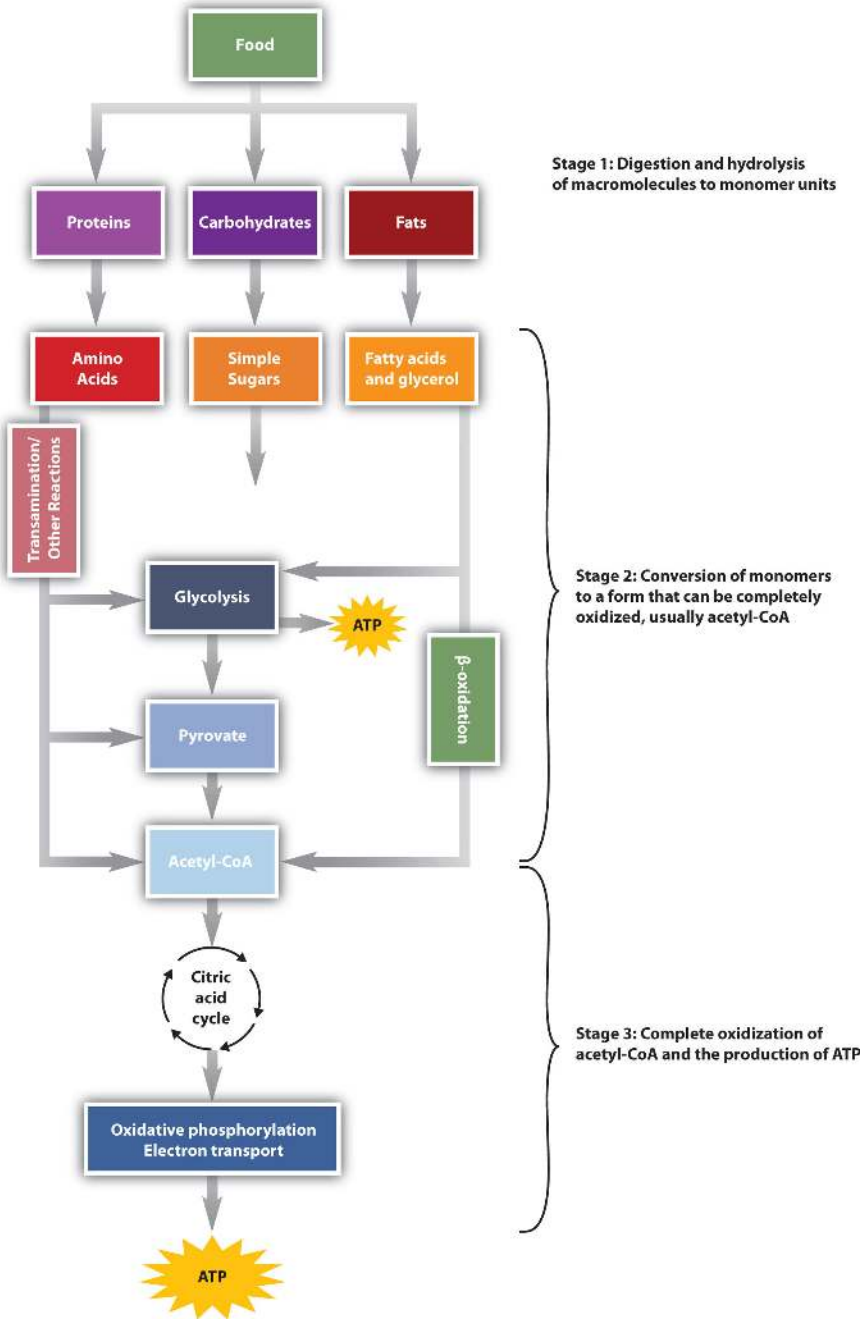


Figure **5.10.1**: Energy Conversions

Summary

The conversion of food into cellular energy (as ATP) occurs in three stages. Catabolism of monosaccharides, fatty acids, and amino acids released in stage I of catabolism occurs in stages II and III of catabolism. Through Acetyl-CoA these steps converge into the citric acid cycle.

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5.11: Citric Acid Cycle

Learning Objectives

- Describe the reactions of the citric acid cycle.
- Describe the function of the citric acid cycle and identify the products produced.

Citric Acid Cycle

Acetyl-CoA is a two-carbon molecule that enters a circular pathway of reactions known collectively as the citric acid cycle (or Krebs cycle). The cycle brings about the oxidation of the two carbons of acetyl-CoA to carbon dioxide and water. It was first proposed by Hans Krebs in 1937. He was awarded the 1953 Nobel Prize in Physiology or Medicine. Acetyl-CoA's entrance into the citric acid cycle is the beginning of stage III of catabolism. The citric acid cycle produces guanosine triphosphate (GTP), three reduced nicotinamide adenine dinucleotide (NADH), reduced flavin adenine dinucleotide ($FADH_2$), and metabolic intermediates required for the synthesis of more ATP molecules.

At first glance, the citric acid cycle appears rather complex. All the reactions, however, are familiar types in organic chemistry: hydration, oxidation, decarboxylation, and hydrolysis. Each reaction of the citric acid cycle is numbered. The two carbon atoms of Acetyl-CoA are highlighted in red. Each intermediate in the cycle is a carboxylic acid, existing as an anion at physiological pH of around 7. All the reactions occur within the mitochondria, which are small organelles within the cells of plants and animals. This part of the cell is known as the **powerhouse of the cell** since this is where most of the ATP is produced.

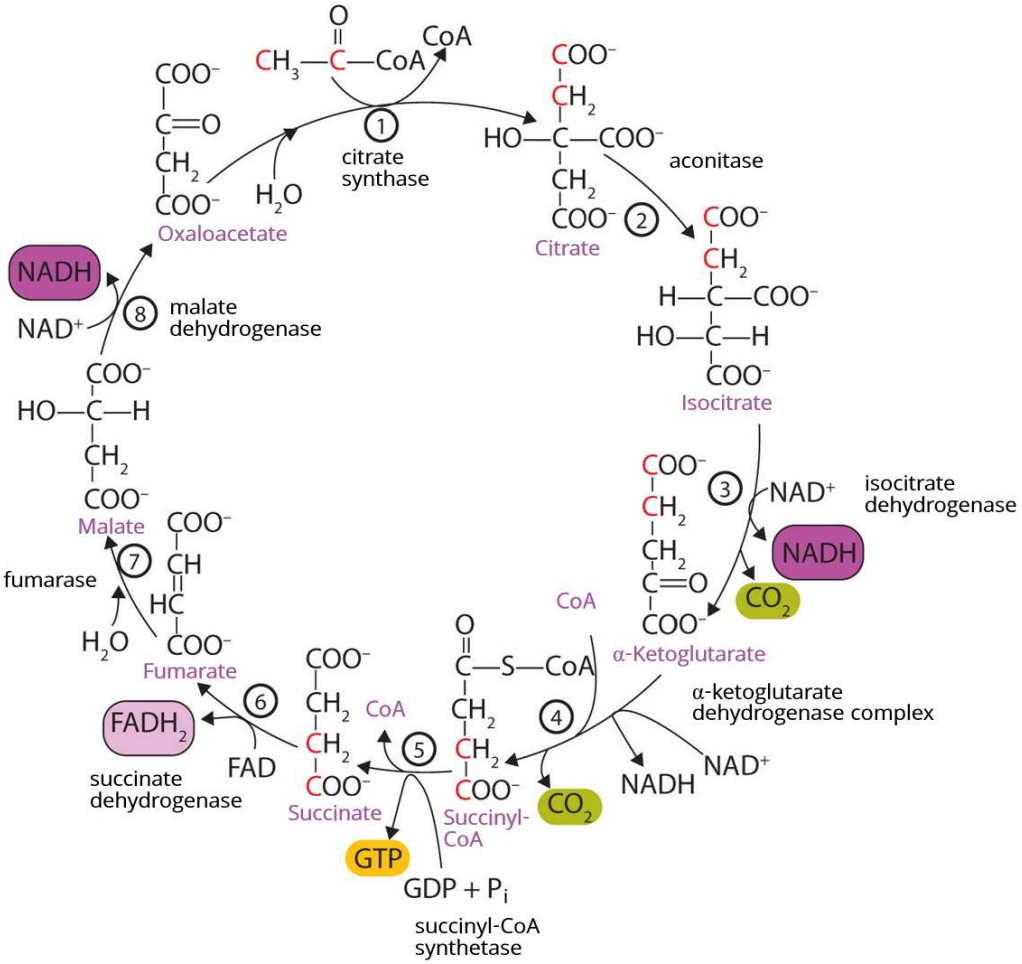


Figure 5.11.1: The Citric Acid Cycle

Step 1

In the first reaction, acetyl-CoA enters the citric acid cycle, and the two-carbons of the acetyl group reacts with a four-carbon anion called oxaloacetate, yielding a six-carbon anion called citrate. This step also involves addition of water. Note that this step releases coenzyme A. The reaction is catalyzed by *citrate synthase*.

Step 2

In this step, *aconitase* catalyzes the isomerization of citrate to isocitrate. In this reaction, a tertiary alcohol, which cannot be oxidized, is converted to a secondary alcohol, which can be oxidized in the next step. Citrate and isocitrate are isomers.

Step 3

Isocitrate then undergoes a reaction known as oxidative decarboxylation because the 2°-alcohol is oxidized and the molecule is shortened by one carbon atom with the release of carbon dioxide (decarboxylation). The reaction is catalyzed by *isocitrate dehydrogenase*, and the product of the reaction is α-ketoglutarate. The oxidizing agent of the 2°-alcohol is the coenzyme nicotinamide adenine dinucleotide (NAD^+) which is reduced to NADH.

Step 4

The fourth step is another oxidative decarboxylation. This time α-ketoglutarate is converted to succinyl-CoA, and another molecule of NAD^+ is reduced to NADH. The *α-ketoglutarate dehydrogenase* catalyzes this reaction.

Note

So far, in the first four steps, two carbon atoms have entered the cycle as an acetyl group, and two carbon atoms have been released as molecules of carbon dioxide. The remaining reactions of the citric acid cycle use the four carbon atoms of the succinyl group to resynthesize a molecule of oxaloacetate, which is the compound needed to combine with an incoming acetyl group and begin another round of the cycle.

Step 5

In the fifth reaction, the energy released by the hydrolysis of the high-energy thioester bond of succinyl-CoA is used to make guanosine triphosphate (GTP) from guanosine diphosphate (GDP) and inorganic phosphate (P_i) in a reaction catalyzed by *succinyl-CoA synthetase*. This step is the only reaction in the citric acid cycle that directly forms a high-energy phosphate compound GTP. GTP is considered equivalent in energy to ATP.

Step 6

Succinate dehydrogenase then catalyzes the removal of two hydrogen atoms from succinate, forming fumarate. This oxidation-reduction reaction uses flavin adenine dinucleotide (FAD), rather than NAD^+, as the oxidizing agent.

Step 7

In step 7, a molecule of water is added to the double bond of fumarate to form malate in a hydration reaction catalyzed by *fumarase*.

Step 8

One cycle is completed with the oxidation of malate to oxaloacetate, catalyzed by *malate dehydrogenase*. This is the third oxidation-reduction reaction. It uses NAD^+ as the oxidizing agent to oxidize the 2° alcohol in malate to a ketone.

Note

Oxaloacetate can accept two carbons from the reaction with acetyl-CoA, allowing the cycle to begin again.

Summary

The citric acid cycle consists of eight step and are made of familiar types of reactions in organic chemistry: hydration, oxidation, decarboxylation, and hydrolysis. The harvest of the citric acid cycle consists of 3NADH, $FADH_2$, and GTP.

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5.12: End of Chapter Problems

Chemical Reactions

1. Which of the following is a chemical reaction?

a. Freezing liquid Mercury.
b. Adding yellow to blue to make green.
c. Cutting a piece of paper into two pieces.
d. Dropping a sliced orange into a vat of sodium hydroxide.
e. Filling a balloon with natural air.
f. Painting a wall blue
g. A bicycle rusting
h. Ice cream melting
i. Scratching a key across a desk
j. Making a sand castle

2. Which of the following is not a chemical reaction?

a. Shattering glass with a baseball
b. Corroding Metal
c. Fireworks exploding
d. Lighting a match
e. Baking a cake
f. Frying an egg
g. Slicing carrots
h. A macbook falling out of a window
i. Creating ATP in the human body
j. Dropping a fizzy tablet into a glass of water

Balancing and Classifying Chemical Equations

1. Balance the following chemical equations.

- $Hg(\ell)+O_2(g)\rightarrow HgO(s)$
- $C_2H_4(g)+O_2(g)\rightarrow CO_2(g)+H_2O(g)$
- $Mg(NO_3)_2(s)+Li(s)\rightarrow Mg(s)+LiNO_3(s)$
- $N_2(g)+O_2(g)\rightarrow NO_2(g)$
- $Pb(NO_3)_2(aq)+FeCl_3(aq)\rightarrow Fe(NO_3)_3(aq)+PbCl_2(s)$
- $C_6H_{14}(l)+O_2(g)\rightarrow CO_2(g)+H_2O(g)$
- $C_7H_{16}(l)+O_2(g)\rightarrow CO_2(g)+H_2O(g)$
- $C_8H_{18}(l)+O_2(g)\longrightarrow CO_2(g)+H_2O(g)$
- $Pb(NO_3)_2(aq)+NaCl(aq)\rightarrow NaNO_3(aq)+PbCl_2(s)$
- $KNO_3(s) \rightarrow KNO_2(s) + O_2(g)$
- $SiO_2(s) + HF(aq) \rightarrow SiF_4(aq) + H_2O(l)$
- $SO_2 + O_2 \rightarrow SO_3$
- $NO + O_2 \rightarrow NO_2$
- $K + Cl_2 \rightarrow KCl$
- $CH_4 + Cl_2 \rightarrow CH_2Cl_2 + HCl$
- $Al + CuO \rightarrow Al_2O_3 + Cu$
- $Mg + P_4 \rightarrow Mg_3P_2$
- $CaCl_2 + AgNO_3 \rightarrow AgCl + Ca(NO_3)_2$
- $(NH_4)_2SO_4 + CaBr_2 \rightarrow NH_4Br + CaSO_4$

- $Cl_2 + NaBr \rightarrow NaCl + Br_2$
- $BiCl_3 + H_2O \rightarrow Bi_2O_3 + HCl$

2. Classify each reaction in question 1. Your choices are addition, decomposition, single replacement, double replacement, and combustion reactions.

3. The oxyacetylene torch used for welding burns acetylene gas (C_2H_2). Balance the equation for the incomplete combustion of acetylene.

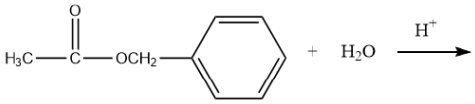
$C_2H_2(g) + O_2(g) \rightarrow CO(g) + H_2O(g)$

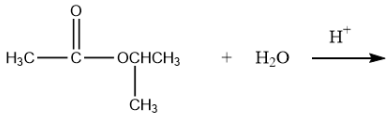
4. One source of acid rain is the addition reaction between sulfur dioxide, oxygen, and water to form sulfuric acid (H_2SO_4). Balance the equation for this reaction.

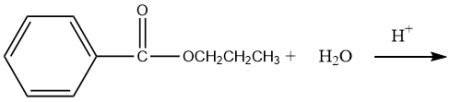
$SO_2(g) + O_2(g) + H_2O\ (g) \rightarrow H_2SO_4(aq)$

Reactions with water

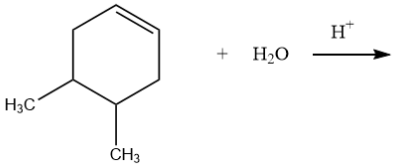
Draw the product of each of the following reactions. Classify each reaction. Your choices are hydration, dehydration, or hydrolysis.

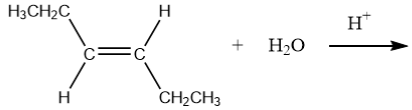
1. 

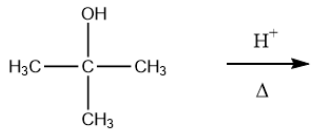
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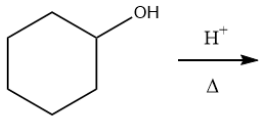
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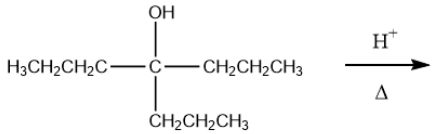
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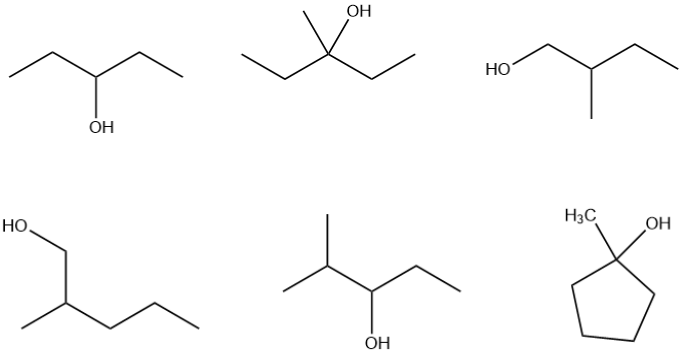
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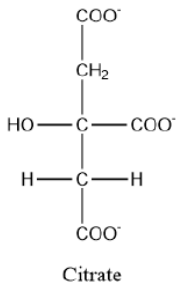
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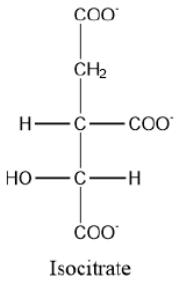
Classification of Alcohols

Identify each alcohol as being 1º, 2º, or 3º.





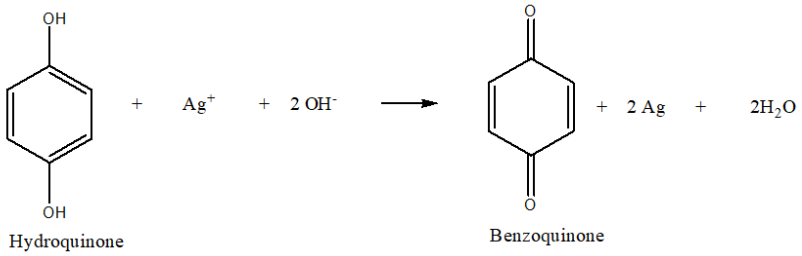
Citrate



Isocitrate

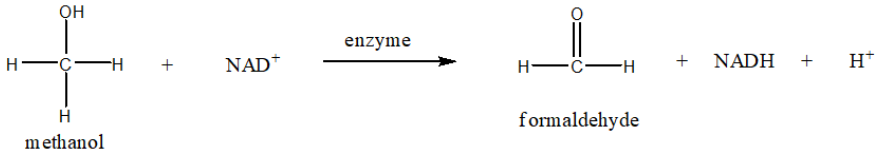
Oxidations and Reductions

1. One of the reactions used to develop photographic film is shown below. Hydroquinone, a developer reacts with silver ions to form silver metal, which gives the black color in a film negative.



a. Is hydroquinone oxidized or is it reduced?
b. What is the function of Ag^+ in this reaction?
c. What is the reducing agent? (Hydroquinone or Ag^+)

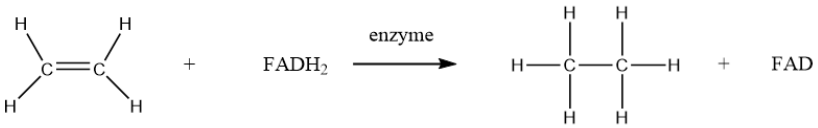
2. The toxic effects of methanol (CH_3OH) are due to its enzymatic conversion into formaldehyde (CH_2O) in the liver.

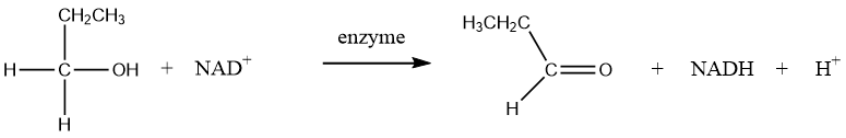


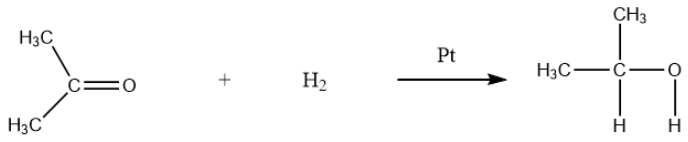
a. In this reaction is methanol oxidized or reduced?

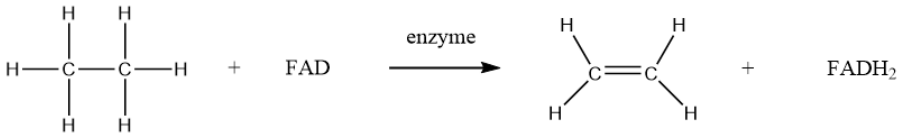
b. Identify the oxidizing agent.

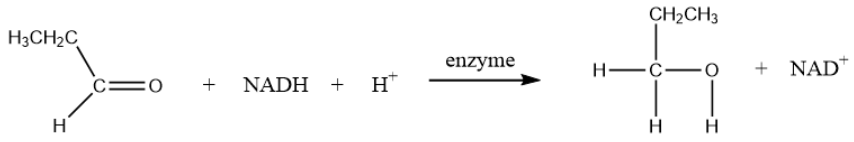
3. To which reaction types do the following reactions belong? You may choose from oxidation or reduction. Identify the agent and the catalyst. Identify the functional group/class in the reactant and product.

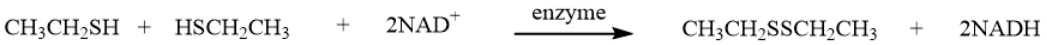
a.


b.


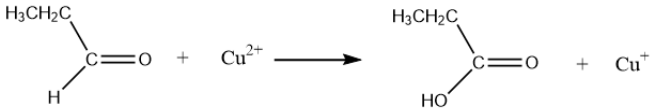
c.


d.


e.




f.



g.

Heat of Reaction

1. Label each reaction below as exothermic or endothermic? What is the ΔH of the reaction?

a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(\ell) + 213$ kcal

b. $N_2(g) + O_2(g) + 45$ kcal $\rightarrow 2NO(g)$

c. $H_2(g) + F_2(g) \rightarrow 2HF\ (g) + 130$ kcal

d. $2C(s) + H_2(g) + 5.3$ kcal $\rightarrow C_2H_2(g)$

2. Draw energy diagrams for reactions 1a & b. Label the x- and y-axis, reactant, products, E_a and ΔH.

3. What is a catalyst?

4. One of the enzyme-catalyzed reactions involved in step 1 of the citric acid cycle is the conversion oxaloacetate to citrate. The word equation for the reaction is shown below.

$$\text{Oxaloacetate} + \text{Acetyl-CoA} + H_2O \rightarrow \text{Citrate} + \text{CoA} + 3.3 \text{ kcal}$$

Answer the following questions based on this reaction.

a. What is the value for ΔH for the reaction?

b. Is the reaction exothermic or endothermic?

c. How would you expect decreasing the concentration of oxaloacetate to affect the reaction rate?

d. How would removing the enzyme affect the reaction rate?

e. The normal body temperature is around 98 ºF. How would the reaction rate be affected if the person suffered hypothermia?

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CHAPTER OVERVIEW

6: Solutions

6.1: Classifying Matter According to Its Composition
6.2: Solutions
6.3: The Effect of Temperature on Solubility of Solids, Liquids, and Gases
6.4: Solubility Classification for Biochemical Compounds
6.5: Solution Concentrations and Dilutions
6.6: Osmosis and Diffusion
6.7: Hemodialysis
6.8: End of Chapter Problems

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6.1: Classifying Matter According to Its Composition

Learning Objectives

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

Matter can be classified into two broad categories: pure substances and mixtures (Figure **6.1.1**). A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

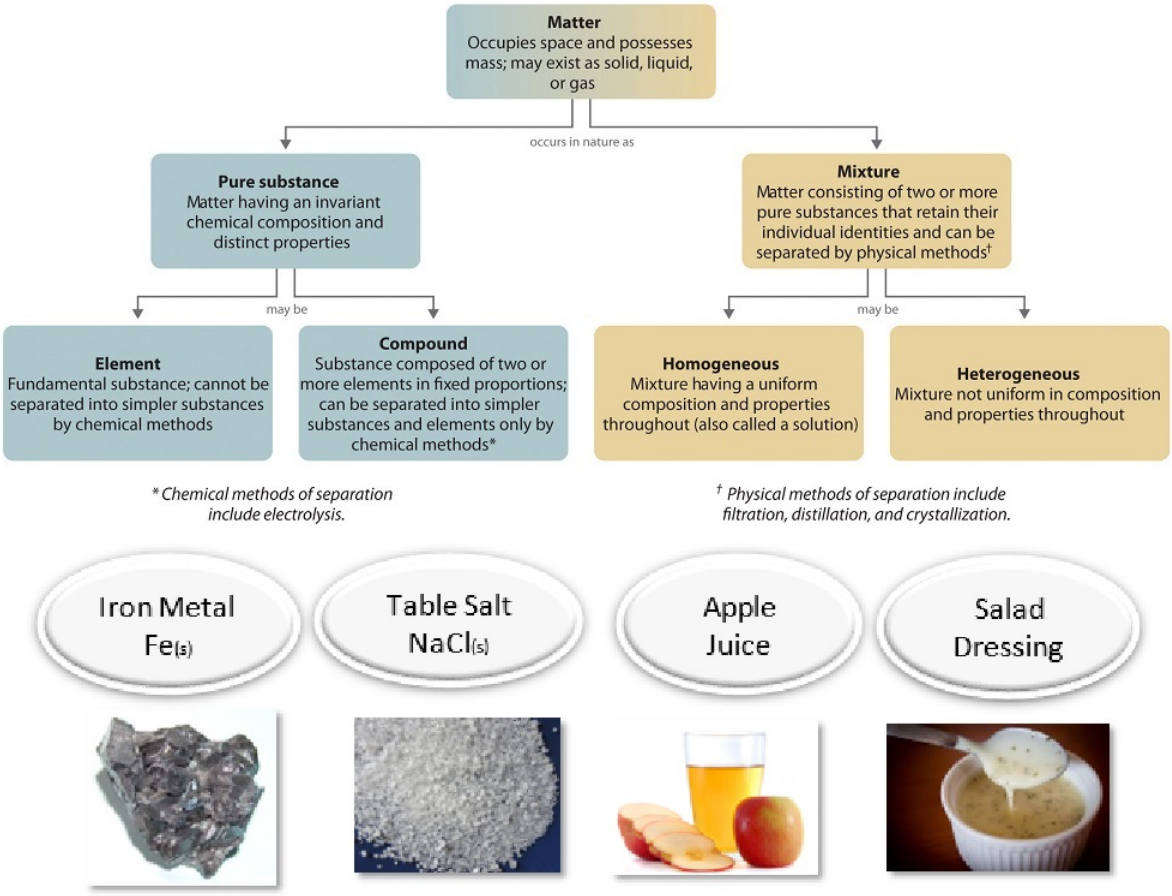


Figure **6.1.1**: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

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

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

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

Phase

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

Example 6.1.1

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

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

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

a. **A)** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
B) Because the composition of the solution is uniform throughout, it is ***a homogeneous mixture***.
b. **A)** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
B) Because its composition is not uniform throughout, orange juice is ***a heterogeneous mixture***.
c. **A)** Aluminum oxide is a single, chemically ***pure compound***.
d. **A)** Selenium is one of the known ***elements***.

Exercise 6.1.1

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

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

Answer a:

homogeneous mixture (solution)

Answer b:

element

Answer c:

heterogeneous mixture

Answer d:

compound

Example 6.1.2

What category does each example of matter belong to?

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

Solution

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

Exercise 6.1.2

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a:

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

Answer b:

element

Answer c:

heterogeneous mixture.

Summary

Matter can be classified into two broad categories: pure substances and mixtures. A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Mixtures are physical combinations of two or more elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

Vocabulary

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

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6.2: Solutions

Learning Objectives

- To understand what causes solutions to form.

A solution is another name for a homogeneous mixture. A *mixture* is a material composed of two or more substances. In a solution, the combination is so intimate that the different substances cannot be differentiated by sight, even with a microscope. Compare, for example, a mixture of salt and pepper and another mixture consisting of salt and water. In the first mixture, we can readily see individual grains of salt and the flecks of pepper. A mixture of salt and pepper is not a solution. However, in the second mixture, no matter how carefully we look, we cannot see two different substances. Salt dissolved in water is a solution.

The major component of a solution, called the **solvent**, is typically the same phase as the solution itself. Each minor component of a solution (and there may be more than one) is called the **solute**. It is the component dissolves in solution. For example, in a solution of salt in water, the solute is salt, and solvent is water.

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table **6.2.1** lists some common types of solutions, with examples of each.

Table **6.2.1**: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	saltwater
solid	gas	H_2 gas absorbed by Pt metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

What causes a solution to form? The simple answer is that the solvent and the solute must have similar intermolecular interactions. When this is the case, the individual particles of solvent and solute can easily mix so intimately that each particle of solute is surrounded by particles of solute, forming a solution. However, if two substances have very different intermolecular interactions, large amounts of energy are required to force their individual particles to mix intimately, so a solution does not form.

Thus two alkanes like heptane, C_7H_{16}, and hexane, C_6H_{14}, are completely miscible in all proportions. The C_7H_{16} and C_6H_{14} molecules are both nonpolar, so they interact with each other via London's Dispersion Forces. They are however insoluble in polar water and hence **hydrophobic** (water-fearing).

For a similar reason, methanol, CH_3OH, is completely miscible with water. In this case both molecules are polar and can form hydrogen bonds among themselves, and with each other. CH_3OH dipoles can align with H_2O dipoles, and CH_3OH molecules can hydrogen bond to H_2O molecules, and so methanol is **hydrophilic** (water-loving).

This process leads to a simple rule of thumb: ***like dissolves like*.** Solvents that are very polar will dissolve solutes that are very polar or even ionic. Solvents that are nonpolar will dissolve nonpolar solutes. Thus water, being polar, is a good solvent for ionic compounds and polar solutes like ethanol (C_2H_5OH). However, water does not dissolve nonpolar solutes, such as many oils, greases, and hydrocarbons.

We use the word **soluble** to describe a solute that dissolves in a particular solvent, and the word **insoluble** for a solute that does not dissolve in a solvent. Thus, we say that sodium chloride is soluble in water but insoluble in hexane (C_6H_{14}). If the solute and the solvent are both liquids and soluble in any proportion, we use the word **miscible**, and the word **immiscible** if they are not.

✔ Example 6.2.1

Water is considered a polar solvent. Which substances should dissolve in water?

1. methanol (CH_3OH)
2. sodium sulfate (Na_2SO_4)
3. octane (C_8H_{18})

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

1. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
2. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
3. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

? Exercise 6.2.1

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Answer

Octane only.

✔ Example 6.2.2

Predict which of the following compounds will be most soluble in water:

a. $\underset{\text{Ethanol}}{CH_3CH_2OH}$
b. $\underset{\text{Hexanol}}{CH_3CH_2CH_2CH_2CH_2CH_2OH}$

Solution

Since ethanol contains an **OH** group, it can hydrogen bond to water. Although the same is true of hexanol, the OH group is found only at one end of a fairly large molecule. The rest of the molecule can be expected to behave much as though it were a nonpolar alkane. This substance should thus be much less soluble than the first. Experimentally we find that ethanol is completely miscible with water, while only 0.6 g hexanol dissolves in 100 g water.

6.2.2Exercise

Would I_2 be more soluble in CCl_4 or H_2O?

Answer

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar and H_2O is polar, so I_2 would be expected to be more soluble in CCl_4.

Electrolytes

A characteristic property of ionic compounds dissolved in water to form a solution is their **electrical conductivity**. The figure below shows three experiments in which two electrodes that are connected to a light bulb are placed in beakers containing three different substances.

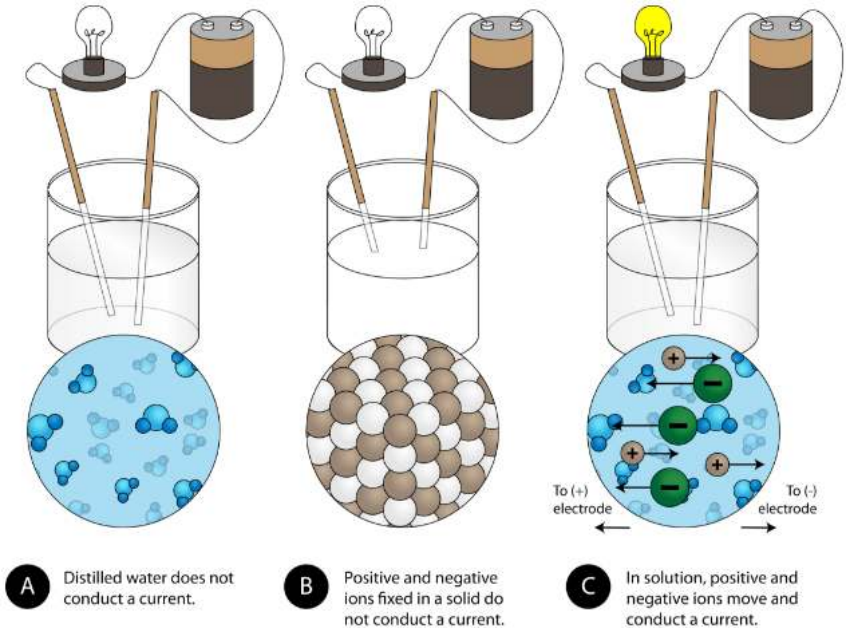


Figure 3.6.3: *(A) Distilled water does not conduct electricity. (B) A solid ionic compound also does not conduct. (C) A water solution of an ionic compound conducts electricity well.*

In the first beaker, distilled water does not conduct a current because water is a molecular compound. In the second beaker, solid sodium chloride also does not conduct a current. Despite being ionic and thus composed of charged particles, the solid crystal lattice does not allow the ions to move between the electrodes. Mobile charged particles are required for the circuit to be complete and the light bulb to light up. In the third beaker, the NaCl has been dissolved into the distilled water. Now the crystal lattice has been broken apart and the individual positive and negative ions can move. Cations move to one electrode, while anions move to the other, allowing electricity to flow. Ionic compounds conduct an electric current when dissolved in water. Such a solution is known as an electrolyte.

Key Takeaway

- Solutions form because a solute and a solvent experience similar intermolecular interactions.

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6.3: The Effect of Temperature on Solubility of Solids, Liquids, and Gases

Learning Objectives

- Describe how temperature affects solubility of different types of solute.

Table salt (NaCl) readily dissolves in water. In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. The solubilities vary widely. NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

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. Shown below Figure **6.3.1** is the **solubility curve**, a graph of the solubility vs. temperature.

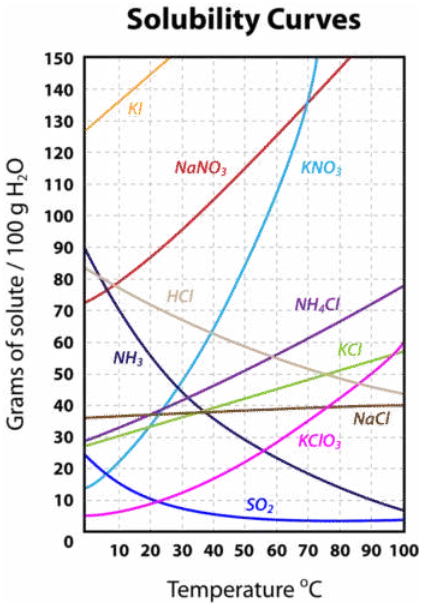


Figure **6.3.1**: A solubility curve is a graph of the solubility of a substance as a function of temperature.

Notice how the temperature dependence of $\mathbf{NaCl}$ is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of $\mathbf{NaCl}$. The curve for $\mathbf{KNO_3}$, on the other hand, is very steep, and so an increase in temperature dramatically increases the solubility of $\mathbf{KNO_3}$. The solubility of liquids also increases as temperature increases.

Several substances such as $\mathbf{HCl}$, $\mathbf{NH_3}$, and $\mathbf{SO_2}$ have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. 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 to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills.

Solubility curves can be used to determine if a given solution is *saturated* or *unsaturated*. Suppose that $\mathbf{80\ g}$ of $\mathbf{KNO_3}$ is added to $\mathbf{100\ g}$ of water at $\mathbf{30°C}$. According to the solubility curve in Figure **6.3.1**, approximately $\mathbf{48\ g}$ of $\mathbf{KNO_3}$ will dissolve at $\mathbf{30°C}$. This means that the solution will be saturated since $\mathbf{48\ g}$ is less than $\mathbf{80\ g}$. We can also determine that there will be $\mathbf{80 - 48 = 32\ g}$ of undissolved $\mathbf{KNO_3}$ remaining at the bottom of the container. In a second scenario, suppose that this saturated solution is heated to $\mathbf{60\,°C}$. According to the curve, the solubility of $\mathbf{KNO_3}$ at $\mathbf{60\,°C}$ is about $\mathbf{107\ g}$. The solution, in this case, is unsaturated since it contains only the original $\mathbf{80\ g}$ of dissolved solute. Suppose in a third case, that the solution is cooled all the way down to $\mathbf{0°C}$. The solubility at $\mathbf{0°C}$ is about $\mathbf{14\ g}$, meaning that $\mathbf{80 - 14 = 66\ g}$ of the $\mathbf{KNO_3}$ will recrystallize.

Supersaturated Solutions

Some solutes, such as sodium acetate, do not recrystallize easily. Suppose an exactly saturated solution of sodium acetate is prepared at 50 °C. As it cools back to room temperature, no crystals 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 in 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.

Precipitation from Supersaturated Solutions

Recrystallization of excess solute from a supersaturated solution usually gives off energy as heat. Commercial heat packs containing supersaturated sodium acetate ($NaC_2H_3O_2$) take advantage of this phenomenon. You can probably find them at your local drugstore.



Video* 6.3.1*: Watered-down sodium acetate trihydrate. Needle crystal is truly wonderful structures

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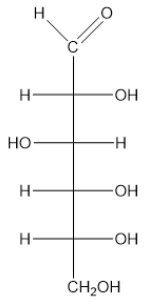
6.4: Solubility Classification for Biochemical Compounds

Learning Objectives

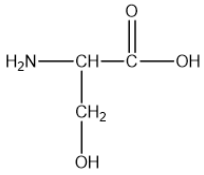
- Explain the terms **hydrophilic**, **hydrophobic**, and **amphipathic** and with examples of compounds that belong to each category.

The biochemical compounds found in living organisms may be placed into one of the three solubility classes. Those soluble in water are **hydrophilic**, those insoluble in water are **hydrophobic**, and those in between are classified as **amphipathic**.

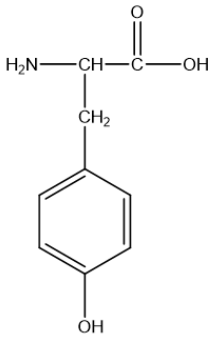
Hydrophilic



D-Glucose



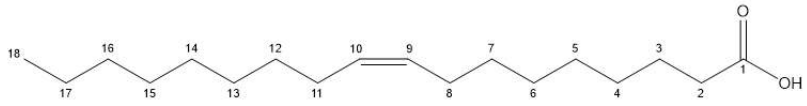
Serine



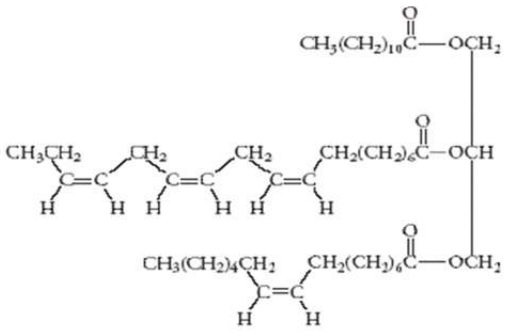
Tyrosine

Glucose and other monosaccharides are classified as hydrophilic. Monosaccharides contain a large number of alcohol functional groups that are able to hydrogen bond to water. The aldehyde functional group in glucose is also polar due to the polar carbonyl and therefore is able to hydrogen bond to water. Amino acids are building blocks of proteins. Two examples of amino acids shown here are serine and tyrosine. They have the functional groups carboxylic acid and 1°-amine that are able to hydrogen bond to water molecules and so they are soluble in water.

Hydrophobic

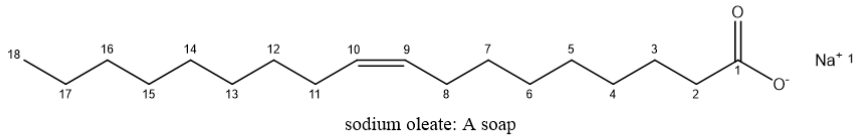


oleic acid: a fatty acid

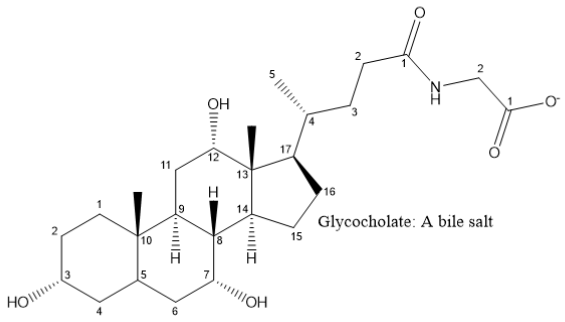


Triglyceride

Oleic acid is a fatty acid with 18 carbon atoms. One end of the molecule has a polar carboxylic acid however the rest of the molecule is a chain of carbon atoms that is nonpolar. Therefore, oleic acid and other fatty acids are hydrophobic. Triglycerides (oils and fats) similarly are hydrophobic due to the large nonpolar areas in the molecule.



sodium oleate: A soap

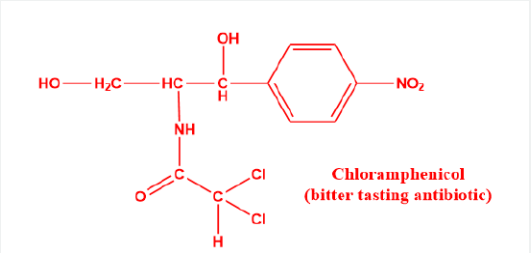


Glycocholate: A bile salt

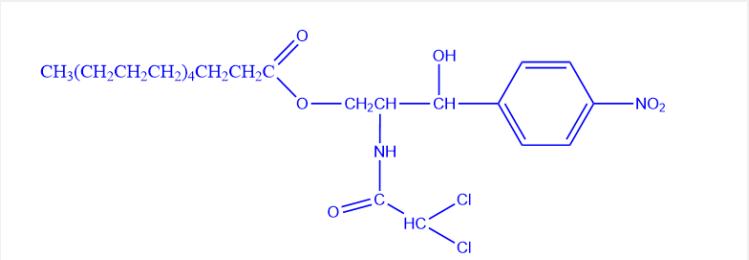
Soaps and bile salts carry a negative charge on the oxygen atom making the polar end very hydrophilic. This is enough to make the molecules amphipathic. Amphipathic compounds have a hydrophilic part and a hydrophobic part.

Prodrugs

A drug is a chemical agent which can affect living processes. Occasionally, a drug is administered as an inactive (a ***pro-drug***) compound. The inactive compound in the body is converted to a fully active drug. In our example of the antibiotic chloramphenicol (in red), a prodrug chloramphenicol palmitate (in blue) is used to overcome problems with solubility of the drug in water.



Chloramphenicol (bitter tasting antibiotic)



Chloramphenicol palmitate (ester) is not bitter in taste so it is administered as a prodrug.

Chloramphenicol has a bitter taste when given orally in solution form. To overcome the bitter taste, it is converted to an ester of palmitic acid to make it water-insoluble and tasteless. The hydrophobic areas of the palmitic acid makes the prodrug-insoluble in saliva and hence almost tasteless. Once in the body, it is undergoes ester hydrolysis to chloramphenicol (drug) and palmitic acid.

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6.5: Solution Concentrations and Dilutions

Learning Outcomes

- Define concentration and use the terms concentrated and dilute to describe the relative concentration of a solution.
- Calculate (w/v) percent concentration.
- Determine concentration in molarity.
- Complete dilution calculations.

Solution Concentration

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.

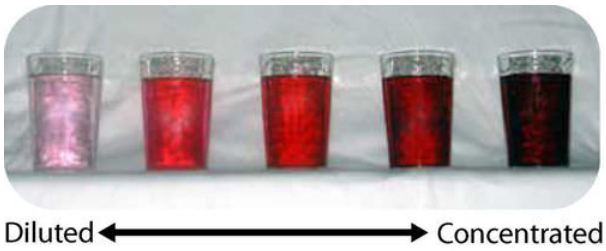


Figure **6.5.1**: Solutions of a red dye in water from the most dilute (on the left) to the most concentrated (on the right).

Weight/Volume Percent

One way to describe the concentration of a solution is by the percent of the solute in a solution. Percent means "parts per hundred". It means the number of parts of solute present every 100 parts of solution. The most commonly used percent measurement is weight/volume percent. The weight/volume percent unit %(w/v), is commonly used for intravenous (IV) fluids. It is calculated by dividing the grams of the solute by the milliliters of the solution and expressing the result as a percent.

For example, if a solution is prepared from $10.g\ \mathrm{NaCl}$ in enough water to make a $150\ \mathrm{mL}$ solution, the weight/volume percent is

$$\text{weight/volume percent} = \frac{\text{grams solute}}{\text{volume solution}} \times 100\% \tag{6.5.1}$$

$$= \frac{10.\ \text{g NaCl}}{150\ \text{mL solution}} \times 100\% \tag{6.5.2}$$

$$= 6.7\%(w/v) \tag{6.5.3}$$

Molarity

A concentration unit based on moles of solute is molarity. 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.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}} \tag{6.5.4}$$

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 6.5.1

A solution is prepared by dissolving $42.23\ \text{g}$ of NH_4Cl into enough water to make $500.0\ \text{mL}$ of solution. Calculate its molarity.

Solution

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

Known

- Mass of $\text{NH}_4\text{Cl} = 42.23\ \text{g}$
- 1 mole of $\text{NH}_4\text{Cl} = 53.50\ \text{g}$
- Volume of solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$

Unknown

- Molarity $=?\ \text{M}$

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.

$$42.23\ \text{g}\ \text{NH}_4\text{Cl} \times \frac{1\ \text{mol}\ \text{NH}_4\text{Cl}}{53.50\ \text{g}\ \text{NH}_4\text{Cl}} = 0.7893\ \text{mol}\ \text{NH}_4\text{Cl} \tag{6.5.5}$$

$$\frac{0.7893\ \text{mol}\ \text{NH}_4\text{Cl}}{0.5000\ \text{L}} = 1.579\ \text{M} \tag{6.5.6}$$

Step 3: Think about your result.

The molarity is $1.579\ \text{M}$, meaning that a liter of the solution would contain 1.579 moles of NH_4Cl. Having four significant figures is appropriate.

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. Dilution is adding more solvent (water) to a solution. Dilution is important in health care because some drugs must be diluted to proper concentration before being administered. The number of moles of the solute does not change, but the total volume of the solution increases.

An useful equation to use when doing dilutions is

$$C_1 \times V_1 = C_2 \times V_2 \tag{6.5.7}$$

The concentration can be in any other unit as long as C_1 and C_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 C_2.

$$C_2 = \frac{C_1 \times V_1}{V_2} = \frac{2.0\ \text{M} \times 100.\ \text{mL}}{500.\ \text{mL}} = 0.40\ \text{M}\ \text{HCl} \tag{6.5.8}$$

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 6.5.2

Nitric acid (HNO_3) 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 $\mathrm{HNO_3}$ $(C_1) = 16\ \mathrm{M}$
- $V_2 = 8.00\ \mathrm{L}$
- $C_2 = 0.50\ \mathrm{M}$

Unknown

- Volume of stock $\mathrm{HNO_3}$ $(V_1) = ?\ \mathrm{L}$

The unknown in the equation is V_1, the necessary volume of the concentrated stock solution.

Step 2: Solve.

$$V_1 = \frac{C_2 \times V_2}{V_1} = \frac{0.50\ \mathrm{M} \times 8.00\ \mathrm{L}}{16\ \mathrm{M}} = 0.25\ \mathrm{L} = 250\ \mathrm{mL} \tag{6.5.9}$$

Step 3: Think about your result.

$250\ \mathrm{mL}$ of the stock $\mathrm{HNO_3}$ solution needs to be diluted with water to a final volume of $8.00\ \mathrm{L}$. The dilution from $16\ \mathrm{M}$ to $0.5\ \mathrm{M}$ is a factor of 32.

Contributors and Attributions

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6.6: Osmosis and Diffusion

Learning Outcomes

- Define osmosis and diffusion.
- Distinguish among hypotonic, hypertonic, and isotonic solutions and predict behavior of blood cells in different solution types.

Fish cells, like all cells, have membranes. A fish that lives in salt water will have somewhat salty water inside itself. When a fish that lives in salt water is put in freshwater, the fish will die, because its cells to swell. What will happen to a freshwater fish in the ocean?

Osmosis

Imagine you have a cup that has $100\ \text{mL}$ water, and you add $15\ \text{g}$ 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 $100\ \text{mL}$ of water, and you add $45\ \text{g}$ 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. A semipermeable membrane has 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 on one side of the membrane and the hypotonic solution is 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 called **hemolysis** and potential (bursting) of the cell. In an isotonic solution, the flow of water in and out of the cell is happening at the same rate.

Two common isotonic solutions used in health care are **5.5 %(w/v) glucose solution and 0.95 %(w/v)**. In isotonic solutions the cells maintain their healthy state and do not undergo crenation or hemolysis.

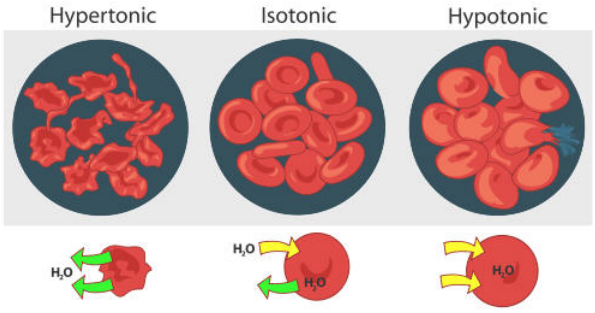


Figure **6.6.1**: Red blood cells in hypertonic, isotonic, and hypotonic solutions.

Osmosis is the diffusion of water molecules across a semipermeable membrane from an area of lower concentration of solute (i.e., higher concentration of water) to an area of higher concentration of solute (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 Cell 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 Cell Membrane

The cell 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 membrane are large molecules like proteins.

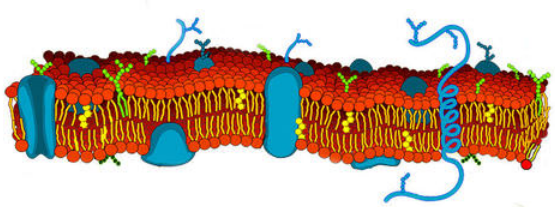


Figure 6.6.2: Plasma membranes are primarily made up of phospholipids (orange). The hydrophilic ("water-loving") head and two hydrophobic ("water-hating") tails are shown. The phospholipids form a bilayer (two layers). The middle of the bilayer is an area without water. There can be water on either side 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

Contributors and Attributions

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6.7: Hemodialysis

Learning Objectives

- Apply the process of diffusion to hemodialysis.

To Your Health: Dialysis

The main function of the kidneys is to filter the blood to remove wastes and extra water, which are then expelled from the body as urine. Some diseases rob the kidneys of their ability to perform this function, causing a buildup of waste materials in the bloodstream. If a kidney transplant is not available or desirable, a procedure called dialysis can be used to remove waste materials and excess water from the blood.

In *hemodialysis*, a patient's blood is passed though a length of tubing that travels through an *artificial kidney machine* (also called a *dialysis machine*). A section of tubing composed of a semipermeable membrane is immersed in a solution of sterile water, glucose, amino acids, and certain electrolytes. This solution called the dialysis solution is an isotonic solution that is formulated to contain many of the same solutes that is in the blood. The concentration of waste molecules such as urea are high in the blood of the patient and so the wastes move via diffusion from the blood to the dialysis solution. The dialysis solution is continually replenished. Red and white blood cells are too large to pass through the membrane, so they remain in the blood. After being cleansed in this way, the blood is returned to the body.



Figure **6.7.3**: A patient undergoing hemodialysis depends on osmosis to cleanse the blood of waste products that the kidneys are incapable of removing due to disease. from Wikipedia.

Dialysis is a continuous process, as the diffusion of waste materials and excess water takes time. Typically, 5–10 lb of waste-containing fluid is removed in each dialysis session, which can last 2–8 hours and must be performed several times a week. Although some patients have been on dialysis for 30 or more years, dialysis is always a temporary solution because waste materials are constantly building up in the bloodstream. A more permanent solution is a kidney transplant.

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6.8: End of Chapter Problems

Classification of Matter

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

a. filtered tea

b. freshly squeezed orange juice

c. aluminum oxide, Al_2O_3

d. selenium

e. white wine

f. mercury

g. ranch-style salad dressing

h. table sugar (sucrose)

2. What category does each example of matter (solid/liquid/gas) belong to?

a. saltwater

b. soil

c. water

d. oxygen

e. coffee grounds

f. hydrogen

g. a boiled egg

h. steam

3. Which of the following substances are elements?

sodium

milk

gold

water

air

liquid nitrogen

paper

electricity

neon

carbon

wood

concrete

Solution and Solubility

1. Water is considered a polar solvent. Which substances should dissolve in water?

a. methanol (CH_3OH)

b. sodium sulfate (Na_2SO_4)

c. octane (C_8H_{18})

2. Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)

b. sodium sulfate (Na_2SO_4)

c. octane (C_8H_{18})

3. Predict which of the following compounds will be most soluble in water:

Ethanol: CH_3CH_2OH or Hexanol: $CH_3CH_2CH_2CH_2CH_2CH_2OH$

Formic acid: HCOOH or Pentanoic acid: $CH_3CH_2CH_2CH_2COOH$

Isopropyl alcohol: $CH_3CH(OH)CH_3$ or Propylene glycol: $CH_3CH(OH)CH_2OH$

4. What happens to solubility of carbon dioxide gas (CO_2) in water when

a. the temperature is increased.

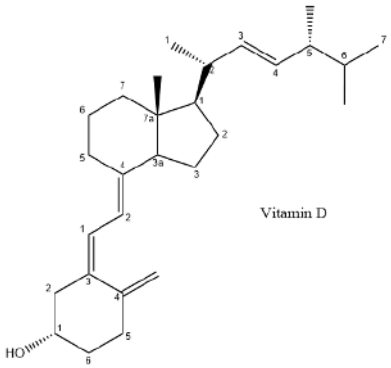
b. the temperature is decreased.

5. Explain why an opened bottle of carbonated water keeps its fizz longer if kept in a refrigerator than if kept at room temperature.

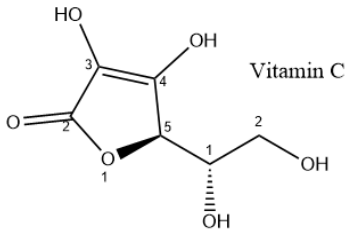
6. On a hot day, it is often easier to catch fish by casting your line into a deep, cool part of a lake than into a shallow, warm spot. One reason that fish gather in cool water may be related to the levels of oxygen dissolved in the water. Explain.

7. One concern about global warming is that increase in water temperature may lead to the death of many fish. How may this be related to the solubility of O_2 in water?

8. Vitamin D is a water-insoluble vitamin, while vitamin C is water-soluble. Why?

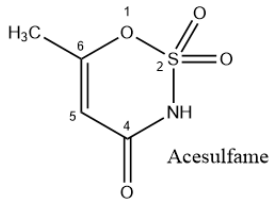


Vitamin D

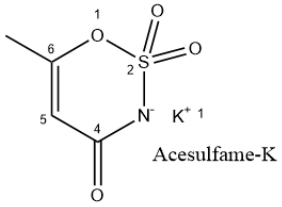


Vitamin C

9. Acesulfame-K (shown below) is the potassium salt of acesulfame. Acesulfame-K is used as a artificial sweetener since it has a greater solubility in water compared to acesulfame. Explain why.



Acesulfame



Acesulfame-K

Solution Concentration

Use the following relationships in the problems where needed:

1 Liter = 1000 mL

1 gram = 1000 mg

Use the periodic table for formula weight and molecular weight of the compounds.

1. Potassium iodide is used to treat iodine deficiencies. What is the %(w/v) of a 75 mL solution containing 2.0 g KI?

2. A solution is prepared by dissolving 42.23 g of NH_4Cl into enough water to make 500.0 mL of solution. Calculate its molarity.

3. Calculate the molarity of each:

    a. 0.33 mol of NaCl in 2.0 L of solution.

    b. 55.0 g of NaCl in 125 mL of solution.

4. Calculate the molarity of each:

    a. 1.75 mol of glucose ($C_6H_{12}O_6$) in 15.2 L of solution.

    b. 270 mg of glucose in 1.00 mL of solution.

5. If 15.0 g of $CaCl_2$ is present in 250 mL of aqueous solution, what is the concentration in terms of the following?

    a. molarity

    b. weight/volume percent

6. If 0.30 mg of sodium fluoride is present in 1.0 L of aqueous solution, what is the concentration in terms of the following?

    a. molarity

    b. weight/volume percent

7. 100 mL of blood serum contains 5.0 mg of thyroxine, a hormone released by the thyroid gland. What is the weight/volume percent concentration?

Dilutions

1. Nitric acid (HNO_3) is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is 16 M. How much of the stock solution of nitric acid needs to be used to make 8.00 L of a 0.50 M solution?

2. You have a 0.50 L of a 0.24 M aqueous glucose solution. You add enough water to give a final volume of 1.0 L. What is the new concentration?

3. You begin with 25 mL of a 1.8 M aqueous LiCl solution and add enough water to a final volume of 350 mL. What is the new concentration?

4. A 5.0% (w/v) solution of ethyl alcohol is diluted from 50.0 mL to 75.0 mL. What is the new weight/volume percent?

5. Calculate the final volume required to prepare 0.020 % (w/v) KI starting with 10. mL of 2.0% (w/v) KI.

Osmosis and Diffusion

1. Define the terms osmosis and diffusion

2. What do the terms isotonic, hypertonic and hypotonic mean?

3. What is crenation and hemolysis? Why do they happen?

4. To make pickles, you soak cucumbers in a concentrated salt solution called brine. Describe how this is related to osmosis.

5. Pickled cucumber tastes salty. Describe how this is related to diffusion.

6. The hypertonic solution made from dissolving Epsom salt ($MgSO_4$) is used as a home remedy that can help reduce the swelling and discomfort of hemorrhoids.

a. How does this hypertonic solution reduce the swelling?'

b. Will pure water have the same effect? Explain.

7. Why is it important for dialyzing solution to be isotonic with blood?

8. Explain the process of hemodialysis.

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CHAPTER OVERVIEW

7: Carbohydrates - An Introduction

7.1: Introduction to Carbohydrates
7.2: Important Pentoses and Hexoses
7.3: Cyclic Structures of Monosaccharides
7.4: Disaccharides
7.5: Polysaccharides of Glucose
7.6: Digestion of Carbohydrates
7.7: Glycolysis
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7.9: Gluconeogenesis
7.10: Ketone Bodies
7.11: End of Chapter Problems

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7.1: Introduction to Carbohydrates

Learning Objectives

- To recognize carbohydrates and classify them as mono-, di-, or polysaccharides.
- To recognize monosaccharides based on functional groups and the number of carbon atoms.

All carbohydrates consist of carbon, hydrogen, and oxygen atoms and are polyhydroxy aldehydes or ketones. Examples of carbohydrates include starch, cellulose (fiber), sucrose (the sweet-tasting compound called sugar). The term *carbohydrate* had its origin in a misinterpretation of the molecular formulas of many of these substances. For example, because its formula is $C_6H_{12}O_6$, glucose was once thought to be a "carbon hydrate" with the structure $C_6{\cdot}6H_2O$.

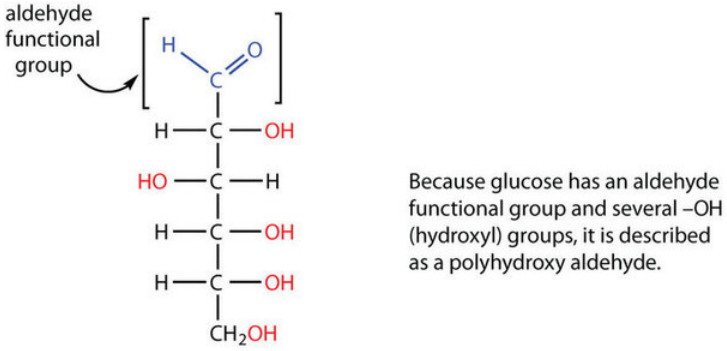


Figure 7.1.1: Glucose

Green plants are capable of synthesizing glucose ($C_6H_{12}O_6$) from carbon dioxide (CO_2) and water (H_2O) by using solar energy in the process known as photosynthesis:

$$6\,CO_2 + 6\,H_2O + 686\ \text{kcal} \rightarrow C_6H_{12}O_6 + 6\,O_2 \tag{7.1.1}$$

(The 686 kcal come from solar energy.) Plants can use the glucose for energy or convert it to larger carbohydrates, such as starch or cellulose. Starch provides energy for later use, perhaps as nourishment for a plant's seeds, while cellulose is the structural material of plants. We can gather and eat the parts of a plant that store energy—seeds, roots, tubers, and fruits—and use some of that energy ourselves. Carbohydrates are also needed for the synthesis of nucleic acids and many proteins and lipids.

Animals, including humans, cannot synthesize carbohydrates from carbon dioxide and water and are therefore dependent on the plant kingdom to provide these vital compounds. We use carbohydrates not only for food (about 60%–65% by mass of the average diet) but also for clothing (cotton), shelter (wood), fuel (wood), and paper (wood).

Classification of Carbohydrates

The simplest carbohydrates—those that cannot be hydrolyzed to produce even smaller carbohydrates—are called monosaccharides. Two or more monosaccharides can link together to form chains that contain from two to several hundred or thousand monosaccharide units. Prefixes are used to indicate the number of such units in the chains. Disaccharide molecules have two monosaccharide units, trisaccharide molecules have three units, and so on. Chains with many monosaccharide units joined together are called polysaccharides. Polysaccharides, disaccharides, and trisaccharides can be hydrolyzed back to their constituent monosaccharides.

Classification of Monosaccharides

The classification of monosaccharides are based on the functional group present and the number of carbon atoms. Monosaccharides such as glucose are called aldoses since they contain an aldehyde functional group. Glucose also contain six carbon atoms and hence it is a hexose. Therefore glucose is an aldohexose since it is an aldehyde monosaccharide with six carbon atoms.

Fructose is a ketose since it contains a ketone functional group. Fructose also contain six carbon atoms and hence it is a hexose. Therefore fructose is an ketohexose since it is an ketone monosaccharide with six carbon atoms.

Ribose and 2-Deoxyribose are monosaccharides found in nucleic acids RNA and DNA. Both monosaccharides are aldopentoses.

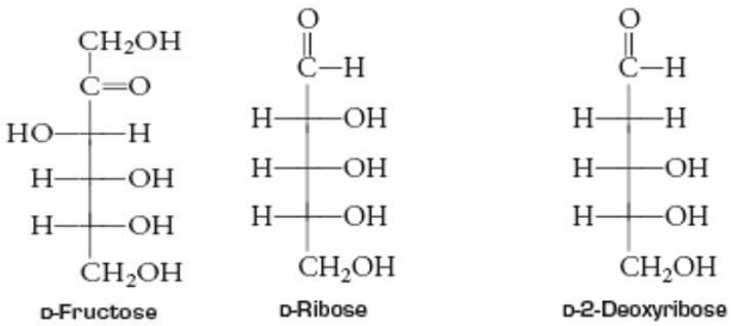
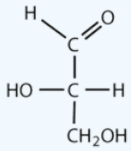
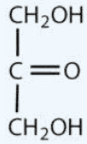


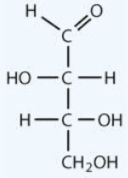
Figure 7.1.2 Monosaccharides are compounds that cannot be hydrolyzed. They will not react with water to form two or more smaller compounds.

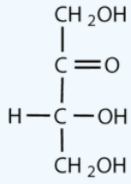
Example 7.1.1

Classify the following compounds based on the functional group and the # of carbon atoms.

a. 

b. 

c. 

d. 

Solution

a. Aldotriose.
b. Ketotriose.
c. Aldotetrose.
d. Ketotetrose.

Summary

Carbohydrates are an important group of biological molecules that includes sugars and starches. Photosynthesis is the process by which plants use energy from sunlight to synthesize carbohydrates. A monosaccharide is the simplest carbohydrate and cannot be hydrolyzed to produce a smaller carbohydrate molecule. Disaccharides contain two monosaccharide units, and polysaccharides contain many monosaccharide units.

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7.2: Important Pentoses and Hexoses

Learning Objectives

- To identify the structures of D-Ribose and D-2-Deoxyribose and describe how they differ from each other.
- To identify the structures of D-glucose, D-galactose, and D-fructose and describe how they differ from each other.

Although a variety of monosaccharides are found in living organisms, two pentoses and three hexoses are particularly abundant.

Pentoses

The two pentoses are D-Ribose and D-2-deoxyribose. Both are aldopentoses.

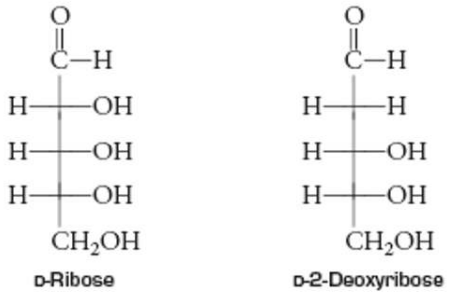


Figure 7.2.1: Structures of Two Important Pentoses.

Ribose and Deoxyribose

Both aldopentoses are incorporated into larger biomolecules. Ribose is in the biochemical oxidizing agent NAD^+ and RNA. Deoxyribose is in DNA. In Ribose all OH groups are on the right. The structural difference between the two monosaccharides is that 2-deoxyribose lacks an alcohol (OH) on carbon atom 2, hence the "2-deoxy" in the name.

Hexoses

The three hexoses are D-glucose, D-galactose, and D-fructose (Figure 7.2.1). Glucose and galactose are both aldohexoses, while fructose is a ketohexose.

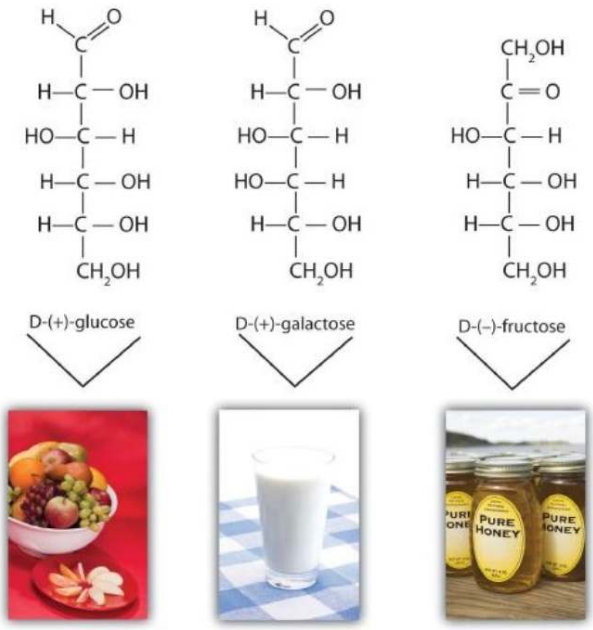


Figure 7.2.2: Structures of Three Important Hexoses. Each hexose is pictured with a food source in which it is commonly found. Source: Photos © Thinkstock.

Glucose

D-Glucose, generally referred to as simply glucose, is the most abundant sugar found in nature is an aldohexose. Most of the carbohydrates we eat are eventually converted to glucose by hydrolysis. In a series of biochemical reactions, glucose is used to produce energy for our cells. It is also known by three other names: *dextrose*, from the fact that it rotates plane-polarized light in a clockwise (dextrorotatory) direction; *corn sugar* because in the United States cornstarch is used in the commercial process that produces glucose from the hydrolysis of starch; and *blood sugar* because it is the carbohydrate found in the circulatory system of animals. Normal blood sugar values range from 70 to 105 mg glucose/dL plasma, and normal urine may contain anywhere from a trace to 20 mg glucose/dL urine.

Galactose

D-Galactose is also an aldohexose. It is released when lactose, a disaccharide found in milk, is hydrolyzed. The galactose needed by the human body for the synthesis of lactose is obtained by the conversion of D-glucose to D-galactose. Galactose is also an important constituent of the glycolipids that occur in the brain and the myelin sheath of nerve cells. For this reason it is also known as *brain sugar*. The structure of D-galactose is shown in Figure 7.2.2. Notice that the configuration differs from that of glucose only at the fourth carbon atom.

Fructose

D-Fructose, shown in Figure 7.2.2, is the most abundant ketohexose. Note that from the third through the sixth carbon atoms, its structure is the same as that of glucose. It occurs, along with glucose and sucrose, in honey (which is 40% fructose) and sweet fruits. Fructose (from the Latin *fructus*, meaning “fruit”) is also referred to as *levulose* because it has a specific rotation that is strongly levorotatory (−92.4°). It is the sweetest sugar, being 1.7 times sweeter than sucrose, although many nonsugars are several hundred or several thousand times as sweet (Table 7.2.1).

Table 7.2.1: The Relative Sweetness of Some Compounds (Sucrose = 100)

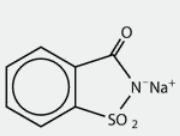
Compound	Relative Sweetness
lactose	16
maltose	32
glucose	74
sucrose	100
fructose	173
aspartame	18,000
acesulfame K	20,000
saccharin	30,000
sucralose	60,000

Looking Closer: Artificial Sweeteners

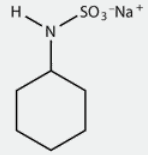
Although sweetness is commonly associated with mono- and disaccharides, it is not a property found only in sugars. Several other kinds of organic compounds have been synthesized that are far superior as sweetening agents. These so-called high-intensity or artificial sweeteners are useful for people with diabetes or other medical conditions that require them to control their carbohydrate intake. The synthetic compounds are noncaloric or used in such small quantities that they do not add significantly to the caloric value of food.

The first artificial sweetener—saccharin—was discovered by accident in 1879. It is 300 times sweeter than sucrose, but it passes through the body unchanged and thus adds no calories to the diet. After its discovery, saccharin was used until it was banned in the early 1900s. However, during the sugar-short years of World War I, the ban was lifted and was not reinstated at the war's end. One drawback to the use of saccharin is its bitter, metallic aftertaste. The initial solution to this problem was to combine saccharin with cyclamate, a second artificial sweetener discovered in 1937.

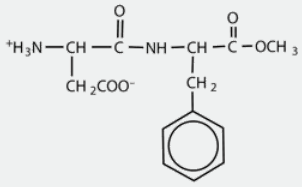
In the 1960s and 1970s, several clinical tests with laboratory animals implicated both cyclamate and saccharin as carcinogenic (cancer-causing) substances. The results from the cyclamate tests were completed first, and cyclamate was banned in the United States in 1969. Then a major study was released in Canada in 1977 indicating that saccharin increased the incidence of bladder cancer in rats. The US Food and Drug Administration (FDA) proposed a ban on saccharin that raised immediate public opposition because saccharin was the only artificial sweetener still available. In response, Congress passed the Saccharin Study and Labeling Act in 1977, permitting the use of saccharin as long as any product containing it was labeled with a consumer warning regarding the possible elevation of the risk of bladder cancer. Today this warning is no longer required; moreover, the FDA is currently reviewing the ban on cyclamate, as 75 additional studies and years of usage in other countries, such as Canada, have failed to show that it has any carcinogenic effect.



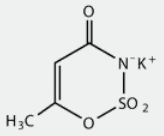
Saccharin



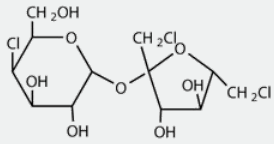
Cyclamate



Aspartame



Acesulfame K



Sucralose

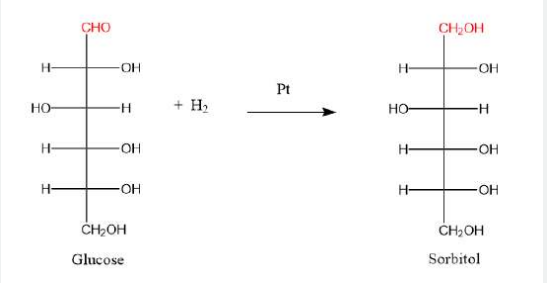
A third artificial sweetener, aspartame, was discovered in 1965. This white crystalline compound is about 180 times sweeter than sucrose and has no aftertaste. It was approved for use in 1981 and is used to sweeten a wide variety of foods because it blends well with other food flavors. Aspartame is not used in baked goods, however, because it is not heat stable.

In the body (or when heated), aspartame is initially hydrolyzed to three molecules: the amino acids aspartic acid and phenylalanine and an alcohol methanol. Repeated controversy regarding the safety of aspartame arises partly from the fact that the body metabolizes the released methanol to formaldehyde. It should be noted, though, that a glass of tomato juice has six times as much methanol as a similar amount of a diet soda containing aspartame. The only documented risk connected to aspartame use is for individuals with the genetic disease *phenylketonuria* (PKU); these individuals lack the enzyme needed to metabolize the phenylalanine released when aspartame is broken down by the body. Because of the danger to people with PKU, all products containing aspartame must carry a warning label.

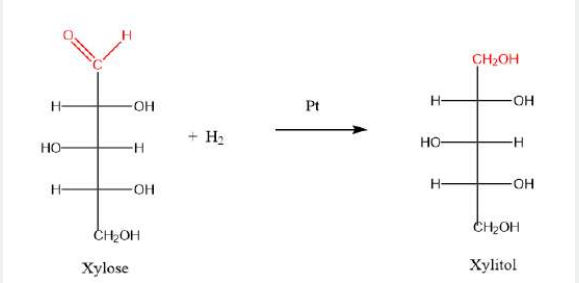
Acesulfame K, discovered just two years after aspartame (1967), was approved for use in the United States in 1988. It is 200 times sweeter than sugar and, unlike aspartame, is heat stable. It has no lingering aftertaste.

One of the newest artificial sweeteners to gain FDA approval (April 1998) for use in the United States is sucralose, a white crystalline solid approximately 600 times sweeter than sucrose. Sucralose is synthesized from sucrose and has three chlorine atoms substituted for three OH groups. It is noncaloric because it passes through the body unchanged. It can be used in baking because it is heat stable.

Alcohol sugars such as sorbitol and xylitol are made by the reduction reaction of glucose and xylose. Both glucose and xylose are monosaccharides.



Glucose Sorbitol



Xylose Xylitol

Sorbitol is used to sweeten food for diabetics. This alcohol sugar is not absorbed well by the intestines and therefore has minimal effects on the blood levels of glucose and insulin. The aldehyde functional group in glucose gains two hydrogen atoms

from H_2 to form sorbitol (a 1°-alcohol).

Xylitol is used for sugarless gums. The aldehyde functional group in xylose gains two hydrogen atoms from H_2 to form xylitol (a 1°-alcohol).

All of the extensive clinical studies completed to date have indicated that these artificial sweeteners approved for use in the United States are safe for consumption by healthy individuals in **moderate** amounts.

Summary

Two pentoses in living organisms are the aldopentoses D-ribose and D-2-deoxyribose.

Three abundant hexoses in living organisms are the aldohexoses D-glucose and D-galactose and the ketohexose D-fructose.

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7.3: Cyclic Structures of Monosaccharides

Learning Objectives

- Describe the reaction between and an alcohol and an aldehyde or ketone to form furanose and pyranose rings in monosaccharides.
- Define what is meant by anomers and describe how they are formed.
- Explain what is meant by mutarotation.

Reaction of Aldehyde and Ketone with Alcohol

So far we have represented monosaccharides as linear molecules, but many of them also adopt cyclic structures. This conversion occurs because of the ability of aldehydes and ketones to react with alcohols:

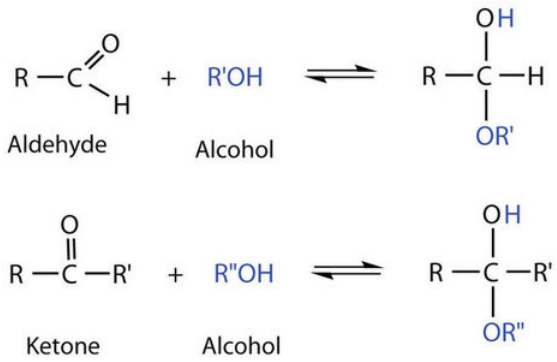


Figure **7.3.1**: Reaction of aldehyde and ketone with alcohol.

The aldehyde reacts with the OH group on the fifth carbon atom of the monosaccharide as in Figure 7.3.2. Cyclic compounds containing five or six carbon atoms in the ring are the most stable. Therefore, monosaccharides rings consisting of five or six carbon atoms are the most common. Rings that contain five carbons are called **furanose** and those that contain six carbons are called **pyranose**.

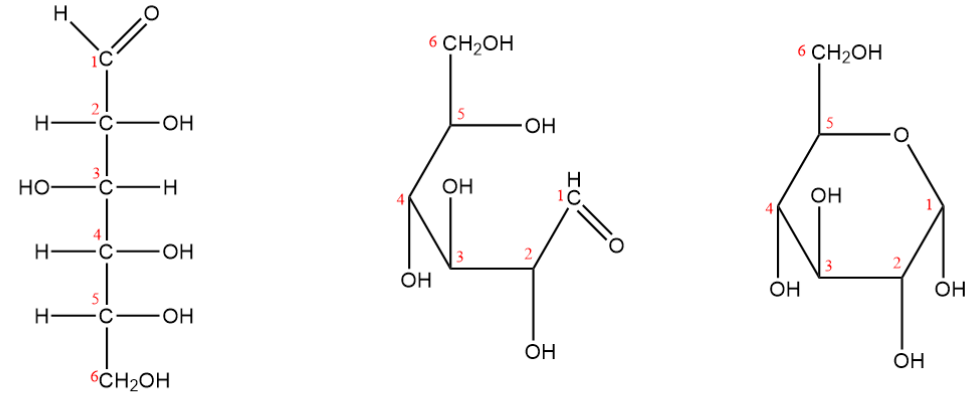


Figure **7.3.2**: Cyclization of Glucose: a. Glucose is represented as straight chain, b. Open form of glucose prior to cyclization, c. By reacting the OH group on the fifth carbon atom with the aldehyde group, the cyclic monosaccharide is produced.

Cyclic Structures of Monosaccharides

When a straight-chain monosaccharide, such as glucose, forms a cyclic pyranose structure, the carbonyl oxygen atom may be pushed either up or down, giving rise to two isomers, as shown in Figure **7.3.3**. The structure shown on the left side of Figure **7.3.3**, with the OH group on the first carbon atom projected downward, represent what is called the *alpha (α) form*. The structures on the right side, with the OH group on the first carbon atom pointed upward, is the *beta (β) form*. These two isomers of a cyclic monosaccharide are known as anomers; they differ in structure around the anomeric carbon—that is, the carbon atom that was the carbonyl carbon atom in the straight-chain form.

It is possible to obtain a sample of crystalline glucose in which all the molecules have the α structure or all have the β structure. The α form melts at 146°C, while the β form melts at 150°C. When the sample is dissolved in water, however, a mixture is soon produced containing both anomers as well as the straight-chain form, in dynamic equilibrium (part (a) of Figure **7.3.3**). You can start with a pure crystalline sample of glucose consisting entirely of either anomer, but as soon as the molecules dissolve in water, they open to form the carbonyl group and then reclose to form either the α or the β anomer. The opening and closing repeats continuously in an ongoing interconversion between anomeric forms and is referred to as mutarotation (Latin *mutare*, meaning "to change"). At equilibrium, the mixture consists of about 36% α-D-glucose, 64% β-D-glucose, and less than 0.02% of the open-chain aldehyde form.

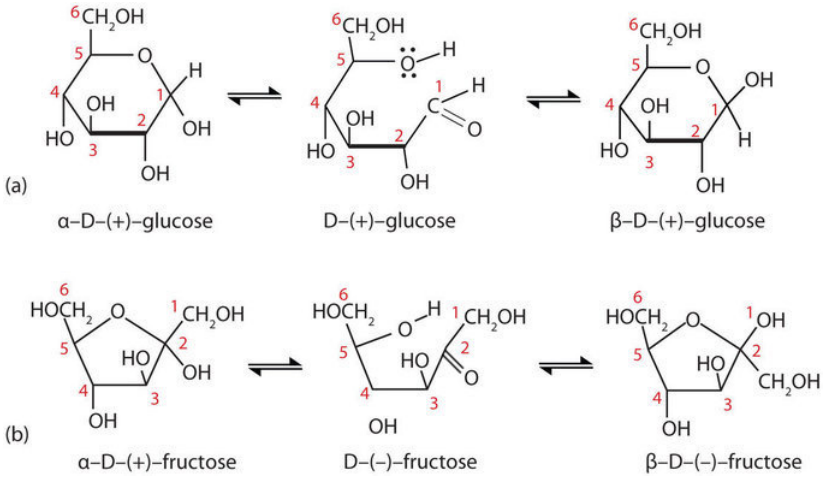


Figure **7.3.3**: Monosaccharides. In an aqueous solution, monosaccharides exist as an equilibrium mixture of three forms. The interconversion between the forms is known as *mutarotation*, which is shown for D-glucose (a) and D-fructose (b).

Even though only a small percentage of the molecules are in the open-chain aldehyde form at any time, the solution will nevertheless exhibit the characteristic reactions of an aldehyde.

Note

The difference between the α and the β forms of sugars may seem trivial, but such structural differences are often crucial in biochemical reactions. This explains why we can get energy from the starch in potatoes and other plants but not from cellulose, even though both starch and cellulose are polysaccharides composed of glucose molecules linked together.

Pentose sugars such as ribose and 2-deoxyribose also exist in furanose cyclic forms shown below:

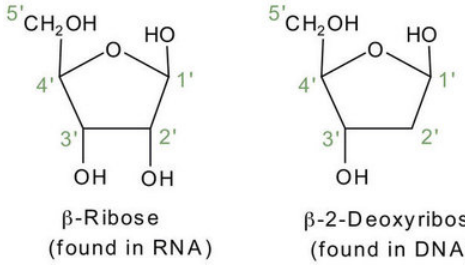


Figure **7.3.4**: Pentose monosaccharides in their cyclic furanose form.

Summary

Monosaccharides that are pentoses or hexoses form cyclic structures in aqueous solution. Two cyclic isomers can form from each straight-chain monosaccharide; these are known as anomers. In an aqueous solution, an equilibrium mixture forms between the two anomers and the straight-chain structure of a monosaccharide in a process known as mutarotation.

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7.4: Disaccharides

Learning Objectives

- Identify the structures of sucrose, lactose, cellobiose, and maltose
- Identify the monosaccharides that are needed to form sucrose, lactose, cellobiose, and maltose

Previously, you learned that monosaccharides can form cyclic structures by the reaction of the carbonyl group with an OH group. These cyclic molecules can in turn react with each other and form disaccharides. So disaccharides ($C_{12}H_{22}O_{11}$) are sugars composed of two monosaccharide units that are joined by a carbon–oxygen-carbon linkage known as a **glycosidic linkage**. This linkage is formed from the reaction of the anomeric carbon of one cyclic monosaccharide with the OH group of a second monosaccharide via dehydration followed by condensation reaction.

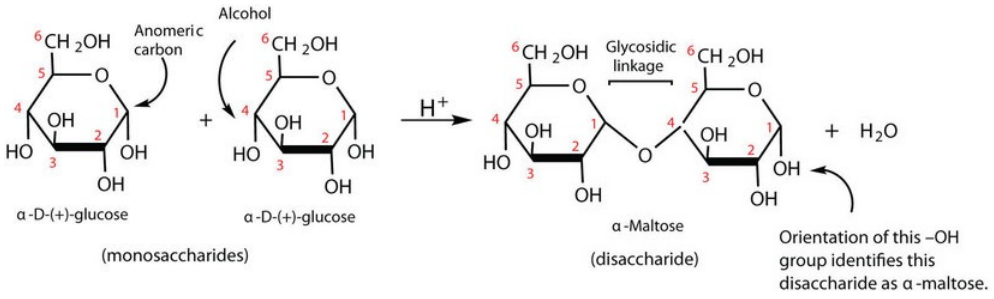


Figure **7.4.1**: Dehydration followed by condensation links two glucose molecules in maltose.

The disaccharides differ from one another in their monosaccharide constituents and in the specific type of glycosidic linkage connecting them. There are four common disaccharides: maltose, cellobiose, lactose, and sucrose. All four are white crystalline solids at room temperature and are soluble in water. We'll consider each disaccharide in more detail.

Maltose

Maltose occurs to a limited extent in sprouting grain. It is formed most often by the partial hydrolysis of polysaccharides such as starch and glycogen. In the manufacture of beer, maltose is liberated by the action of malt (germinating barley) on starch; for this reason, it is often referred to as *malt sugar*. Maltose is about 30% as sweet as sucrose. The human body is unable to metabolize maltose or any other disaccharide directly from the diet because the molecules are too large to pass through the cell membranes of the intestinal wall. Therefore, an ingested disaccharide must first be broken down by hydrolysis into its two constituent monosaccharide units.

In the body, such hydrolysis reactions are catalyzed by an enzyme *maltase*. The same reactions can be carried out in the laboratory with dilute acid as a catalyst, although in that case the rate is much slower, and high temperatures are required. Whether it occurs in the body or a glass beaker, the hydrolysis of maltose produces two molecules of D-glucose.

$$\text{maltose} \xrightarrow{\text{H}^+ \text{ or maltase}} 2\ \text{D-glucose}$$

In maltose the two glucose molecules is linked in such a way as to leave one anomeric carbon that can open to form an aldehyde group. The glucose units in maltose are joined in a *head-to-tail* fashion through an α-linkage from the first carbon atom of one glucose molecule to the fourth carbon atom of the second glucose molecule that is, an α-(1→4)-glycosidic linkage; see Figure **7.4.2**. The bond from the anomeric carbon of the first monosaccharide unit is directed downward, which is why this is known as an α-glycosidic linkage. The OH group on the anomeric carbon of the second glucose can be in either the α or the β position, as shown in Figure **7.4.2**.

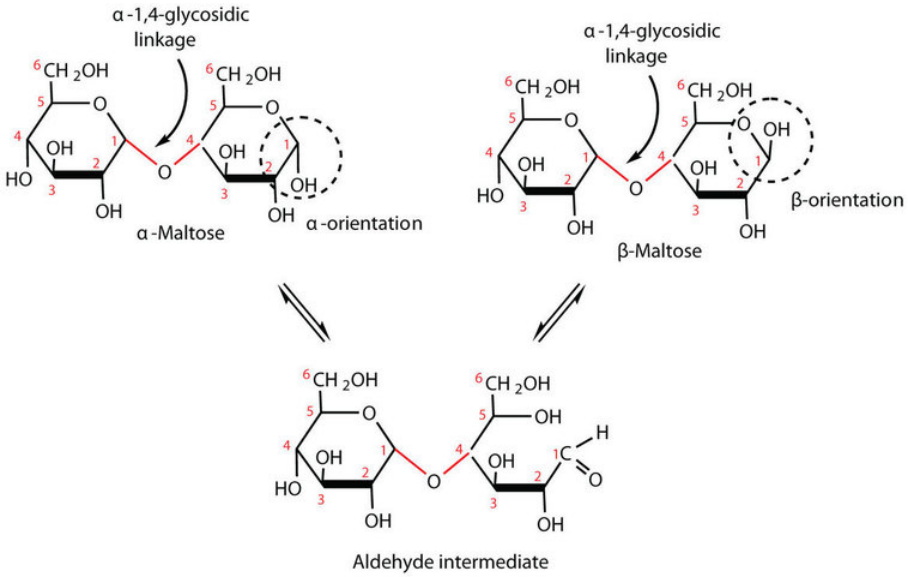


Figure **7.4.2**: An Equilibrium Mixture of Maltose Isomers

Cellobiose

In cellobiose the two glucose molecules is linked in such a way as to leave one anomeric carbon that can open to form an aldehyde group. The glucose units in cellobiose are joined in a *head-to-tail* fashion through an β-linkage from the first carbon atom of one glucose molecule to the fourth carbon atom of the second glucose molecule that is, an β-(1→4)-glycosidic linkage. The bond from the anomeric carbon of the first monosaccharide unit is directed upward, which is why this is known as an β-glycosidic linkage. The OH group on the anomeric carbon of the second glucose can be in either the α or the β position.

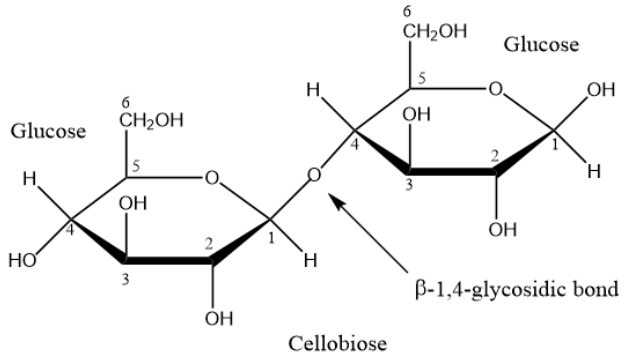


Figure **7.4.3**: In cellobiose, two glucose molecules are connected by β-(1→4)-glycosidic linkage.

Cellobiose is formed by the partial hydrolysis of the polysaccharide cellulose. The hydrolysis of cellobiose also produces two molecules of D-glucose. however humans do not produce the enzyme necessary to catalyse this hydrolysis reactions and so we are unable to use cellobiose as a source of food and energy.

Lactose

Lactose is known as *milk sugar* because it occurs in the milk of humans, cows, and other mammals. In fact, the natural synthesis of lactose occurs only in mammary tissue, whereas most other carbohydrates are plant products. Human milk contains about 7.5% lactose, and cow's milk contains about 4.5%. This sugar is one of the lowest ranking in terms of sweetness, being about one-sixth as sweet as sucrose. Lactose is produced commercially from whey, a by-product in the manufacture of cheese. It is important as an infant food and in the production of penicillin.

Lactose is a disaccharide composed of one molecule of D-galactose and one molecule of D-glucose joined by a β-(1→4)-glycosidic bond (the bond from the anomeric carbon of the first monosaccharide unit being directed upward). The two

monosaccharides are obtained from lactose by acid hydrolysis or the catalytic action of the enzyme *lactase*:

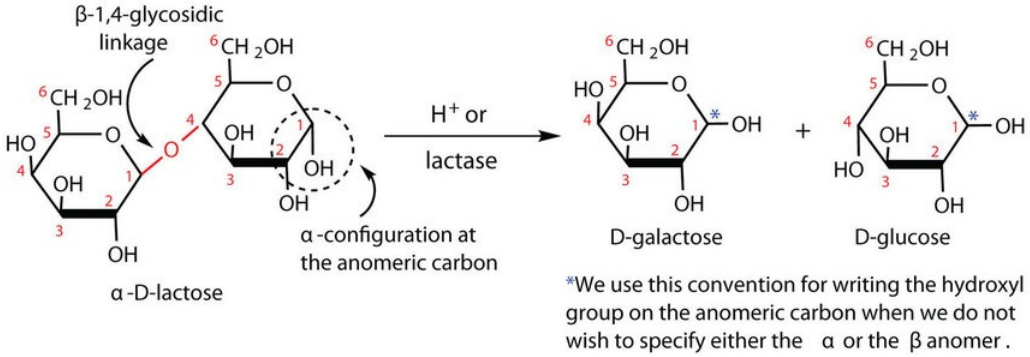


Figure **7.4.4**: Lactose is made of a galactose and a glucose connected by β-(1→4)-glycosidic bond.

Many adults and some children suffer from a deficiency of lactase. These individuals are said to be lactose intolerant because they cannot digest the lactose found in milk. A more serious problem is the genetic disease galactosemia, which results from the absence of an enzyme needed to convert galactose to glucose. Certain bacteria can metabolize lactose, forming lactic acid as one of the products. This reaction is responsible for the "souring" of milk.

Sucrose

Sucrose, probably the largest-selling pure organic compound in the world, is known as *beet sugar*, *cane sugar*, *table sugar*, or simply *sugar*. Most of the sucrose sold commercially is obtained from sugar cane and sugar beets (whose juices are 14%–20% sucrose) by evaporation of the water and recrystallization. The dark brown liquid that remains after the recrystallization of sugar is sold as molasses.

The sucrose molecule is unique among the common disaccharides in having an α,β-(1↔2)-glycosidic (head-to-head) linkage. Because this glycosidic linkage is formed by the OH group on the anomeric carbon of α-D-glucose and the OH group on the anomeric carbon of β-D-fructose, it ties up the anomeric carbons of both glucose and fructose.

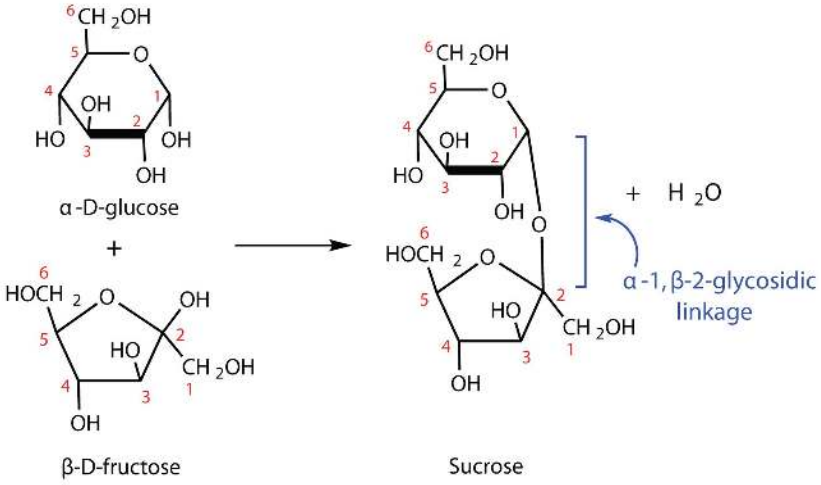


Figure **7.4.5**: Sucrose is made of a fructose and a glucose unit connected by α,β-(1↔2)-glycosidic (head-to-head) linkage.

This linkage gives sucrose certain properties that are quite different from those of maltose and lactose. As long as the sucrose molecule remains intact, neither monosaccharide "uncyclizes" to form an open-chain structure. Thus, sucrose is incapable of mutarotation and exists in only one form both in the solid state and in solution. In addition, sucrose does not undergo reactions that are typical of aldehydes and ketones.

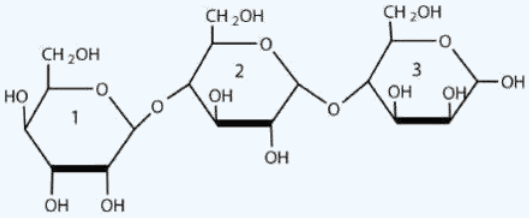
The hydrolysis of sucrose in dilute acid or through the action of the enzyme *sucrase* gives an equimolar mixture of glucose and fructose. This 1:1 mixture is referred to as *invert sugar*. The hydrolysis reaction has several practical applications. Sucrose readily

recrystallizes from a solution, but invert sugar has a much greater tendency to remain in solution. In the manufacture of jelly and candy and in the canning of fruit, the recrystallization of sugar is undesirable. Therefore, conditions leading to the hydrolysis of sucrose are employed in these processes. Moreover, because fructose is sweeter than sucrose, the hydrolysis adds to the sweetening effect. Bees carry out this reaction when they make honey.

The average American consumes more than 100 lb of sucrose every year. About two-thirds of this amount is ingested in soft drinks, pre-sweetened cereals, and other highly processed foods. The widespread use of sucrose is a contributing factor to obesity and tooth decay. Carbohydrates such as sucrose, are converted to fat when the caloric intake exceeds the body's requirements, and sucrose causes tooth decay by promoting the formation of plaque that sticks to teeth.

✔ Example 7.4.1

For this trisaccharide, identify the two glycosidic linkages.



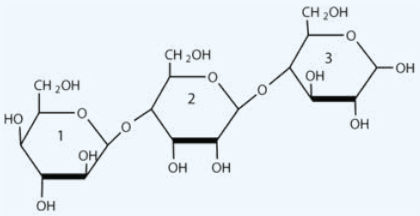
Solution

a. The glycosidic linkage between sugars 1 and 2 is β-(1→4) because the bond is directed up from the anomeric carbon.

b. The glycosidic linkage between sugars 2 and 3 is α-(1→4) because the bond is directed down from the anomeric carbon.

? Exercise 7.4.1

For this trisaccharide, identify the two glycosidic linkages.



To Your Health: Lactose Intolerance and Galactosemia

Lactose makes up about 40% of an infant's diet during the first year of life. **Lactose intolerance** is caused by the deficiency of the enzyme lactase, the enzyme that catalyzes the hydrolysis of the β-(1→4) glycosidic bond in lactose. Infants and small children have one form of the enzyme lactase in their small intestines and can digest the sugar easily; however, adults usually have a less active form of the enzyme, and about 70% of the world's adult population has some deficiency in its production. As a result, many adults experience a reduction in the ability to hydrolyze lactose to galactose and glucose in their small intestine. For some people the inability to synthesize sufficient enzyme increases with age. Up to 20% of the US population suffers some degree of lactose intolerance.

In people with lactose intolerance, some of the unhydrolyzed lactose passes into the colon, where it tends to draw water from the interstitial fluid into the intestinal lumen by osmosis. At the same time, intestinal bacteria may act on the lactose to produce organic acids and gases such as CO_2, H_2, and CH_4. The buildup of water and bacterial decay products leads to abdominal distention, bloating, cramps, and diarrhea, which are symptoms of the condition.

The symptoms disappear if milk or other sources of lactose are excluded from the diet or consumed only sparingly. Alternatively, many food stores now carry special brands of milk that have been pretreated with lactase to hydrolyze the lactose. Cooking or fermenting milk causes at least partial hydrolysis of the lactose, so some people with lactose intolerance

are still able to enjoy cheese, yogurt, or cooked foods containing milk. The most common treatment for lactose intolerance, however, is the use of lactase preparations (e.g., Lactaid), which are available in liquid and tablet form at drugstores and grocery stores. These are taken orally with dairy foods—or may be added to them directly—to assist in their digestion.

This same type of digestion problems is associated with some larger oligosaccharides found in peas and beans. These oligosaccharides contain galactose involve in α-(1→6) glycosidic bond. Humans do not produce the enzyme necessary to hydrolyze such bonds. So when foods containing these carbohydrates with galactose are consumed, intestinal bacteria break them down to produce gases. The enzymes needed to convert galactose to glucose is also missing. The condition is known as **Galactosemia** in which the blood galactose level is markedly elevated, and galactose is found in the urine. A person with galactosemia experiences a lack of appetite, weight loss, diarrhea, and jaundice. The disease may result in impaired liver function, cataracts, mental retardation, and even death. If galactosemia is recognized in early infancy, its effects can be prevented by the exclusion of milk and all other sources of galactose from the diet. As a child with galactosemia grows older, he or she usually develops an alternate pathway for metabolizing galactose, so the need to restrict milk is not permanent. The incidence of galactosemia in the United States is 1 in every 65,000 newborn babies.

Summary

Maltose is composed of two molecules of glucose joined by an α-(1→4)-glycosidic linkage. It is a disaccharide that is found in sprouting grain. Lactose is composed of a molecule of galactose joined to a molecule of glucose by a β-(1→4)-glycosidic linkage. It is a sugar that is found in milk. Sucrose is composed of a molecule of glucose joined to a molecule of fructose by an α,β-(1↔2) glycosidic linkage. It is found in sugar cane and sugar beets. Cellobiose derived from the hydrolysis of the polysaccharide cellulose, is composed of two molecules of glucose joined by a β-(1→4) glycosidic linkage. Cellobiose cannot be used as a source of food for humans.

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7.5: Polysaccharides of Glucose

Learning Objectives

- To compare and contrast the structures and properties of starch, glycogen, and cellulose.

The polysaccharides are the most abundant carbohydrates in nature and serve a variety of functions, such as energy storage or as components of plant cell walls. Polysaccharides are very large polymers composed of tens to thousands of monosaccharides joined together by glycosidic linkages. The three most abundant polysaccharides are starch, glycogen, and cellulose. These three are referred to as *homopolysaccharides* because each yields only one type of monosaccharide (glucose) after complete hydrolysis. *Heteropolysaccharides* may contain sugar acids, amino sugars, or noncarbohydrate substances in addition to monosaccharides. Heteropolysaccharides are common in nature (gums, pectins, and other substances) but will not be discussed further in this textbook. The polysaccharides are not sweet tasting, and do not undergo mutarotation.

Starch

Starch is the most important source of carbohydrates in the human diet and accounts for more than 50% of our carbohydrate intake. It is found in plants and are abundant in seeds (especially the cereal grains) and tubers, where they serve as a storage form of carbohydrates. The breakdown of starch to glucose nourishes the plant during periods of reduced photosynthetic activity. We often think of potatoes as a "starchy" food, yet other plants contain a much greater percentage of starch (potatoes 15%, wheat 55%, corn 65%, and rice 75%). Commercial starch is a white powder.

Starch is a mixture of two components: **amylose** and **amylopectin**. Natural starches consist of about 10%–30% amylose and 70%–90% amylopectin.

Amylose

Amylose is a polysaccharide (part (a) of Figure 7.5.1) composed entirely of glucose units joined by the α-(1⟶4)-glycosidic linkages we saw in maltose. Experimental evidence indicates that amylose is coiled like a spring, with six glucose monomers per turn (part (b) of Figure 7.5.1). When coiled in this fashion, amylose has just enough room in its core to accommodate an iodine molecule. The characteristic blue-violet color that appears when starch is treated with iodine is due to the formation of the amylose-iodine complex. This color test is sensitive enough to detect even minute amounts of starch in solution.

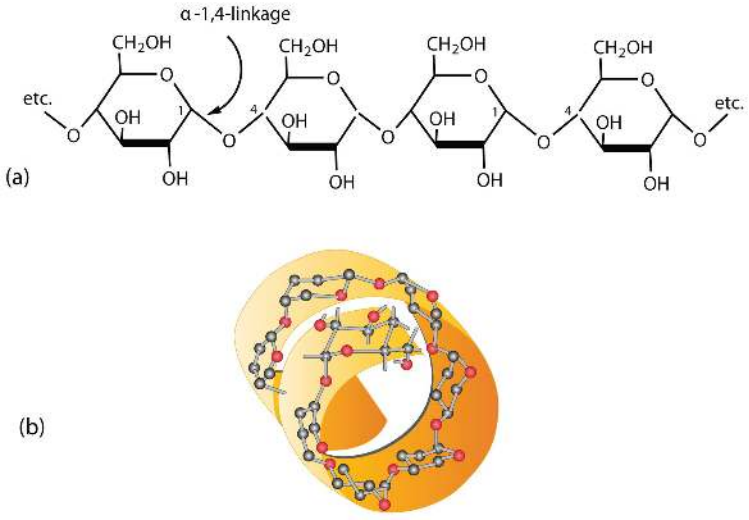


Figure 7.5.1: Amylose. (a) Amylose is a linear chain of α-D-glucose units joined together by α-1,4-glycosidic bonds. (b) Because of hydrogen bonding, amylose acquires a spiral structure that contains six glucose units per turn.

Amylopectin

Amylopectin is a branched-chain polysaccharide composed of glucose units linked primarily by α-(1⟶4)-glycosidic bonds but with occasional α-(1⟶6)-glycosidic bonds, which are responsible for the branching. A molecule of amylopectin may contain many

thousands of glucose units with branch points occurring about every 25–30 units (Figure **7.5.2**). The helical structure of amylopectin is disrupted by the branching of the chain, so instead of the deep blue-violet color amylose gives with iodine, amylopectin produces a less intense reddish brown.

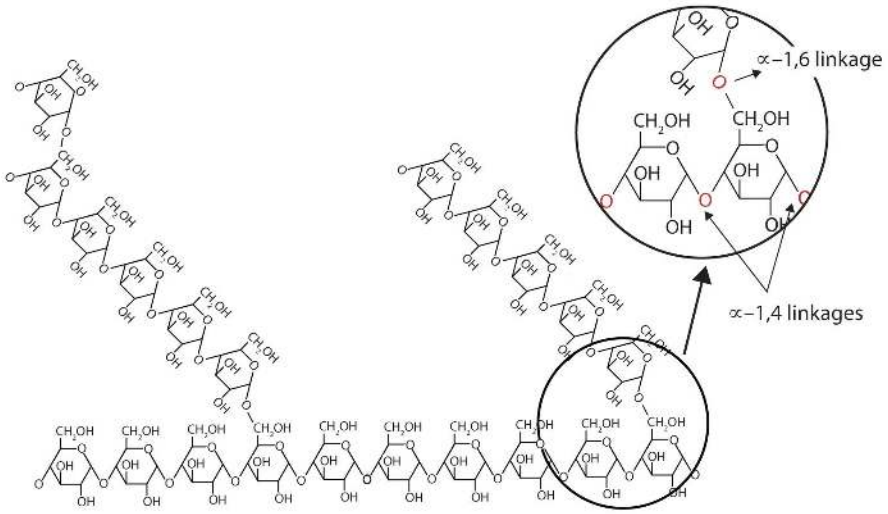


Figure **7.5.2**: Representation of the Branching in Amylopectin. These branch points occur in the polysaccharide glycogen too.

Dextrins are glucose polysaccharides of intermediate size. The shine and stiffness imparted to clothing by starch are due to the presence of dextrins formed when clothing is ironed. Because of their characteristic stickiness with wetting, dextrins are used as adhesives on stamps, envelopes, and labels; as binders to hold pills and tablets together; and as pastes. Dextrins are more easily digested than starch and are therefore used extensively in the commercial preparation of infant foods.

The complete hydrolysis of starch yields, in successive stages, glucose:

$$\text{starch} \rightarrow \text{dextrins} \rightarrow \text{maltose} \rightarrow \text{glucose}$$

In the human body, several enzymes known collectively as amylases hydrolyze the α-(1⟶4)-glycosidic bonds in starch sequentially into usable glucose units. The hydrolysis of α-(1⟶6)-glycosidic bonds in amylopectin is catalyzed by debranching enzymes.

Glycogen

Glycogen is the energy reserve carbohydrate of animals. Practically all mammalian cells contain some stored carbohydrates in the form of glycogen, but it is especially abundant in the liver (4%–8% by weight of tissue) and in skeletal muscle cells (0.5%–1.0%). Like starch in plants, glycogen is found as granules in liver and muscle cells. When fasting, animals draw on these glycogen reserves during the first day without food to obtain the glucose needed to maintain metabolic balance.

Glycogen is structurally quite similar to amylopectin, although glycogen is more highly branched (8–12 glucose units between branches) and the branches are shorter. When treated with iodine, glycogen gives a reddish brown color. Glycogen can be broken down into its D-glucose subunits by the same enzymes that catalyze the breakdown of starch. In animals, the enzyme phosphorylase catalyzes the breakdown of glycogen to phosphate esters of glucose.

> *About 70% of the total glycogen in the body is stored in muscle cells. Although the percentage of glycogen (by weight) is higher in the liver, the much greater mass of skeletal muscle stores a greater total amount of glycogen.*

Cellulose

Cellulose, a fibrous carbohydrate found in all plants, is the structural component of plant cell walls. Because the earth is covered with vegetation, cellulose is the most abundant of all carbohydrates, accounting for over 50% of all the carbon found in the

vegetable kingdom. Cotton fibrils and filter paper are almost entirely cellulose (about 95%), wood is about 50% cellulose, and the dry weight of leaves is about 10%–20% cellulose. The largest use of cellulose is in the manufacture of paper and paper products. Although the use of noncellulose synthetic fibers is increasing, rayon (made from cellulose) and cotton still account for over 70% of textile production.

Like amylose, cellulose is a polymer of glucose connected by 1 to 4 glycosidic bonds. It differs, however, in that the glucose units are joined by β-(1⟶4)-glycosidic linkages, producing a **sheet** structure (part (a) of Figure **7.5.3**). This extreme linearity allows a great deal of hydrogen bonding between OH groups on adjacent chains, causing them to pack closely into fibers (part (b) of Figure **7.5.3**). As a result, cellulose exhibits little interaction with water or any other solvent. Cotton and wood, for example, are completely insoluble in water and have considerable mechanical strength. Because cellulose does not have a helical structure, it does not bind to iodine to form a colored product.

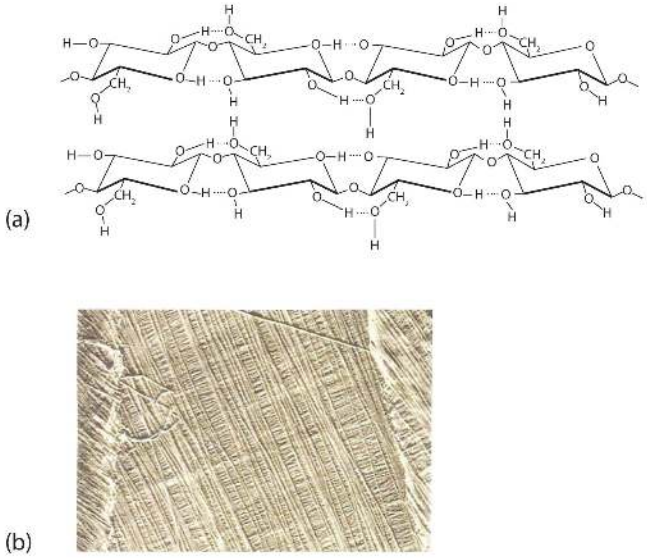


Figure **7.5.3**: Cellulose. (a) There is extensive hydrogen bonding in the structure of cellulose. (b) In this electron micrograph of the cell wall of an algae, the wall consists of successive layers of cellulose fibers in parallel arrangement.

Cellulose yields glucose after complete acid hydrolysis, yet humans are unable to metabolize cellulose as a source of glucose. Our digestive juices lack cellulase enzyme that can hydrolyze the β-glycosidic linkages found in cellulose, so although we can eat potatoes, we cannot eat grass. However, certain microorganisms can digest cellulose because they make the enzyme cellulase, which catalyzes the hydrolysis of cellulose. The presence of these microorganisms in the digestive tracts of herbivorous animals (such as cows, horses, and sheep) allows these animals to degrade the cellulose from plant material into glucose for energy. Termites also contain cellulase-secreting microorganisms and thus can subsist on a wood diet.

Career Focus: Certified Diabetes Educator

Certified diabetes educators come from a variety of health professions, such as nursing and dietetics, and specialize in the education and treatment of patients with diabetes. A diabetes educator will work with patients to manage their diabetes. This involves teaching the patient to monitor blood sugar levels, make good food choices, develop and maintain an exercise program, and take medication, if required.



A certified diabetes educator at Naval Medical Center Portsmouth (left) and a registered dietician at the medical center (center), provide nutritional information to a diabetes patient and her mother at the Diabetes Boot Camp.

Diabetes educators also work with hospital or nursing home staff to improve the care of diabetic patients. Educators must be willing to spend time attending meetings and reading the current literature to maintain their knowledge of diabetes medications, nutrition, and blood monitoring devices so that they can pass this information to their patients.

Summary

Starch is a storage form of energy in plants. It contains two polymers composed of glucose units: amylose (linear) and amylopectin (branched). Glycogen is a storage form of energy in animals. It is a branched polymer composed of glucose units. It is more highly branched than amylopectin. Cellulose is a structural polymer of glucose units found in plants.

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7.6: Digestion of Carbohydrates

Learning Objectives

- Describe digestion of carbohydrates.

Introduction to Metabolism

Metabolism is all reactions that take place in a living thing. It is divided into two parts, **catabolism** and **anabolism**. Catabolism involves compounds being broken down into smaller ones. This results in release of energy (exothermic). Anabolism involves synthesis of larger compounds from smaller ones. This requires energy (endothermic).

ATP is a key player in metabolism. Energy released during catabolism is used to drive the formation of ATP. ATP can be hydrolyzed to ADP and a phosphate ion. The energy released during the hydrolysis of ATP is used for anabolism. Digestion of food is an example of catabolism. Starch is hydrolyzed with water and the aid of enzymes to split into its building blocks glucose.

Digestion of Carbohydrates

The first stage of catabolism is **digestion**, in which food molecules are broken down into smaller molecules. Carbohydrate digestion begins in the mouth (Figure **7.6.1**) where salivary α-amylase attacks the α-glycosidic linkages in starch, the main carbohydrate ingested by humans. Hydrolysis of the glycosidic linkages with water produces a mixture of dextrins, maltose, and glucose. The α-amylase mixed into the food remains active as the food passes through the esophagus, but it is rapidly inactivated in the acidic environment (low pH) of the stomach.

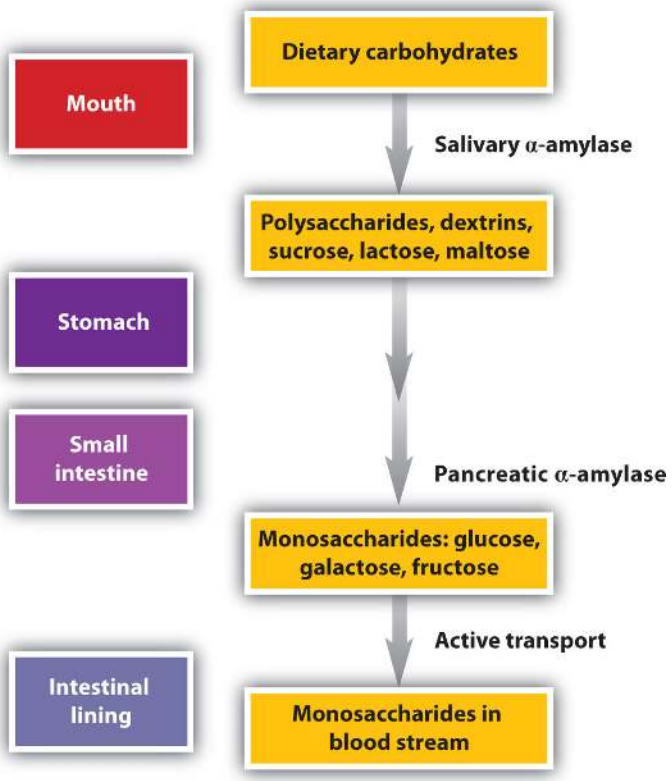


Figure **7.6.1**: The Principal Events and Sites of Carbohydrate Digestion

The primary site of carbohydrate digestion is the small intestine. Starch is digested to glucose by the action of amylase [hydrolyzes the α-(1⟶4) glycosidic bond] and debranching enzymes [hydrolyzes the α-(1⟶6) glycosidic bond]. Dextrins too are converted to maltose. Maltose is then hydrolyzed into two glucose molecules with the help of the enzyme maltase. Disaccharides such as sucrose and lactose are digested/hydrolyzed in the small intestine, where they are acted on by the enzymes sucrase and lactase,

respectively. The major products of the complete hydrolysis of disaccharides and polysaccharides are three monosaccharide units: glucose, fructose, and galactose. These are absorbed through the wall of the small intestine into the bloodstream.

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7.7: Glycolysis

Learning Objectives

- Explain the purpose and products of glycolysis.
- Describe the use and formation of ATP during glycolysis.

Introduction to Glycolysis

Glycolysis is a 10-step linear catabolic process in which glucose is converted into pyruvate via ten steps that are enzyme catalyzed.. The glucose molecules come from largely from the digestion of starch. The overall reaction is shown below.

$$\text{Glucose} + 2\,\text{NAD}^+ + 2\,\text{ADP} + 2\,\text{Pi} \longrightarrow 2\,\text{pyruvate} + 2\,\text{NADH} + 2\,\text{ATP}$$

Glycolysis is the only metabolic cycle that we will discuss that does not occur in the mitochondria; it occurs in the cytoplasm of the cell. This is an old cycle in terms of evolution and is common to animals, plants, fungi, and bacteria. Glycolysis can be divided into 2 stages: the 6-carbon stage and the 3-carbon stage.

There are also two phases of glycolysis:

- the "preparatory phase" because it requires an input of energy in the form of 2 ATP's and
- the "pay off phase" because energy is released in the form of 4 ATP's.

The end result of glycolysis is two pyruvate molecules which can be converted to acetyl-CoA under aerobic conditions which then feeds into the citric acid cycle (also known as the Kreb's cycle). The structures of glycolysis intermediates can be found in the following diagram:

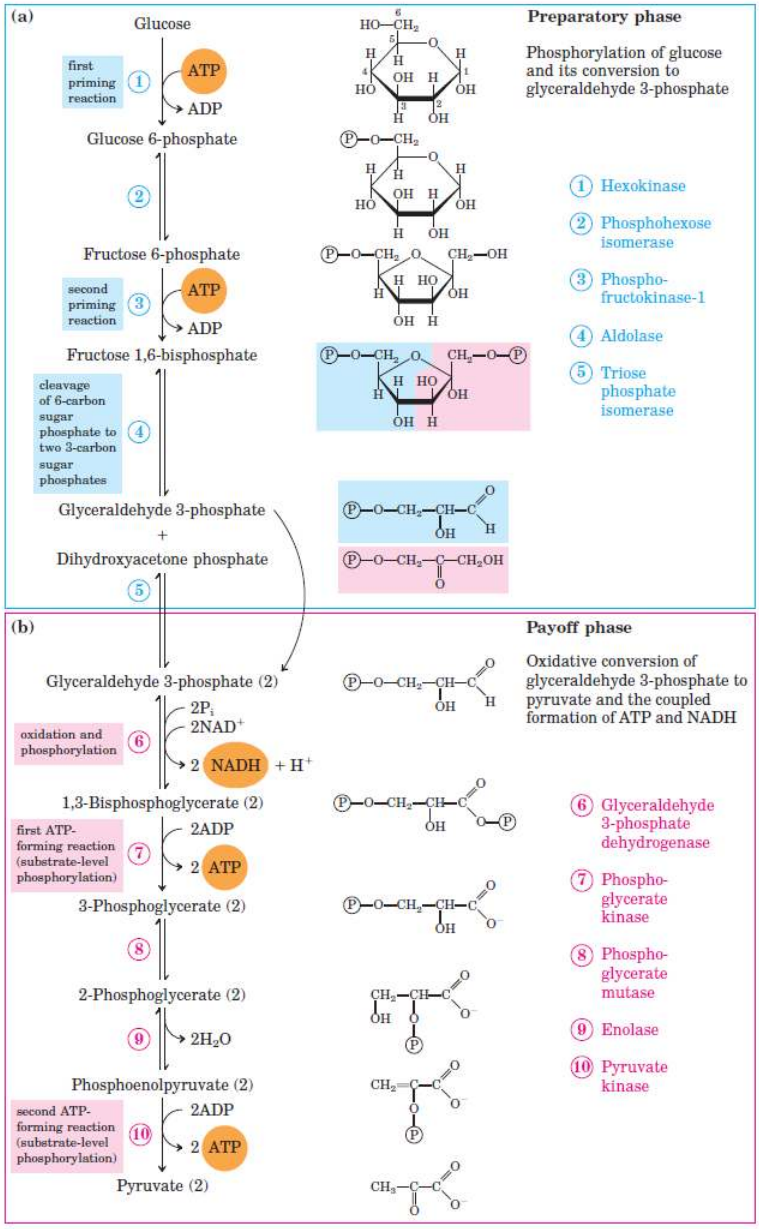


Figure 1: Glycolysis pathway.

6-Carbon Stage:

Step 1: Glucose is phosphorylated at the sixth carbon by ATP via the enzyme hexokinase to yield glucose-6-phosphate. This step is requires energy from the hydrolysis of one ATP. Therefore this step is said to be coupled to ATP conversion to ADP which is an exothermic process. In a coupled reaction an exothermic reaction provides energy needed by an endothermic one. This step also serves to trap glucose in the cell by converting glucose into a charged compound. Glucose-6-phosphate is an anion at the phosphate end.

Step 2: This is an isomer forming reaction. A pyranose sugar is converted to a furanose sugar; glucose (an aldose) is converted into fructose (a ketose). This step is catalyzed by the enzyme phosphoglucose isomerase.

Step 3: This is another phosphorylation reaction like step 1. Fructose-6-phosphate is phosphorylated at the 1-carbon position by ATP via the enzyme phosphofructokinase. Like step 1 this step requires energy and must be coupled to ATP hydrolysis.

Step 4: This step takes a 6-carbon molecule and splits it into two different 3-carbon molecules. Only one of these 3-carbon molecules is a useful intermediate in glycolysis. The useful molecule is glyceraldehyde-3-phosphate. The other 3-carbon molecule, dihydroxyacetone phosphate, is not and must be converted into glyceraldehyde-3-phosphate in step 5.

Note

The 4 steps discussed so far forms the first phase of glycolysis which requires an input of energy in the form of ATP (adenosine triphosphate) and releases none of the energy stored in glucose.

3-Carbon Stage:

Step 5: This an isomer forming reaction. It converts dihydroxyacetone phosphate into glyceraldehyde-3-phosphate. This reaction is reversible and at equilibrium greatly favors (96%) the dihydroxyacetone phosphate; the reaction proceeds readily because of efficient removal of glyceraldehyde-3-phosphate by subsequent steps (removing a product shifts equilibrium to right—Le Chatelier Principle).

Note

From here, everything should multiplied by 2.

Step 6: This is a combination of oxidation-phosphorylation step that generates a high-energy phosphate bond and a NADH. Glyceraldehyde 3-phosphate is both oxidized and phosphorylated in a reaction catalyzed by *glyceraldehyde-3-phosphate dehydrogenase*, an enzyme that requires nicotinamide adenine dinucleotide (NAD^+) as the oxidizing agent and inorganic phosphate as the phosphate donor. In the reaction, NAD^+ is reduced to nicotinamide adenine dinucleotide (NADH), and 1,3-bisphosphoglycerate is formed.

Step 7: 1,3-bisphosphoglycerate has a high-energy phosphate bond joining a phosphate group to carbon atom 1. The high-energy phosphate bond is broken and that energy is harnessed to generate an ATP. The phosphate group is transferred directly to a molecule of ADP, thus forming ATP and 3-phosphoglycerate. The enzyme that catalyzes the reaction is *phosphoglycerate kinase*, which, requires magnesium ions to function. This is the first reaction to produce ATP in the pathway.

Step 8: In the next reaction, the phosphate group on 3-phosphoglycerate is transferred from the OH group of C3 to the OH group of C2, forming 2-phosphoglycerate in a reaction catalyzed by *phosphoglyceromutase*.

Step 9: A dehydration reaction catalyzed by *enolase*, forms phosphoenolpyruvate. An alcohol is converted into an alkene with loss of water. This generates another high-energy phosphate bond.

Step 10: The high-energy phosphate bond is cleaved in phosphoenolpyruvate and the energy is harnessed to produce another ATP. The reaction is catalyzed by *pyruvate kinase*, which requires both magnesium and potassium ions to be active.This final step generates pyruvate which is used in subsequent metabolic cycles.

Note

In the second phase of glycolysis, four molecules of ATP are produced per molecule of glucose. Because glucose is split to yield two molecules in step 4, each step in the "Pay Off" phase occurs twice per molecule of glucose. The two pyruvate molecules can be converted to acetyl-CoA under aerobic conditions which then feeds into the **Citric Acid Cycle**.

To Your Health: Diabetes

Although medical science has made significant progress against diabetes , it continues to be a major health threat. Some of the serious complications of diabetes are as follows:

- It is the leading cause of lower limb amputations in the United States.
- It is the leading cause of blindness in adults over age 20.
- It is the leading cause of kidney failure.

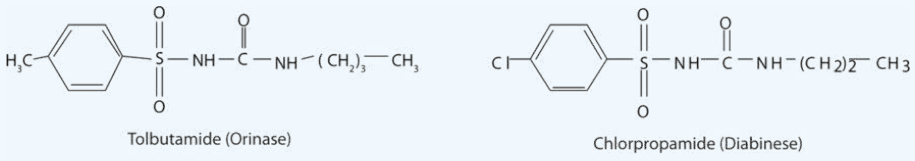
- It increases the risk of having a heart attack or stroke by two to four times.

Because a person with diabetes is unable to use glucose properly, excessive quantities accumulate in the blood and the urine. Other characteristic symptoms are constant hunger, weight loss, extreme thirst, and frequent urination because the kidneys excrete large amounts of water in an attempt to remove excess sugar from the blood.

There are two types of diabetes. In immune-mediated diabetes, insufficient amounts of insulin are produced. This type of diabetes develops early in life and is also known as *Type 1 diabetes*, as well as insulin-dependent or juvenile-onset diabetes. Symptoms are rapidly reversed by the administration of insulin, and Type 1 diabetics can lead active lives provided they receive insulin as needed. Because insulin is a protein that is readily digested in the small intestine, it cannot be taken orally and must be injected at least once a day.

In Type 1 diabetes, insulin-producing cells of the pancreas are destroyed by the body's immune system. Researchers are still trying to find out why. Meanwhile, they have developed a simple blood test capable of predicting who will develop Type 1 diabetes several years before the disease becomes apparent. The blood test reveals the presence of antibodies that destroy the body's insulin-producing cells.

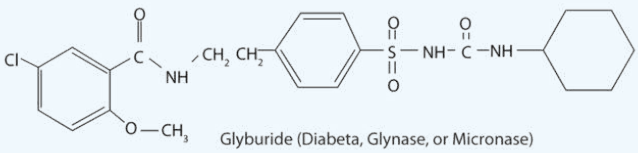
Type 2 diabetes, also known as noninsulin-dependent or adult-onset diabetes, is by far the more common, representing about 95% of diagnosed diabetic cases. (This translates to about 16 million Americans.) Type 2 diabetics may produce sufficient amounts of insulin. But the insulin-producing cells in the pancreas do not release enough of it, or it is not used properly because of defective insulin receptors or a lack of insulin receptors on the target cells. In many of these people, the disease can be controlled with a combination of diet and exercise alone. For some people who are overweight, losing weight is sufficient to bring their blood sugar level into the normal range, after which medication is not required if they exercise regularly and eat wisely.



Tolbutamide (Orinase)

Chlorpropamide (Diabinese)

First-Generation Antidiabetic Drugs



Glyburide (Diabeta, Glynase, or Micronase)

Second-Generation Antidiabetic Drugs

Those who require medication may use oral antidiabetic drugs that stimulate the islet cells to secrete insulin. First-generation antidiabetic drugs stimulated the release of insulin. Newer second-generation drugs, such as glyburide, do as well, but they also increase the sensitivity of cell receptors to insulin. Some individuals with Type 2 diabetes do not produce enough insulin and thus do not respond to these oral medications; they must use insulin. In both Type 1 and Type 2 diabetes, the blood sugar level must be carefully monitored and adjustments made in diet or medication to keep the level as normal as possible (70–120 mg/dL).

References

1. Garrett, H., Reginald and Charles Grisham. Biochemistry. Boston: Twayne Publishers, 2008.
2. Raven, Peter. Biology. Boston: Twayne Publishers, 2005.

Problems

1. What is the net yield of glycolysis as far as ATP?
2. Why is the priming phase necessary?

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7.8: Fate of Pyruvate

Learning Objectives

- Describe how the presence or absence of oxygen determines what happens to the pyruvate produced in glycolysis.

Metabolism of Pyruvate

The presence or absence of oxygen determines the fates of the pyruvate produced in glycolysis. When plenty of oxygen is available (aerobic conditions), pyruvate is first converted to acetyl-CoA. However, in the absence of oxygen (that is, under anaerobic conditions), the fate of pyruvate is different in different organisms. In vertebrates, pyruvate is converted to lactate, while other organisms, such as yeast, convert pyruvate to ethanol and carbon dioxide. These possible fates of pyruvate are summarized in Figure 7.8.1.

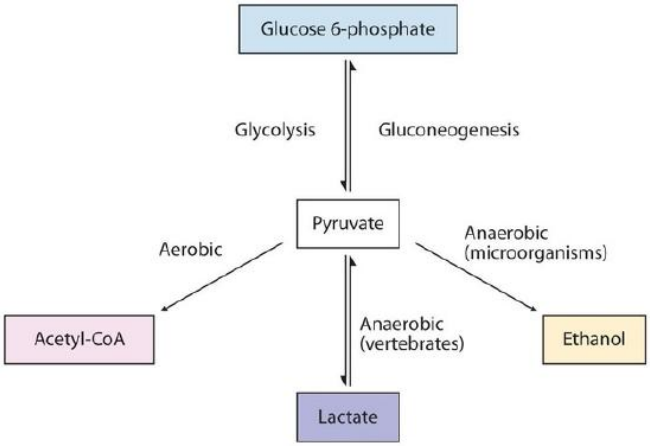


Figure 7.8.1: Metabolic Fates of Pyruvate

Aerobic Respiration

If oxygen is available, aerobic respiration will go forward. The pyruvate molecules produced at the end of glycolysis are transported into mitochondria. In order for pyruvate, the product of glycolysis, to enter the next pathway, it must undergo changes.

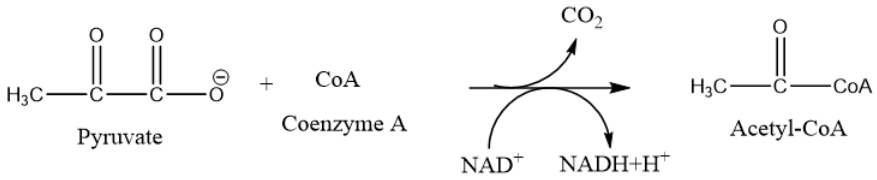


Figure 7.8.2: Metabolic Fates of Pyruvate During Aerobic Conditions

In the mitochondrial matrix, pyruvate will be transformed into a two-carbon compound by removing a molecule of carbon dioxide. This also produces NADH. An activating group called coenzyme A (CoA), which is made from vitamin B5 reacts with the two-carbon molecule. The resulting compound is called **acetyl-CoA** as shown in figure 7.8.2. Acetyl-CoA can be used in a variety of ways by the cell, but its major function is to start the citric acid cycle.

Anaerobic Respiration or Fermentation

Fermentation is the process of producing ATP in the absence of oxygen. 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 exercise.

Note: The difference between the structures of lactic acid and lactate is in the presence or absence of a H^+ ion. At physiological pH the lactic acid is present as lactate ion.

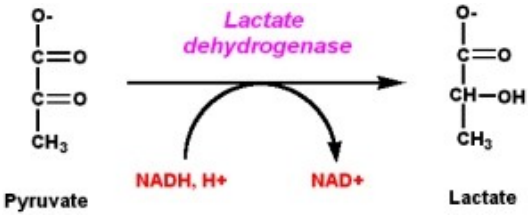


Figure 7.8.3: Metabolic Fates of Pyruvate During Anaerobic Conditions

Lactic acid fermentation converts the 3-carbon pyruvate to the 3-carbon lactate 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 oxidizing agent must be regenerated to allow ATP production to continue. To achieve this, NADH acts as a reducing agent, regenerating NAD^+. Lactate is formed by the reduction of pyruvate as shown in Figure 7.8.3.

In the human body, the lactate ions are sent in blood to the liver, where it can be used to manufacture glucose through the process of gluconeogenesis. You may have noticed this type of fermentation in your own muscles, because muscle fatigue and pain are associated with lactate ions. Lactate ions 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.

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.



Alcoholic Fermentation in Yeast

The purpose of fermentation in yeast is the same as that in muscle and bacteria, to replenish the supply of NAD^+ for glycolysis. This process occurs in two steps as shown in figure 7.8.4. It is used to make bread, wine, and biofuels.

- Alcoholic fermentation consists of pyruvate being first converted into acetaldehyde by the enzyme pyruvate decarboxylase and releasing CO_2.
- In the second step acetaldehyde is reduced to ethanol using alcohol dehydrogenase and producing NAD^+ in the process. It is this recycled NAD^+ that can be used to continue on with glycolysis.

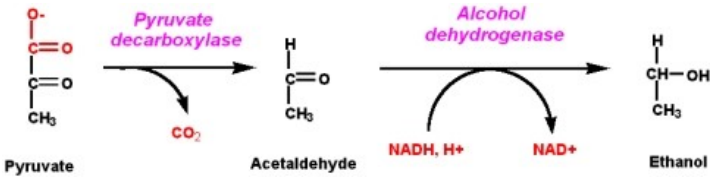


Figure 7.8.4: Metabolic Fates of Pyruvate in Yeast.

Ethanol is produced by alcoholic fermentation of the glucose in corn or other plants. This type of fermentation also explains why bread dough rises. Yeasts in bread dough use alcoholic fermentation and produce carbon dioxide gas. The gas forms bubbles in the dough, which cause the dough to expand. The bubbles also leave small holes in the bread after it bakes, making the bread light and

fluffy. The small holes in bread are formed by bubbles of carbon dioxide gas. The carbon dioxide gas was produced by alcoholic fermentation carried out by yeast.



Summary

- In the presence of oxygen, pyruvate is converted to acetyl-CoA which then enters the citric acid cycle to produce more ATP.
- In the absence of oxygen, pyruvate is converted to lactate, and NADH is re-oxidized to NAD^+.
- In alcoholic fermentation, pyruvic acid changes to alcohol and carbon dioxide. This type of fermentation is carried out by yeasts and some bacteria.

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7.9: Gluconeogenesis

The anabolic counterpart to glycolysis is gluconeogenesis. This occurs mostly in the cells of the liver and kidney. It is an anabolic pathway that synthesizes glucose from non-carbohydrate sources such as lactate, amino acids, oxaloacetate, glycerol and pyruvate. Amino acids are derived from protein digestion and glycerol is obtained from digestion of triglycerides (fats). This is an important pathway because the brain and red blood cells can only use glucose as fuel. Pyruvate is the end product of glycolysis and oxaloacetate can be diverted from the citric acid cycle. This pathway is generally only active during periods of intense exercise or starvation. While glycolysis and gluconeogenesis share many common intermediates, they are not exact opposites. The two precess are summarized below in Figure **7.9.1**.

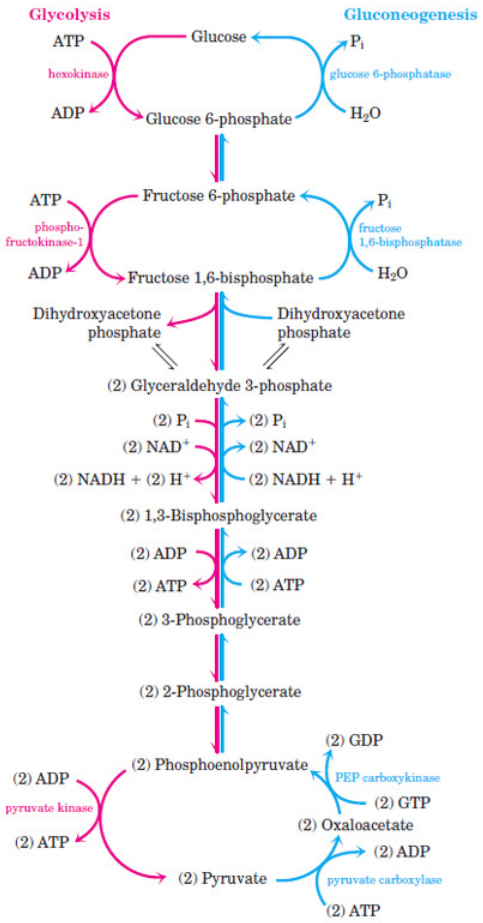


Figure **7.9.1**: Glycolysis and gluconeogenesis pathway. Glycolysis in red and gluconeogenesis in blue. (Adapted from Lehninger et al., 2005)

In seven of the eleven reactions of gluconeogenesis (starting from pyruvate), many of the same enzymes are used as in glycolysis, but the reaction directions are reversed.

- There are 3 steps in glycolysis that are so exothermic in the forward direction that they are practically impossible in the reverse direction.
  - Phosphoenolpyruvate (PEP) is much higher in energy than pyruvate and therefore the last step of glycolysis cannot simply be reversed (too much energy required). To accomplish this conversion, Gluconeogenesis proceeds via oxaloacetate. Recall that oxaloacetate is an intermediate in the Citric Acid Cycle. This conversions to oxaloacetate use one ATP. The oxaloacetate intermediate is then converted to phosphoenolpyruvate (PEP) which requires one GTP.
  - The conversion of 3-phosphoglycerate to 1,3-bisphosphoglycerate is coupled to ATP
  - The conversion of 1,3-bisphosphoglycerate to glyceraldehyde-3-phosphate requires NADH.

- Steps 9 and 11 of Gluconeogenesis are reverse of steps 1 and 3 of glycolysis. These steps involve the conversion of fructose 1,6-bisphosphate to fructose-6-phosphate and conversion of glucose-6-phosphate to glucose. The compounds are identical in both processes, but the enzymes are different and no ATP is required.

> **Note**
>
> Remember the 3-carbon stage of Gluconeogenesis is also multiplied by 2

The overall reaction of conversion of pyruvate to glucose is shown below:

$$\mathbf{2\,pyruvate + 4\,ATP + 2\,GTP + 2\,NADH \longrightarrow glucose + 4\,ADP + 2\,GDP + 2\,NAD^{+}}$$

The reaction requires the equivalent of $\mathbf{6\,ATP}$ and $\mathbf{2\,NADH}$ and hence is an anabolic process (build up). GTP is considered to be equivalent to ATP in terms of energy.

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7.10: Ketone Bodies

Learning Objectives

- Describe the structure and function of ketone bodies.

Acetyl-CoA obtained from pyruvate during aerobic conditions is oxidized by the citric acid cycle. Other sources of acetyl-CoA are lipid and protein metabolism. Some of the acetyl-CoA is used to synthesize a group of compounds known as ketone bodies. These compounds are acetoacetate, 3-hydroxybutyrate, and acetone. Two acetyl-CoA molecules combine, via several steps to form acetoacetate, which is reduced to 3-hydroxybutyrate, and a small amount is decarboxylated to carbon dioxide and acetone as shown in figure 7.10.1.

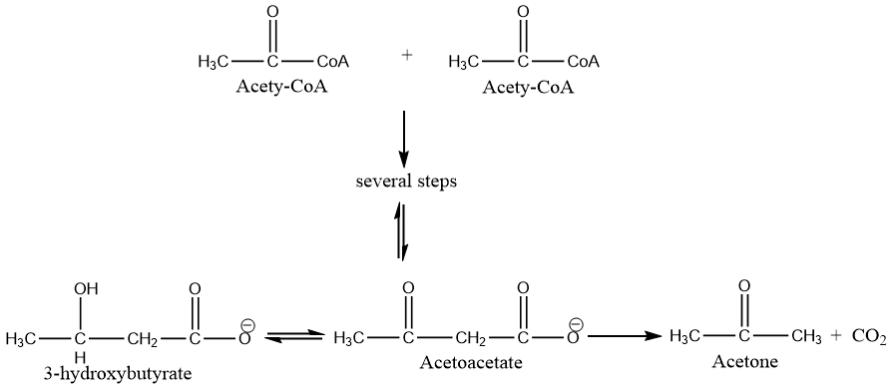


Figure **7.10.1**: Acetyl-CoA is converted to ketone bodies if not used used by the citric acid cycle.

The acetoacetate and 3-hydroxybutyrate are synthesized by the liver are released into the blood for use as a metabolic fuel (to be converted back to acetyl-CoA) by other tissues, particularly the kidney and the heart. Thus, during prolonged starvation, ketone bodies provide about 70% of the energy requirements of the brain. Under normal conditions, the kidneys excrete about 20 mg of ketone bodies each day, and the blood levels are maintained at about 1 mg of ketone bodies per 100 mL of blood.

In starvation or diabetes mellitus in which cells do not receive sufficient amounts of carbohydrates (glucose), the rate of fatty acid oxidation increases to provide energy. This leads to an increase in the concentration of acetyl-CoA. The increased acetyl-CoA cannot be oxidized by the citric acid cycle because of a decrease in the concentration of oxaloacetate, which is diverted to glucose synthesis (gluconeogenesis). In response, the rate of ketone body formation in the liver increases further, to a level much higher than can be used by other tissues. The excess ketone bodies accumulate in the blood and the urine, a condition referred to as *ketosis*. When the acetone in the blood reaches the lungs, its volatility causes it to be expelled in the breath. The sweet smell of acetone, a characteristic of ketosis, is frequently noticed on the breath of severely diabetic patients.

Two of the three kinds of ketone bodies are weak acids, their presence in the blood in excessive amounts overwhelms the blood buffers and causes a marked decrease in blood pH (to 6.9 from a normal value of 7.4). This decrease in pH leads to a serious condition known as *acidosis*. One of the effects of acidosis is a decrease in the ability of hemoglobin to transport oxygen in the blood. In moderate to severe acidosis, breathing becomes labored and very painful. The body also loses fluids and becomes dehydrated as the kidneys attempt to get rid of the acids by eliminating large quantities of water. The lowered oxygen supply and dehydration lead to depression; even mild acidosis leads to lethargy, loss of appetite, and a generally run-down feeling. Untreated patients may go into a coma. At that point, prompt treatment is necessary if the person's life is to be saved.

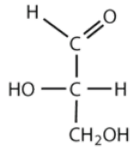
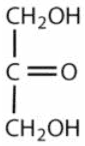
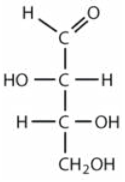
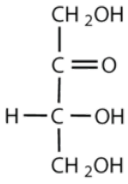
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7.11: End of Chapter Problems

Classification of Monosaccharides

What is the classification of the following monosaccharides in terms of the # of carbon atoms and functional groups.

1. 
2. 
3. 
4. 

Common Monosaccharides

1. Name two common pentoses and draw their structure in the Fischer Projection formula.

2. Name three common hexoses and draw their structure in the Fischer Projection formula.

3. How does an alpha anomer differ from a beta anomer?

4. How does a pyranose differ from a furanose?

5. Draw each molecule.

a. α-glucopyranose

b. α-galactopyranose

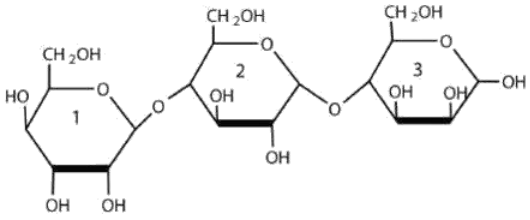
c. β-fructofuranose

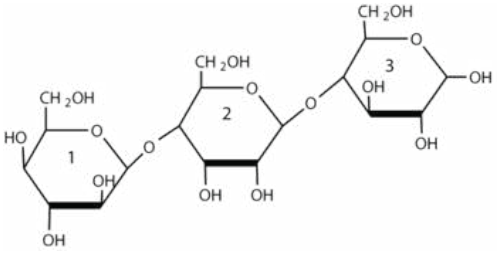
d. β-ribofuranose

e. β-2-deoxyribofuranose

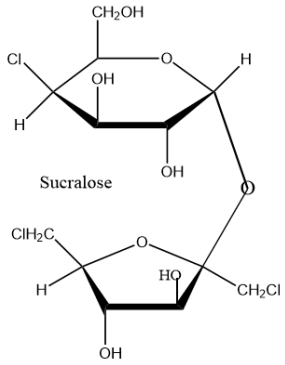
Oligosaccharides

1. For the trisaccharides, identify the glycosidic linkages.

a. 

b. 

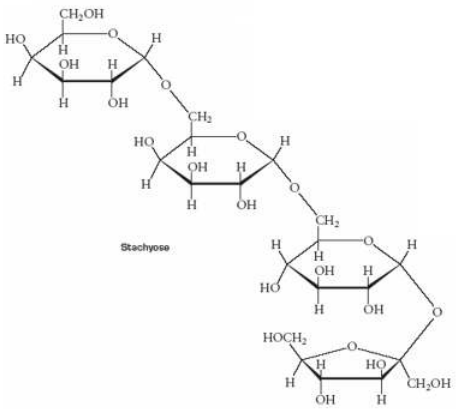
2. Sucralose is an artificial sweetener that is 600 times sweeter than sucrose and does have a structure that resembles a disaccharide. This sweetener does not add calories to your diet. It is stable at high temperature can be widely used for cooking and baking



a. Identify the glycosidic linkage.

b. What are two monosaccharides derivatives this disaccharide is made of?

3. Stachyose is an indigestible oligosaccharide present in peas and beans. This oligosaccharide contain galactopyranose residues involved in α-(1→6) glycosidic bond. Humans do not produce the α-galactosidase enzyme necessary to hydrolyze this bond. So when foods containing these carbohydrates are eaten, intestinal bacteria break this oligosaccharide down and produce gases. Products like "beano" contain an enzyme that catalyse the hydrolysis of the α-(1-6) glycosidic bond and may be taken with a meal to prevent discomfort.



a. Identify the type of glycosidic bonds in stachyose.

b. Identify the monosaccharides in stachyose.

c. Label the rings as pyranose or furanose.

Polysaccharides

1. Compare the structures of amylose and amylopectin. How are they similar and how are they different?

2. Compare the structures of amylose and cellulose. How are they similar and how are they different?

3. Compare the structures of glycogen and amylopectin. How are they similar and how are they different?

4. Why can horses use cellulose as a food source?

5. Why can humans use starch as a food source, but not cellulose?

Digestion

1. Where in the digestive system does the digestion of each of the following take place?

a. starch b. maltose c. sucrose d. lactose

2. What monosaccharides are formed when each of the following undergoes digestion and are broken down to its building blocks?

a. starch b. maltose c. sucrose d. lactose

3. Name the glycosidic bond hydrolyzed when each of the following undergoes digestion?

a. starch b. maltose c. sucrose d. lactose

4. Name the enzyme that catalyzes a reaction involved in the digestion of the following?

a. starch b. maltose c. sucrose d. lactose

5. Name three common disaccharides in our diet and name the monosaccharides formed when they are hydrolyzed in the small intestine?

6. Only a small fraction of the starch in food is digested once it enters the small intestine. Why?

7. What are the two glycosidic bonds that must be hydrolyzed to digest amylopectin to individual glucose molecules? Does the same enzyme hydrolyze both types of glycosidic bonds?

Glucose Metabolism

1. What is the role of glycolysis?

2. What is the net change in ATP and NADH from the passage of one glucose molecule through glycolysis?

3. In what part of the cell does glycolysis take place?

4. Glycolysis is described as a process in which energy is invested "up front" in exchange for greater return of energy later on. Which steps in glycolysis involve investment of energy?

5. What is the fate of the pyruvate ion under anaerobic conditions in humans? What is the purpose of this conversion?

6. What is the fate of the pyruvate ion during fermentation in yeast cells? What is the purpose of this conversion?

7. What is the fate of the pyruvate ion during aerobic conditions? What is the purpose of this conversion? What part of the cell does this conversion happen?

8. What is gluconeogenesis? What is the role of this pathway?

9. Which compound is common to the gluconeogenesis pathway and the citric acid cycle?

10. What are the three steps in glycolysis that cannot be reversed in gluconeogenesis?

11. What are ketone bodies? How are they produced?

12. A "low carb" diet craze swept the United States in 2004.

a. Explain how this diet can lead to overproduction of ketone bodies.

b. One effect of this diet was that people had "sickeningly sweet" bad breath (keto breath). What molecule is responsible for the bad breath?

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CHAPTER OVERVIEW

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8.1: Acid and Base Neutralization

Learning Objectives

- Write and understand acid-base neutralization reactions.
- Write the neutralization reactions between carboxylic acids and bases.

Neutralization Reactions

This is a reaction that happens when an acid, such as $\mathbf{HCl}$, is mixed with a base, such as $\mathbf{NaOH}$.

$$\mathbf{HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)}$$

When an acid and a base are combined, water and a salt are the products. Salts are ionic compounds containing a positive ion other than H^+ and a negative ion other than the hydroxide ion, OH^-. Double displacement reactions of this type are called **neutralization reactions**.

When a strong acid and a strong base are combined in the proper amounts, when $[H^+]$ equals $[OH^-]$, a neutral solution results in which pH = 7. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

One of the most concentrated acids in the body is stomach acid, which can be approximated as a 0.05 M hydrochloric acid solution. Special cells in the stomach wall secrete this acid, along with special enzymes, as part of the digestion process. In a laboratory, a 0.05 M solution of hydrochloric acid would dissolve some metals. How does the stomach survive the presence of such a reactive acid?

The stomach has several mechanisms for withstanding this chemical onslaught. First, the lining of the stomach is coated with a thin layer of mucus that contains some bicarbonate ions (HCO_3^-). These react with the hydrochloric acid to produce water, carbon dioxide, and harmless chloride ions.

$HCl(aq) + HCO_3^-(aq) \rightarrow HOH(l) + CO_2(g) + Cl^-(aq)$

If any acid penetrates through the mucus, it can attack the surface layer of stomach cells, called the *gastric epithelium*. Cells in the gastric epithelium are being constantly shed, so damaged cells are quickly removed and replaced with healthy cells.



However, if the gastric epithelium is destroyed faster than it can be replaced, the acid may reach the wall of the stomach, resulting in ulcers. If an ulcer grows large enough, it can expose blood vessels in the stomach wall, causing bleeding. In extreme situations, the loss of blood through a severe ulcer can threaten a person's health.

Ulcers can also result from the presence of a certain bacterium—*Helicobacter pylori*—in the stomach. The mechanism for this ulcer formation is not the same as that for ulcers caused by stomach acid and is not completely understood. However, there are two main treatments for ulcers: (1) *antacids* to react chemically with excess hydrochloric acid in the stomach and (2) *antibiotics* to destroy the *H. pylori* bacteria in the stomach. Some of the common antacids used to neutralize excess stomach acid are milk of magnesia, $Mg(OH)_2$; Tums, $CaCO_3$; and alka-selzer $NaHCO_3$.

The balanced double displacement reaction between milk of magnesia is shown below. The hydrochloric acid in the stomach is neutralized by the base in milk of magnesia.

$$Mg(OH)_2(s) + 2HCl\ (aq) \rightarrow MgCl_2(aq) + 2HOH(l)$$

With tums a neutralization reaction shown in equation 1 between stomach acid and calcium carbonate happens. This double displacement reaction produces carbonic acid, H_2CO_3 which then decomposes to produce carbon dioxide (CO_2) and water (HOH) as shown in equation 2.

$$CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCl_2(aq) + H_2CO_3(aq)\ \text{Equation 1}$$

$$H_2CO_3(aq) \rightarrow CO_2(g) + HOH(l)\ \text{Equation 2}$$

The two equations may be combined and written as one:

$$CaCO_3(s) + 2HCl\ (aq) \rightarrow CaCl_2(aq) + CO_2(g) + HOH(l)$$

Salt solutions as a result of neutralization reaction do not always have a pH of 7, however. The ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. As a general concept, if a strong acid is mixed with a weak base, the resulting solution will be slightly acidic. If a strong base is mixed with a weak acid, the solution will be slightly basic.

Neutralization of Carboxylic Acids

Carboxylic acids are more acidic than other functional groups in organic chemistry. So carboxylic acids will react with bases such as sodium hydroxide (NaOH), sodium carbonate (Na_2CO_3), and sodium bicarbonate ($NaHCO_3$) to form water and a **carboxylic acid salt.** Shown below are the neutralization reactions of acetic acid (CH_3COOH), a component in vinegar.

$$CH_3COOH + NaOH(aq) \rightarrow CH_3COO^-Na^+(aq) + H_2O$$

$$2CH_3COOH + Na_2CO_3(aq) \rightarrow 2CH_3COO^-Na^+(aq) + H_2O + CO_2(g)$$

$$CH_3COOH + NaHCO_3(aq) \rightarrow CH_3COO^-Na^+(aq) + H_2O + CO_2(g)$$

In these reactions, the carboxylic acids neutralize basic compounds. With solutions of carbonate and bicarbonate ions, they also form carbon dioxide gas.

Carboxylic acid salts are named in the same manner as inorganic salts: the name of the cation is followed by the name of the organic anion. The name of the anion is obtained by dropping the *-ic* ending of the acid name and replacing it with the suffix *-ate*. This rule applies whether we are using common names or International Union of Pure and Applied Chemistry (IUPAC) names:

$CH_3COO^-Li^+$	$CH_3CH_2CH_2COO^-K^+$	$C_6H_5COO^-Na^+$
Lithium acetate (lithium ethanoate)	Potassium butyrate (potassium butanoate)	Sodium benzoate



Video: Equimolar (~0.01 M) and equivolume solutions of **HCl** and **NaOH** are combined to make salt water. Neutralization Reaction

Example 8.1.1: Propionic Acid + Calcium Hydroxide

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($CH_3CH_2CO_2H$) with aqueous calcium hydroxide [$Ca(OH)_2$].

Solution

Double Displacement Reaction

Steps	Reaction
Write the unbalanced equation. This is a double displacement reaction, so the cations and anions swap to create the water and the salt.	$CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + H_2O(l)$
Balance the equation. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of propionic acid, $CH_3CH_2CO_2H$, to provide H^+ ions.	$\underline{2}CH_3CH_2CO_2H(aq) + Ca(OH)_2(aq) \rightarrow (CH_3CH_2CO_2)_2Ca(aq) + \underline{2}H_2O(l)$

Exercise 8.1.1

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

Answer

$$\mathbf{Ba(OH)_2\,(s) + 2\,CH_3CO_2H\,(aq) \longrightarrow Ba(CH_3CO_2)_2\,(aq) + 2\,H_2O\,(l)}$$

Contributions & Attributions

This page was constructed from content via the following contributor(s) and edited (topically or extensively) by the LibreTexts development team to meet platform style, presentation, and quality:

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8.2: Acids-Properties and Examples

Learning Objectives

- Examine properties of acids.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits from hard water build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits of hard water and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

Properties of Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid. Ascorbic acid is better known as vitamin C. Carbonated sodas contain carbonic acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicators. Two common indicators in the laboratory are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless. Dye extracted from purple cabbage is an excellent indicator to qualitatively determine the acid or base properties of household chemicals.
4. Acids react with active metals to yield hydrogen gas. Metals will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g) \tag{8.2.1}$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

Common Acids

Common **binary acids** are formed from monoatomic anions of the halogen family (group 7A). The anions fluoride (F^-), chloride (Cl^-), bromide (Br^-) and iodide (I^-) combine with one hydrogen ion, (H^+) to form hydrofluoric acid, hydrochloric acid, hydrobromic acid, and hydroiodic acid. The acids are called binary because they are made of two elements.

hydrofluoric acid HF (aq)

hydrochloric acid HCl (aq)

hydrobromic acid HBr(aq)

hydroiodic acid HI (aq)

Common **ternary acids** are formed from polyatomic anions such as sulfate (SO_4^{2-}), nitrate (NO_3^-), carbonate (CO_3^{2-}) and phosphate (PO_4^{3-}). The anions sulfate and carbonate combine with two hydrogen ions and form sulfuric acid and carbonic acid. The nitrate ion combine with one hydrogen ion and form nitric acid, and the phosphate anion combine with three hydrogen ions to form phosphoric acid. The acids are called ternary because they are made of three elements.

Nitric acid HNO_3(aq)

Sulfuric acid H_2SO_4(aq)

Carbonic acid H_2CO_3 (aq)

Phosphoric acid H_3PO_4(aq)

There are other common acids based on the carboxylic acid functional group. Examples are acetic acid in vinegar, citric acid in citrus fruits, and ascorbic acid in vitamin C. Here's another partial list of some common acids, along with some chemical formulas and uses.

Table **8.2.1**: Common Acids and Their Uses

Chemist Name	Common Name	Uses
hydrochloric acid, HCl	muriatic acid (used in pools) and stomach acid is HCl	Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning.
sulfuric acid, H_2SO_4		Used in car batteries, and in the manufacture of fertilizers.
nitric acid, HNO_3		Used in the manufacture of fertilizers, explosives and in extraction of gold.
acetic acid, $HC_2H_3O_2$	vinegar	Main ingredient in vinegar.
carbonic acid, H_2CO_3	responsible for the "fizz" in carbonated drinks	As an ingredient in carbonated drinks.
citric acid, $C_6H_8O_7$		Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions.
acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$	aspirin	The active ingredient in aspirin.

Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. A colorless, highly pungent solution of hydrogen chloride (HCl) in water. Hydrochloric acid is usually prepared by treating $\mathbf{HCl}$ with water.

$$\mathbf{HCl\,(g) + H_2O\,(l) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{8.2.2}$$

Hydrochloric acid can therefore be used to prepare chloride salts. Hydrochloric acid is a strong acid, since it is completely dissociated in water. Hydrochloric acid is the preferred acid in titration for determining the amount of bases.

Sulfuric Acid

Sulfuric acid is a highly corrosive strong mineral acid with the molecular formula $\mathbf{H_2SO_4}$. Sulfuric acid is a diprotic acid and has a wide range of applications including use in domestic acidic drain cleaners,[as an electrolyte in lead-acid batteries, and in various cleaning agents. It is also a central substance in the chemical industry.



Figure **8.2.1**: Drops of concentrated sulfuric acid rapidly decompose a piece of cotton towel by dehydration. (CC BY-SA 3.0; Toxic Walker).

The hydration of sulfuric acid is highly exothermic and the affinity of it for water is sufficiently strong. So sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating property, removing water ($\mathbf{H_2O}$) from

other compounds including sugar and other carbohydrates and producing carbon, heat, steam. Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas (Equation 8.2.3).

$$\mathrm{Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2} \tag{8.2.3}$$

Nitric Acid

Nitric acid ($\mathrm{HNO_3}$) is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. Nitric acid can be made by reacting nitrogen dioxide ($\mathrm{NO_2\,(g)}$) with water.

$$\mathrm{3\,NO_2\,(g) + H_2O\ (l) \longrightarrow 2\,HNO_3\,(ag) + NO\,(g)} \tag{8.2.4}$$

Nitric acid reacts with most metals, but the details depend on the concentration of the acid and the nature of the metal. Dilute nitric acid behaves as a typical acid in its reaction with most metals (e.g., nitric acid with magnesium, manganese or zinc will liberate $\mathrm{H_2}$ gas):

$$\mathrm{Mg + 2\,HNO_3 \longrightarrow Mg(NO_3)_2 + H_2} \tag{8.2.5}$$

$$\mathrm{Mn + 2\,HNO_3 \longrightarrow Mn(NO_3)_2 + H_2} \tag{8.2.6}$$

$$\mathrm{Zn + 2\,HNO_3 \longrightarrow Zn(NO_3)_2 + H_2} \tag{8.2.7}$$

Nitric acid is a corrosive acid and a powerful oxidizing agent. The major hazard posed by it is chemical burns as it carries out acid hydrolysis with proteins (amide) and fats (ester) which consequently decomposes living tissue (Figure 8.2.2). Concentrated nitric acid stains human skin yellow due to its reaction with the keratin



Figure 8.2.2: Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).

Carbonic Acid

Carbonic acid is a chemical compound with the chemical formula $\mathrm{H_2CO_3}$ and is also a name sometimes given to solutions of carbon dioxide in water (carbonated water), because such solutions contain small amounts of $\mathrm{H_2CO_3\,(aq)}$. Carbonic acid, which is a weak acid, forms two kinds of salts: the carbonates and the bicarbonates. In geology, carbonic acid causes limestone to dissolve, producing calcium bicarbonate—which leads to many limestone features such as stalactites and stalagmites. Carbonic acid is a polyprotic acid, specifically it is diprotic, meaning it has two protons which may dissociate from the parent molecule.

When carbon dioxide dissolves in water it produces carbonic acid:

$$\mathrm{CO_2 + H_2O \rightleftharpoons H_2CO_3} \tag{8.2.8}$$

The reaction can be pushed to favor the reactants to generate $\mathrm{CO_2\,(g)}$ from solution, which is key to the bubbles observed in carbonated beverages (Figure 8.2.3).



Figure 8.2.3: A glass of sparkling water. (CC BY-SA 3.0; Nevit Dilmen).

Formic Acid

Formic acid (HCO_2H) is the simplest carboxylic acid and is an important intermediate in chemical synthesis and occurs naturally, most notably in some ants. The word "formic" comes from the Latin word for ant, formica, referring to its early isolation by the distillation of ant bodies. Formic acid occurs widely in nature as its conjugate base formate.

Citric Acid

Citric acid ($C_6H_8O_7$) is a weak organic tricarboxylic acid that occurs naturally in citrus fruits. The citrate ion is an intermediate in the citric acid cycle (Krebs cycle), a central metabolic pathway for animals, plants and bacteria. Because it is one of the stronger edible acids, the dominant use of citric acid is used as a flavoring and preservative in food and beverages, especially soft drinks.

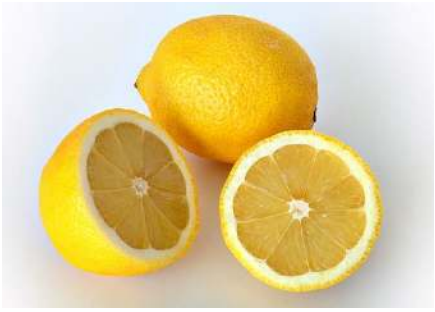


Figure 8.2.4: Lemons, oranges, limes, and other citrus fruits possess high concentrations of citric acid (CC BY-SA 2.5; André Karwath).

Acetylsalicylic Acid

Acetylsalicylic acid (also known as aspirin) is a medication used to treat pain, fever, and inflammation. Aspirin, in the form of leaves from the willow tree, has been used for its health effects for at least 2,400 years.

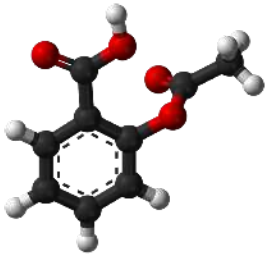


Figure 8.2.5: Ball-and-stick model of the aspirin molecule. (Public Domain; Ben Mills).

Aspirin is a white, crystalline, weakly acidic substance.

Summary

A brief summary of key aspects of several acids commonly encountered by students was given. Acids are a distinct class of compounds because of the properties of their aqueous solutions.

Contributions & Attributions

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8.3: Bases-Properties and Examples

Learning Objectives

- Examine properties and nomenclature of bases.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the **HCl**(aq) in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink. The indicator extracted from purple cabbage can qualitatively determine the pH of household chemicals. In basic solutions the color of the indicator from the purple cabbage varies from bluish green to yellow.
4. Bases react with acids to produce a salt and water.



Figure **8.3.1**: Phenolphthalein indicator in presence of base.

Bases are less common in foods. However, **antacids** are products which combat excess stomach acid, are comprised of bases such as magnesium hydroxide, or sodium hydrogen carbonate, or calcium carbonate. They are also present in many household cleaning products. Ammonia cleaner is a base. Sodium hydroxide is found in drain cleaner.

Nomenclature and Formulae of Common Bases

Common bases are ionic compounds of polyatomic ions such as hydroxide (OH^-), carbonate (CO_3^{2-}) and bicarbonate (HCO_3^-) anions.

Group 1A cations such as sodium ion (Na^+), potassium ion (K^+), or lithium ion (Li^+) combine with one hydroxide ion to form an ionic compound which functions as a base. Group 2A cations such as magnesium ion (Mg^{2+}), calcium ion (Ca^{2+}), or barium ion (Ba^{2+}) combine with two hydroxide ions to form an ionic compound that is a base.

Sodium hydroxide NaOH

Potassium hydroxide KOH

Magnesium hydroxide $Mg(OH)_2$

Barium hydroxide $Ba(OH)_2$

The polyatomic anion carbonate combines with two group 1A metal cations and one group 2A metal cation to form an ionic compound that functions as a base.

Lithium Carbonate Li_2CO_3

Sodium Carbonate Na_2CO_3

Calcium carbonate $CaCO_3$

Magnesium Carbonate $MgCO_3$

The polyatomic anion hydrogencarbonate also called the bicarbonate anion combines with one group 1A metal cations and two group 2A metal cation to form an ionic compound that functions as a base.

Sodium bicarbonate $NaHCO_3$

Potassium bicarbonate $KHCO_3$

Calcium bicarbonate $Ca(HCO_3)_2$

Magnesium bicarbonate $Mg(HCO_3)_2$

Various common bases and corresponding uses are given in Table **8.3.1**.

Table **8.3.1**: Common Bases and Corresponding Uses

Some Common Bases	Uses
sodium hydroxide, NaOH (lye or caustic soda)	Used in the manufacture of soaps and detergents, and as the main ingredient in oven and drain cleaners.
potassium hydroxide, KOH (lye or caustic potash)	Used in the production of liquid soaps and soft soaps. Used in alkaline batteries.
magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia)	Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater.
calcium hydroxide, $Ca(OH)_2$ (slaked lime)	Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil.
aluminum hydroxide	Used in water purification and as an ingredient in antacids.
ammonia, NH_3	Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers.

Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula **NaOH**. It is a white solid ionic compound consisting of sodium cations $\mathbf{Na^+}$ and hydroxide anions $\mathbf{OH^-}$.



Figure **8.3.1**: Sample of sodium hydroxide as pellets in a watch glass. (Public Domain; Walkerma).

Dissolution of solid sodium hydroxide in water is a highly exothermic reaction

$$\mathbf{NaOH\,(s)\ \rightarrow Na^+\,(aq) + OH^-\,(aq)} \tag{8.3.1}$$

The resulting solution is usually colorless and odorless and feels slippery when it comes in contact with skin.

Potassium Hydroxide

Potassium hydroxide is an inorganic compound with the formula **KOH**, and is commonly called caustic potash. Along with sodium hydroxide (NaOH), this colorless solid is a prototypical strong base. It has many industrial and niche applications, most of which exploit its corrosive nature and its reactivity toward acids. Its dissolution in water is strongly exothermic.

$$\mathrm{KOH\,(s)\ \rightarrow K^+\,(aq) + OH^-\,(aq)} \tag{8.3.2}$$

Concentrated aqueous solutions are sometimes called *potassium lyes*.

Magnesium Hydroxide

Magnesium hydroxide is the inorganic compound with the chemical formula $\mathrm{Mg(OH)_2}$. Magnesium hydroxide is a common component of antacids, such as milk of magnesia, as well as laxatives.



Figure 8.3.1: Bottle of Antacid tablets. (CC BY 2.,5; Midnightcomm).

It is a white solid with low solubility in water. Combining a solution of many magnesium salts with basic water induces precipitation of solid $\mathrm{Mg(OH)_2}$. However, a weak concentration of dissociated ions can be found in solution:

$$\mathrm{Mg(OH)_2\,(s) \rightleftharpoons Mg^{2+}\,(aq) + 2\,OH^-\,(aq)} \tag{8.3.3}$$

Calcium Hydroxide

Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $\mathrm{Ca(OH)_2}$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime. Calcium hydroxide is used in many applications, including food preparation. Limewater is the common name for a saturated solution of calcium hydroxide.

Calcium hydroxide is relatively insoluble in water, but is large enough that its solutions are basic according to the following reaction:

$$\mathrm{Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2\,OH^-\,(aq)} \tag{8.3.4}$$

Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula $\mathrm{NH_3}$ and is a colorless gas with a characteristic pungent smell. It is the active product of “smelling salts,” and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form.

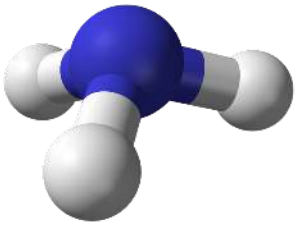


Figure 8.3.1: Ball-and-stick model of the ammonia molecule. (Public Domain; Ben Mills).

In aqueous solution, ammonia acts as a base, acquiring hydrogen ions from $\mathrm{H_2O}$ to yield ammonium and hydroxide ions:

$$\mathrm{NH_3\,(g) + H_2O\,(l) \rightleftharpoons NH_4{}^+\,(aq) + OH^-\,(aq)} \tag{8.3.5}$$

Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

Summary

- A brief summary of properties of bases was given.
- The properties of bases mostly contrast those of acids.
- Bases have many, varied uses.

Contributions & Attributions

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8.4: Brønsted-Lowry Theory of Acids and Bases

Learning Objectives

- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Understand amphoteric properties of ions and compounds.
- Identify conjugate acid-base pairs in an acid-base reaction.

There are three major theories of acids or bases. The theory developed by Svante Arrhenius in 1883, the Arrhenius definition, states that an acid produces H^+ in solution and a base produces OH^-. Later, two more sophisticated and general theories were proposed. These theories are the Brønsted-Lowry and Lewis definitions of acids and bases. This section covers the Brønsted-Lowry theory.

The Brønsted-Lowry Theory of Acids and Bases

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, a Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

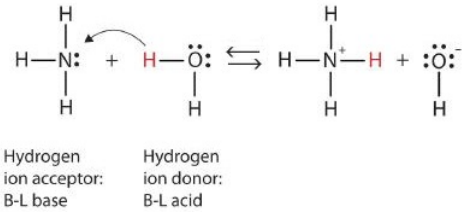
> *A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.*

Example 1:

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$\mathrm{NH_3\,(aq) + H_2O\,(\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)} \tag{8.4.1}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:



Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule accepts the hydrogen ion, is the Brønsted-Lowry base.

Example 2:

In this example water and HCN are reactants, while cyanide ion and hydronium ion are products.

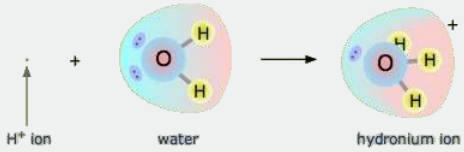
$$\mathrm{HCN(aq) + H_2O(\ell) \rightleftharpoons CN^-(aq) + H_3O^+(aq)}$$

The water molecule has accepted a hydrogen ion from the HCN molecule which in turn donates a proton to water. Because the water molecule accepts a hydrogen ion from HCN, it is the Brønsted-Lowry base, while the HCN molecule donates the hydrogen ion is the Brønsted-Lowry acid.

The Hydronium Ion

Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes,

H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

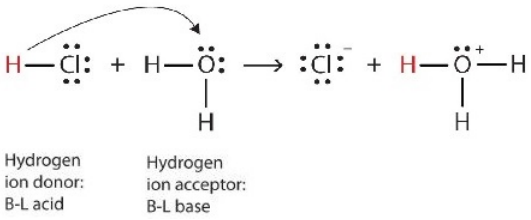


Example 3:

With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$\mathbf{HCl\,(g) + H_2O\ (\ell)\ \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{8.4.2}$$

We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base.

Amphoteric or Amphiprotic Compounds

Water (H_2O) is an interesting compound. Depending on the circumstances, H_2O can act as either as a Brønsted-Lowry acid or a Brønsted-Lowry base. In example 1, the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid. In example 2, the water molecule accepts a hydrogen ion from HCN, it is the Brønsted-Lowry base. In example 3, as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base.

A substance that can either donate or accept a proton, depending on the circumstances, is called amphoteric or amphiprotic.

Water is not the only substance that can react as an acid in some cases or a base in others, but it is certainly the most common example—and the most important one. Other examples are the bicarbonate ion (HCO_3^-), the hydrogen phosphate (HPO_4^{2-}) and the dihydrogenphosphate ion ($H_2PO_4^-$).

A water molecule can act as an acid or a base even in a sample of pure water. About 6 in every 100 million (6 in 10^8) water molecules undergo the following reaction:

$$H_2O(\ell) + H_2O(\ell) \rightarrow H_3O^+(aq) + OH^-(aq)$$

This process is called the **autoionization of water** and occurs in every sample of water, whether it is pure or part of a solution. Autoionization occurs to some extent in any amphiprotic liquid. (For comparison, liquid ammonia undergoes autoionization as well, but only about 1 molecule in a million billion (1 in 10^{15}) reacts with another ammonia molecule.)

- A Brønsted-Lowry acid is a proton (hydrogen ion) donor.
- A Brønsted-Lowry base is a proton (hydrogen ion) acceptor.

8.4.1 Example

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule, just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry

acid and base in the reactant side.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$C_6H_5NH_2\,(aq) + H_2O(\ell) \rightleftharpoons C_6H_5NH_3^+\,(aq) + OH^-\,(aq)$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

8.4.1 Exercise

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in the reactant side of this chemical equation.

$$H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+ \tag{8.4.3}$$

Answer

Brønsted-Lowry acid: $H_2PO_4^-$; Brønsted-Lowry base: H_2O

8.4.2 Exercise

Which of the following compounds is a Bronsted-Lowry base?

a. HCl
b. HPO_4^{2-}
c. H_3PO_4
d. NH_4^+
e. $CH_3NH_3^+$

Answer

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H^+. This eliminates HCl, H_3PO_4, NH_4^+ and $CH_3NH_3^+$ because they are Bronsted-Lowry acids. They all give away protons. In the case of HPO_4^{2-}, consider the following equation:

$$HPO_4^{2-}\,(aq) + H_2O\,(l) \rightarrow PO_4^{3-}\,(aq) + H_3O^+\,(aq)$$

Here, it is clear that HPO_4^{2-} is the acid since it donates a proton to water to make H_3O^+ and PO_4^{3-}. Now consider the following equation:

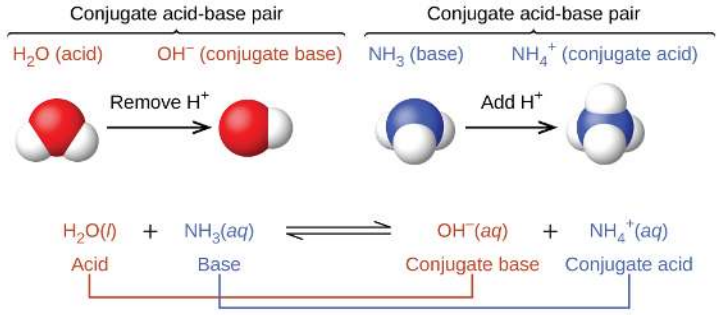
$$HPO_4^{2-}\,(aq) + H_2O\,(l) \rightarrow H_2PO_4^- + OH^-\,(aq)$$

In this case, HPO_4^{2-} is the base since it accepts a proton from water to form $H_2PO_4^-$ and OH^-. Thus, HPO_4^{2-} is an acid and base together, making it amphoteric.

Since HPO_4^{2-} is the only compound from the options that can act as a base, the answer is **(b) HPO_4^{2-}.**

Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, OH^-, and the conjugate acid of ammonia, NH_4^+:



In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are NH_4^+/NH_3 and H_2O/OH^-.

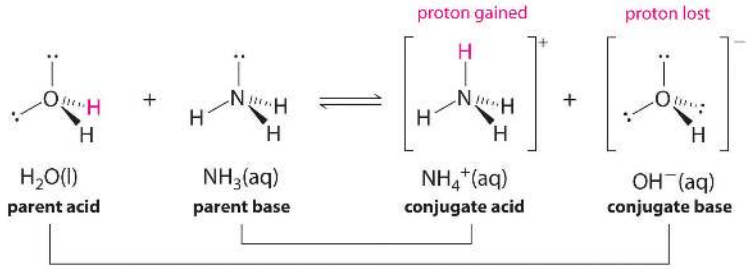


Figure 8.4.1. The pairing of parent acids and bases with conjugate acids and bases.

8.4.2 Example

Identify the conjugate acid-base pairs in this equilibrium.

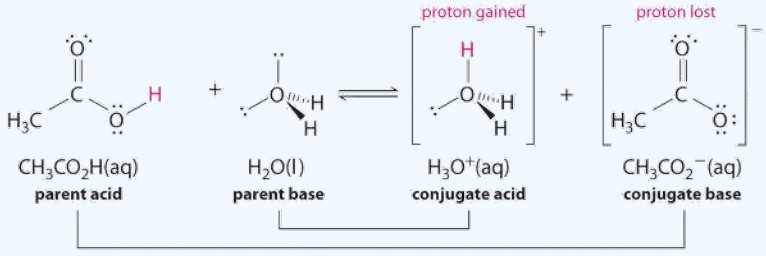
$$CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-$$

Solution

Similarly, in the reaction of acetic acid with water, acetic acid **donates** a proton to water, which acts as the base. In the reverse reaction, H_3O^+ is the acid that donates a proton to the acetate ion, which acts as the base.

Once again, we have two conjugate acid-base pairs:

- the parent acid and its conjugate base ($CH_3CO_2H/CH_3CO_2^-$) and
- the parent base and its conjugate acid (H_3O^+/H_2O).



8.4.3 Example

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H_2O and OH^-, where H_2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base.

The other pair consists of $(CH_3)_3N$ and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and $(CH_3)_3N$ is the conjugate base.

8.4.3 Exercise

Identify the conjugate acid-base pairs in this equilibrium.

$$\mathbf{NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-}$$

Answer

H_2O (acid) and OH^- (base); NH_2^- (base) and NH_3 (acid)

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8.5: Equilibrium

Learning Objectives

- Define *chemical equilibrium* and the equilibrium constant, ***K***.
- Recognize chemical equilibrium as a dynamic process.

Consider the following reaction occurring in a closed container (so that no material can go in or out):

$$H_2 + I_2 \rightarrow 2HI$$

This is the reaction between elemental hydrogen and elemental iodine to make hydrogen iodide. The way the equation is written, we are led to believe that the reaction goes to completion, that all the H_2 and the I_2 react to make HI.

However, this is not the case. The reverse chemical reaction is also taking place:

$$2HI \rightarrow H_2 + I_2$$

It acts to undo what the first reaction does. Eventually, the reverse reaction proceeds so quickly that it matches the speed of the forward reaction. When that happens, the reaction has reached **chemical equilibrium**, the point at which the forward and reverse processes balance each other's progress.

Because two opposing processes are occurring at once, it is conventional to represent an equilibrium using a double arrow, like this:

$$H_2 + I_2 \rightleftharpoons 2HI$$

The double arrow implies that the reaction is going in both directions. Note that the reaction must still be balanced.

Example 8.5.1

Write the equilibrium equation that exists between calcium carbonate as a reactant and calcium oxide and carbon dioxide as products.

Solution

As this is an equilibrium situation, a double arrow is used. The equilibrium equation is written as follows:

$$CaCO_3 \rightleftharpoons CaO + CO_2$$

Exercise 8.5.1

Write the equilibrium equation between elemental hydrogen and elemental oxygen as reactants and water as the product.

Answer

$$2H_2 + O_2 \rightleftharpoons 2H_2O$$

One thing to note about equilibrium is that the reactions do not stop; both the forward reaction and the reverse reaction continue to occur. They both occur at the same rate, so any overall change by one reaction is canceled by the reverse reaction. We say that chemical equilibrium is *dynamic*, rather than static.The concentration of reactants and products does not change over time.

Both reactions are occurring simultaneously, the equilibrium can be written backward. For example, representing an equilibrium as

$$H_2 + I_2 \rightleftharpoons 2HI$$

is the same thing as representing the same equilibrium as

$$2HI \rightleftharpoons H_2 + I_2$$

The reaction must be at equilibrium for this to be the case.

The Equilibrium Constant

In the mid 1860s, Norwegian scientists C. M. Guldberg and P. Waage noted a peculiar relationship between the amounts of reactants and products in an equilibrium. No matter how many reactants they started with, a certain ratio of reactants and products was achieved at equilibrium. Today, we call this observation the **law of mass action**. It relates the amounts of reactants and products at equilibrium for a chemical reaction. For a general chemical reaction occurring in solution,

$$aA + bB \rightleftharpoons cC + dD$$

the **equilibrium constant**, also known as K_{eq}, is defined by the following expression:

$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

where [A] is the molar concentration of species A at equilibrium, and so forth. The coefficients *a*, *b*, *c*, and *d* in the chemical equation become exponents in the expression for K_{eq}. The K_{eq} is a characteristic numerical value for a given reaction at a given temperature; that is, each chemical reaction has its own characteristic K_{eq}. The concentration of each reactant and product in a chemical reaction at equilibrium is *related*; the concentrations cannot be random values, but they depend on each other. The numerator of the expression for K_{eq} has the concentrations of every product (however many products there are), while the denominator of the expression for K_{eq} has the concentrations of every reactant, leading to the common *products over reactants* definition for the K_{eq}.

Example 8.5.2

Balance the reaction equation and then write the corresponding equilibrium constant expression:

$CO(g) + O_2(g) \rightleftharpoons CO_2(g)$

Solution

The balanced chemical equation is:

$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)$

The equilibrium constant expression is:

$$K_{eq} = \frac{[CO_2]^2}{[CO]^2\,[O_2]}$$

Many chemical reactions involve substances in the solid or liquid phases. For example, a particular chemical reaction is represented as follows:

$$2\,NaHCO_3\,(s) \rightleftharpoons Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(l) \qquad (8.5.1)$$

This chemical equation includes all three phases of matter. This kind of equilibrium is called a **heterogeneous equilibrium** because there is more than one phase present.

The rule for heterogeneous equilibria is as follows: *Do not include the concentrations of pure solids and pure liquids in* K_{eq} *expressions. See example 3 and 4 below.*

Example 8.5.3

Write the equilibrium constant expression for the following reaction:

$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

Solution

The equilibrium constant expression is:

$$K_{eq} = \frac{[CO]\ [H_2]}{[H_2O]}$$

The concentration of the solid C(s) is the number of moles present in a given volume, a value that never changes at a given temperature, so the concentration of solid reactants or products are omitted in the equilibrium constant expression.

Example 8.5.4

Write the equilibrium constant expression for the following reaction:

$H_2SO_4(aq) + H_2O(l) \rightleftharpoons HSO_4^-(aq) + H_3O^+(aq)$

Solution

The equilibrium constant expression is:

$$K_{eq} = \frac{[HSO_4^-]\ [H_3O^+]}{[H_2SO_4]}$$

The solvent (H_2O) is left out of the equilibrium constant expression because the concentration of the solvent is very high and does not change significantly during a reaction.

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8.6: Acid and Base Strength

Learning Objectives

- Describe the difference between strong and weak acids and bases.

Strong and Weak Acids

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider $\mathbf{HCl(aq)}$. When $\mathbf{HCl}$ is dissolved in $\mathbf{H_2O}$, it completely dissociates into H^+ (aq) and Cl^-(aq) ions; all the HCl molecules become ions:

$$\mathbf{HCl \xrightarrow{100\,\%} H^+(aq) + Cl^-(aq)} \tag{8.6.1}$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. $HC_2H_3O_2$ is an example of a weak acid:

$$HC_2H_3O_2 \xrightarrow{\sim 5\%} H^+(aq) + C_2H_3O_2^-(aq) \tag{8.6.2}$$

Because this dissociation does not go 100% to completion, it is more appropriate to write it as **a reversible reaction**:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq) \tag{8.6.3}$$

As it turns out, there are very few strong acids, which are given in Table **8.6.1**. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

> *Any acid that dissociates 100% into ions is called a strong acid. If it does not dissociate 100%, it is a weak acid.*

Table **8.6.1**: Strong Acids and Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	RbOH
H_2SO_4	CsOH
$HClO_3$	
$HClO_4$	

Strong and Weak Bases

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it is a **weak base**. There are very few strong bases (Table **8.6.1**); any base not listed is a weak base. All strong bases are OH^- compounds. So a base based on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base. Bases that are salts of carbonate and bicarbonate ions are also weak bases

8.6.1 Example : Identifying Strong and Weak Acids and Bases

Identify each acid or base as strong or weak.

a. HCl
b. $Mg(OH)_2$

c. C_5H_5N

Solution

a. Because HCl is listed in Table **8.6.1**, it is a strong acid.

b. Because $Mg(OH)_2$ is not listed in Table **8.6.1**, it is a weak base.

c. The nitrogen in C_5H_5N would act as a proton acceptor and therefore can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

8.6.1 Exercise

Identify each acid or base as strong or weak.

a. **RbOH**

b. **HNO_2**

Answer a

strong base

Answer b

weak acid

Chemical Equilibrium in Weak Acids and Bases

The behavior of **weak acids and bases** illustrates a key concept in chemistry. Does the chemical equation describing the ionization of a weak acid or base just stop when the acid or base is done ionizing? Actually, no. Rather, the reverse process, the reformation of the molecular form of the acid or base occurs, ultimately at the same rate as the ionization process. For example, the ionization of the weak acid $HC_2H_3O_2(aq)$ is as follows:

$$HC_2H_3O_2(aq)+H_2O(\ell)\rightarrow H_3O^+(aq)+C_2H_3O^-(aq)$$

The reverse process also begins to occur:

$$H_3O^+(aq)+C_2H_3O^-(aq)\rightarrow HC_2H_3O_2(aq)+H_2O(\ell)$$

Eventually, there is a balance between the two opposing processes, and no additional change occurs. The chemical reaction is better represented at this point with a double arrow:

$$HC_2H_3O_2(aq)+H_2O(\ell)\rightleftharpoons H_3O^+(aq)+C_2H_3O^-(aq)$$

The ⇌ implies that **both the forward and reverse reactions** are occurring, and their effects cancel each other out. The process at this point is considered to be at **chemical equilibrium** (or equilibrium). It is important to note that the processes do not stop. They balance out each other so that there is no further net change; that is, chemical equilibrium is a *dynamic equilibrium*.

Acid Ionization Constant, *K*a

The ionization for a general weak acid, HA, can be written as follows:

$$HA(aq)\rightleftharpoons H^+(aq)+A^-(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid ionization constant (*K*a)** is the equilibrium constant for the ionization of an acid.

$$K_a=\frac{[H^+]\,[A^-]}{[HA]}$$

The acid ionization represents the fraction of the original acid that has been ionized in solution. Therefore, the numerical value of *K*a is a reflection of the strength of the acid. Weak acids with relatively higher *K*a values are stronger than acids with relatively lower *K*a values. Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the *K*a value approaches infinity. For this reason, *K*a values are generally reported for weak acids only.

The table below is a listing of acid ionization constants for several acids. Note that polyprotic acids have a distinct ionization constant for each ionization step, with each successive ionization constant being smaller than the previous one.

Table **8.6.2**: Acid Ionization Constants at 25°C

Name of Acid	Ionization Equation	K_a
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$	very large
	$HSO_4 \rightleftharpoons H^+ + SO_4^{2-}$	1.3×10^{-2}
Hydrofluoric acid	$HF \rightleftharpoons H^+ + F^-$	7.1×10^{-4}
Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$	4.5×10^{-4}
Benzoic acid	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.5×10^{-5}
Acetic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$	4.2×10^{-7}
	$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.8×10^{-11}

Key Takeaways

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.

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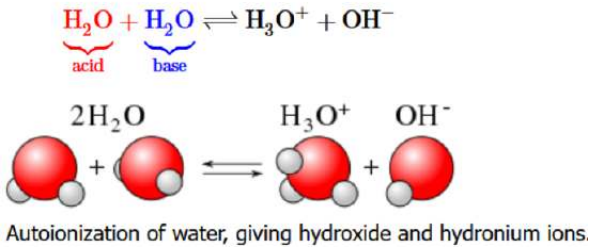
8.7: Ionization of Water and The pH Scale

Learning Objectives

- Define and use the ion product constant for water to calculate concentrations of H_3O^+ and OH^- in aqueous solutions.
- Define the pH scale and use it to describe acids and bases.

Ionization of Water

As we have already seen, H_2O can act as an acid or a base. Within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:



Autoionization of water, giving hydroxide and hydronium ions.

Similar to a weak acid, the autoionization of water is an equilibrium process, and is more properly written as follows:

$$H_2O(\ell)+H_2O(\ell) \rightleftharpoons H_3O^+(aq)+OH^-(aq)$$

The equilibrium constant for the autoionization of water is referred to as the ion-product for water and is given the symbol *Kw*.

$$Kw = [H_3O^+][OH^-]$$

The **ion-product constant for water (*Kw*)** is the mathematical product of the concentration of hydronium ions and hydroxide ions. Note that H_2O is not included in the ion-product expression because it is a pure liquid. The value of *Kw* is very small, in accordance with a reaction that favors the reactants. At 25°C, the experimentally determined value of *Kw* in pure water is 1.0×10^{-14}.

$$Kw = [H_3O^+][OH^-] = 1.0\times10^{-14}$$

In a sample of pure water, the concentrations of hydronium and hydroxide ions are equal to one another. Pure water or any other aqueous solution in which this ratio holds is said to be *neutral*.

$$[H_3O^+] = [OH^-] = 1.0\times10^{-7}$$

The product of these two concentrations is 1.0×10^{-14}

$$[H_3O^+]\times[OH^-]=(1.0\times10^{-7})(1.0\times10^{-7})=1.0\times10^{-14}$$

- For acids, the concentration of H_3O^+ is greater than 1.0×10^{-7} M
- For bases, the concentration of OH^- is greater than 1.0×10^{-7} M.

No matter whether the aqueous solution is an acid, a base, or neutral the *Kw* remains constant.

- For acidic solutions, $[H_3O^+]$ is greater than $[OH^-]$.
- For basic solutions, $[OH^-]$ is greater than $[H_3O^+]$.
- For neutral solutions, $[H_3O^+] = [OH^-] = 1.0\times10^{-7}$M

NOTE: It is rare to truly have pure water. Water exposed to air will usually be slightly acidic because dissolved carbon dioxide gas, or carbonic acid, decreases the pH slightly below 7. Alternatively, dissolved minerals, like calcium carbonate (limestone), can make water slightly basic.

Knowing the amount of acid and base in solutions is extremely important for a wide variety of applications ranging from brewing beer or wine to studying the effects of river and creek water acidification to health and medicine. Scientists are good at calculating

and measuring the concentration of hydronium in a solution, however, there is a more convenient way to make comparisons between solutions – *the pH scale*.

The pH Scale

One qualitative measure of the strength of an acid or a base solution is the pH scale, which is based on the concentration of the hydronium in aqueous solution.

$$pH = -\log[H_3O^+]$$

Figure **8.7.1** illustrates this relationship, along with some examples of various solutions. Because hydronium ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$, the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the *negative* log of $[H_3O^+]$, which will give a positive value for pH. Note that $[H_3O^+]$ may be addressed as $[H^+]$.

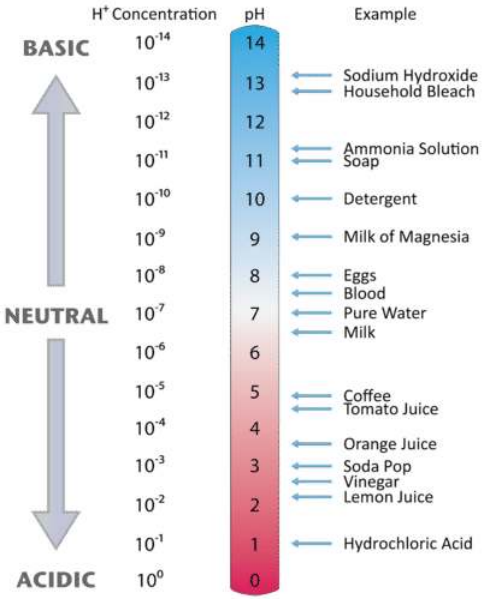


Figure **8.7.1**: The relationship between $[H^+]$ and ***pH*** *values for several common materials.*

A neutral (neither acidic nor basic) solution has a pH of 7. A pH below 7 means that a solution is acidic, with lower values of pH corresponding to increasingly acidic solutions. A pH greater than 7 indicates a basic solution, with higher values of pH corresponding to increasingly basic solutions. Thus, given the pH of several solutions, you can state which ones are acidic, which ones are basic, and which are more acidic or basic than others. These are summarized in table below.

Table **8.7.1**: Acidic, Basic and Neutral pH Values

pH scale

Classification	Relative Ion Concentrations	pH at 25 °C
acidic	$[H_3O^+] > [OH^-]$	pH < 7
neutral	$[H_3O^+] = [OH^-]$	pH = 7
basic	$[H_3O^+] < [OH^-]$	pH > 7

Example 8.7.1

Find the pH, given the $[H_3O^+]$ of the following:

a. 1×10^{-3} M
b. 2.5×10^{-11} M
c. 4.7×10^{-9} M

Solution

pH = - log $[H_3O^+]$

Substitute the known quantity into the equation and solve. Use a scientific calculator for b and c.

a. pH = - log $[\underline{1} \times 10^{-3}]$ = 3.$\underline{0}$ (1 decimal place since 1 has 1 significant figure)
b. pH = - log $[\underline{2.5} \times 10^{-11}]$ = 10.$\underline{60}$ (2 decimal places since 2.5 has 2 significant figures)
c. pH = - log $[\underline{4.7} \times 10^{-9}]$ = 8.$\underline{33}$ (2 decimal places since 4.7 has 2 significant figures)

Note on significant figures:

Because the number(s) before the decimal point in the pH value relate to the power on 10, the number of digits *after* the decimal point (underlined) is what determines the number of significant figures in the final answer.

8.7.1 Exercise

Find the pH, given $[H_3O^+]$ of the following:

a. 5.8×10^{-4} M
b. 1.0×10^{-7} M

Answer

a. 3.24

b. 7.00

Figure **8.7.2** lists the pH of several common solutions. The most *acidic* among the listed solutions is 1 M HCl with the lowest pH value (0.0): battery acid is the next most acidic solution with a pH value of 0.3. The most *basic* is 1M NaOH solution with the highest pH value of 14.0. Notice that some biological fluids (gastric juices and blood) are not neutral. You may also notice that many food products are slightly acidic. They are acidic because they contain solutions of weak acids. If the acid components of these foods were strong acids, the food would likely be inedible.

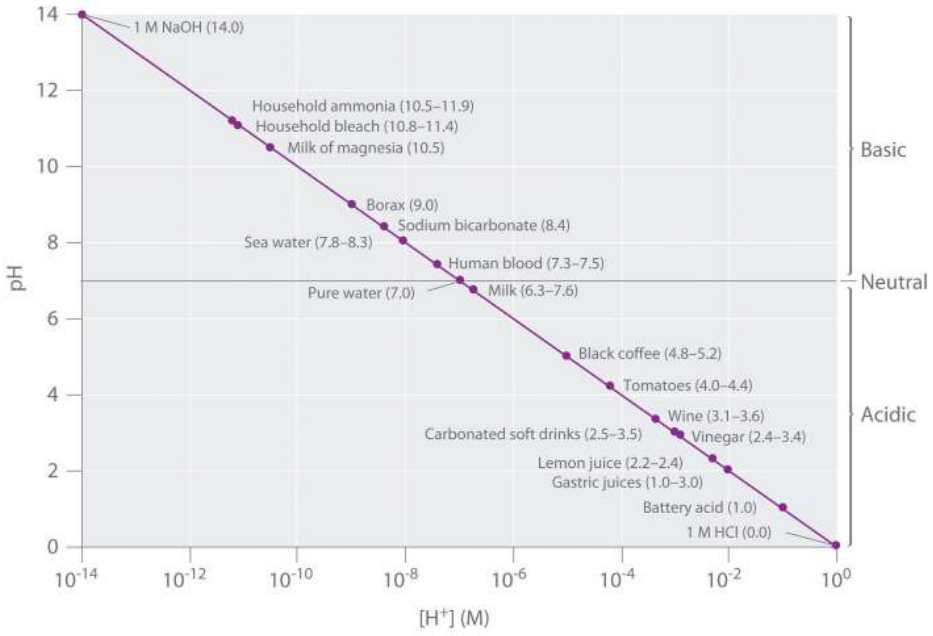


Figure **8.7.2**: A Plot of pH versus $[H^+]$ for Some Common Aqueous Solutions. Although many substances exist in a range of pH values (indicated in parentheses), they are plotted using typical values.

Example 8.7.2

Label each solution as acidic, basic, or neutral based only on the stated ***pH***.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Solution

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

8.7.2 Exercise

Identify each substance as acidic, basic, or neutral based only on the stated ***pH***.

a. human blood with ***pH*** = 7.4
b. household ammonia with ***pH*** = 11.0
c. cherries with ***pH*** = 3.6

Answer

a. slightly basic
b. basic
c. acidic

Measuring pH

Tools have been developed that make the measurement of pH simple and convenient (Figures 8.7.3). For example, pH paper consists of strips of paper impregnated with one or more acid–base indicators, which are intensely colored organic molecules whose colors change dramatically depending on the pH of the solution. Placing a drop of a solution on a strip of pH paper and comparing its color with standards give the solution's approximate pH. A more accurate tool, the pH meter, uses a glass electrode, a device whose voltage depends on the H^+ ion concentration (Figure **8.7.3**).



Figure **8.7.3**: Ways to measure pH. (left) These pH strips can measure pH in a series of ranges by putting sample on a strip and comparing its color change with colors on the box that correspond to a certain pH. The strip on the left measures pH 0-7 and shows results of a strong acid sample; the center strip is pH range 5-10 and shows results of a 6.97 buffer solution sample; the strip on the right measures a broad range (pH 1-14) and shows results of a 10% bleach water solution sample. This image can be enlarged by clicking on it. (right) These three meters and probes can measure pH (left) and electrical conductivity (center and right). Photos by Monica Bruckner (courtesy of the Science Education Resource Center (SERC) at Carleton College).

Acid Rain

Normal rainwater has a pH between 5 and 6 due to the presence of dissolved CO_2 which forms carbonic acid:

$$\mathbf{H_2O\,(l) + CO_2\,(g) \rightleftharpoons H_2CO_3\,(aq)} \tag{8.7.1}$$

$$H_2CO_3\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-\,(aq)$$

Acid rain is rainwater that has a pH of less than 5, due to a variety of nonmetal oxides, including CO_2, SO_2, SO_3, NO, and NO_2 being dissolved in the water and reacting with it to form not only carbonic acid, but sulfuric acid and nitric acid. The formation and subsequent ionization of sulfuric acid are shown here:

$$\mathbf{H_2O\,(l) + SO_3\,(g) \rightleftharpoons H_2SO_4\,(aq)} \tag{8.7.2}$$

$$H_2SO_4(aq) + H_2O\,(l) \rightarrow H_3O^+(aq) + HSO_4^-\,(aq)$$

Carbon dioxide is naturally present in the atmosphere because we and most other organisms produce it as a waste product of metabolism. Carbon dioxide is also formed when fires release carbon stored in vegetation or when we burn wood or fossil fuels. Sulfur trioxide in the atmosphere is naturally produced by volcanic activity, but it also stems from burning fossil fuels, which have traces of sulfur, and from the process of “roasting” ores of metal sulfides in metal-refining processes. Oxides of nitrogen are formed in internal combustion engines where the high temperatures make it possible for the nitrogen and oxygen in air to chemically combine.

Acid rain is a particular problem in industrial areas where the products of combustion and smelting are released into the air without being stripped of sulfur and nitrogen oxides. In North America and Europe until the 1980s, it was responsible for the destruction of forests and freshwater lakes, when the acidity of the rain actually killed trees, damaged soil, and made lakes uninhabitable for all but the most acid-tolerant species. Acid rain also corrodes statuary and building facades that are made of marble and limestone (Figure **8.7.4**). Regulations limiting the amount of sulfur and nitrogen oxides that can be released into the

atmosphere by industry and automobiles have reduced the severity of acid damage to both natural and manmade environments in North America and Europe. It is now a growing problem in industrial areas of China and India.

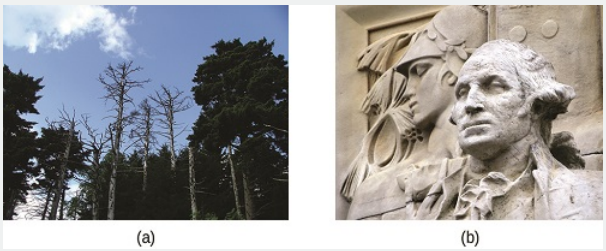


Figure 8.7.4: (a) Acid rain makes trees more susceptible to drought and insect infestation, and depletes nutrients in the soil. (b) It also is corrodes statues that are carved from marble or limestone. (credit a: modification of work by Chris M Morris; credit b: modification of work by “Eden, Janine and Jim”/Flickr)

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8.8: Le Châtelier’s Principle

Learning Objectives

- Define Le Chatelier's principle.
- Predict the direction of shift for an equilibrium under stress with changes in concentrations.

Once equilibrium is established, the reaction is over, right? Not exactly. An experimenter has some ability to affect the equilibrium.

Chemical equilibria can be shifted by changing the conditions that the system experiences. We say that we "stress" the equilibrium. When we stress the equilibrium, the chemical reaction is no longer at equilibrium, and the reaction starts to move back toward equilibrium in such a way as to decrease the stress. The formal statement is called **Le Châtelier's principle**: If an equilibrium is stressed, then the reaction shifts to reduce the stress.

There are several ways to stress an equilibrium. One way is to add or remove a product or a reactant in a chemical reaction at equilibrium.

When additional **reactant is added**, the **equilibrium** shifts to reduce this stress: it **makes more product**.

When additional **product is added**, the **equilibrium** shifts to reactants to reduce the stress: it **makes more reactant**.

If reactant or product is removed, the equilibrium shifts to make more product or reactant, respectively, to make up for the loss.

To **make more product, the rate of the forward reaction increases**. To **make more reactant, the rate of the reverse reaction increases**.

Example 8.8.1

Given this reaction at equilibrium:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

In which direction—toward reactants or toward products-—does the reaction shift if the equilibrium is stressed by each change?

1. H_2 is added.
2. NH_3 is added.
3. NH_3 is removed.

Solution

1. If H_2 is added, there is now more reactant, so the reaction will shift toward products to reduce the added H_2. (Rate of forward reaction increases/shifts right)
2. If NH_3 is added, there is now more product, so the reaction will shift toward reactants to reduce the added NH_3. (Rate of reverse reaction increases/shifts left)
3. If NH_3 is removed, there is now less product, so the reaction will shift toward products to replace the product removed. (Rate of forward reaction increases/shifts right)

Exercise 8.8.1

Given this reaction at equilibrium:

$$CO(g) + Br_2(g) \rightleftharpoons COBr_2(g)$$

In which direction—toward reactants or toward products—does the reaction shift if the equilibrium is stressed by each change?

1. Br_2 is removed.
2. $COBr_2$ is added.

Answers

1. toward reactants (Rate of reverse reaction increases/shifts left)

1. toward reactants (Rate of reverse reaction increases/shifts left)
2. toward reactants (Rate of reverse reaction increases/shifts left)

It is worth noting that when reactants or products are added or removed, *the value of the* K_{eq} *does not change*. The chemical reaction simply shifts, in a predictable fashion, to reestablish concentrations so that the K_{eq} expression reverts to the correct value.

A **catalyst** is a substance that increases the speed of a reaction. Overall, a catalyst is not a reactant and is not used up, but it still affects how fast a reaction proceeds by lowering the activation energy. When a catalyst is used in an equilibrium equation the lowered activation energy speeds up both the forward and reverse reactions. The overall result is that a catalyst has no effect on an equilibrium or on the value of K_{eq}.

Chemistry is Everywhere: Equilibria in the Garden

Hydrangeas are common flowering plants around the world. Although many hydrangeas are white, there is one common species (*Hydrangea macrophylla*) whose flowers can be either red or blue, as shown in the accompanying figure. How is it that a plant can have different colored flowers like this?



Figure **8.8.1** Garden Equilibria © Thinkstock. This species of hydrangea has flowers that can be either red or blue. Why the color difference?

Interestingly, the color of the flowers is due to the acidity of the soil that the hydrangea is planted in. An astute gardener can adjust the pH of the soil and actually change the color of the flowers. However, it is not the H^+ or OH^- ions that affect the color of the flowers. Rather, it is the presence of aluminum that causes the color change.

The solubility of aluminum in soil, and the ability of plants to absorb it, is dependent upon the acidity of the soil. If the soil is relatively acidic, the aluminum is more soluble, and plants can absorb it more easily. Under these conditions, hydrangea flowers are blue, as Al^{3+} ions interact with anthocyanin pigments in the plant. In more basic soils, aluminum is less soluble, and under these conditions the hydrangea flowers are red. Gardeners who change the pH of their soils to change the color of their hydrangea flowers are therefore employing Le Chatelier's principle: the amount of acid in the soil changes the equilibrium of aluminum solubility, which in turn affects the color of the flowers.

Key Takeaways

- Le Chatelier's principle addresses how an equilibrium shifts when the conditions of an equilibrium are changed.
- Catalysts do not affect the position of an equilibrium; they help reactions achieve equilibrium faster.

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8.9: Blood Buffers

Learning Objectives

- To define *buffer* and describe how it reacts with an acid or a base.
- Understand the pH of blood is kept in range by the carbonic acid and bicarbonate ion buffer system.

Buffers

Weak acids are relatively common, in the foods we eat. But we occasionally come across a strong acid, such as stomach acid, that has a strongly acidic pH of 1 to 2. By definition, strong acids and bases can produce a relatively large amount of hydronium or hydroxide ions and, as a consequence, have a marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(aq)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9. Such a drop in pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

The mechanism involves a **buffer**, a solution that resists dramatic changes in pH. A *buffer* (or *buffered*) solution is one that resists a large change in its pH when small amounts of H_3O^+ or OH^- ions are added or removed owing to a neutralizing reaction taking place in the same solution.

Composition of Buffers

Buffers contain either a **weak acid plus its conjugate base** or a **weak base plus its conjugate acid**. For example, a buffer can be composed of dissolved **acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate** ($NaC_2H_3O_2$). Sodium acetate is a salt that dissociates into sodium ions and acetate ions in solution. For as long as acetic acid and acetate ions are present in significant amounts a solution, this can resist dramatic pH changes. Another example of a buffer is a solution containing **ammonia (NH_3, a weak base) and ammonium chloride** (NH_4Cl). Ammonium chloride is also a salt that dissociates into ammonium ions and chloride ions in solution. The presence of ammonium ions with ammonia molecules satisfies the requisite condition for a buffer solution.

How Buffers Work

The essential component of a buffer system is a **conjugate acid-base pair** whose concentration is fairly high in relation to the concentrations of added H_3O^+ or OH^-. Let us use an acetic acid–sodium acetate buffer to demonstrate how buffers work. If a strong base, a source of OH^-(aq) ions is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$HC_2H_3O_{2(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(\ell)} + C_2H_3O^-_{2(aq)} \tag{8.9.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

> Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid, a source of H_3O^+ ions is added to the buffer solution, the H_3O^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if hydronium ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid.

$$H_3O^+(aq) + C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq) + H_2O(l)$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure 8.9.1 illustrates both actions of a buffer.

Figure 8.9.1 shows acetic acid HAc and its conjugate base, the acetate ion Ac^-. The idea is that this conjugate pair "pool" will be available to gobble up any small ($\leq 10^{-3}$ *M*) addition of H_3O^+ or OH^-. For simplicity in the diagram, H_3O^+ is addressed as H^+, acetic acid as HAc, and acetate ion as Ac^-.

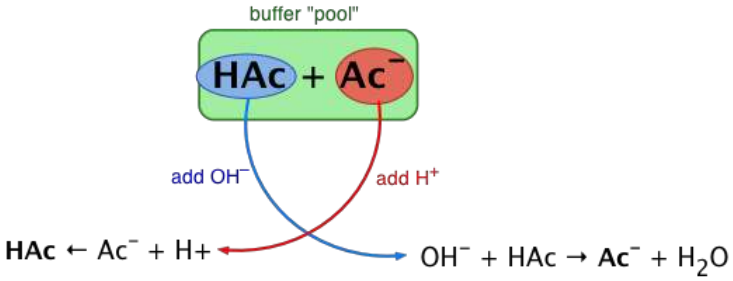


Figure **8.9.1**: How acetic acid (HAc)/acetate (Ac^-) buffer works.

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted and depleted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain capacity. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:

$$NH_{3(aq)} + H^+_{(aq)} \rightarrow NH^+_{4(aq)} \tag{8.9.2}$$

while the ammonium ion [NH_4^+(aq)] can react with any hydroxide ions introduced by strong bases:

$$NH^+_{4(aq)} + OH^-_{(aq)} \rightarrow NH_{3(aq)} + H_2O_{(\ell)} \tag{8.9.3}$$

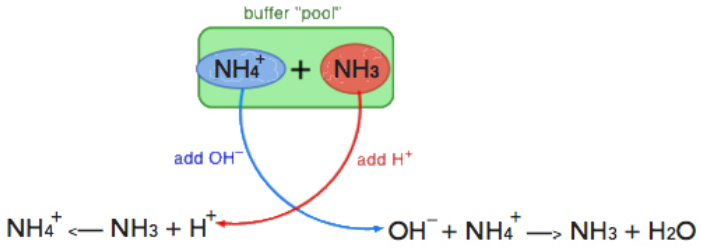


Figure **8.9.2**: How NH_4^+/NH_3 buffer works.

Buffer solutions are essential components of all living organisms.

- Our blood is buffered to maintain a pH of 7.35 to 7.45 that must remain unchanged as metabolically-generated CO_2 (carbon dioxide) is added and then removed by our lungs.
- Buffers in the oceans, in natural waters such as lakes and streams, and within soils help maintain their environmental stability against acid rain and increases in atmospheric CO_2.
- Many industrial processes, such as brewing, require buffer control, as do research studies in biochemistry and physiology that involve enzymes, are active only within certain pH ranges.

The pH in living systems is maintained by buffer systems.

Table **8.9.1**: pH in Living Systems

Compartment	pH
Gastric Acid	1
Lysosomes	4.5
Granules of Chromaffin Cells	5.5
Human Skin	5.5
Urine	6
Neutral H_2O at 37 °C	6.81
Cytosol	7.2
Cerebrospinal Fluid	7.3
Blood	7.35-7.45
Mitochondrial Matrix	7.5
Pancreas Secretions	8.1

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [H_2CO_3 is another way to write CO_2 dissolved in water]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on $H_2PO_4^-$ and HPO_4^{2-}.

Medicine: The Buffer System in Blood

The pH of the blood is maintained between 7.35 and 7.45 by buffers in the body. If blood pH goes above or below that range, it can destroy cells in the blood. The pH of the blood is kept constant by several buffer solutions. One important buffer solution is made of carbonic acid, H_2CO_3 (weak acid), and bicarbonate anion, HCO_3^- (conjugate base). The equilibrium equation is shown below:

$$H_2CO_3\,(aq) + H_2O\,(l) \rightleftharpoons HCO_3^-\,(aq) + H_3O^+\,(aq)$$

When acid (H_3O^+) enters the blood, it reacts with the HCO_3^- anion and forms water and carbonic acid as follows:

$$H_3O^+ + HCO_3^- \rightarrow H_2O + H_2CO_3$$

When base (OH^-) is added, it reacts with the carbonic acid in the buffer and produces bicarbonate ion and water as follows:

$$OH^- + H_2CO_3 \rightarrow HCO_3^- + H_2O$$

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 an excess of acetyl-CoA which is then converted to ketone bodies. This results in the decrease in blood pH. Buffer 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.

Lactic acid is produced in our muscles when we exercise. As the lactic acid enters the bloodstream, it is neutralized by the HCO_3^- ion, producing H_2CO_3. An enzyme then accelerates the breakdown of the excess carbonic acid to carbon dioxide and water, which can be eliminated by breathing. In fact, in addition to the regulating effects of the carbonate buffering system on the pH of blood, the body uses breathing to regulate blood pH. If the pH of the blood decreases too far, an increase in breathing removes CO_2 from the blood through the lungs driving the equilibrium reaction such that [H_3O^+] is lowered. If the blood is too alkaline, a lower breath rate increases CO_2 concentration in the blood, driving the equilibrium reaction the other way, increasing [H_3O^+] and restoring an appropriate pH.

Career Focus: Blood Bank Technology Specialist

At this point in this text, you should have the idea that the chemistry of blood is fairly complex. Because of this, people who work with blood must be specially trained to work with it properly.

A blood bank technology specialist is trained to perform routine and special tests on blood samples from blood banks or transfusion centers. This specialist measures the pH of blood, types it (according to the blood's ABO+/− type, Rh factors, and other typing schemes), tests it for the presence or absence of various diseases, and uses the blood to determine if a patient has any of several medical problems, such as anemia. A blood bank technology specialist may also interview and prepare donors to give blood and may actually collect the blood donation.

Blood bank technology specialists are well trained. Typically, they require a college degree with at least a year of special training in blood biology and chemistry. In the United States, training must conform to standards established by the American Association of Blood Banks.

Food and Drink App: The Acid That Eases Pain

Although medicines are not exactly "food and drink," we do ingest them, so let's take a look at an acid that is probably the most common medicine: acetylsalicylic acid, also known as aspirin. Aspirin is well known as a pain reliever, antipyretic (fever reducer) and anti-inflammatory.

The structure of aspirin is shown in the accompanying figure. The acid part is circled; it is the H atom in that part that can be donated as aspirin acts as a Brønsted-Lowry acid. Acetylsalicylic acid is a weak acid. Some people consume relatively large amounts of aspirin daily, its acidic nature can cause problems in the stomach lining, despite the stomach's defenses against its own stomach acid.

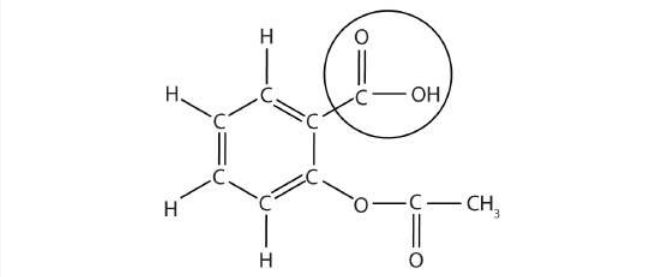


Figure 8.9.4 The Molecular Structure of Aspirin. The circled atoms are the acid part of the molecule.

Because the acid properties of aspirin may be problematic, many aspirin brands offer a "buffered aspirin" form of the medicine. In these cases, the aspirin also contains a buffering agent-usually MgO-that regulates the acidity of the aspirin to minimize its acidic side effects.

As useful and common as aspirin is, it was formally marketed as a drug starting in 1899. The US Food and Drug Administration (FDA), the governmental agency charged with overseeing and approving drugs in the United States, wasn't formed until 1906. Some have argued that if the FDA had been formed before aspirin was introduced, aspirin may never have gotten approval due to its potential for side effects-gastrointestinal bleeding, ringing in the ears, Reye's syndrome (a liver problem), and some allergic reactions. However, recently aspirin has been touted for its effects in lessening heart attacks and strokes, so it is likely that aspirin is here to stay.

Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

Concept Review Exercise

Explain how a buffer prevents large changes in pH.

Answer

A buffer has components that react with both strong acids and strong bases to resist sudden changes in pH.

Exercises

1. Describe a buffer. What two related chemical components are required to make a buffer?
2. Can a buffer be made by combining a strong acid with a strong base? Why or why not?
3. Which solute combinations can make a buffer? Assume all are aqueous solutions.
    1. HCl and NaCl
    2. HNO_2 and $NaNO_2$
    3. NH_4NO_3 and HNO_3
    4. NH_4NO_3 and NH_3
4. Which solute combinations can make a buffer? Assume all are aqueous solutions.
    1. H_3PO_4 and Na_3PO_4
    2. $NaHCO_3$ and Na_2CO_3
    3. $NaNO_3$ and $Ca(NO_3)_2$
    4. HN_3 and NH_3
5. For each combination in Exercise 3 that is a buffer, write the chemical equations for the reactions of the buffer components when a strong acid and a strong base is added.
6. For each combination in Exercise 4 that is a buffer, write the chemical equations for the reaction of the buffer components when a strong acid and a strong base is added.
7. The complete phosphate buffer system is based on four substances: H_3PO_4, $H_2PO_4^-$, HPO_4^{2-}, and PO_4^{3-}. What different buffer solutions can be made from these substances?
8. Explain why NaBr cannot be a component in either an acidic or a basic buffer.
9. Explain why $Mg(NO_3)_2$ cannot be a component in either an acidic or a basic buffer.

Answers

1. A buffer resists sudden changes in pH. It has a weak acid or base and a salt of that weak acid or base.
2. No. Combining a strong acid and a strong base will produce salt and water. Excess strong acid or strong base will not act as a buffer.
3. 1. not a buffer
    2. buffer
    3. not a buffer
    4. buffer
4. 1. not a buffer

    2. buffer

    3. not a buffer

    4. not buffer
5. 3b: strong acid: $H^+ + NO_2^- \rightarrow HNO_2$; strong base: $OH^- + HNO_2 \rightarrow H_2O + NO_2^-$; 3d: strong acid: $H^+ + NH_3 \rightarrow NH_4^+$; strong base: $OH^- + NH_4^+ \rightarrow H_2O + NH_3$
6. 4b: strong acid: $H^+ + CO_3^{2-} \rightarrow HCO3^-$; strong base: $OH^- + HCO_3^- \rightarrow H_2O + CO_3^{2-}$;
7. Buffers can be made by combining H_3PO_4 and $H_2PO_4^-$, $H_2PO_4^-$ and HPO_4^{2-}, and HPO_4^{2-} and PO_4^{3-}.
8. NaBr splits up into two ions in solution, Na^+ and Br^-. Na^+ will not react with any added base knowing that NaOH is a strong base. Br- will not react with any added acid knowing that HBr is a strong acid. Because NaBr will not react with any added base or acid, it does not resist change in pH and is not a buffer.
9. $Mg(NO_3)_2$ includes two types of ions, Mg^{2+} and NO_3^-. $Mg(OH)_2$ is strong base and completely dissociates (100% falls apart), so Mg^{2+} will not react with any added base (0% combines with OH^-). HNO_3 is strong acid and completely dissociates (100%

falls apart), so NO_3^- will not react with any added acid (0% combines with H^+). Because $Mg(NO_3)_2$ will not react with any added base or acid, it does not resist change in pH and is not a buffer.

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8.10: Compensation of pH via Organs

Learning Objectives

- Describe the role of lungs and the kidneys in maintaining a stable blood serum pH with the buffers

The pH of your blood normally ranges between 7.35 and 7.45. A blood pH below the normal range is called **acidosis.** A blood pH above this range is called **alkalosis.** Either one is potentially fatal. Why? Enzymes catalyze chemical reactions in the body. The enzymes are made of proteins which are chains of amino acids connected via peptide bonds. Some of the amino acids in the proteins contain acidic or basic functional groups. Changes in the pH of the blood serum can change the nature of the functional groups in the side chain through neutralization reactions. This in turn will change the intermolecular forces that affect protein folding or shape of the protein. The change in the shape of the protein will adversely affect the function of the enzyme or protein.

To maintain a constant serum pH the body relies on two approaches.

- One involves the use of buffers to minimize the changes in hydronium ion concentrations. This was discussed in the earlier section 8.9.
- The other is to control the hydronium ion concentrations by the involvement of the organs such as lungs and the kidneys.

Carbonic Acid and Bicarbonate Ion Buffer System

The most important buffer system in blood is formed from carbonic acid (H_2CO_3) and its conjugate base, the bicarbonate ion (HCO_3^-). Carbon dioxide (CO_2) is a product of metabolism. Some of the CO_2 dissolves in the blood serum to generate carbonic acid as shown in equation 1. The carbonic acid then acts as a Bronsted-Lowry acid to set up an equilibrium as shown in equation 2.

$$\mathrm{H_2O\,(l) + CO_2\,(g) \rightleftharpoons H_2CO_3\,(aq)} \quad (1)$$

$$\mathrm{H_2CO_3\,(aq) + H_2O\,(l) \rightleftharpoons HCO_3^-\,(aq) + H_3O^+\,(aq)} \quad (2)$$

Acidosis

During conditions of acidosis, when the pH falls below 7.35, the concentration of H_3O^+ increases, the equilibrium in equation 2 shifts left increasing the concentration of H_2CO_3. Equilibrium in equation 1 then shifts left as a result of increasing H_2CO_3 concentrations, increasing the concentration of CO_2. The CO_2 is then expelled out by **increased respiration** to lower CO_2. This is how the **lungs** compensate for acidosis.

The kidney play, a role during conditions of **acidosis,** when the pH falls below 7.35. As the concentration of H_3O^+ increases, the equilibrium in the equation shifts left, decreasing the concentration of HCO_3^- ion. The kidneys can then remove the excess H_3O^+ into the urine or add HCO_3^- ion into the blood stream, thus correcting the pH. This is how the kidneys compensate for acidosis.

The role of the organs to compensate for acidosis are summarized in figure 8.10.1 below.

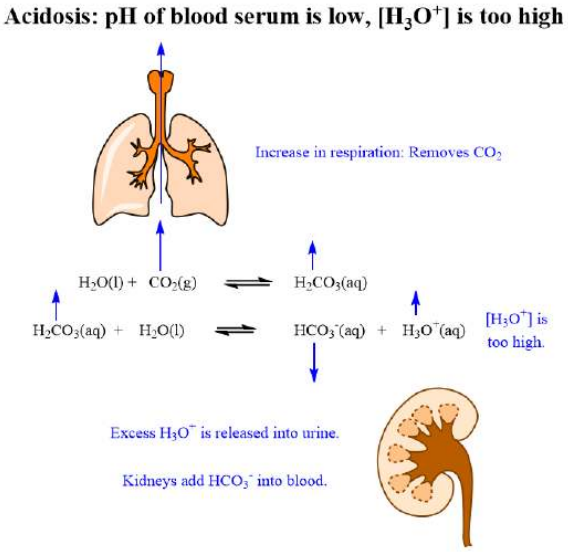


Figure **8.10.1**: Role of organs to compensate for acidosis.

Alkalosis

During conditions of alkalosis, when the pH increases above 7.45, the relative concentration of H_3O^+ decreases, the equilibrium in equation 2 shifts right decreasing the concentration of H_2CO_3. When the concentration of H_2CO_3 decreases the equilibrium in equation 1 shifts right decreasing the concentration of CO_2. The respiration is then decreased to conserve CO_2. This is compensation for alkalosis by the lungs.

$$\mathbf{H_2O\,(l) + CO_2\,(g) \rightleftharpoons H_2CO_3\,(aq)} \quad (1)$$

$$\mathbf{H_2CO_3\,(aq) + H_2O\,(l) \rightleftharpoons HCO_3^-\,(aq) + H_3O^+\,(aq)} \quad (2)$$

The kidney play, a role during conditions of **alkalosis,** when the pH is above 7.35. As the concentration of H_3O^+ decreases, the equilibrium in equation 2 shifts right increasing the concentration of HCO_3^- ion. The kidneys can then add H_3O^+ into the blood stream and remove excess the HCO_3^- ion through the urine, correcting the pH. This is compensation for alkalosis by the kidneys.

The role of the organs to compensate for alkalosis are summarized in figure 8.10.2 below.

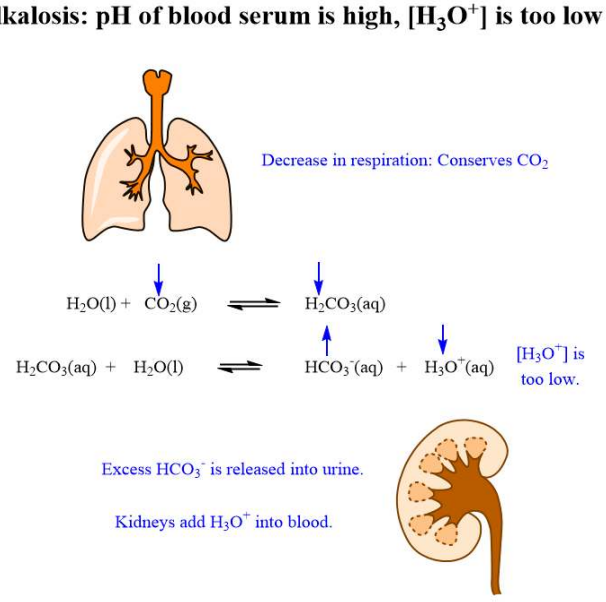


Figure **8.10.2**: Role of organs to compensate for alkalosis.

Dihydrogenphosphate and the Hydrogenphosphate Buffer System

The second buffer system is the dihydrogenphosphate ($H_2PO_4^-$) and its conjugate base the (HPO_4^{2-}). The mixture of anions $H_2PO_4^-$ / HPO_4^{2-} sets up the following acid-base equilibrium. Note the absence of gases in this equilibrium and the presence of three ions in the equilibrium.

$$\mathbf{H_2PO_4^-\,(aq) + H_2O\,(l) \rightleftharpoons HPO_4^{2-}\,(aq) + H_3O^+\,(aq)}$$

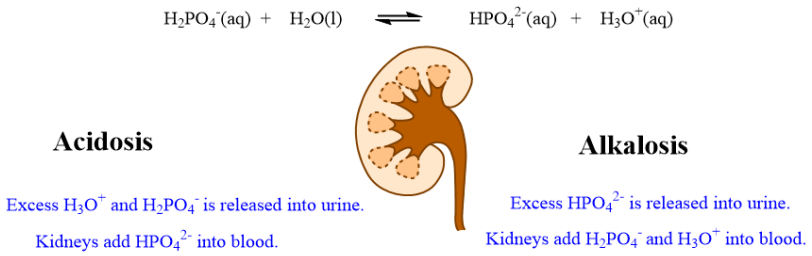
Acidosis

During conditions of acidosis, when the pH falls below 7.35, the concentration of H_3O^+ increases, the equilibrium in equation shifts left increasing the concentration of $H_2PO_4^-$ and decreasing the concentrations of HPO_4^{2-}. The kidneys remove excess H_3O^+ and $H_2PO_4^-$ via urine and add HPO_4^{2-} to the blood. This is another compensation for acidosis by the kidneys using a second buffer system.

Alkalosis

During conditions of **alkalosis**, when the pH is above 7.45, the relative concentration of H_3O^+ decreases, the equilibrium in the equation shifts right increasing the concentration of HPO_4^{2-} ion and decreasing the concentrations of $H_2PO_4^-$ ion. The kidneys can then add H_3O^+ and $H_2PO_4^-$ into the blood stream and remove excess the HPO_4^{2-} ion through the urine, correcting the pH. This is compensation for alkalosis by the kidneys.

The role of the kidneys to compensate for acidosis and alkalosis are summarized in the diagram below.



Note

Even with all these controls, the pH of blood can still sometimes fall outside the normal range. The acid/base disorders are classified as being either respiratory or metabolic in nature. Respiratory disorders occur when exhaling does not remove CO_2 from the blood at the same rate that it is produced in cells. Metabolic disorders arise from the inability to remove acids produced by metabolism, inability to control concentrations of HCO_3^-, ingestion of compounds that are acids or bases or compounds that can be converted to acids or bases. The following lists conditions that can lead to either acidosis or alkalosis.

Acidosis:

- Lung disease, asthma
- Excessive alcohol consumption
- Ketosis, ketoacidosis (starvation, diabetic conditions)
- Holding your breath too long
- Hypoventilation (not enough CO_2 is exhaled)

Alkalosis:

- Excessive use of antacids
- Anxiety, fever
- Hyperventilation (CO_2 is blown off faster than it is produced in the cells)

Mild cases of acidosis can lead to light-headedness. Severe cases can lead to coma and death. Symptoms of alkalosis include headaches, nervousness, cramps. In severe cases of alkalosis convulsions and death results.

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8.11: End of Chapter Problems

Neutralization Reaction

Write a double replacement reaction of

a. stomach acid with milk of magnesia

b. stomach acid with tums

c. vinegar and egg shell

d. hydrochloric acid and sodium hydroxide.

Acids and Bases

1. Give the formula of the following.

Phosphoric acid

Carbonic acid

Hydroiodic acid

Hydrofluoric acid

potassium hydroxide

Calcium bicarbonate

Barium hydroxide

Lithium carbonate

2. Give the name of the following:

HCl (aq)

HBr (aq)

HNO_3 (aq)

H_2SO_4 (aq)

NaOH

$KHCO_3$

Na_2CO_3

$CaCO_3$

Bronsted-Lowry Acids and Bases

1. What is the formula and name for the conjugate base of each of the following acids?

H_2CO_3

H_2SO_4

HCO_3^-

$H_2PO_4^-$

HSO_4^-

HPO_4^{2-}

H_3PO_4

HNO_3

2. What is the formula and name for the conjugate acid of each of the following bases?

F^-

NH_3

CN^-

NO_2^-

OH^-

CH_3NH_2

Cl^-

PO_4^{3-}

3. Identify the Bronsted-Lowry acids and bases for the forward and reverse reaction of each.

$F^-(aq) + HCl(aq) \rightleftharpoons HF(aq) + Cl^-(aq)$

$CH_3CO_2H(aq) + NO_3^-(aq) \rightleftharpoons CH_3CO_2^-(aq) + HNO_3(aq)$

$PO_4^{3-}(aq) + NH_4^+(aq) \rightleftharpoons NH_3(aq) + HPO_4^{2-}(aq)$

$HCN(aq) + H_2PO_4^-(aq) \rightleftharpoons CN^-(aq) + H_3PO_4(aq)$

$NH_4^+(aq) + SO_4^{2-}(aq) \rightleftharpoons NH_3(aq) + HSO_4^{2-}(aq)$

$CN^-(aq) + HI(aq) \rightleftharpoons HCN(aq) + I^-(aq)$

$OH^-(aq) + HNO_3(aq) \rightleftharpoons HOH(l) + NO_3^-(aq)$

$HPO_4^{2-}(aq) + CO_3^{2-}(aq) \rightleftharpoons PO_4^{3-}(aq) + HCO_3^-(aq)$

Equilibrium

1. True or False

a. The concentration of the reactants is always equal to the concentration of the products.

b. The reactants are not converted to products.

c. The rate of the forward reaction is equal to the rate of the reverse reaction.

d. The forward and reverse reactions stop.

e. The concentration of the reactants and products does not change.

f. The rate of the forward reaction is greater than the rate of the reverse reaction.

2. Balance the following chemical equations and write the equilibrium constant expression.

$CH_4(g) + H_2O(g) \rightleftharpoons H_2(g) + CO(g)$

$H_2(g) + N_2(g) \rightleftharpoons N_2H_4(g)$

$SO_2(g) + O_2(g) \rightleftharpoons SO_3(g)$

$CS_2(g) + H_2(g) \rightleftharpoons CH_4(g) + H_2S(g)$

$PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$

3. Balance the following chemical equations and write the equilibrium constant expression.

$C(s) + CO_2(g) \rightleftharpoons CO(g)$

$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

$H_2SO_4(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$

4. For each reaction, which is greater at equilibrium, the concentration of reactants or products?

Hint: Write an equilibrium constant expression and determine based on the value of the K_{eq} provided.

$HPO_4^{2-}(aq) + CN^-(aq) \rightleftharpoons PO_4^{3-}(aq) + HCN(aq)\ K_{eq} = 8.6 \times 10^{-4}$

$H_3PO_4(aq) + CN^-(aq) \rightleftharpoons H_2PO_4^-(aq) + HCN(aq)\ K_{eq} = 1.5 \times 10^{7}$

$HCO_3^-(aq) + HF(aq) \rightleftharpoons F^-(aq) + H_2CO_3(aq)\ K_{eq} = 1.5 \times 10^{3}$

$NH_4^+(aq) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq) + NH_3(aq)\ K_{eq} = 1.5 \times 10^{-3}$

Strong and Weak Acids

1. Based on the Ka value provided., indicate which acid is stronger of the pair? Which has a lower pH?

a. HClO, $K_a = 3.0 \times 10^{-8}$

b. $C_2O_4H^-$, $K_a = 6.4 \times 10^{-5}$

2. Based on the Ka value provided., indicate which acid is stronger of the pair? Which has a lower pH?

a. $C_2O_4H_2$, $K_a = 5.9 \times 10^{-2}$

b. $C_6O_2H_8$, $K_a = 1.7 \times 10^{-5}$

3. Based on the Ka value provided., indicate which acid is stronger of the pair? Which has a lower pH?

a. NH_4^+, $K_a = 5.6 \times 10^{-10}$

b. HPO_4^{2-}, $K_a = 4.2 \times 10^{-13}$

The pH Scale

Find the pH, given the $[H_3O^+]$ of the following. Label the solution as acidic, basic, or neutral.

a. 1×10^{-3} M

b. 2.5×10^{-11} M

c. 4.7×10^{-9} M

d. 5.8×10^{-4} M

e. 1.0×10^{-7} M

f. 3.9×10^{-2}

g. 7.0×10^{-5}

h. 6.8×10^{-7}

Le Chatelier's principle

1. For the equilibrium below, what is the effect (give shift) of

$S_2(g) + C(s) \rightleftharpoons CS_2(g)$

a. increasing $[CS_2]$

b. decreasing $[CS_2]$

c. decreasing $[S_2]$

2. The enzyme carbonic anhydrase catalyzes the conversion of CO_2 and H_2O into H_2CO_3.

$CO_2(g) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

Give shift for b thru f.

a. Write the equilibrium constant expression.

b. What is the effect of doubling the amount of carbonic anhydrase on the equilibrium?

c. What is the effect of increasing [CO_2]?

d. What is the effect of decreasing [CO_2]?

e. What is the effect of increasing [H_2CO_3]?

f. What is the effect of decreasing [H_2CO_3]?

3. The following reaction happens in step 1 of the citric acid cycle:

Oxaloacetate + Acetyl-CoA + H_2O ⇌ Citrate + CoA

Give shift for a thru c.

a. What is the effect of decreasing [oxaloacetate]?

b. What is the effect of increasing [oxaloacetate]?

c. What is the effect of increasing [Acetyl-CoA]?

4. The following reaction is from step 5 of glycolysis:

Dihydroxyacetone phosphate ⇌ Glyceraldehyde-3-phosphate

a. What is the effect of removing [Glyceraldehyde-3-phosphate]?

b. What is the effect of adding [Dihydroxyacetone phosphate]?

Buffers

1. What is a buffer solution? What is the composition of a buffer solution? Give an example.

2. Explain with chemical equations, how the H_2CO_3/HCO_3^- buffer system helps maintain blood serum at a constant pH.

3. What is the effect of changes in pH on an enzyme structure? Please use the picture of a partial protein strand to show the effect using chemical equations.

4. Severe cases of acidosis may lead to coma and death. Explain with equilibrium equations & Le Chatelier's principle how the lungs compensate for acidosis.

5. A patient has symptoms of light-headedness associated with acidosis. Explain how the kidney's compensate for acidosis. Give equilibrium equations and apply Le Chatelier's principle to explain your answer.

6. Due to excessive use of antacids a patient shows symptoms of alkalosis. Explain how the lungs try to compensate for alkalosis. Give equilibrium equations and apply Le Chatelier's principle to explain your answer.

7. Symptoms of alkalosis include nervousness and cramps. Explain with chemical equations & equilibrium principles how the kidneys help compensate for alkalosis.

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CHAPTER OVERVIEW

9: Proteins - An Introduction

9.1: Properties of Carboxylic Acids and Amines
9.2: Amino Acids
9.3: The Peptide Bond
9.4: Oxidation of Thiols
9.5: Protein Structure
9.6: Denaturation of Proteins
9.7: Digestion of Proteins
9.8: End of Chapter Problems

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9.1: Properties of Carboxylic Acids and Amines

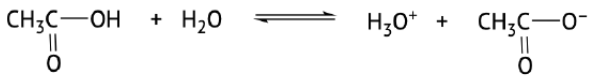
Learning Objectives

- Recognize the weak acid property of carboxylic acids.
- Recognize the weak base property of amines.

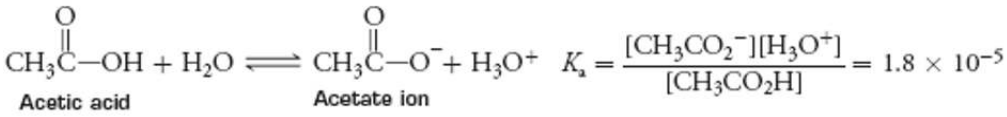
Amino acids are building blocks of proteins. All amino acids are made of two important functional groups. The functional groups are carboxylic acids and amines. We will start by looking at properties of each of these functional groups.

Carboxylic Acids are Weak Organic Acids

Carboxylic acids are named such because they tend to be more acidic than other functional groups in organic chemistry. In dilute aqueous solutions, they act as *weak acids* that partially dissociate to produce the corresponding **carboxylate anion** and hydronium cation (H_3O^+). Carboxylate anions are named by replacing the *-ic acid* ending from the carboxylic acid with *-ate*, see example below.

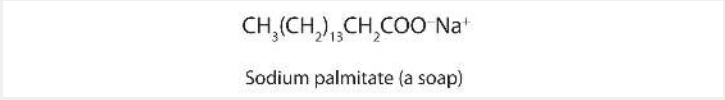


The extent of dissociation of these weak acids in water is described by K_a values. Remember that a compound with a smaller K_a value will be a weaker acid.



Note

The anion of the carboxylic acid is more soluble in water compared to the carboxylic acid. The salts of long-chain carboxylic acids are called soaps. The fatty acid palmitic acid is classified as hydrophobic. Sodium palmitate on the other hand is amphipathic.



Note To Your Health: Organic Salts as Preservatives

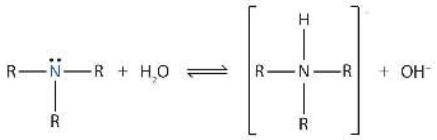
Some organic salts are used as preservatives in food products. They prevent spoilage by inhibiting the growth of bacteria and fungi. Calcium and sodium propionate, for example, are added to processed cheese and bakery goods; sodium benzoate is added to cider, jellies, pickles, and syrups; and sodium sorbate and potassium sorbate are added to fruit juices, sauerkraut, soft drinks, and wine. Look for them on ingredient labels the next time you shop for groceries.



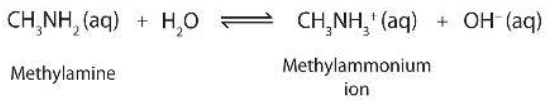
Amines as Weak Organic Bases

Ammonia (NH_3) acts as a base because the nitrogen atom has a lone pair of electrons that can accept a proton. Amines also have a lone electron pair on their nitrogen atoms and can accept a proton from water to form substituted ammonium (NH_4^+) ions and

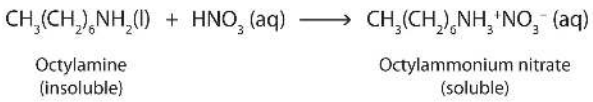
hydroxide (OH^-) ions:



As a specific example, methylamine reacts with water to form the methylammonium ion and the OH^- ion. Amines accept H^+ from water and form a 4° ammonium ion.



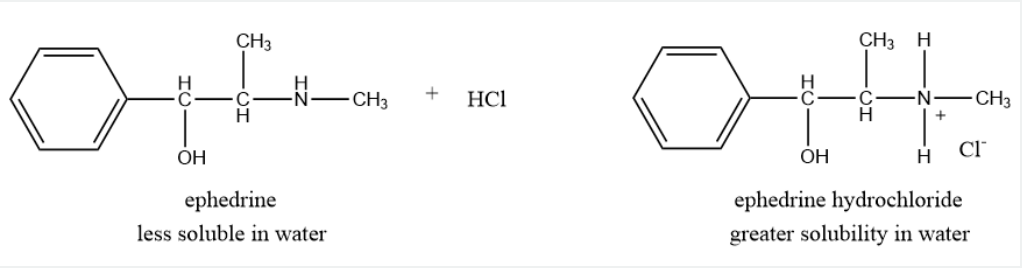
Nearly all amines, including those that are not very soluble in water, will react with strong acids to form salts soluble in water. Amines accept H^+ from acids form a 4° ammonium ion.



Amine salts are named like other salts: the name of the cation is followed by the name of the anion. The 4° ammonium salts have a greater water solubility compared to amines.

Note To Your Health: Amine Salts as Drugs

Salts of amines are ionic and their solubility are those characteristic of salts. Many drugs that are amines are converted to salts to increase their solubility in aqueous solution.The decongestant ephedrine is sold as ephedrine hydrochloride or ephedrine sulfate. The amine salt drugs are easier to administer due to their greater water solubility.



ephedrine
less soluble in water

ephedrine hydrochloride
greater solubility in water

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9.2: Amino Acids

Learning Objectives

- To recognize amino acids and classify them based on the characteristics of their side chains.

The proteins in all living species, from bacteria to humans, are polymers constructed from the same set of 20 amino acids, so called because each contains an amino group attached to a carboxylic acid. The amino acids in proteins are α-amino acids, which means the amino group is attached to the α-carbon of the carboxylic acid. Humans can synthesize only about half of the needed amino acids; the remainder must be obtained from the diet and are known as essential amino acids.

Amino Acid Structure

Every amino acid contains an amino group, ($–NH_2$), a carboxyl group, (–COOH), and a **side chain** or R group, which are all attached to the alpha-carbon (the one directly bonded to the carboxyl functional group). Therefore, amino acids are commonly called alpha-amino acids. The figure 9.2.1 shows the structure of a generic α-amino acid.

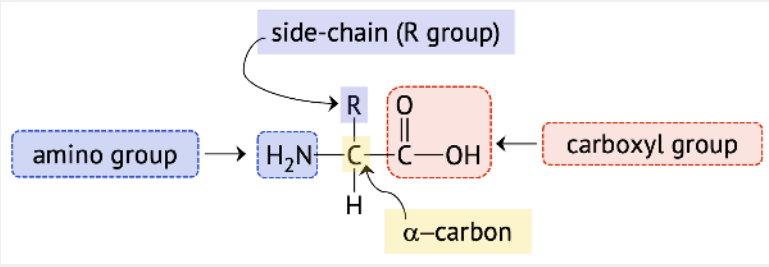


Figure **9.2.1**: An α-amino acid. The three parts to an amino acid include the amino group, the carboxyl group, and the side-chain or R group all attached to the α-carbon. The R group is the part that distinguishes one amino acid from the next.

Recall from earlier that a carboxylic acid is a weak acid (H^+ ion donor) and an amine is a weak base (H^+ ion acceptor). Both functional groups are present in the same molecule and so the structure of the amino acid allows it to act as both an acid and a base in the same molecule. The properties of amino acids reflect this acid-base property. The amino acids are soluble in water, colorless, nonvolatile, crystalline solids, melting and decomposing at temperatures above 200°C. These melting temperatures are more like those of inorganic salts than those of amines or organic acids and indicate that the structures of the amino acids in the solid state and in neutral solution are best represented as having both a negatively charged group and a positively charged group. Such a species is known as a **zwitterion**.

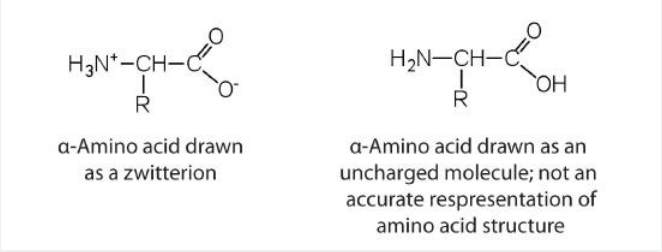


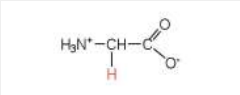
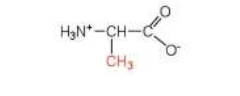
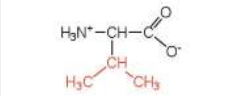
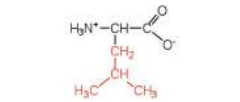
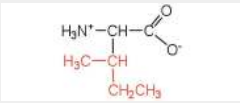
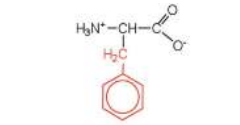
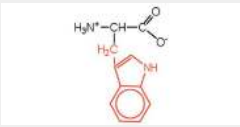
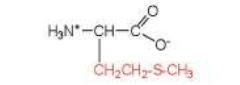
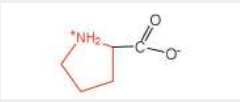
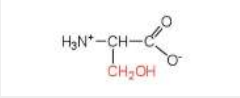
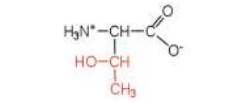
Figure **9.2.2**: Zwitterion vs Uncharged Molecule.

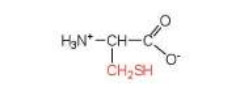
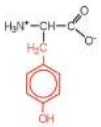
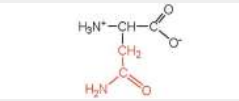
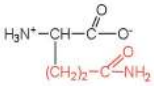
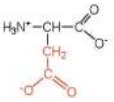
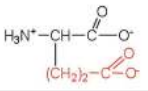
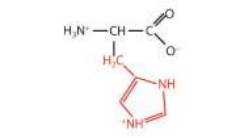
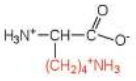
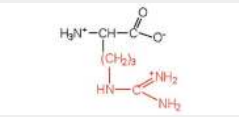
Classification

In addition to the amino and carboxyl groups, amino acids have a side chain or R group attached to the α-carbon. Each amino acid has unique characteristics arising from the size, shape, solubility, and ionization properties of its R group. As a result, the side chains of amino acids exert a profound effect on the structure and biological activity of proteins. Although amino acids can be classified in various ways, one common approach is to classify them according to whether the functional group on the side chain at neutral pH is nonpolar, polar neutral, negatively charged, or positively charged. The nonpolar and the polar neutral amino acids exist as zwitterions at physiological pH. Amino acids are overall negatively charged when they have a carboxylic acid in their sidechains and are also called polar acidic amino acids. Amino acids are overall positively charged when they have an amine functional group in their side chain and are addressed as polar basic.

The structures and names of the 20 amino acids, their one- and three-letter abbreviations, and some of their distinctive features are given in Table **9.2.1**.

Table **9.2.1**: Common Amino Acids Found in Proteins

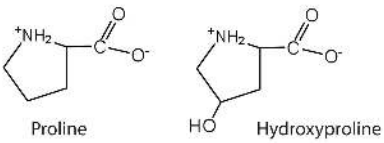
Common Name	Abbreviation	Structural Formula at physiological pH	Molecular Weight (amu)	Distinctive Feature
Amino acids with a nonpolar side chain (nonpolar amino acids)				
glycine	gly (G)		75	–
alanine	ala (A)		89	—
valine	val (V)		117	a branched-chain amino acid
leucine	leu (L)		131	a branched-chain amino acid
isoleucine	ile (I)		131	an essential amino acid because most animals cannot synthesize branched-chain amino acids
phenylalanine	phe (F)		165	also classified as an aromatic amino acid
tryptophan	trp (W)		204	also classified as an aromatic amino acid
methionine	met (M)		149	side chain has a sulfide functional group and functions as a methyl group donor
proline	pro (P)		115	contains a secondary amine group
Amino acids with a polar but neutral side chain (polar neutral amino acids)				
serine	ser (S)		105	found at the active site of many enzymes
threonine	thr (T)		119	named for its similarity to the sugar threose

Common Name	Abbreviation	Structural Formula at physiological pH	Molecular Weight (amu)	Distinctive Feature
cysteine	cys (C)		121	oxidation of two cysteine molecules yields disulfide bridge in proteins
tyrosine	tyr (Y)		181	also classified as an aromatic amino acid
asparagine	asn (N)		132	contains the amide of aspartic acid
glutamine	gln (Q)		146	contains the amide of glutamic acid
Amino acids with a negatively charged side chain (polar acidic/polar negative amino acids)				
aspartic acid	asp (D)		132	carboxylic acid functional groups are ionized at physiological pH; also known as aspartate
glutamic acid	glu (E)		146	carboxylic acid functional groups are ionized at physiological pH; also known as glutamate
Amino acids with a positively charged side chain (polar basic/polar basic amino acids)				
histidine	his (H)		155	at physiological pH the side chain exists as a cation
lysine	lys (K)		147	at physiological pH the side chain exists as a cation
arginine	arg (R)		175	at physiological pH the side chain exists as a cation

Note: interesting Facts About Amino Acids

The first amino acid to be isolated was asparagine in 1806. It was obtained from protein found in asparagus juice (hence the name). Glycine, the major amino acid found in gelatin, was named for its sweet taste (Greek *glykys*, meaning “sweet”). Glutamic acid is named as such because it was first isolated from gluten. The crystalline salt of glutamic acid is called monosodium glutamate (MSG), which naturally occurs in some foods but is also added to as a savory or "umami" flavor

enhancer. In some cases an amino acid found in a protein is actually a derivative of one of the common 20 amino acids (one such derivative is hydroxyproline). The modification occurs *after* the amino acid has been assembled into a protein.



Summary

Amino acids can be classified based on the characteristics of their distinctive side chains as nonpolar, polar neutral, negatively charged, or positively charged.

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9.3: The Peptide Bond

Learning Outcomes

- Explain how a peptide bond is formed from individual amino acids.
- Identify the N-terminus and C-terminus of a polypeptide or protein.

Our body has an intricate mechanism for the manufacture of proteins. Scientists have to use 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.

The Peptide Bond

A **peptide bond** is formed by a combination of amino acids in which the amine group of one amino acid has undergone a reaction with the carboxylic acid of another amino acid. The reaction is a dehydration-condensation reaction, forming an amide group (CO–NH), shown below. The result is elimination of a water molecule and condensation of two amino acids. The molecule formed is called a dipeptide.

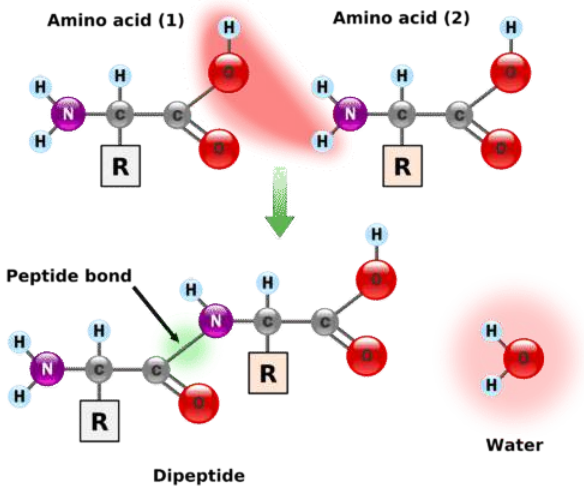


Figure 9.3.1: Amino acids undergo condensation to form a molecule called a dipeptide. The C–N bond is called a peptide bond.

A chain consisting of only two amino acid units is called a *dipeptide*; a chain consisting of three is a *tripeptide*. By convention, peptide and protein structures are depicted with the amino acid whose amino group is free (the N-terminal end) on the left and the amino acid with a free carboxyl group (the C-terminal end) to the right as shown in figure 9.3.2.

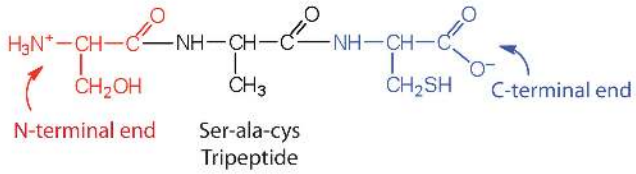


Figure 9.3.2: A tripeptide showing the two terminus.

The sequence of amino acids in a chain is called an amino acid sequence. 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 and Ala-Trp-Gly are isomers (figure 9.3.3) and have different biological functions and properties.

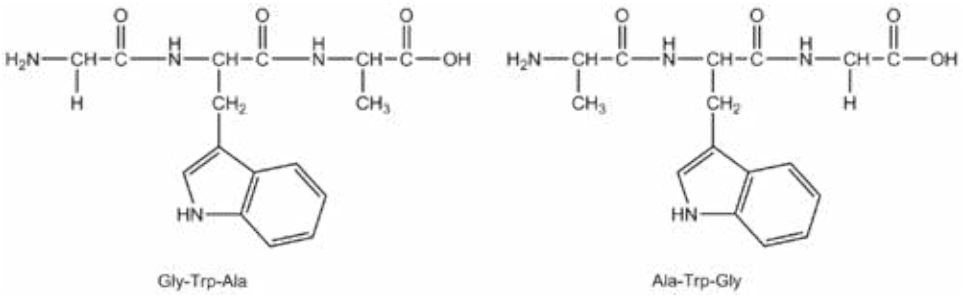


Figure **9.3.3**: Two tripeptides that are isomers of each other.

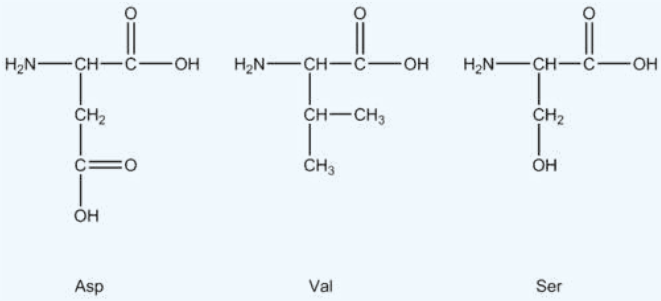
The general term *polypeptide* refers to an amino acid chain of unspecified length. However, chains of about 50 amino acids or more are usually called proteins. In its physiologically active form, a protein may be composed of one or more polypeptide chains.

Example 9.3.1

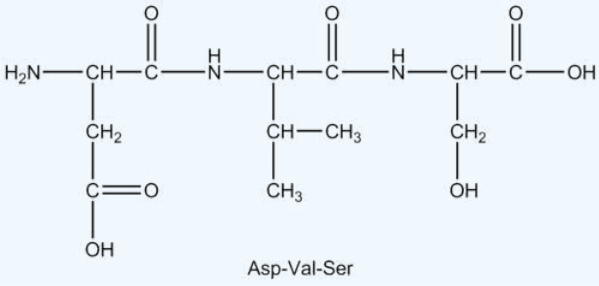
Draw the polypeptide Asp-Val-Ser.

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.



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.



Example 9.3.2

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

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9.4: Oxidation of Thiols

Learning Objectives

- The oxidation of thiols gives disulfides.

Sulfur is in the same group (6A) of the periodic table as oxygen, the two elements have some similar properties. We might expect sulfur to form organic compounds related to those of oxygen, and indeed it does. Thiols, which are sulfur analogs of alcohols, have the general formula RSH. Methanethiol, has the formula CH_3SH. Ethanethiol is the most common odorant for liquid propane (LP) gas.

The oxidation of thiols are however very different from alcohols. Primary alcohols are oxidized to aldehydes and secondary alcohols are oxidized to ketones. Tertiary alcohols cannot be oxidized. The oxidation of thiols gives compounds called disulfides. Two thiol molecules are required for the oxidation to form the disulfide. The loss of a hydrogen atom by each thiol is evidence that oxidation has taken place. The biological oxidizing agent NAD^+ removes the hydrogen atoms as shown in figure 9.4.1 with ethanethiol.

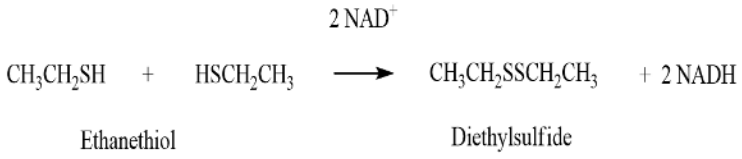


Figure **9.4.1**: Oxidation of Thiols.

The amino acid cysteine [$HSCH_2CH(NH_2)COOH$] contains the thiol functional group which can be oxidized as shown in figure 9.4.2. Disulfide linkages (–S–S–) between protein chains are extremely important in protein structure.

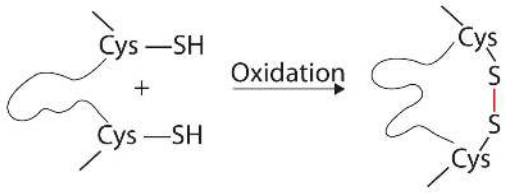


Figure **9.4.2**: Disulfide Bridge

Thioethers or sulfides, are sulfur analogs of ethers, have the form general formula RSR′. An example is dimethylsulfide (CH_3SCH_3), which is responsible for the sometimes unpleasant odor of cooking cabbage and related vegetables. Note that methionine [$CH_3SCH_2CH_2CH(NH_2)COOH$] has a thioether functional group.

Summary

The oxidation of two thiol functional groups leads to the formation of the disulfide functional group. Disulfide linkages (–S–S–) between protein chains are extremely important in tertiary structure of the protein.

Concept Review Exercises

1. What is the functional group of a thiol? Write the condensed structural formula for ethanethiol.
2. What is the functional group of a disulfide? Write the condensed structural formula for dipropyl disulfide.

Answers

1. SH; CH_3CH_2SH
2. –S–S–; $CH_3CH_2CH_2SSCH_2CH_2CH_3$

Exercises

1. A common natural gas odorant is *tert*-butyl mercaptan. What is its condensed structural formula?
2. Write the equation for the oxidation of ethanethiol to diethyl disulfide.

Answer

1. $(CH_3)_3CSH$

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9.5: Protein Structure

Learning Objectives

- Describe the four levels of protein structure.
- Identify the types of attractive interactions that hold proteins in their most stable three-dimensional structure.

Each of the thousands of naturally occurring proteins has its own characteristic amino acid composition and sequence that result in a unique three-dimensional shape. Proteins are compounds of high molecular mass consisting largely or entirely of chains of amino acids. Hemoglobin is a transport protein found in the red blood cells. It transports O_2 from the lungs to the cells. Collagen is a structural protein found in tendon, cartilage, bone, blood vessels, and teeth. It make up one-third of the total protein in the human body. Collagen is strong, water insoluble and resistant to chemical change.

Levels of Protein Structure

The structure of proteins is generally described as having four organizational levels.

Primary Structure

The first of these is the primary structure, which is the sequence of amino acids in a protein's polypeptide chain or chains, beginning with the free amino group and maintained by the peptide bonds connecting each amino acid to the next.

For peptides and proteins to be physiologically active, it is not enough that they incorporate certain amounts of specific amino acids. The order, or *sequence*, in which the amino acids are connected is also of critical importance. Bradykinin is a nine-amino acid peptide produced in the blood that has the following amino acid sequence:

arg-pro-pro-gly-phe-ser-pro-phe-arg

This peptide lowers blood pressure, stimulates smooth muscle tissue, increases capillary permeability, and causes pain. When the order of amino acids in bradykinin is reversed,

arg-phe-pro-ser-phe-gly-pro-pro-arg

the peptide resulting from this synthesis shows none of the activity of bradykinin.

The primary structure of insulin, composed of 51 amino acids, is shown in Figure **9.5.1**.

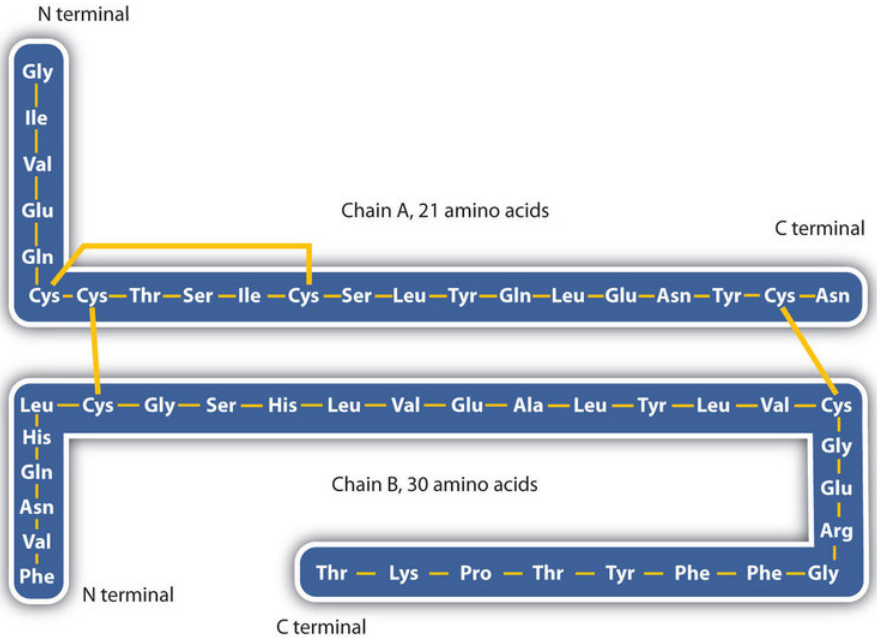


Figure **9.5.1**: Primary Structure of Human Insulin. Human insulin, whose amino acid sequence is shown here, is a hormone that is required for the proper metabolism of glucose.

Just as millions of different words are spelled with our 26-letter English alphabet, millions of different proteins are made with the 20 common amino acids. However, just as the English alphabet can be used to write gibberish, amino acids can be put together in the *wrong sequence* to produce nonfunctional proteins. Although the correct sequence is ordinarily of utmost importance, it is not always absolutely required. Just as you can sometimes make sense of incorrectly spelled English words, a protein with a small percentage of "incorrect" amino acids may continue to function. However, it rarely functions as well as a protein having the correct sequence. There are also instances in which seemingly minor errors of sequence have disastrous effects. For example, in some people, every molecule of hemoglobin (a protein in the blood that transports oxygen) has a single incorrect amino acid unit out of about 300 (a single valine replaces a glutamic acid). That "minor" error is responsible for sickle cell anemia, an inherited condition that usually is fatal.

Secondary Structure

A protein molecule is not a random tangle of polypeptide chains. Instead, the chains are arranged in unique but specific conformations. The term secondary structure refers to the fixed arrangement of the polypeptide backbone. On the basis of X ray studies, Linus Pauling and Robert Corey postulated that certain proteins or portions of proteins twist into a spiral or a helix. This helix is stabilized by *intrachain* hydrogen bonding between the carbonyl oxygen atom of one amino acid and the amide hydrogen atom four amino acids up the chain (located on the next turn of the helix) and is known as a right-handed **α-helix**. X ray data indicate that this helix makes one turn for every 3.6 amino acids, and the side chains of these amino acids project outward from the coiled backbone (Figure **9.5.2**). The α-keratins, found in hair and wool, are exclusively α-helical in conformation. Some proteins, such as gamma globulin, chymotrypsin, and cytochrome c, have little or no helical structure. Others, such as hemoglobin and myoglobin, are helical in certain regions but not in others.

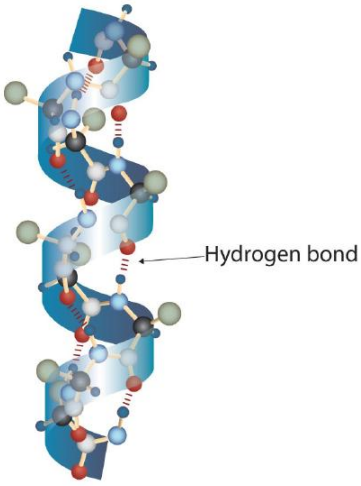


Figure **9.5.2** A Ball-and-Stick Model of an α-Helix. This ball-and-stick model shows the intrachain hydrogen bonding between carbonyl oxygen atoms and amide hydrogen atoms. Each turn of the helix spans 3.6 amino acids. Note that the side chains (represented as green spheres) point out from the helix.

Another common type of secondary structure, called the ***β-sheet*** *conformation*, is a sheetlike arrangement in which two or more extended polypeptide chains (or separate regions on the same chain) are aligned side by side. The aligned segments can run either parallel or antiparallel—that is, the N-terminals can face in the same direction on adjacent chains or in different directions—and are connected by *interchain* hydrogen bonding (Figure **9.5.3**). The β-sheet is particularly important in structural proteins, such as silk fibroin. It is also seen in portions of many enzymes, such as carboxypeptidase A and lysozyme.

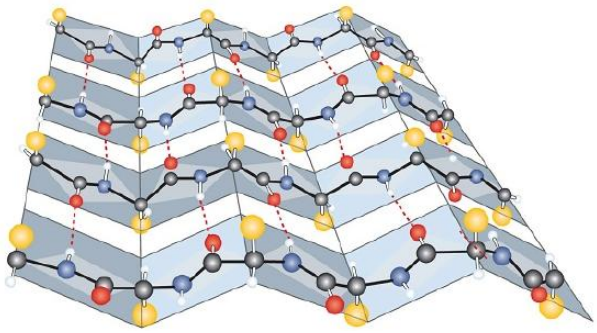


Figure 9.5.3: A Ball-and-Stick Model of the β-Pleated Sheet Structure in Proteins. The side chains extend above or below the sheet and alternate along the chain. The protein chains are held together by interchain hydrogen bonding.

Tertiary Structure

Tertiary structure refers to the unique three-dimensional shape of the protein as a whole, which results from the folding and bending of the protein backbone. The tertiary structure is intimately tied to the proper biochemical functioning of the protein. This unique folding conformation leads to a native (biologically acitive) molecule. Figure 9.5.4 shows a depiction of the three-dimensional structure of insulin.

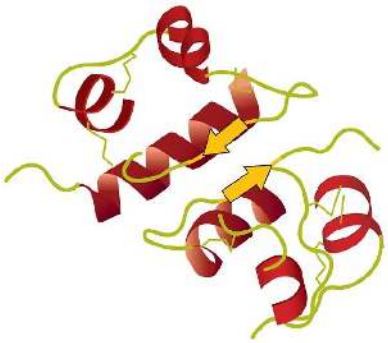
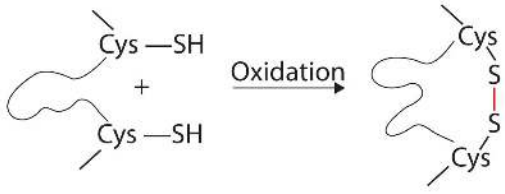


Figure 9.5.4: A Ribbon Model of the Three-Dimensional Structure of Insulin. The spiral regions represent sections of the polypeptide chain that have an α-helical structure, while the broad arrows represent β-pleated sheet structures.

Four major types of attractive interactions determine the shape and stability of the tertiary structure of proteins.

1. **Ionic bonding or Salt Bridges:** Ionic bonds result from electrostatic attractions between positively and negatively charged side chains of amino acids. For example, the mutual attraction between an aspartic acid carboxylate ion and a lysine ammonium ion helps to maintain a particular folded area of a protein (part (a) of Figure 9.5.5).
2. **Hydrogen Bonding:** Hydrogen bonding forms between a highly electronegative oxygen atom or a nitrogen atom and a hydrogen atom attached to another oxygen atom or a nitrogen atom, such as those found in polar amino acid side chains shown below in part (b) of Figure 9.5.5.
3. **Disulfide Bonds:** Two cysteine amino acid units may be brought close together as the protein molecule folds. Subsequent oxidation of the thiol functional group and linkage of the sulfur atoms leads to the formation of intrachain disulfide linkages part (c) of Figure 9.5.5. The disulfide bridges have a strong stabilizing effect on the tertiary structure.



4. **London's forces:** These forces arise between nonpolar parts of a molecule. London's forces are weak but can be important when other types of interactions are either missing or minimal part (d) of Figure 9.5.5). This is the case with fibroin, the major

protein in silk, in which a high proportion of amino acids in the protein have nonpolar side chains. The term *hydrophobic interaction* is often used as a synonym for dispersion forces. Because nonpolar groups cannot engage in hydrogen bonding, the protein folds in such a way that these groups are buried in the interior part of the protein structure, minimizing their contact with water. This folding rule is "nonpolar in, polar out".

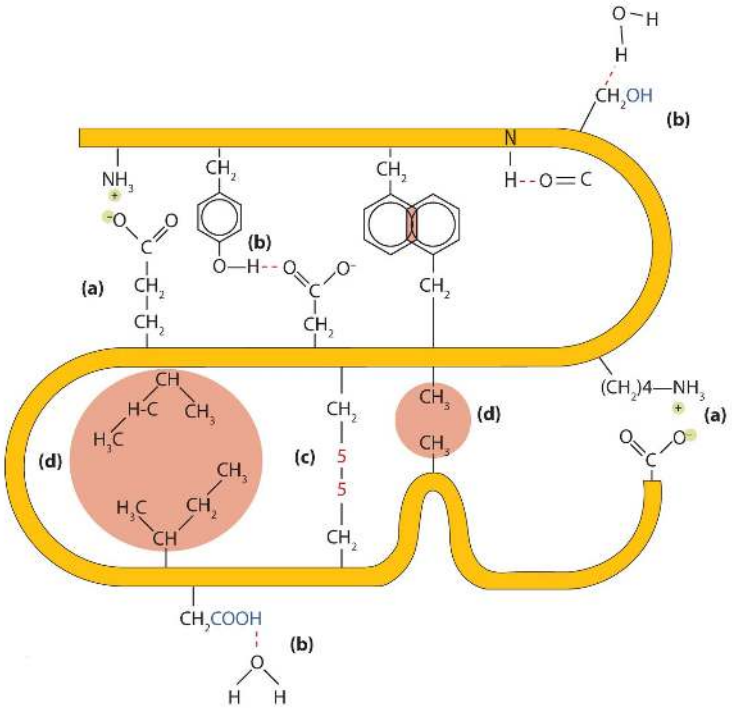


Figure **9.5.5**: Tertiary Protein Structure Interactions. Four interactions stabilize the tertiary structure of a protein: (a) ionic bonding (salt bridges), (b) hydrogen bonding, (c) disulfide linkages, and (d) dispersion forces.

A **prosthetic group** is a nonpeptide component tightly bound to a polypeptide chain. The **heme** group in hemoglobin is an example of a prosthetic group. Proteins that require a prosthetic group for biological activity are called **complex proteins** or **conjugated proteins**. Simple proteins do not require a prosthetic group. At the center of heme is an iron(II) ion (Fe^{2+}) which binds to an oxygen molecule.

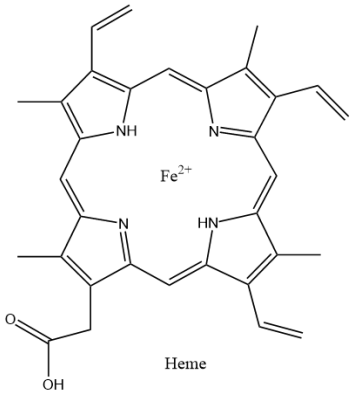


Figure **9.5.6**:The heme is the prosthetic group part of the protein hemoglobin.

Quaternary Structure

When a protein contains more than one polypeptide chain, each chain is called a *subunit*. The arrangement of multiple subunits represents a fourth level of structure, the quaternary structure of a protein. Hemoglobin, with four polypeptide chains or subunits, is a good example of a protein having quaternary structure as shown in Figure **9.5.7**. The quaternary structure of a protein is produced and stabilized by the same kinds of interactions that maintain the tertiary structure such as the salt bridges, disulfide bonds, and

IMF's such as hydrogen bonding and hydrophobic effect. Since there are four subunits in hemoglobin, there are four heme prosthetic groups with four Fe^{2+} ions. Each subunit can bind to an oxygen molecule (O_2). So hemoglobin can bind to and carry four oxygen molecules.

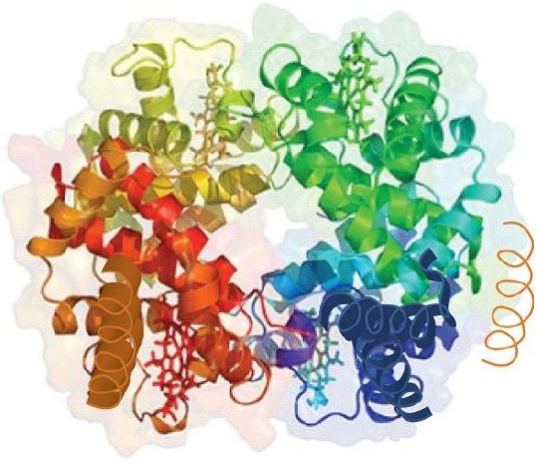


Figure **9.5.7** The Quaternary Structure of Hemoglobin. Hemoglobin is a protein that transports oxygen throughout the body.

Source: Image from the RCSB PDB (www.pdb.org(opens in new window)) of PDB ID 1I3D (R.D. Kidd, H.M. Baker, A.J. Mathews, T. Brittain, E.N. Baker (2001) Oligomerization and ligand binding in a homotetrameric hemoglobin: two high-resolution crystal structures of hemoglobin Bart's (gamma(4)), a marker for alpha-thalassemia. Protein Sci. 1739–1749).

A schematic representation of the four levels of protein structure is in Figure **9.5.8**.

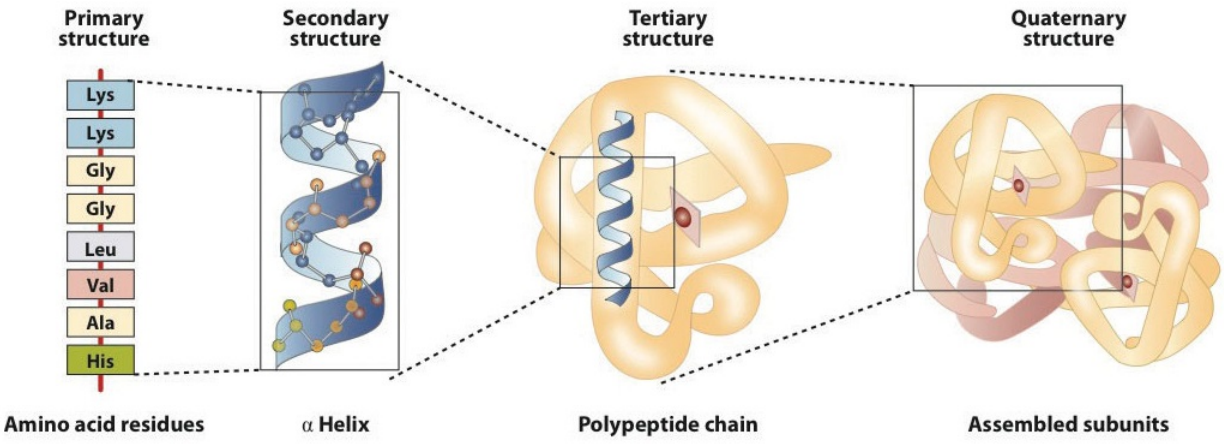


Figure **9.5.8**: Levels of Structure in Proteins

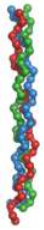
The *primary structure* consists of the specific amino acid sequence. The resulting peptide chain can twist into an α-helix, which is one type of *secondary structure*. This helical segment is incorporated into the *tertiary structure* of the folded polypeptide chain. The single polypeptide chain is a subunit that constitutes the *quaternary structure* of a protein, such as hemoglobin that has four polypeptide chains.

Fibrous and Globular Proteins

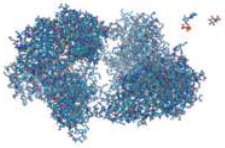
Once proteins form and have developed all levels of their structure, they can be classified as either fibrous or globular. The two major structural classifications of proteins are based on general qualities: whether the protein is (1) fiber-like and insoluble or (2) globular and soluble. Some proteins, such as those that compose hair, skin, muscles, and connective tissue, are fiber-like.

The fibrous proteins are insoluble in water and usually serve structural, connective, and protective functions. Examples of fibrous proteins are keratins, collagens, myosins, and elastins. Hair and the outer layer of skin are composed of keratin. Connective tissues

contain collagen. Myosins are muscle proteins and are capable of contraction and extension. Elastins are found in ligaments and the elastic tissue of artery walls.



Globular proteins, the other major class, are soluble in aqueous media. In these proteins, the chains are folded so that the molecule as a whole is roughly spherical. Familiar examples include egg albumin from egg whites and serum albumin in blood. Serum albumin plays a major role in transporting fatty acids and maintaining a proper balance of osmotic pressures in the body. Hemoglobin and myoglobin, which are important for binding oxygen, are also globular proteins.



Summary

Proteins can be divided into two categories: fibrous, which tend to be insoluble in water, and globular, which are more soluble in water. A protein may have up to four levels of structure. The primary structure consists of the specific amino acid sequence. The resulting peptide chain can form an α-helix or β-pleated sheet (or local structures not as easily categorized), which is known as secondary structure. These segments of secondary structure are incorporated into the tertiary structure of the folded polypeptide chain. The quaternary structure describes the arrangements of subunits in a protein that contains more than one subunit. Four major types of attractive interactions determine the shape and stability of the folded protein: ionic bonding, hydrogen bonding, disulfide linkages, and dispersion forces.

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9.6: Denaturation of Proteins

Learning Objectives

- Define denaturation of proteins.
- Identify ways in which proteins are denatured.

Protein Denaturation

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. Denaturation often leads to the precipitation or coagulation of a protein. Anyone who has fried an egg has observed denaturation. The clear egg white turns opaque as the albumin denatures and coagulates. This precess cannot be reversed.

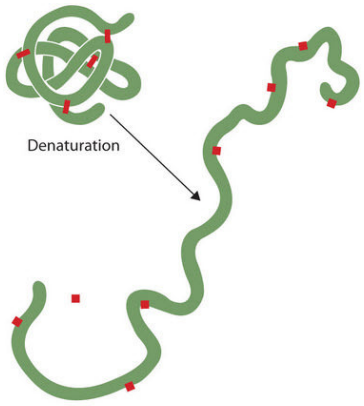


Figure **9.6.1**: Denaturation of a Protein.

There are a variety of ways to denature proteins including those below.

- Heat above 50°C
- Treatment with acids and bases (changes in pH)
- Exposure to Alcohol
- Concentrated Brine Solution (i.e. NaCl)
- Reaction with Reducing agents and Oxidizing agents
- Soaps and Detergents
- Heavy metal ions such as Ag^{+}, Cu^{2+}, Pb^{2+} and Hg^{2+}
- Agitation

The table 9.6.1 lists the denaturing agents and the disruption of chemical bonds or IMF's.

Table **9.6.1**: Protein Denaturation Methods

Method	Effect on Protein Structure
Heat above 50°C, ultraviolet (UV) radiation, or agitation	Heat, UV radiation, Agitation supplies kinetic energy to protein molecules, causing their atoms to vibrate more rapidly and disrupting IMF's such as hydrogen bonding and Londons dispersion forces.
Acids or Bases	Acids react with amino acid side chains that are negatively charged and bases react with the amino acid side chains that are positively charged resulting in the disruption of the salt bridges.
Use of organic compounds, such as isopropyl alcohol (rubbing alcohol)	These compounds are capable of engaging in intermolecular hydrogen bonding with protein molecules, disrupting intramolecular hydrogen bonding within the protein.
Salts of heavy metal ions, such as mercury, copper, silver, and lead ions	These ions form strong bonds with the carboxylate anions of aspartic acid and glutamic acid or SH groups of cysteine, disrupting ionic bonds and disulfide linkages.
Soaps and Detergents	These agents disrupt the hydrophobic interactions in the protein since soaps and detergents are amphipathic.

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.

The primary structures of proteins are quite sturdy. In general, fairly vigorous conditions are needed to hydrolyze peptide bonds. So the 1° structure is not disrupted on denaturation. At the secondary, tertiary, and quaternary levels, however, proteins are quite vulnerable to denaturation, 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.

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9.7: Digestion of Proteins

Learning Objectives

- List the steps of protein digestion.

Protein digestion begins in the stomach (Figure **9.7.1**), where the action of gastric juice hydrolyzes about 10% of the peptide bonds. Gastric juice is a mixture of water (more than 99%), inorganic ions, hydrochloric acid, and various enzymes and other proteins.

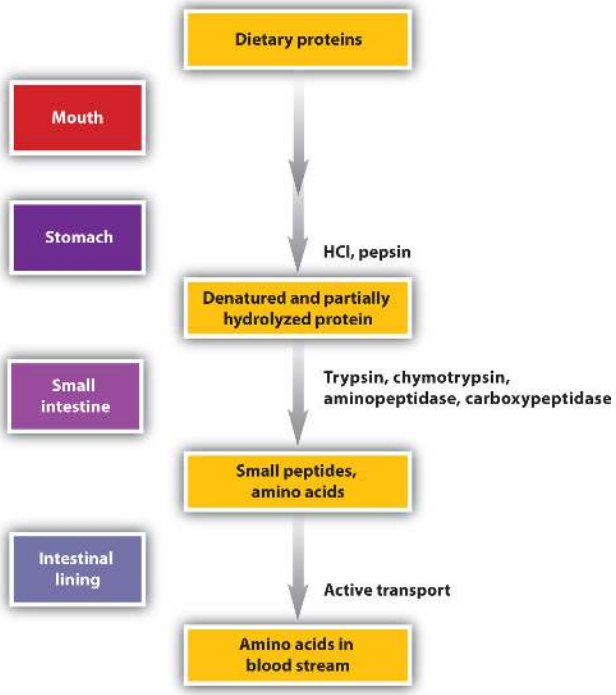


Figure **9.7.1**: The Principal Events and Sites of Protein Digestion

The hydrochloric acid (HCl) in gastric juice is secreted by glands in the stomach lining. The pH of freshly secreted gastric juice is about 1.0, but the contents of the stomach may raise the pH to between 1.5 and 2.5. HCl(aq) helps to denature food proteins; that is, it unfolds the protein molecules to expose their chains to more efficient enzyme action. The principal digestive enzyme precursor of gastric juice is pepsinogen, an inactive enzyme produced in cells located in the stomach wall. When food enters the stomach after a period of fasting, pepsinogen is converted to its active form—pepsin—in a series of steps initiated by the drop in pH. Pepsin catalyzes the hydrolysis of peptide linkages within protein molecules. It has a fairly broad specificity but acts preferentially on linkages involving the aromatic amino acids tryptophan, tyrosine, and phenylalanine, as well as methionine and leucine.

Protein digestion is completed in the small intestine. Pancreatic juice, carried from the pancreas via the pancreatic duct, contains inactive enzymes such as trypsinogen and chymotrypsinogen. They are activated in the small intestine as follows (Figure **9.7.2**): The intestinal mucosal cells secrete the proteolytic enzyme enteropeptidase, which converts trypsinogen to trypsin; trypsin then activates chymotrypsinogen to chymotrypsin (and also completes the activation of trypsinogen). Both of these active enzymes catalyze the hydrolysis of peptide bonds in protein chains. Chymotrypsin preferentially attacks peptide bonds involving the carboxyl groups of the aromatic amino acids (phenylalanine, tryptophan, and tyrosine). Trypsin attacks peptide bonds involving the carboxyl groups of the basic amino acids (lysine and arginine). Pancreatic juice also contains procarboxypeptidase, which is cleaved by trypsin to carboxypeptidase. The latter is an enzyme that catalyzes the hydrolysis of peptide linkages at the free carboxyl end of the peptide chain, resulting in the stepwise liberation of free amino acids from the carboxyl end of the polypeptide.

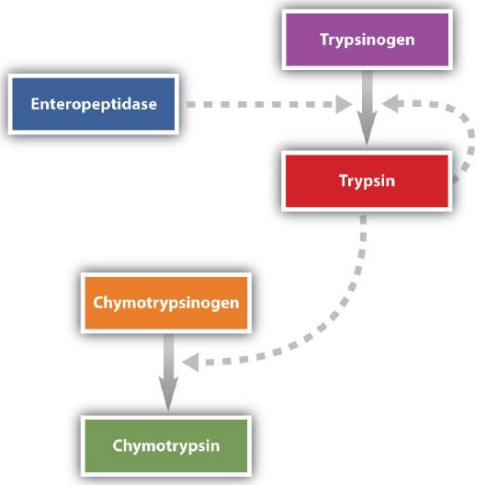


Figure **9.7.2**: Activation of Some Pancreatic Enzymes in the Small Intestine

Aminopeptidases in the intestinal juice remove amino acids from the N-terminal end of peptides and proteins possessing a free amino group. Figure **9.7.3** illustrates the specificity of these protein-digesting enzymes. The amino acids that are released by protein digestion are absorbed across the intestinal wall into the circulatory system, where they can be used for protein synthesis.

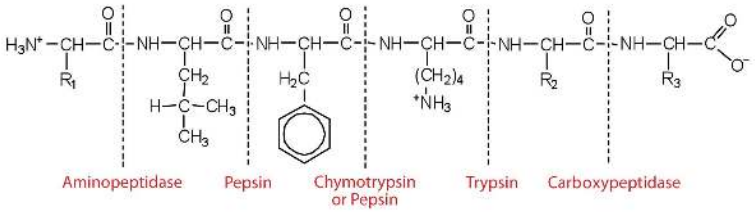


Figure **9.7.3**: Hydrolysis of a Peptide by Several Peptidases

This diagram illustrates where in a peptide the different peptidases we have discussed would catalyze hydrolysis the peptide bonds.

Summary

During digestion, the peptide bonds in proteins undergo hydrolysis to amino acids. The digestion process of proteins begins in the stomach and continues in the small intestine.

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9.8: End of Chapter Problems

Amino acids

1. For the four amino acids below:

Ile

Arg

Asn

Glu

a. What is the net charge of each amino acid at physiological pH?

b. What is the classification of each amino acid?

2. Identify an amino acid that has in its side chain a (an) ----------.

a. thiol

b. phenol

c. sulfide

d. alcohol

e. amide

f. aromatic ring

3. Which nonpolar amino acid has a side chain that can participate in hydrogen bonding?

Polypeptides

1. Draw the tripeptide Asp-Val-Ser. Label the terminuses in this tripeptide.

2. List all of the possible polypeptides that can be formed from the three amino acids, cysteine (Cys), leucine (Leu), and arginine (Arg).

3. Which of the 20 amino acids can participate in a disulfide bridge? What level of protein structure involves this covalent bonds between two sulfur atoms?

Protein structure

1. What is the primary (1°) structure of a protein

2. Describe the two types of secondary (2°) structure of proteins.

3. Describe the tertiary structure of a protein. What are the chemical bonds or IMF's responsible for maintaining the 3° structure?

4. Describe the quaternary structure of hemoglobin. What are the chemical bonds or IMF's responsible for maintaining the 4° structure?

5. Name pairs of amino acids that have side chains that can participate in salt bridge formation at pH=7?

6. a. Draw the polypeptide fragment shown below in its line bond formula or condensed structural formula at the physiological pH = 7.

Leu-Ala-Trp-Phe-Gly

b. Label C-terminus & N-terminus of the polypeptide fragment. Circle the peptide bonds.

c. In an aqueous environment, will the peptide fragment more likely be buried inside or located on the surface? Why?

d. What is most likely to change the biological activity: Replacement of the Leu with Ile residue or the replacement of Leu with Glu? Please discuss why.

7. What distinguishes a protein:

a. that has a 4° structure from one that does not?

b. a simple protein form a conjugated (complex) protein?

Denaturation

1. What is denaturation? List some of the ways to denaturation a protein.

2. Discuss what happens when a native protein is subjected to the following agents? Identify the IMF's or chemical bonds that are disrupted.

a. Alcohol

b. Soaps and detergents

c. Changes in pH

d. Heavy metal ions such as Pb^{2+}

e. Heat or Agitation

3. True or False: A denatured globular protein ----------.

a. contains no peptide bond

b. has an abnormal 1° structure

c. is biologically inactive

Digestion of Proteins

1. What part of the digestive system is involved in the digestion of proteins?

2. Name the functional group hydrolyzed when a protein undergoes digestion.

3. Name an enzyme that catalyzes the digestion of proteins.

4. What products are formed when a protein undergoes digestion?

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CHAPTER OVERVIEW

10: DNA and RNA - An Introduction

10.1: Nucleotides - The Building Blocks of Nucleic Acids
10.2: Polynucleotides
10.3: Nucleic Acid Structure
10.4: Replication of DNA
10.5: Transcription
10.6: Translation and the Genetic Code
10.7: Mutations and Genetic Diseases
10.8: End of Chapter Problems

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10.1: Nucleotides - The Building Blocks of Nucleic Acids

Learning Objectives

- Describe the makeup of nucleotides that are building blocks of nucleic acids DNA and RNA.

Nucleic acids

Nucleic acids are biopolymers that carry the codes for the synthesis of proteins and pass them on from generation to generation, i.e., they are genetic materials. In other words, nucleic acids are the instruction manual for biochemical reactions taking place in living things.

Nucleotides

Nucleotides are the building blocks, i.e., the repeat units or monomers of nucleic acids.

Nucleotides are composed of three sub-units:

1. a 5-carbon carbohydrate,
2. a base that is an aromatic compound containing nitrogen, and
3. an anion of phosphoric acid, i.e., phosphate (PO_4^{3-}).

5-Carbon Monosaccharide

There are two 5-carbon carbohydrates i.e., ribose and deoxyribose, found in nucleic acids, as shown in the figure 10.1.1 below. Both are in furanose (five-membered cyclic) form with hydroxyl ($-OH$) group at anomeric (C#1') in a β configuration. The number labels have prime symbols on them, i.e. 1', 2', 3' etc. to distinguish them from regular numbers 1, 2, 3, etc that are used for the nitrogen bases of the nucleic acids. The only difference between ribose and deoxyribose is that hydroxy group at C#2' is missing in the latter, which gives the deoxy prefix to its name.

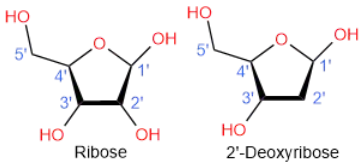


Figure 10.1.1: The two monosaccharides in nucleic acids.

Type of nucleic acids

There are two types of nucleic acids, one contains ribose in its nucleotides is called **ribonucleic acid** (**RNA**) and the other contains deoxyribose in its nucleotides is called **deoxyribonucleic acid** (**DNA**).

Nitrogen bases

Nitrogen bases in nucleic acids are derivatives of two aromatic compounds; purine and pyrimidine, shown in figure 10.1.2.

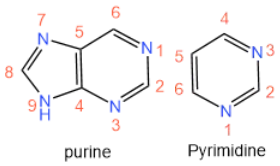


Figure 10.1.2: Purine and Pyrimidine.

Cyclic compounds that contain atoms other than C's in the cycle are called heterocyclic compounds. Purine and pyrimidine are heterocyclic aromatic compounds because they contain N's in the cycles. Purine is bicyclic containing a six-membered ring with

two N's fused with a five-membered ring with two N's. Pyrimidine is a six-membered cycle with two N's. Both purine and pyrimidine are planar molecules like benzene.

There are five nitrogen bases found in nucleic acids: two are purines, i.e., adenine (A) and guanine (G), and three are pyrimidines, i.e., cytosine (C), thymine (T), and uracil (U), as shown in figure 10.1.3.

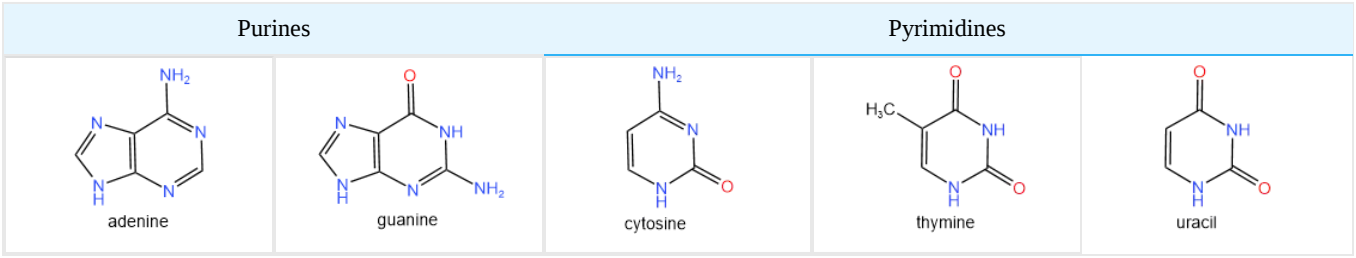


Figure 10.1.3: The five bases in nucleic acids.

DNA contains four nitrogen bases, i.e., adenine, guanine, cytosine, and thymine. RNA also contains four nitrogen bases i.e., adenine, guanine, cytosine, and uracil. Note that the first three, i.e., adenine, guanine, and cytosine are common in DNA and RNA, but the fourth, i.e, thymine in DNA is replaced with uracil in RNA.

Nucleosides

When a monomer sugar like ribose or deoxyribose reacts with an amine, the –OH group at C#1' is replaced with a N of the amine. The product is called N-glycoside, and the C–N bond in the N-glycoside is called an **N-glycosidic bond**, as shown in the following example.

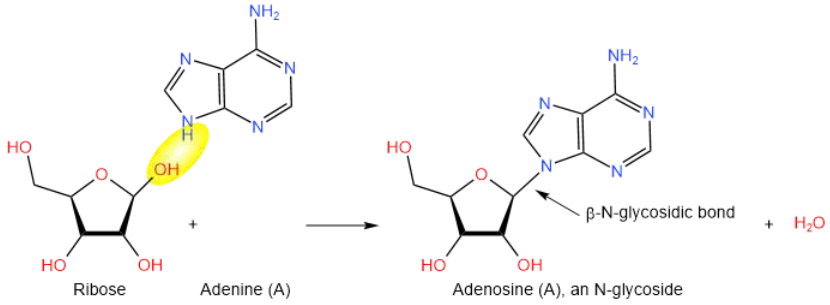


Figure 10.1.4:Purines and pyrimidines connect with the anomeric carbon of the sugar to form the N-glycoside.

The N-glycosides are named by using the name of purine but ending with -osine e.g., adenine becomes adenosine and guanine becomes gunosine, or by using the name of pyrimidine ending with -idine, e.g., cytosine becomes cytidine and uracil become uridine. One letter abbreviation of the nitrogen bases remains the same for the corresponding nucleoside, e.g., adenine (A) and adenosine (A). The four nucleosides found in RNA are shown in figure 10.1.5 with their names and one-letter abbreviations.

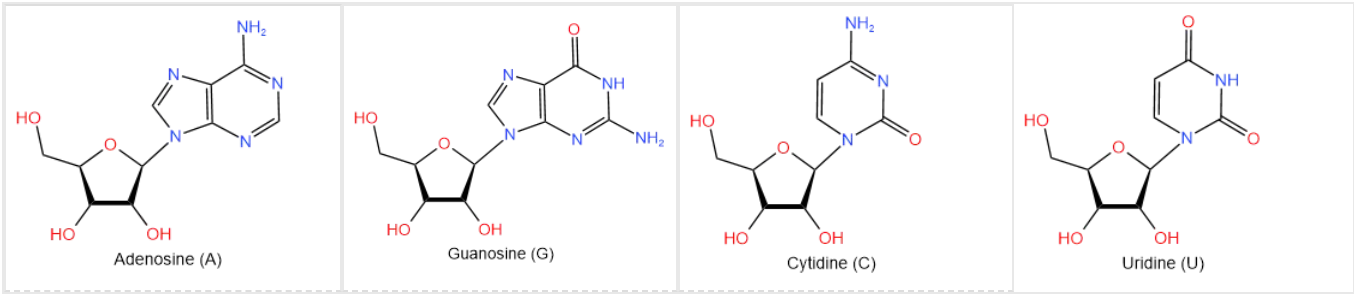


Figure 10.1.5: The four nucleosides of RNA.

In DNA there is deoxyribose in place of ribose. So the names of nucleosides found in DNA begin with the deoxy- prefix. The four nucleosides found in DNA are shown in figure 10.1.6 with their names and one-letter abbreviations.

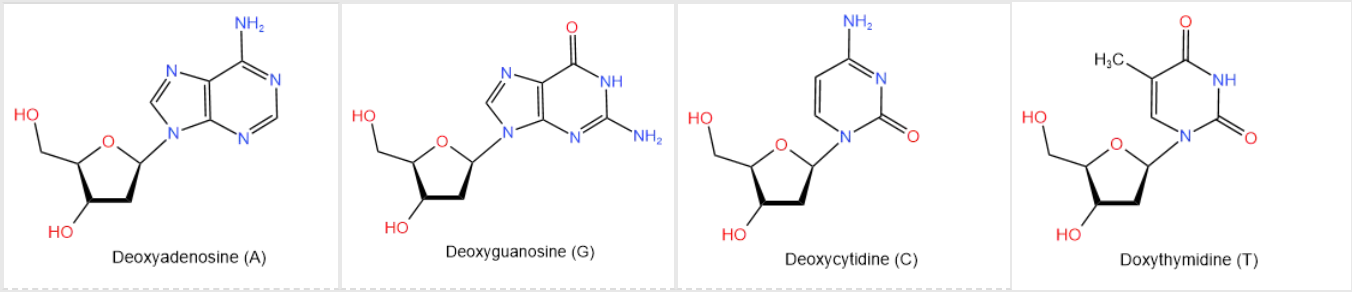
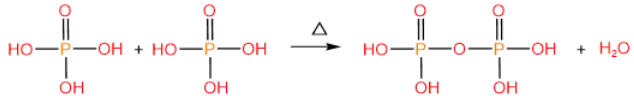


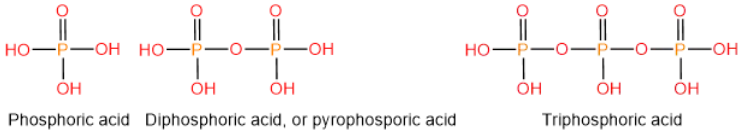
Figure **10.1.6**: The four nucleosides of DNA.

Phosphate

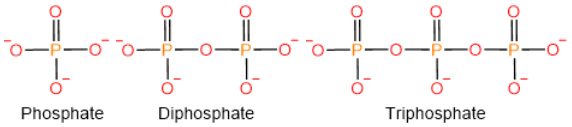
Phosphoric acid (H_3PO_4) is an oxyacid with three acidic protons, i.e., it is triprotic acid. Two phosphoric acids condense and form diphosphoric acid, with the elimination of water.



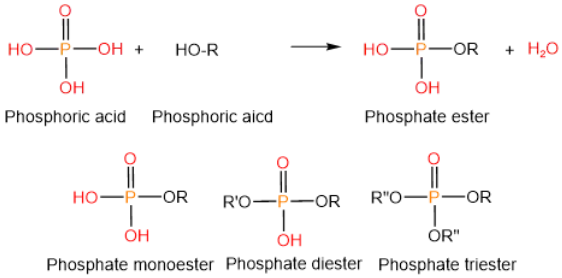
Similarly, three phosphoric acids can condense to form triphosphoric acid. Phosphoric acid, diphosphoric acid, and triphosphoric acid are shown below.



Under physiological conditions at pH ~7.4, the phosphoric acids loose proton and exist as phosphate anions as shown below.



Phosphoric acids also condense with alcohols (**R–OH**) and form mono-, di-, or triesters, as illustrated below.



Nucleotide

Phosphate esters of nucleosides are called nucleotides. The nucleotides found in nucleic acids are formed when **–OH** group at **C**#5 of ribose or deoxyribose makes ester with a phosphate. Examples of monphosphate, diphosphate, and triphosphate esters of adenosine are shown in figure 10.1.7.

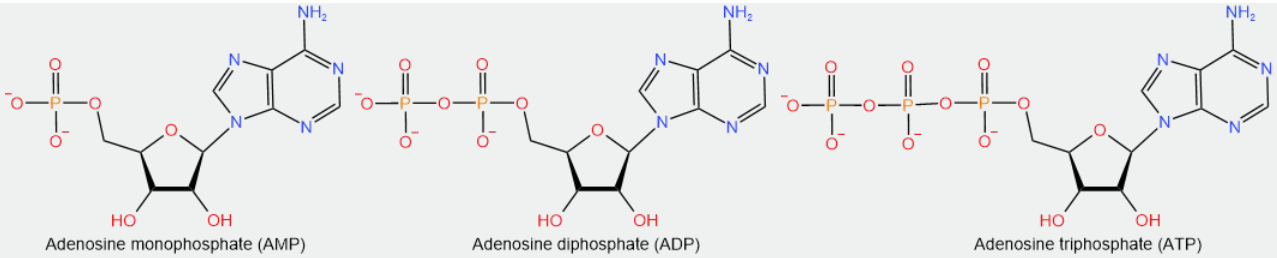


Figure 10.1.7:Phosphate esters of nucleosides are called nucleotides.

The names of the nucleotides are derived by using the name of the corresponding nucleoside with monophosphate, diphosphate, or triphosphate added at the end, depending on whether the phosphate group is monophosphate, diphosphate, or triphosphate. Abbreviation of the nucleotide starts with the one-letter abbreviation of the nitrogen base followed by MP, DP, or TP depending on whether the phosphate group is monophosphate, diphosphate, or triphosphate, as shown in the example structures above.

The nucleotides like ADP, and ATP are high-energy molecules that are energy-rich molecules in biochemical systems. For example, when ATP converts to ADP by releasing a phosphate, energy is released that is used to do work or to synthesize other compounds that need energy for the synthesis. Synthesis of nucleic acids begins with triphosphate esters that convert into monophosphate esters when they incorporate in the nucleic acid polymer. The energy released in the conversion of triphosphate to monophosphate is used to power an endothermic reaction.

The four nucleotides found in RNA are shown in figure 10.8.1 with their names and abbreviations.

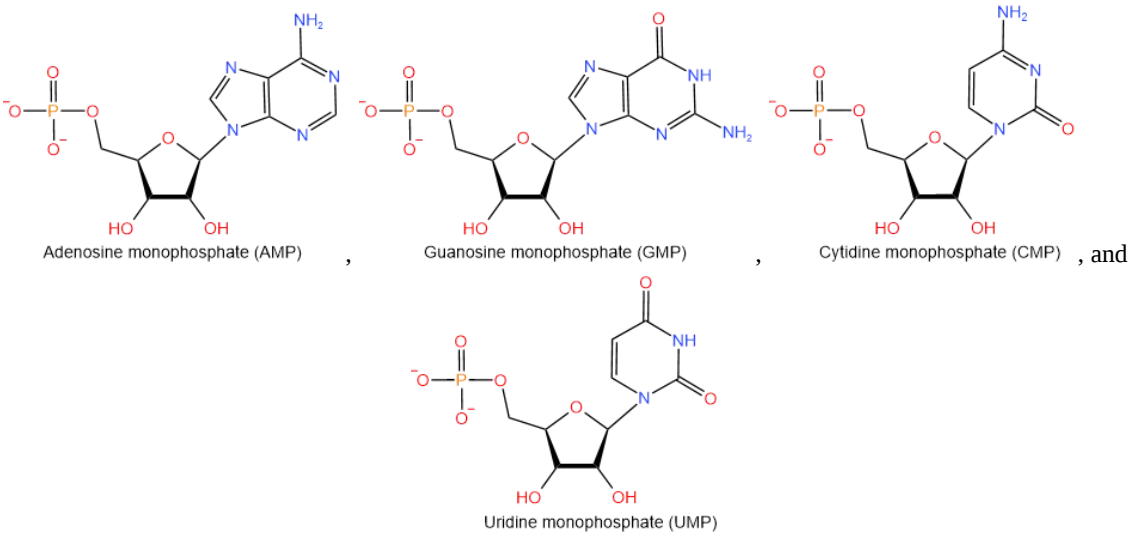


Figure 10.1.8: The four nucleotides found in RNA.

The nucleotides found in DNA are named similarly to those of RNA, i.e., start with the name of the corresponding nucleoside and end with monophosphate, diphosphate, or triphosphate. The abbreviations of the nucleotides found in DNA start with a small alphabet *d* representing the deoxy- prefix of the nucleotides. The four nucleotides found in DNA are shown in figure 10.1.9 with their names and abbreviations.

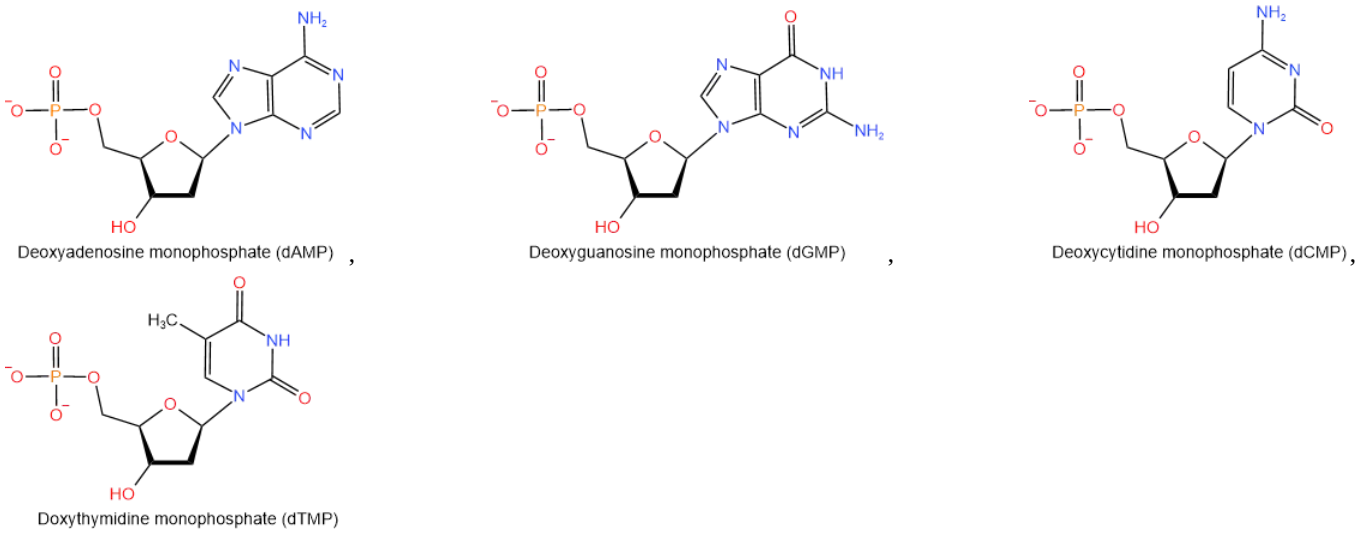


Figure **10.1.9**: The four nucleotides found in DNA.

Summary

Nucleotides are composed of an anion of phosphoric acid, a pentose sugar (ribose or deoxyribose), and a nitrogen-containing base (adenine, cytosine, guanine, thymine, or uracil). Ribonucleotides contain ribose, while deoxyribonucleotides contain deoxyribose.

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10.2: Polynucleotides

Learning Objectives

- Explain the structure of a polynucleotide
- Explain the primary structure of DNA and RNA.

DNA Polynucleotide

Deoxyribonucleic acid (DNA) is the nucleic acid that stores genetic information. Nucleotides are joined together through the phosphate group of one nucleotide linked to the OH group on the third carbon atom of the sugar of a second nucleotide. This bond that connects one nucleotide to another is called the **3',5'-phosphodiester bond**. The process is repeated to produce a long nucleic acid chain. The backbone of the chain consists of alternating phosphate and sugar units (2-deoxyribose in DNA). The purine and pyrimidine bases branch off this sugar-phosphate backbone. The structure of a DNA polynucleotide is shown in figure 10.2.1.

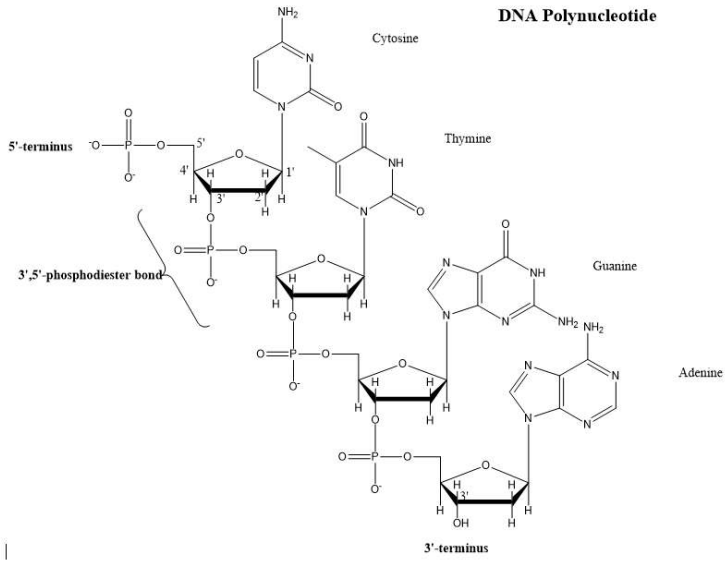


Figure **10.2.1**: A DNA polynucleotide.

There are 20 different kinds of amino acids. However there are only 4 different nucleotides in DNA. For amino acid sequences in the primary structure of the protein, the convention is to write the amino acids in order starting with the N-terminal amino acid. In writing nucleotide sequences for the primary structure of DNA, the convention is to write the one-letter abbreviations for the bases, starting with the nucleotide having a free phosphate group, which is known as the 5′-end, and indicate the nucleotides in order. For DNA, a lowercase *d* is often written in front of the sequence to indicate that the monomers are deoxyribonucleotides. The final nucleotide has a free OH group on the 3′ carbon atom and is called the 3'-end. Note the absence of uracil and presence of thymine in DNA.

The sequence of in the DNA segment shown in figure 10.2.1 is written as: 5′-dC-T-G-A-3′.

RNA Polynucleotide

Ribonucleic acid (RNA) is the nucleic acid responsible for using the genetic information encoded in DNA to produce the proteins found in living organisms. In RNA too nucleotides are joined together through the phosphate group of one nucleotide linked to the OH group on the third carbon atom of the sugar of a second nucleotide. This bond that connects one nucleotide to another is called the 3',5'-phosphodiester bond.The process is repeated to produce a long nucleic acid chain. The backbone of the chain consists of alternating phosphate and sugar units (ribose in RNA). The purine and pyrimidine bases branch off this backbone. The structure of a RNA polynucleotide is shown in figure 10.2.2.

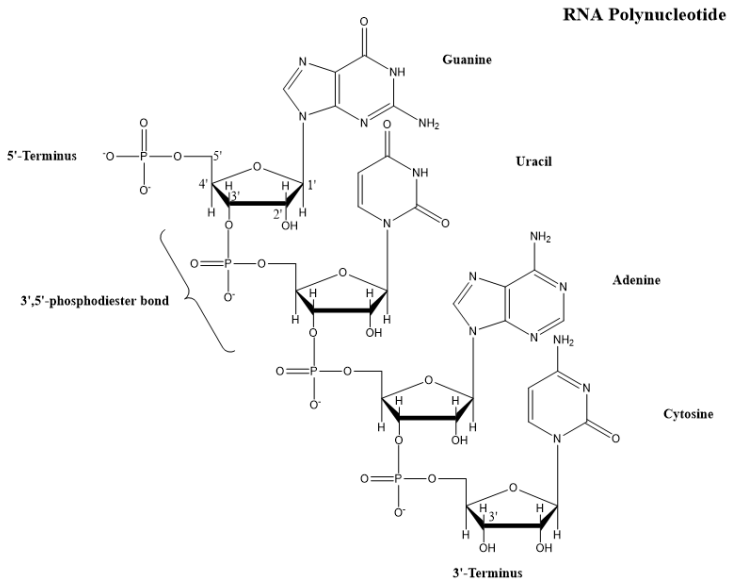


Figure 10.2.2: A RNA polynucleotide.

There are only 4 different nucleotides in RNA. In writing nucleotide sequences for the primary structure of RNA, the convention is to write the one-letter abbreviations for the bases, starting with the nucleotide having a free phosphate group, which is known as the 5′-end, and indicate the nucleotides in order. The final nucleotide has a free OH group on the 3′ carbon atom and is called the 3'-end. Note the absence of thymine and presence of uracil in RNA.

The sequence of in the RNA segment shown in figure 10.2.2 would be written 5′-G-U-A-C-3′.

Summary

Nucleotides are joined together through the phosphate group of one nucleotide linked to the OH group on the third carbon atom of the sugar of a second nucleotide. The process is repeated to produce a long nucleic acid chain. The backbone of the chain consists of alternating phosphate and sugar units. The purine and pyrimidine bases branch off this backbone.

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10.3: Nucleic Acid Structure

Learning Objectives

- Describe the primary structure of nucleotides.
- Describe the secondary structure of DNA and the importance of complementary base pairing.

Nucleic acids are large polymers formed by linking nucleotides together and are found in every cell. Deoxyribonucleic acid (DNA) is the nucleic acid that stores genetic information. If all the DNA in a typical mammalian cell were stretched out end to end, it would extend more than 2 meters. Ribonucleic acid (RNA) is the nucleic acid responsible for using the genetic information encoded in DNA to produce the thousands of proteins found in living organisms.

Primary Structure of Nucleic Acids

Nucleotides are joined together through the phosphate group of one nucleotide connected in an ester linkage to the OH group on the third carbon atom of the sugar unit of a second nucleotide. This unit joins to a third nucleotide, and the process is repeated to produce a long nucleic acid chain (Figure **10.3.1**). The backbone of the chain consists of alternating phosphate and sugar units (2-deoxyribose in DNA and ribose in RNA). The purine and pyrimidine bases branch off this backbone.

> *Each phosphate group has one acidic hydrogen atom that is ionized at physiological pH. This is why these compounds are known as nucleic acids.*

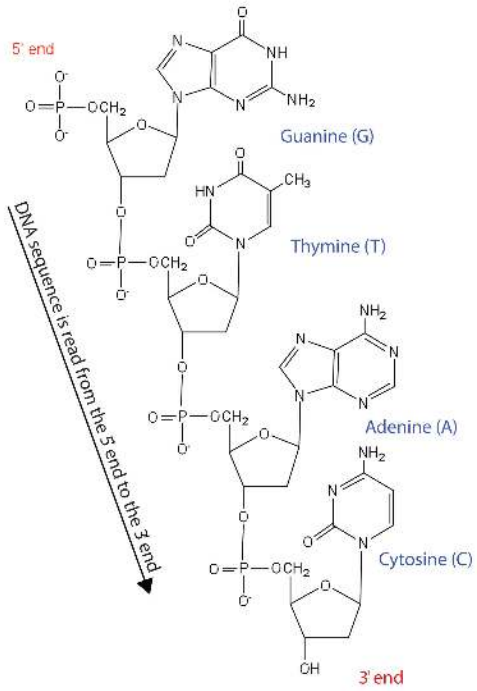


***Figure* 10.3.1** *Structure of a Segment of DNA. A similar segment of RNA would have OH groups on each C2', and uracil would replace thymine.*

Like proteins, nucleic acids have a primary structure that is defined as the sequence of their nucleotides. Unlike proteins, which have 20 different kinds of amino acids, there are only 4 different kinds of nucleotides in nucleic acids. For amino acid sequences in proteins, the convention is to write the amino acids in order starting with the N-terminal amino acid. In writing nucleotide sequences for nucleic acids, the convention is to write the nucleotides (usually using the one-letter abbreviations for the bases, starting with the nucleotide having a free phosphate group, which is known as the 5′ end, and indicate the nucleotides in order. For

DNA, a lowercase *d* is often written in front of the sequence to indicate that the monomers are deoxyribonucleotides. The final nucleotide has a free OH group on the 3′ carbon atom and is called the *3′ end*. The sequence of nucleotides in the DNA segment shown in Figure **10.3.1** would be written 5′-dG-T-A-C-3′, which is often further abbreviated to dGTAC or just GTAC.

Secondary Structure of DNA

The three-dimensional structure of DNA was the subject of an intensive research effort in the late 1940s to early 1950s. Initial work revealed that the polymer had a regular repeating structure. In 1950, Erwin Chargaff of Columbia University showed that the molar amount of adenine (A) in DNA was always equal to that of thymine (T). Similarly, he showed that the molar amount of guanine (G) was the same as that of cytosine (C). Chargaff drew no conclusions from his work, but others soon did.

At Cambridge University in 1953, James D. Watson and Francis Crick announced that they had a model for the secondary structure of DNA. Using the information from Chargaff's experiments (as well as other experiments) and data from the X ray studies of Rosalind Franklin (which involved sophisticated chemistry, physics, and mathematics), Watson and Crick worked with models and finally concluded that DNA is composed of two nucleic acid chains running antiparallel to one another—that is, side-by-side with the 5′ end of one chain next to the 3′ end of the other. Moreover, as their model showed, the two chains are twisted to form a double helix—a structure that can be compared to a spiral staircase, with the phosphate and sugar groups (the backbone of the nucleic acid polymer) representing the outside edges of the staircase. The purine and pyrimidine bases face the inside of the helix, with guanine always paired with cytosine and adenine always paired with thymine. These specific base pairs, referred to as complementary bases, are the steps in the staircase analogy (Figure **10.3.2**).

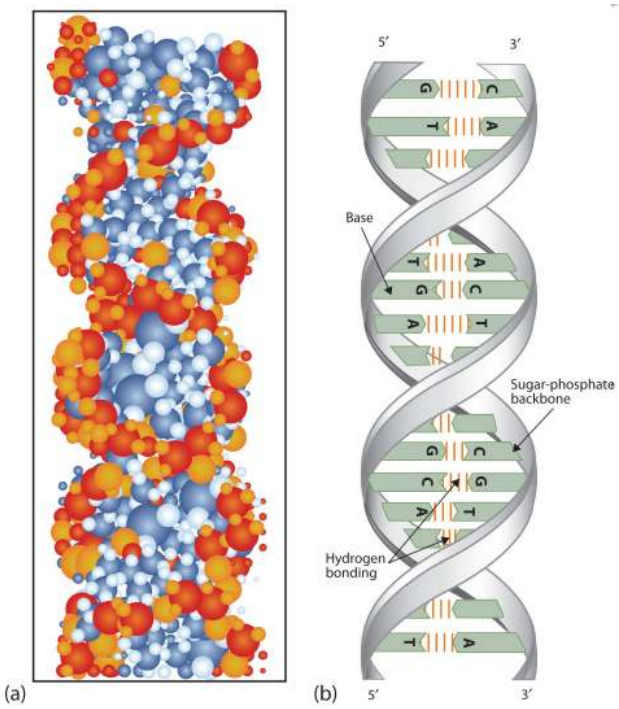


***Figure* 10.3.2** *DNA Double Helix. (a) This represents a computer-generated model of the DNA double helix. (b) This represents a schematic representation of the double helix, showing the complementary bases.*

The structure proposed by Watson and Crick provided clues to the mechanisms by which cells are able to divide into two identical, functioning daughter cells; how genetic data are passed to new generations; and even how proteins are built to required specifications. All these abilities depend on the pairing of complementary bases. Figure **10.3.3** shows the two sets of base pairs and illustrates two things. First, a pyrimidine is paired with a purine in each case, so that the long dimensions of both pairs are identical (1.08 nm).

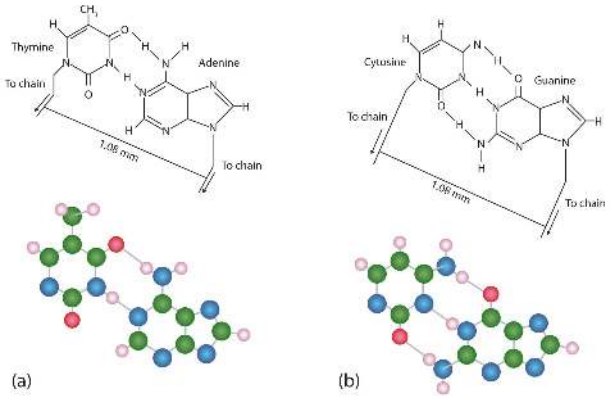


***Figure* 10.3.3** *Complementary Base Pairing. Complementary bases engage in hydrogen bonding with one another: (a) thymine and adenine; (b) cytosine and guanine.*

If two pyrimidines were paired or two purines were paired, the two pyrimidines would take up less space than a purine and a pyrimidine, and the two purines would take up more space, as illustrated in Figure **10.3.4**. If these pairings were ever to occur, the structure of DNA would be like a staircase made with stairs of different widths. For the two strands of the double helix to fit neatly, a pyrimidine must always be paired with a purine. The second thing you should notice in Figure **10.3.3** is that the correct pairing enables formation of three instances of hydrogen bonding between guanine and cytosine and two between adenine and thymine. The additive contribution of this hydrogen bonding imparts great stability to the DNA double helix.

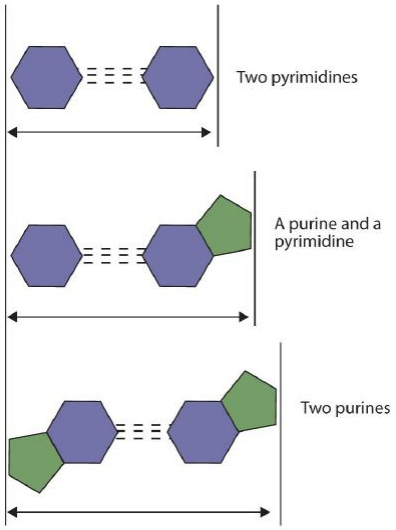


***Figure* 10.3.4** *Difference in Widths of Possible Base Pairs*

Denaturation of DNA

When the two DNA strands of the double helix are subjected to high temperatures or to certain chemicals it can denature the DNA. Denaturation in DNA can break the hydrogen bonds between complementary bases, thus separating the strands into two separate single strands of DNA. This process called DNA denaturation and is analogous to protein denaturation. The single strands can also be put back together as double-stranded DNA, through renaturing by cooling or removing the chemical denaturants, allowing these hydrogen bonds to reform. The ability to artificially manipulate DNA in this way is the basis for several important techniques in biotechnology. Because of the additional hydrogen bonding between the C = G base pair, DNA with a high GC content is more difficult to denature than DNA with a lower GC content.

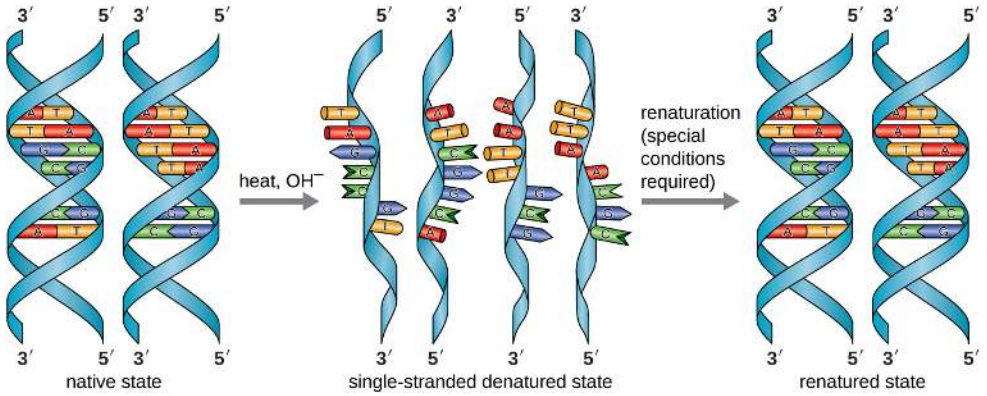


Figure 10.3.5: Denaturation of DNA.

Summary

- DNA is the nucleic acid that stores genetic information. RNA is the nucleic acid responsible for using the genetic information in DNA to produce proteins.
- Nucleotides are joined together to form nucleic acids through the phosphate group of one nucleotide connecting in an ester linkage to the OH group on the third carbon atom of the sugar unit of a second nucleotide.
- Nucleic acid sequences are written starting with the nucleotide having a free phosphate group (the 5′ end).
- Two DNA strands link together in an antiparallel direction and are twisted to form a double helix. The nitrogenous bases face the inside of the helix. Guanine is always opposite cytosine, and adenine is always opposite thymine.

Concept Review Exercises

1. Name the two kinds of nucleic acids.

2. Which type of nucleic acid stores genetic information in the cell?

3. What are complementary bases?

4. Why is it structurally important that a purine base always pair with a pyrimidine base in the DNA double helix?

Answers

1. deoxyribonucleic acid (DNA) and ribonucleic acid (RNA)

2. DNA

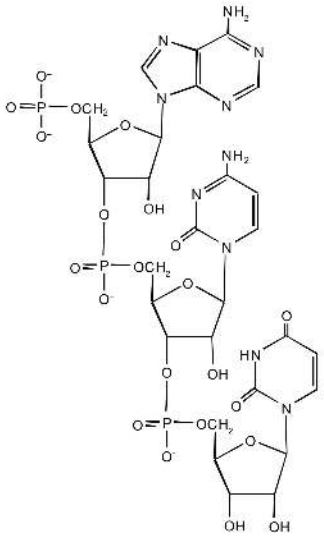
3. the specific base pairings in the DNA double helix in which guanine is paired with cytosine and adenine is paired with thymine

4. The width of the DNA double helix is kept at a constant width, rather than narrowing (if two pyrimidines were across from each other) or widening (if two purines were across from each other).

Exercises

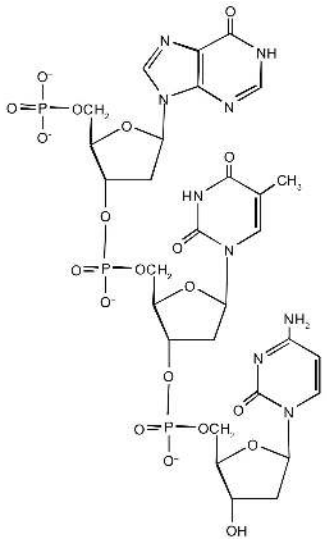
1. For this short RNA segment,

    a. identify the 5′ end and the 3′ end of the molecule.
    b. circle the parts that comprise the backbone of the nucleic acid chain.
    c. write the nucleotide sequence of this RNA segment.



2. For this short DNA segment,

   a. identify the 5′ end and the 3′ end of the molecule.
   b. circle the parts that comprise the backbone of the nucleic acid chain.
   c. write the nucleotide sequence of this DNA segment.



3. Which nitrogenous base in DNA pairs with each nitrogenous base?

   a. cytosine
   b. adenine
   c. guanine
   d. thymine

4. Which nitrogenous base in RNA pairs with each nitrogenous base?

   a. cytosine
   b. adenine
   c. guanine
   d. thymine

5. How many hydrogen bonds can form between the two strands in the short DNA segment shown below?

   5′ ATGCGACTA 3′

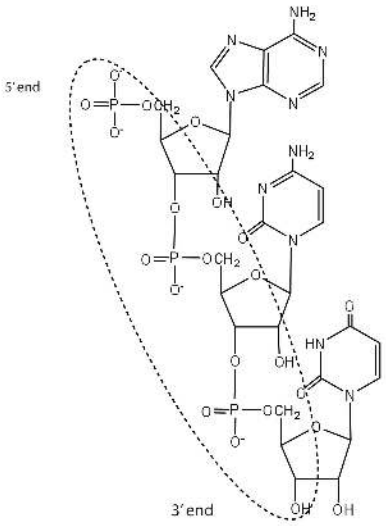
3′ TACGCTGAT 5′

6. How many hydrogen bonds can form between the two strands in the short DNA segment shown below?

   5′ CGATGAGCC 3′

   3′ GCTACTCGG 5′

Answers

1. 

   c. ACU

3. a. guanine
   b. thymine
   c. cytosine
   d. adenine

5. 22 (2 between each AT base pair and 3 between each GC base pair)

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10.4: Replication of DNA

Learning Objectives

- Describe how a new copy of DNA is synthesized in replication

Information Flow

We previously stated that deoxyribonucleic acid (DNA) stores genetic information, while ribonucleic acid (RNA) is responsible for transmitting or expressing genetic information by directing the synthesis of thousands of proteins found in living organisms. But how do the nucleic acids perform these functions? Three processes are required: (1) ***replication***, in which new copies of DNA are made; (2) ***transcription***, in which a segment of DNA is used to produce RNA; and (3) ***translation***, in which the information in RNA is translated into a protein sequence. Together the three steps are addressed as **information flow**. Here we will discuss the first step which is replication.

Replication

New cells are continuously forming in the body through the process of cell division. For this to happen, the DNA in a dividing cell must be copied in a process known as replication. The complementary base pairing of the double helix provides a ready model for how genetic replication occurs. If the two chains of the double helix are pulled apart, disrupting the hydrogen bonding between base pairs, each chain can act as a *template*, for the synthesis of a new complementary DNA chain.

The nucleus contains all the necessary enzymes, proteins, and nucleotides required for this synthesis. A short segment of DNA is "unzipped," so that the two strands in the segment are separated to serve as templates for new DNA. DNA polymerase, an enzyme, binds and moves in the 3' to 5' direction, recognizes each base in a template strand and matches it to the complementary base in a free nucleotide. The enzyme then catalyzes the formation of an ester bond between the 5′ phosphate group of the nucleotide and the 3′ OH end of the new, growing DNA chain. In this way, each strand of the original DNA molecule is used to produce a duplicate of its former partner (Figure **10.4.1**). Whatever information was encoded in the original DNA double helix is now contained in each replicate helix. When the cell divides, each daughter cell gets one of these replicates and thus all of the information that was originally possessed by the parent cell.

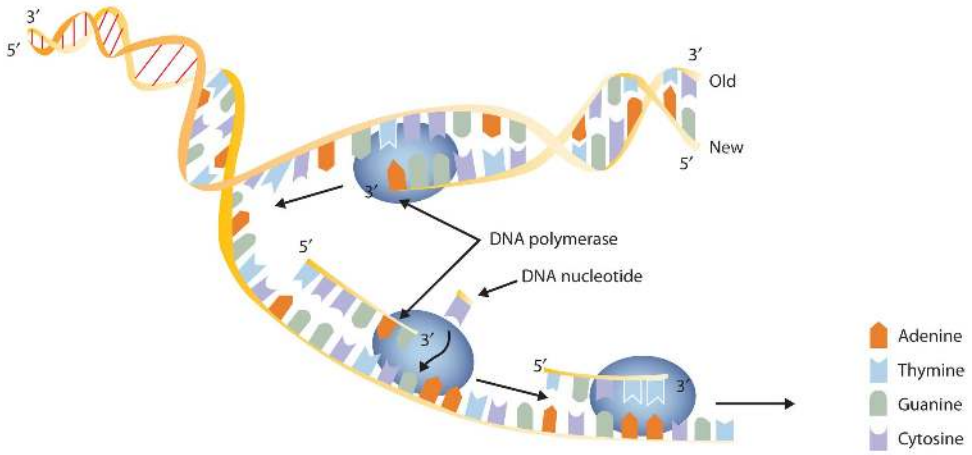


Figure **10.4.1**: A Schematic Diagram of DNA Replication. DNA replication occurs by the sequential unzipping of segments of the double helix. Each new nucleotide is brought into position by DNA polymerase and is added to the growing strand by the formation of a phosphate ester bond. Thus, two double helixes form from one, and each consists of one old strand and one new strand, an outcome called *semiconservative replications*. (This representation is simplified; many more proteins are involved in replication.)

Example 10.4.1

A segment of one strand from a DNA molecule has the sequence 3′-TCCATGAGTTGA-5′. What is the sequence of nucleotides in the opposite, or complementary, DNA chain?

Solution

Knowing that the two strands are antiparallel and that T base pairs with A, while C base pairs with G, the sequence of the complementary strand will be 5′-AGGTACTCAACT-3′.

? Exercise 10.4.1

A segment of one strand from a DNA molecule has the sequence 3′-CCAGTGAATTGCCTAT-5′. List the base sequence in a new DNA strand produced by DNA polymerase. What process is this called? Label the 5’ and 3’ ends.

Answer

A complimentary strand is produced during replication, which means T base pairs with A, while C base pairs with G. The sequence of the new strand is the same as a complementary strand. The new strand is 5′-GGTCACTTAACGGATA-3′.

What do we mean when we say information is encoded in the DNA molecule? An organism’s DNA can be compared to a book containing directions for assembling a model airplane or for knitting a sweater. Letters of the alphabet are arranged into words, and these words direct the individual to perform certain operations with specific materials. If all the directions are followed correctly, a model airplane or sweater is produced.

In DNA, the particular sequences of nucleotides along the chains encode the directions for building an organism. Just as *saw* means one thing in English and *was* means another, the sequence of bases CGT means one thing, and TGC means something different. Although there are only four letters—the four nucleotides—in the genetic code of DNA, their sequencing along the DNA strands can vary so widely that information storage is essentially unlimited.

Summary

- In DNA replication, each strand of the original DNA serves as a template for the synthesis of a complementary strand.
- DNA polymerase is the primary enzyme needed for replication.

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10.5: Transcription

Learning Objectives

- Describe how RNA is synthesized from DNA.
- Identify the different types of RNA and the function of each type of RNA.

Transcription

For the hereditary information in DNA to be useful, it must be "expressed," that is, used to direct the growth and functioning of an organism. DNA expression is the synthesis of RNA, by a template mechanism that is in many ways analogous to DNA replication. Because the RNA that is synthesized is a complementary copy of information contained in DNA, RNA synthesis is referred to as transcription. There are three key differences between replication and transcription:

1. RNA molecules are much shorter than DNA molecules; only a portion of one DNA strand is copied or transcribed to make an RNA molecule.
2. RNA is built from ribonucleotides rather than deoxyribonucleotides.
3. The newly synthesized RNA strand does not remain associated with the DNA sequence it was transcribed from.

The DNA sequence that is transcribed to make RNA is called the *template strand*, while the complementary sequence on the other DNA strand is called the *coding* or *informational strand*. To initiate RNA synthesis, the two DNA strands unwind at specific sites along the DNA molecule. Ribonucleotides are attracted to the uncoiling region of the DNA molecule, beginning at the 3′ end of the template strand, according to the rules of base pairing. Thymine in DNA calls for adenine in RNA, cytosine specifies guanine, guanine calls for cytosine, and adenine requires uracil. RNA polymerase—an enzyme—binds the complementary ribonucleotide and catalyzes the formation of the ester linkage between ribonucleotides, a reaction very similar to that catalyzed by DNA polymerase (Figure **10.5.1**). Synthesis of the RNA strand takes place in the 5′ to 3′ direction, antiparallel to the template strand. Only a short segment of the RNA molecule is hydrogen-bonded to the template strand at any time during transcription. When transcription is completed, the RNA is released, and the DNA helix reforms. The nucleotide sequence of the RNA strand formed during transcription is identical to that of the corresponding coding strand of the DNA, except that U replaces T.

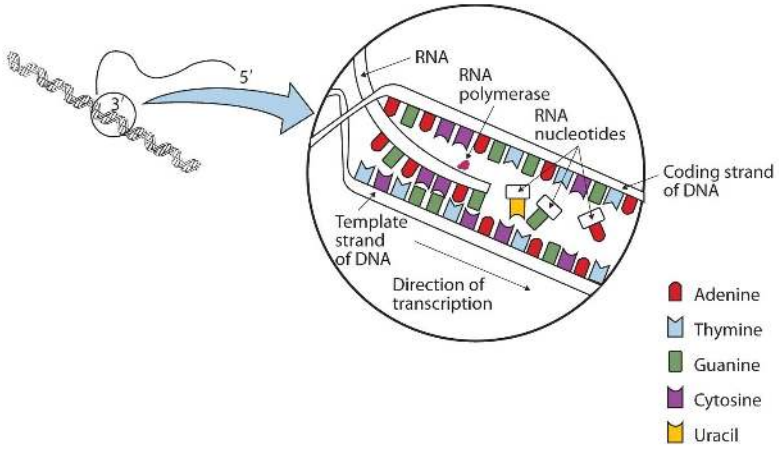


Figure **10.5.1**: A Schematic Diagram of RNA Transcription from a DNA Template. The representation of RNA polymerase is proportionately much smaller than the actual molecule, which encompasses about 50 nucleotides at a time.

Example 10.5.1

A portion of the template strand of a gene has the sequence 3′-TCCATGAGTTGA-5′. What is the sequence of nucleotides in the RNA that is formed from this template?

Solution

Four things must be remembered in answering this question: (1) the DNA strand and the RNA strand being synthesized are antiparallel; (2) RNA is synthesized in a 5′ to 3′ direction, so transcription begins at the 3′ end of the template strand; (3) ribonucleotides are used in place of deoxyribonucleotides; and (4) thymine (T) base pairs with adenine (A), A base pairs with uracil (U; in RNA), and cytosine (C) base pairs with guanine (G).

The sequence is determined to be 5′-AGGUACUCAACU-3′.

? Exercise 10.5.1

A portion of the template strand of a gene has the sequence 3′-CCAGTGAATTGCCTAT-5′. What is the sequence of nucleotides in the RNA that is formed from this template?

Answer

Four things must be remembered in answering this question: (1) the DNA strand and the RNA strand being synthesized are antiparallel; (2) RNA is synthesized in a 5′ to 3′ direction, so transcription begins at the 3′ end of the template strand; (3) ribonucleotides are used in place of deoxyribonucleotides; and (4) thymine (T) base pairs with adenine (A), A base pairs with uracil (U; in RNA), and cytosine (C) base pairs with guanine (G).

The sequence is determined to be 5′-GGUCACUUAACGGAUA-3′.

RNA undergoes modification following transcription called the **post-transcriptional modification**. Modifications to the mRNA involves altering the 3' and 5' ends and removing sections from the middle, whose codes are not needed for protein production.

RNA

Three types of RNA are formed during transcription: *messenger RNA* (mRNA), *ribosomal RNA* (rRNA), and *transfer RNA* (tRNA). These three types of RNA differ in function, size, and percentage of the total cell RNA (Table **10.5.1**). mRNA makes up only a small percent of the total amount of RNA within the cell, primarily because each molecule of mRNA exists for a relatively short time; it is continuously being degraded and resynthesized. The molecular dimensions of the mRNA molecule vary according to the amount of genetic information a given molecule contains. After transcription, which takes place in the nucleus, the mRNA passes into the cytoplasm, carrying the genetic message from DNA to the ribosomes, the sites of protein synthesis. mRNA directly determines the sequence of amino acids during protein synthesis.

Table **10.5.1**: Properties of Cellular RNA in Escherichia coli

Type	Function	Approximate Number of Nucleotides	Percentage of Total Cell RNA
mRNA	codes for proteins	100–6,000	~3
rRNA	component of ribosomes	120–2900	83
tRNA	transfer molecule that transfers the amino acid to the ribosome	75–90	14

Ribosomes are cellular substructures where proteins are synthesized. They contain about 65% rRNA and 35% protein, held together by numerous noncovalent interactions, such as hydrogen bonding, in an overall structure consisting of two globular particles of unequal size.

Molecules of tRNA, which bring amino acids (one at a time) to the ribosomes for the construction of proteins, differ from one another in the kinds of amino acid each is specifically designed to carry. A set of three nucleotides, known as a codon, on the mRNA determines which kind of tRNA will add its amino acid to the growing chain. Each of the 20 amino acids found in proteins has at least one corresponding kind of tRNA, and most amino acids have more than one.

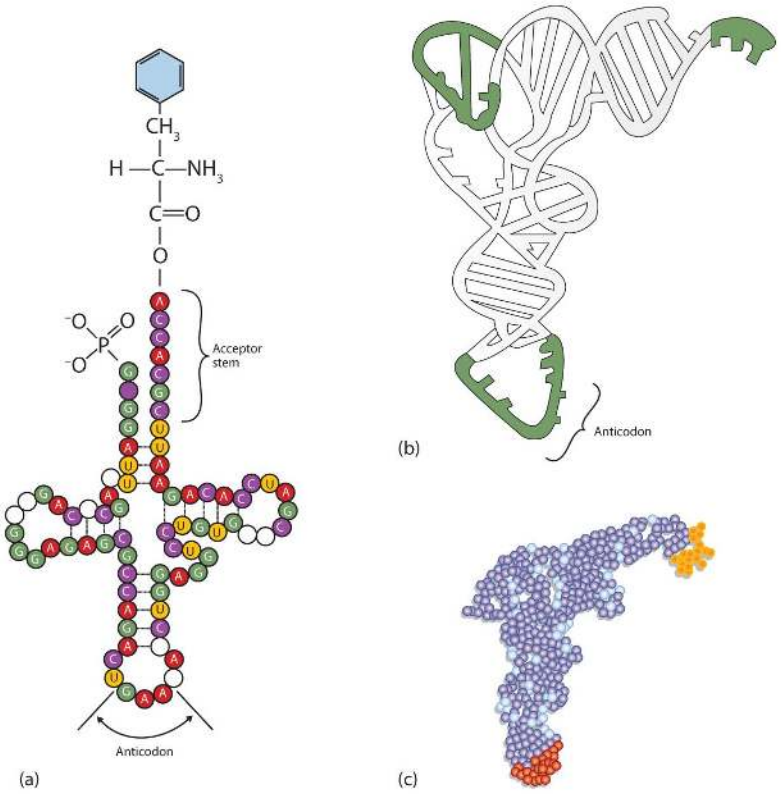


Figure **10.5.2**: Transfer RNA. (a) In the two-dimensional structure of a yeast tRNA molecule for phenylalanine, the amino acid binds to the acceptor stem located at the 3′ end of the tRNA primary sequence. (The nucleotides that are not specifically identified here are slightly altered analogs of the four common ribonucleotides A, U, C, and G.) (b) In the three-dimensional structure of yeast phenylalanine tRNA, note that the anticodon loop is at the bottom and the acceptor stem is at the top right. (c) This shows a space-filling model of the tRNA.

The two-dimensional structure of a tRNA molecule has three distinctive loops, reminiscent of a cloverleaf (Figure **10.5.2**). On one loop is a sequence of three nucleotides that varies for each kind of tRNA. This triplet, called the anticodon, is complementary to and pairs with the codon on the mRNA. At the opposite end of the molecule is the acceptor stem, where the amino acid is attached.

Summary

- In transcription, a segment of DNA serves as a template for the synthesis of an RNA sequence.
- RNA polymerase is the primary enzyme needed for transcription.
- Three types of RNA are formed during transcription: mRNA, rRNA, and tRNA.

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10.6: Translation and the Genetic Code

Learning Objectives

- describe how a protein is synthesized from mRNA.
- describe the characteristics of the genetic code.

Gene is a segment of deoxyribonucleic acid (DNA) carrying the code for a specific polypeptide. Each molecule of messenger RNA (mRNA) is a transcribed copy of a gene that is used by a cell for synthesizing a polypeptide chain. If a protein contains two or more different polypeptide chains, each chain is coded by a different gene. We turn now to the question of how the sequence of nucleotides in a molecule of ribonucleic acid (RNA) is translated into an amino acid sequence.

Genetic Code

How can a molecule containing just 4 different nucleotides specify the sequence of the 20 amino acids that occur in proteins? If each nucleotide coded for 1 amino acid, then obviously the nucleic acids could code for only 4 amino acids. What if amino acids were coded for by groups of 2 nucleotides? There are 4^2, or 16, different combinations of 2 nucleotides (AA, AU, AC, AG, UU, and so on). Such a code is more extensive but still not adequate to code for 20 amino acids. However, if the nucleotides are arranged in groups of 3, the number of different possible combinations is 4^3, or 64. Here we have a code that is extensive enough to direct the synthesis of the primary structure of a protein molecule.



***Video**: NDSU Virtual Cell Animations project animation "Translation". For more information, see VCell, NDSU(opens in new window) [vcell.ndsu.nodak.edu]*

The genetic code can therefore be described as the identification of each group of three nucleotides and its particular amino acid. The sequence of these triplet groups in the mRNA dictates the sequence of the amino acids in the protein. Each individual three-nucleotide coding unit, as we have seen, is called a *codon*.

The Codon Chart

Early experimenters were faced with the task of determining which of the 64 possible codons stood for each of the 20 amino acids. The cracking of the genetic code was the joint accomplishment of several well-known geneticists—notably Har Khorana, Marshall Nirenberg, Philip Leder, and Severo Ochoa—from 1961 to 1964. The genetic dictionary they compiled, summarized in Figure **10.6.1**, shows that out of 64 codons, 61 codons code for amino acids, and 3 codons serve as signals for the termination of polypeptide synthesis (much like the period at the end of a sentence). Notice that only methionine (AUG) and tryptophan (UGG) have single codons. All other amino acids have two or more codons.

First base	Second base: U	Second base: C	Second base: A	Second base: G	Third base
U	Phe	Ser	Tyr	Cys	U
	Phe	Ser	Tyr	Cys	C
	Leu	Ser	Stop	Stop	A
	Leu	Ser	Stop	Trp	G
C	Leu	Pro	His	Arg	U
	Leu	Pro	His	Arg	C
	Leu	Pro	Gln	Arg	A
	Leu	Pro	Gln	Arg	G
A	Ile	Thr	Asn	Ser	U
	Ile	Thr	Asn	Ser	C
	Ile	Thr	Lys	Arg	A
	Met	Thr	Lys	Arg	G
G	Val	Ala	Asp	Gly	U
	Val	Ala	Asp	Gly	C
	Val	Ala	Glu	Gly	A
	Val	Ala	Glu	Gly	G

Figure **10.6.1**: The Genetic Code

Translation

Protein synthesis is accomplished by orderly interactions between mRNA and the other ribonucleic acids (transfer RNA [tRNA] and ribosomal RNA [rRNA]), the ribosome, and more than 100 enzymes. The mRNA formed in the nucleus during transcription is transported across the nuclear membrane into the cytoplasm to the ribosomes—carrying with it the genetic instructions. The process in which the information encoded in the mRNA is used to direct the sequencing of amino acids and thus ultimately to synthesize a protein is referred to as translation.

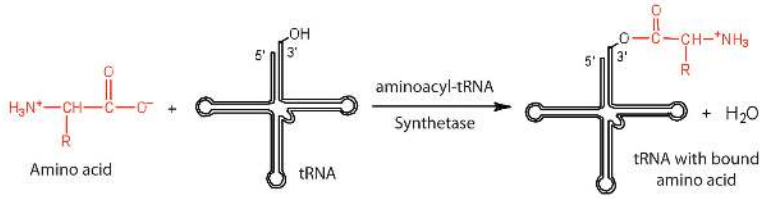


Figure **10.6.2**: Binding of an Amino Acid to Its tRNA

Before an amino acid can be incorporated into a polypeptide chain, it must be attached to its unique tRNA. This crucial process requires an enzyme known as aminoacyl-tRNA synthetase (Figure **10.6.2**). There is a specific aminoacyl-tRNA synthetase for each amino acid. This high degree of specificity is vital to the incorporation of the correct amino acid into a protein. After the amino acid molecule has been bound to its tRNA carrier, protein synthesis can take place. Figure **10.6.3** depicts a schematic stepwise representation of this all-important process.

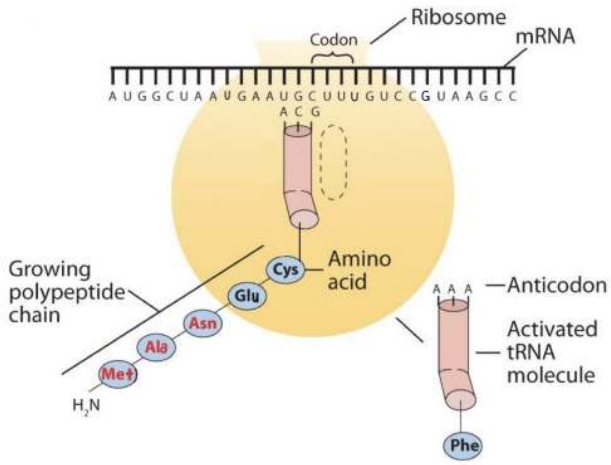


Figure **10.6.3*a***: The Elongation Steps in Protein Synthesis - Protein synthesis is already in progress at the ribosome. The growing polypeptide chain is attached to the tRNA that brought in the previous amino acid (in this illustration, cys.)

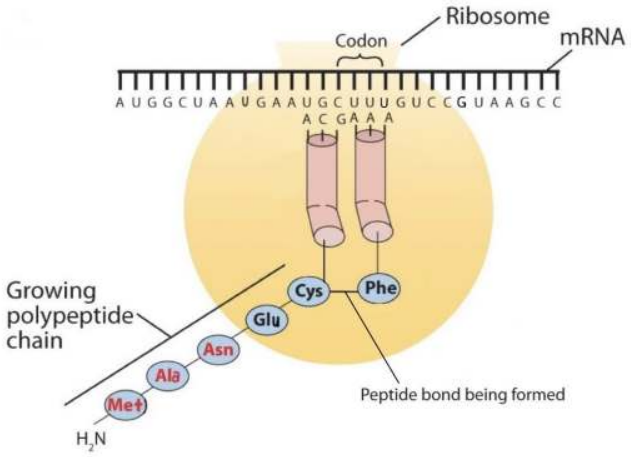


Figure **10.6.3*b***: The Elongation Steps in Protein Synthesis - An activated tRNA, which has the anticodon AAA, binds to the ribosome next to the previous bound tRNA and interacts with the mRNA molecule though base pairing of the codon and anticodon. The amino acid Phe is being incorporated into the polypeptide chain by the formation of a peptide linkage between the carboxyl group of Cys and the amino acid group of the Phe. This reaction is catalyzed by the enzyme peptidyl transferase, a component of the ribosome.

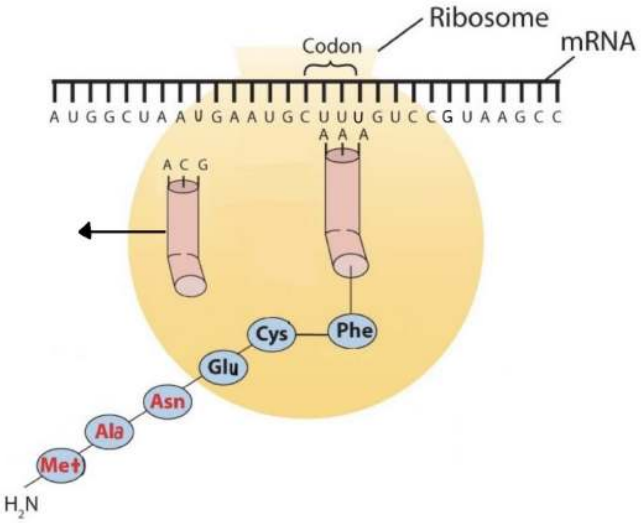


Figure **10.6.3*c***: The Elongation Steps in Protein Synthesis - The Cys-Phe linkage is now complete, and the growing polypeptide chain remains attached to the tRNA for Phe.

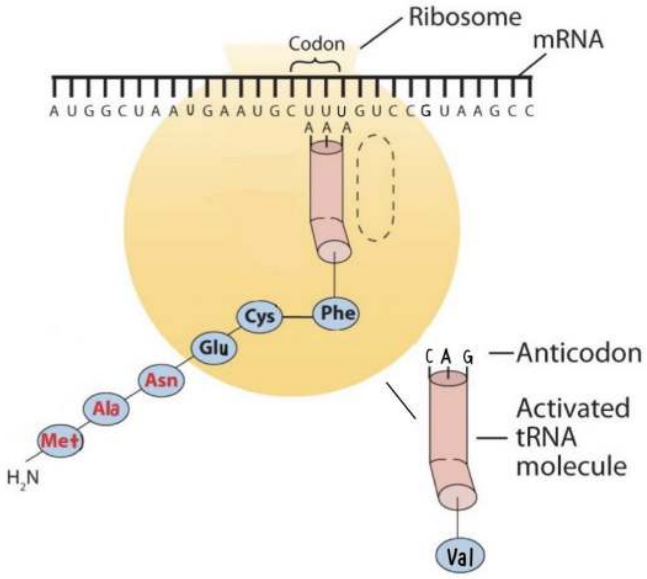


Figure **10.6.3*d***: The Elongation Steps in Protein Synthesis - The ribosome moves to the right along the mRNA strand. This shift brings the next codon, GUC, into its correct position on the surface of the ribosome. Note that an activated tRNA molecule, containing the next amino acid to be attached to the chain is moving to the ribosome. Steps (b)-(d) will be repeated until the ribosome reaches a stop codon.

✔ Example **10.6.1**: Using the Genetic Code

A portion of an mRNA molecule has the sequence 5′-AUGCCACGAGUUGAC-3′. What amino acid sequence does this code for?

Solution

Use Figure **10.6.1** to determine what amino acid each set of three nucleotides (codon) codes for. Remember that the sequence is read starting from the 5′ end and that a protein is synthesized starting with the N-terminal amino acid.

The sequence 5′-AUGCCACGAGUUGAC-3′ codes for met-pro-arg-val-asp.

? Exercise 10.6.1

A portion of an RNA molecule has the sequence 5′-AUGCUGAAUUGCGUAGGA-3′. What amino acid sequence does this code for? What process is this called?

Answer

The sequence 5′-AUGCUGAAUUGCGUAGGA-3′ codes for met-leu-asn-cys-val-gly. The process is translation.

Further experimentation threw much light on the nature of the genetic code, as follows:

1. The code is virtually universal; animal, plant, and bacterial cells use the same codons to specify each amino acid (with a few exceptions).
2. The code is “degenerate”; in all but two cases (methionine and tryptophan). Degenerate means more than one triplet codes for a given amino acid.
3. The first two bases of each codon are most significant; the third base often varies. This suggests that a change in the third base by a mutation may still permit the correct incorporation of a given amino acid into a protein. The third base is sometimes called the “wobble” base.
4. The code is continuous and nonoverlapping; there are *no* nucleotides between codons, and adjacent codons do not overlap.
5. The three termination codons are read by special proteins called release factors, which signal the end of the translation process.
6. The codon AUG codes for methionine and is also the initiation codon. Thus methionine is the first amino acid in each newly synthesized polypeptide. This first amino acid is usually removed enzymatically before the polypeptide chain is completed; the vast majority of polypeptides do not begin with methionine.

After the protein is made it is processed to become active with **post-translational modification**. There is removal of Met at the N-terminus, trimming and cutting of the polypeptide chain, folding of the protein into its proper three-dimensional shape, addition of prosthetic groups and metal ions, formation of the disulfide bridges, and association with other polypeptide chains to form the quaternary structure.

Summary

In translation, the information in mRNA directs the order of amino acids in protein synthesis. A set of three nucleotides (codon) codes for a specific amino acid.

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10.7: Mutations and Genetic Diseases

Learning Objectives

- To describe the causes of genetic mutations and how they lead to genetic diseases.

We have seen that the sequence of nucleotides in a cell's deoxyribonucleic acid (DNA) is what ultimately determines the sequence of amino acids in proteins made by the cell and thus is critical for the proper functioning of the cell. On rare occasions, however, the nucleotide sequence in DNA may be modified either spontaneously (by errors during replication, occurring approximately once for every 10 billion nucleotides) or from exposure to heat, radiation, or certain chemicals. Any chemical or physical change that alters the nucleotide sequence in DNA is called a mutation. When a mutation occurs in an egg or sperm cell that then produces a living organism, it will be inherited by all the offspring of that organism.

Common types of mutations include **substitution** (a different nucleotide is substituted), **insertion** (the addition of a new nucleotide), and **deletion** (the loss of a nucleotide). These changes within DNA are called point mutations because only one nucleotide is substituted, added, or deleted (Figure **10.7.1**). Because an insertion or deletion results in a frame-shift that changes the reading of subsequent codons and, therefore, alters the entire amino acid sequence that follows the mutation, insertions and deletions are usually more harmful than a substitution in which only a single amino acid is altered.

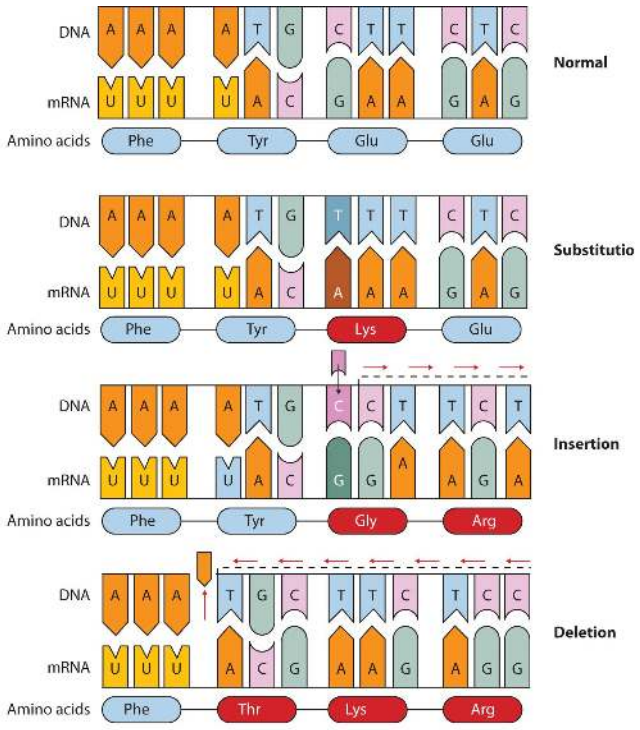


Figure **10.7.1**: Three Types of Point Mutations

The chemical or physical agents that cause mutations are called mutagens. Examples of physical mutagens are ultraviolet (UV) and gamma radiation. Radiation exerts its mutagenic effect either directly or by creating free radicals that in turn have mutagenic effects. Radiation and free radicals can lead to the formation of bonds between nitrogenous bases in DNA. For example, exposure to UV light can result in the formation of a covalent bond between two adjacent thymines on a DNA strand, producing a thymine dimer (Figure **10.7.2**). If not repaired, the dimer prevents the formation of the double helix at the point where it occurs. The genetic disease *xeroderma pigmentosum* is caused by a lack of the enzyme that cuts out the thymine dimers in damaged DNA. Individuals affected by this condition are abnormally sensitive to light and are more prone to skin cancer than normal individuals.

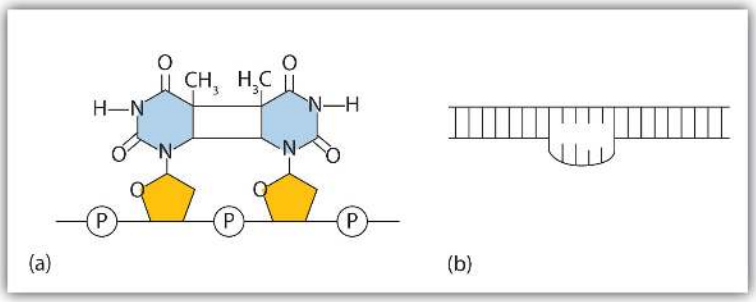


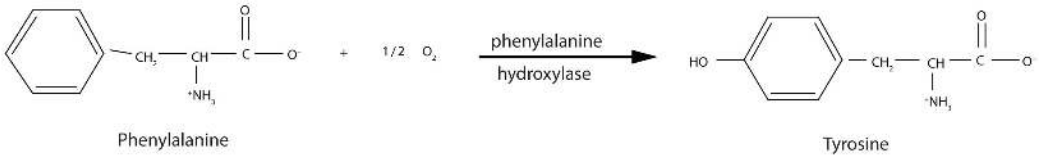
Figure **10.7.2**: An Example of Radiation Damage to DNA. (a) The thymine dimer is formed by the action of UV light. (b) When a defect in the double strand is produced by the thymine dimer, this defect temporarily stops DNA replication, but the dimer can be removed, and the region can be repaired by an enzyme repair system.

Sometimes gene mutations are beneficial, but most of them are detrimental. For example, if a point mutation occurs at a crucial position in a DNA sequence, the affected protein will lack biological activity, perhaps resulting in the death of a cell. In such cases the altered DNA sequence is lost and will not be copied into daughter cells. Nonlethal mutations in an egg or sperm cell may lead to metabolic abnormalities or hereditary diseases. Such diseases are called *inborn errors of metabolism* or genetic diseases. A partial listing of genetic diseases is presented in Figure **10.7.1**, and two specific diseases are discussed in the following sections. In most cases, the defective gene results in a failure to synthesize a particular enzyme.

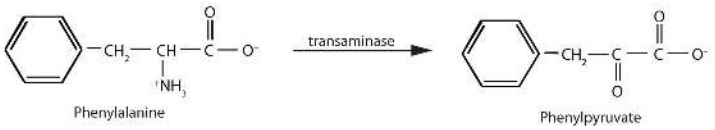
Figure **10.7.1**: Some Representative Genetic Diseases in Humans and the Protein or Enzyme Responsible

Disease	Responsible Protein or Enzyme
alkaptonuria	homogentisic acid oxidase
galactosemia	galactose 1-phosphate uridyl transferase, galactokinase, or UDP galactose epimerase
Gaucher disease	glucocerebrosidase
gout and Lesch-Nyhan syndrome	hypoxanthine-guanine phosphoribosyl transferase
hemophilia	antihemophilic factor (factor VIII) or Christmas factor (factor IX)
homocystinuria	cystathionine synthetase
maple syrup urine disease	branched chain α-keto acid dehydrogenase complex
McArdle syndrome	muscle phosphorylase
Niemann-Pick disease	sphingomyelinase
phenylketonuria (PKU)	phenylalanine hydroxylase
sickle cell anemia	hemoglobin
Tay-Sachs disease	hexosaminidase A
tyrosinemia	fumarylacetoacetate hydrolase or tyrosine aminotransferase
von Gierke disease	glucose 6-phosphatase
Wilson disease	Wilson disease protein

PKU results from the absence of the enzyme phenylalanine hydroxylase. Without this enzyme, a person cannot convert phenylalanine to tyrosine, which is the precursor of the neurotransmitters dopamine and norepinephrine as well as the skin pigment melanin.



When this reaction cannot occur, phenylalanine accumulates and is then converted to higher than normal quantities of phenylpyruvate. The disease acquired its name from the high levels of phenylpyruvate (a phenyl ketone) in urine. Excessive amounts of phenylpyruvate impair normal brain development, which causes severe mental retardation.



PKU may be diagnosed by assaying a sample of blood or urine for phenylalanine or one of its metabolites. Medical authorities recommend testing every newborn's blood for phenylalanine within 24 h to 3 weeks after birth. If the condition is detected, mental retardation can be prevented by immediately placing the infant on a diet containing little or no phenylalanine. Because phenylalanine is plentiful in naturally produced proteins, the low-phenylalanine diet depends on a synthetic protein substitute plus very small measured amounts of naturally produced foods. Before dietary treatment was introduced in the early 1960s, severe mental retardation was a common outcome for children with PKU. Prior to the 1960s, 85% of patients with PKU had an intelligence quotient (IQ) less than 40, and 37% had IQ scores below 10. Since the introduction of dietary treatments, however, over 95% of children with PKU have developed normal or near-normal intelligence. The incidence of PKU in newborns is about 1 in 12,000 in North America.

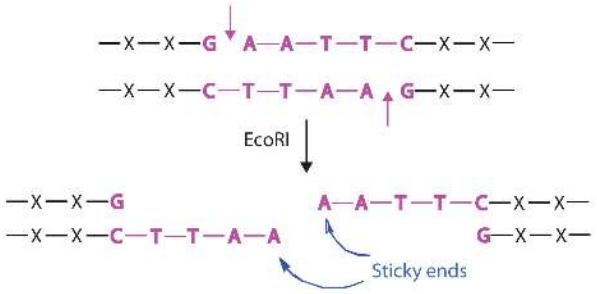
> *Every state in the United States has mandated that screening for PKU be provided to all newborns.*

Several genetic diseases are collectively categorized as *lipid-storage diseases*. Lipids are constantly being synthesized and broken down in the body, so if the enzymes that catalyze lipid degradation are missing, the lipids tend to accumulate and cause a variety of medical problems. When a genetic mutation occurs in the gene for the enzyme hexosaminidase A, for example, gangliosides cannot be degraded but accumulate in brain tissue, causing the ganglion cells of the brain to become greatly enlarged and nonfunctional. This genetic disease, known as *Tay-Sachs disease*, leads to a regression in development, dementia, paralysis, and blindness, with death usually occurring before the age of three. There is currently no treatment, but Tay-Sachs disease can be diagnosed in a fetus by assaying the amniotic fluid (amniocentesis) for hexosaminidase A. A blood test can identify Tay-Sachs carriers—people who inherit a defective gene from only one rather than both parents—because they produce only half the normal amount of hexosaminidase A, although they do not exhibit symptoms of the disease.

Looking Closer: Recombinant DNA Technology

More than 3,000 human diseases have been shown to have a genetic component, caused or in some way modulated by the person's genetic composition. Moreover, in the last decade or so, researchers have succeeded in identifying many of the genes and even mutations that are responsible for specific genetic diseases. Now scientists have found ways of identifying and isolating genes that have specific biological functions and placing those genes in another organism, such as a bacterium, which can be easily grown in culture. With these techniques, known as *recombinant DNA technology*, the ability to cure many serious genetic diseases appears to be within our grasp.

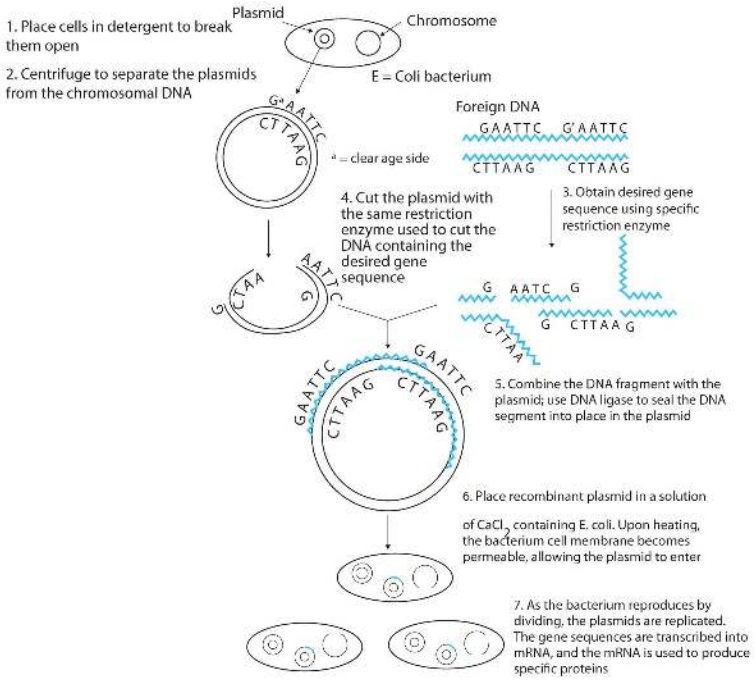
Isolating the specific gene or genes that cause a particular genetic disease is a monumental task. One reason for the difficulty is the enormous amount of a cell's DNA, only a minute portion of which contains the gene sequence. Thus, the first task is to obtain smaller pieces of DNA that can be more easily handled. Fortunately, researchers are able to use *restriction enzymes* (also known as restriction endonucleases), discovered in 1970, which are enzymes that cut DNA at specific, known nucleotide sequences, yielding DNA fragments of shorter length. For example, the restriction enzyme *EcoRI* recognizes the nucleotide sequence shown here and cuts both DNA strands as indicated:



Once a DNA strand has been fragmented, it must be cloned; that is, multiple identical copies of each DNA fragment are produced to make sure there are sufficient amounts of each to detect and manipulate in the laboratory. Cloning is accomplished by inserting the individual DNA fragments into phages (bacterial viruses) that can enter bacterial cells and be replicated. When a bacterial cell infected by the modified phage is placed in an appropriate culture medium, it forms a colony of cells, all containing copies of the original DNA fragment. This technique is used to produce many bacterial colonies, each containing a different DNA fragment. The result is a *DNA library*, a collection of bacterial colonies that together contain the entire genome of a particular organism.

The next task is to screen the DNA library to determine which bacterial colony (or colonies) has incorporated the DNA fragment containing the desired gene. A short piece of DNA, known as a *hybridization probe*, which has a nucleotide sequence complementary to a known sequence in the gene, is synthesized, and a radioactive phosphate group is added to it as a "tag." You might be wondering how researchers are able to prepare such a probe if the gene has not yet been isolated. One way is to use a segment of the desired gene isolated from another organism. An alternative method depends on knowing all or part of the amino acid sequence of the protein produced by the gene of interest: the amino acid sequence is used to produce an approximate genetic code for the gene, and this nucleotide sequence is then produced synthetically. (The amino acid sequence used is carefully chosen to include, if possible, many amino acids such as methionine and tryptophan, which have only a single codon each.)

After a probe identifies a colony containing the desired gene, the DNA fragment is clipped out, again using restriction enzymes, and spliced into another replicating entity, usually a plasmid. *Plasmids* are tiny mini-chromosomes found in many bacteria, such as *Escherichia coli* (*E. coli*). A recombined plasmid would then be inserted into the host organism (usually the bacterium *E. coli*), where it would go to work to produce the desired protein.



Proponents of recombinant DNA research are excited about its great potential benefits. An example is the production of human growth hormone, which is used to treat children who fail to grow properly. Formerly, human growth hormone was available only in tiny amounts obtained from cadavers. Now it is readily available through recombinant DNA technology. Another gene that has been cloned is the gene for epidermal growth factor, which stimulates the growth of skin cells and can be used to speed the healing of burns and other skin wounds. Recombinant techniques are also a powerful research tool, providing enormous aid to scientists as they map and sequence genes and determine the functions of different segments of an organism's DNA.

In addition to advancements in the ongoing treatment of genetic diseases, recombinant DNA technology may actually lead to cures. When appropriate genes are successfully inserted into *E. coli*, the bacteria can become miniature pharmaceutical factories, producing great quantities of insulin for people with diabetes, clotting factor for people with hemophilia, missing enzymes, hormones, vitamins, antibodies, vaccines, and so on. Recent accomplishments include the production in *E. coli* of recombinant DNA molecules containing synthetic genes for tissue plasminogen activator, a clot-dissolving enzyme that can rescue heart attack victims, as well as the production of vaccines against hepatitis B (humans) and hoof-and-mouth disease (cattle).

Scientists have used other bacteria besides *E. coli* in gene-splicing experiments and also yeast and fungi. Plant molecular biologists use a bacterial plasmid to introduce genes for several foreign proteins (including animal proteins) into plants. The bacterium is *Agrobacterium tumefaciens*, which can cause tumors in many plants, but which can be treated so that its tumor-causing ability is eliminated. One practical application of its plasmids would be to enhance a plant's nutritional value by transferring into it the gene necessary for the synthesis of an amino acid in which the plant is normally deficient (for example, transferring the gene for methionine synthesis into pinto beans, which normally do not synthesize high levels of methionine).

> Restriction enzymes have been isolated from a number of bacteria and are named after the bacterium of origin. *EcoRI* is a restriction enzyme obtained from the R strain of *E. coli*. The roman numeral I indicates that it was the first restriction enzyme obtained from this strain of bacteria.

Summary

- The nucleotide sequence in DNA may be modified either spontaneously or from exposure to heat, radiation, or certain chemicals and can lead to mutations.
- Mutagens are the chemical or physical agents that cause mutations.
- Genetic diseases are hereditary diseases that occur because of a mutation in a critical gene.

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10.8: End of Chapter Problems

Building Blocks of Nucleic Acids

1. Name the four bases that are present in DNA.

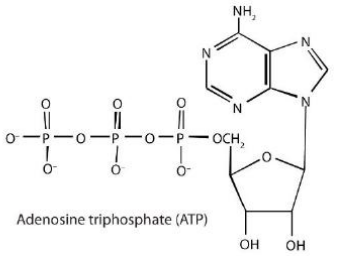
2. Name the four bases that are present in RNA.

3. Name the monosaccharide used to make DNA. Draw in the ß-furanose form.

4. Name the monosaccharide used to make RNA. Draw in the ß-furanose form.

5. Which three components go into making a nucleotide?

6. Shown below is the structure of ATP, a nucleotide triphosphate.



a. How many phosphoester bonds are present in the molecule?

b. How many phosphoanhydride bonds are present?

c. What is the name of the base?

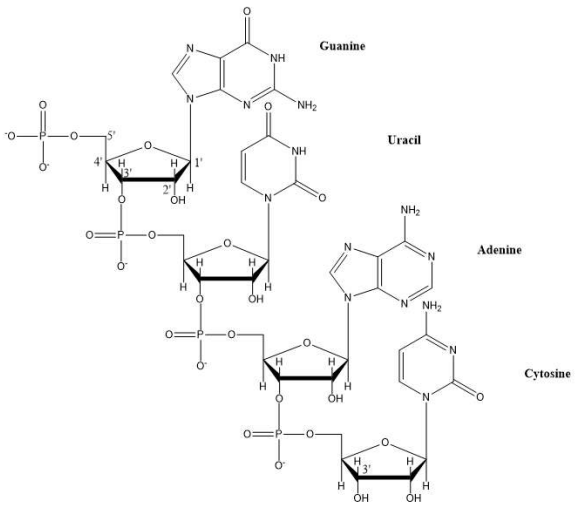
d. What is the name of the monosaccharide?

e. Where is the base attached?

f. What type of bond connects the sugar and the base?

Polynucleotide

Shown below is the structure of a polynucleotide. Answer the following questions based on the structure.



a. Name the monosaccharide in this polynucleotide.

b. Name the bond that connects the heterocyclic amine to the monosaccharide.

c. Name the bond that connects one nucleotide to another.

d. Label the terminus on the polynucleotide with the appropriate nomenclature.

e. Is this polynucleotide a part of RNA or DNA?

f. Write the primary sequence of this nucleotide.

7. Name the two kinds of nucleic acids.

8. Which type of nucleic acid stores genetic information in the cell?

9. What is the primary structure of DNA?

10. What are complementary bases?

11. Why is it structurally important that a purine base always pair with a pyrimidine base in the DNA double helix?

12. Which nitrogenous base in DNA pairs with each nitrogenous base?

cytosine adenine guanine thymine

13. Describe the secondary structure of DNA?

14. What level of structure is disrupted when DNA is denatured?

15. What does it mean to renature DNA?

Information Flow

1. A segment of one strand from a DNA molecule has the sequence 3′-TCCATGAGTTGA-5′. What is the sequence of nucleotides in the opposite, or complementary, DNA chain?

2, A segment of one strand from a DNA molecule has the sequence 3′-CCAGTGAATTGCCTAT-5′. List the base sequence in a new DNA strand produced by DNA polymerase. What process is this called?

3. A portion of the template strand of a gene has the sequence 3′-TCCATGAGTTGA-5′. What is the sequence of nucleotides in the RNA that is formed from this template? What process is this called?

4. A portion of the template strand of a gene has the sequence 3′-CCAGTGAATTGCCTAT-5′. What is the sequence of nucleotides in the RNA that is formed from this template? What process is this called?

5. Name three types of RNA and describe their function.

6. A portion of an mRNA molecule has the sequence 5′-AUGCCACGAGUUGAC-3′. What amino acid sequence does this code for? What process is this called?

7. A portion of an RNA molecule has the sequence 5′-AUGCUGAAUUGCGUAGGA-3′. What amino acid sequence does this code for? What process is this called?

8. The sequence dTATCCT appears in a template strand of DNA.

a. Which base is at the 3'-terminus of this primary structure?

b. What is the sequence in the DNA strand that is complementary to this sequence? Label the terminus.

c. What is the sequence in the RNA strand that is synthesized from this sequence? Label the terminus.

d. What is the sequence of the dipeptide from this strand?

e. Draw the dipeptide in the line bond structure at physiological pH. Label the terminuses.

f. In an aqueous environment will this dipeptide that is part of a larger protein likely to be buried inside the tertiary structure or located on the surface? Explain.

Mutation

1. Define the term mutation.

2. Are all mutations harmful to an individual? Explain.

3. Are all mutations passed on to an offspring?

4. What level of protein structure is affected by a mutation?

5. A mutation produces a DNA polymerase with lower than normal effectiveness at proofreading and repair. Could an individuals health be impacted by this mutation?

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CHAPTER OVERVIEW

11: Lipids - An Introduction

11.1: Fatty Acids
11.2: Ester Formation and Waxes
11.3: Triglycerides- Fats and Oils
11.4: Oxidation and Reduction Reactions of Triglycerides
11.5: Neutralization of Fatty Acids and Hydrolysis of Triglycerides
11.6: Steroids
11.7: Amphipathic Lipids and Membranes
11.8: Transport Across Cell Membrane
11.9: Digestion of Triglycerides
11.10: Fatty Acid Catabolism
11.11: End of Chapter Problems

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11.1: Fatty Acids

This page is a draft and is under active development.

Learning Objectives

- To recognize the structures of common fatty acids and classify them as saturated, monounsaturated, or polyunsaturated.

Fatty acids are carboxylic acids that are structural components of fats and oils. They usually contain an even number of carbon atoms (typically 12–20), and can be classified by the presence and number of carbon-to-carbon double bonds. Note the presence of a nonpolar hydrocarbon chain (tail) and a polar carboxylic acid functional group (head) in the molecule. A fatty acid with one carbon-carbon double bond is a *monounsaturated* fatty acid. A fatty acid with two or more double bonds is a *polyunsaturated* fatty acid.

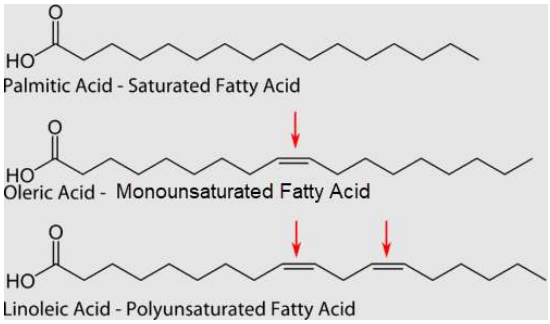


Figure **11.1.1**: Saturated fatty acids have only single bonds while monounsaturated have one double bond and polyunsaturated have more than one double bond.

Saturated Fatty Acids

Saturated fatty acids contain all carbon-carbon *single* bonds. This causes the molecules to form straight chains, as shown in the figure 11.1.2. Palmitic acid is an example of a saturated fatty acid. The hydrocarbon tails of saturated fatty acids are able to be packed together very tightly with one another, interacting through London forces. London forces depends on the surface area of the molecule. The longer the hydrocarbon tail in a fatty acid, the stronger the interactions between molecules. The melting points of saturated fatty acids increases as the molecules get larger. This explains why saturated fatty acids are solids at room temperature. Animals use saturated fatty acids to store energy.

Unsaturated Fatty Acids

In an **unsaturated fatty acids**, the hydrocarbon chain contains at least one carbon-carbon double bond (alkene). The double bonds are usually cis. The presence of a cis double bond causes chains to bend or kink in the hydrocarbon tail (see figure 11.1.2). The bent tails in neighboring unsaturated fatty acids are farther apart from one another compared to saturated fatty acids. This reduction of contact between the tails weakens the london forces between them. So unsaturated fatty acids are liquids at room temperature. For fatty acids with the same number of carbon atoms, the presence of more double bonds lowers the melting point. Plants use unsaturated fatty acids to store energy.

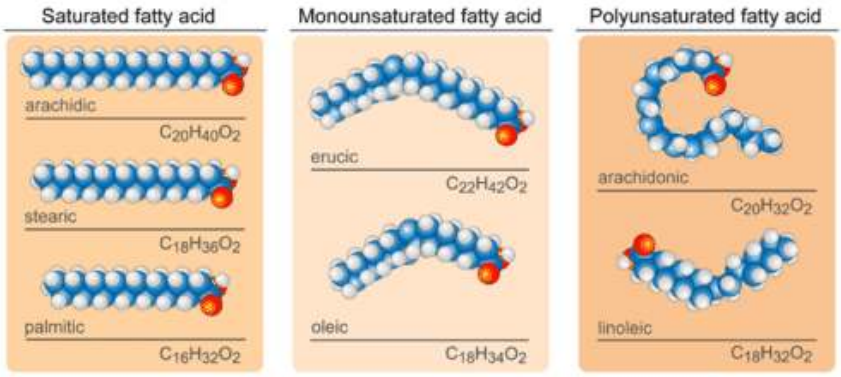


Figure **11.1.2**: Structures of saturated and unsaturated fatty acids.

Essential Fatty Acids

Two polyunsaturated fatty acids—linoleic and linolenic acids—are termed essential fatty acids because humans must obtain them from their diets. Both substances are required for normal growth and development, but the human body does not synthesize them. The average daily diet should contain about 4–6 g of the essential fatty acids.

Common Fatty Acids

Table **11.1.1** lists some common fatty acids. Chemists often use a shorthand notation for fatty acids along with the common name because they are all carboxylic acids with different numbers of carbons and number of double bonds. This notation gives the number of carbons followed by the number of double bonds present. For example: palmitic acid is a 16-carbon saturated fatty acid, would be represented by C16:0 and palmitoleic acid is a 16-carbon monounsaturated fatty acid, would be C16:1.

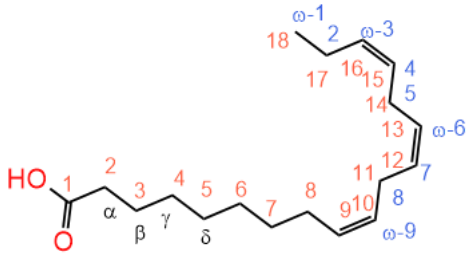
Table **11.1.1**: Some Common Fatty Acids Found in Natural Fats

Name	Number of Carbons	Number of Double Bonds	Condensed Structural Formula	Melting Point (°C)	Shorthand Notation	Source
lauric acid	12	0	$CH_3(CH_2)_{10}COOH$	44	C12:0	palm kernel oil
myristic acid	14	0	$CH_3(CH_2)_{12}COOH$	58	C14:0	oil of nutmeg
palmitic acid	16	0	$CH_3(CH_2)_{14}COOH$	63	C16:0	palm oil
palmitoleic acid	16	1	$CH_3(CH_2)_5CH=CH(CH_2)_7COOH$	0.5	C16:1	macadamia oil
stearic acid	18	0	$CH_3(CH_2)_{16}COOH$	70	C18:0	cocoa butter
oleic acid	18	1	$CH_3(CH_2)_7CH=CH(CH_2)_7COOH$	16	C18:1	olive oil
linoleic acid	18	2	$CH_3(CH_2)_3(CH_2CH=CH)_2(CH_2)_7COOH$	−5	C18:2	canola oil
linolenic acid	18	3	$CH_3(CH_2CH=CH)_3(CH_2)_7COOH$	−11	C18:3	flaxseed
arachidonic acid	20	4	$CH_3(CH_2)_4(CH=CHCH_2)_4(CH_2)_2COOH$	−50	C20:4	liver

ω-3 Fatty acids

Fats containing saturated fatty acids are considered not so healthy as they have a tendency to deposit in arteries forming plaque, i.e., atherosclerosis. It can lead to high blood pressure, rupture of arteries, or heart attack. Saturated fatty acids are more common in animal fats while unsaturated fatty acids which are considered more healthy are more common in vegetable oils.

In omega (ω) labeling the last C, i.e., the $-CH_3$ at the end opposite to carboxyl (−COOH) group is labeled ω-1, the C next to it as ω-2, then ω-3 and so on, as as shown in the figure 11.1.3. Omega (ω) is the last alphabet in Greek, so, ω-C is the last carbon. The (ω) label is used to refer to the position of the first C=C bond from the ω-end. For example, linolenic acid in figure below is an ω-3 fatty acid.



alpha-Linolenic acid

Figure **11.1.3**: Labelling of C's in fatty acids: systematic IUPAC numbers are in red, nonsystematic Greek letter labels are in black, and omega- (ω-#) are in blue. (Copyright; Public domain)

It has been observed that peoples from Alaska eat a higher amount of unsaturated fats and have a low occurrence of atherosclerosis and heart attack. Fats in Alaskan food is primarily from fish, such as salmon, tuna, and herring. Vegetable oils have higher contents of ω-6 fatty acids, such as linoleic acid and arachidonic acid. Fish oil mostly contains ω-3 fatty acids, such as linolenic acid, eicosapentaenoic acid, and docosahexaenoic acid.



salmon



tuna



herring



ω-3 fatty acids in fish

Figure **11.1.4**: Three common fish in the diet of people of Alaska and three common ω-3 fatty acids found in the fish: linolenic acid (top), eicosapentaenoic acid (middle), and docosahexaenoic acid (bottom). (Copyright: Public domain).

Fatty Acid Compositions of Common Fats and Oils

Table 11.1.2 shows the fatty acid compositions of some common fats and oils. The composition of any given fat or oil can vary depending on the plant or animal species it comes from as well as on dietetic and climatic factors. To cite just one example, lard from corn-fed hogs is more highly saturated than lard from peanut-fed hogs. Palmitic acid is the most abundant of the saturated fatty acids, while oleic acid is the most abundant unsaturated fatty acid. Terms such as *saturated fat* or *unsaturated oil* are often used to describe the fats or oils obtained from foods. Saturated fats contain a high proportion of saturated fatty acids, while unsaturated oils contain a high proportion of unsaturated fatty acids. The high consumption of saturated fats is a factor, along with the high consumption of cholesterol, in increased risks of heart disease.

Table **11.1.2**: Average Fatty Acid Composition of Some Common Fats and Oils (%)*

Source	Lauric (C12:0)	Myristic (C14:0)	Palmitic (C16:0)	Stearic (C18:0)	Oleic (C18:1)	Linoleic (C18:2)	Linolenic (C18:3)
Fats							
butter (cow)	3	11	27	12	29	2	1
tallow		3	24	19	43	3	1
lard		2	26	14	44	10	
Oils							
canola oil			4	2	62	22	10
coconut oil†	47	18	9	3	6	2	
corn oil			11	2	28	58	1
olive oil			13	3	71	10	1
peanut oil			11	2	48	32	
soybean oil			11	4	24	54	7

*Totals less than 100% indicate the presence of fatty acids with fewer than 12 carbon atoms or more than 18 carbon atoms.

†Coconut oil is highly saturated. It contains an unusually high percentage of the low-melting C_8, C_{10}, and C_{12} saturated fatty acids.

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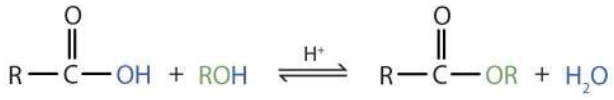
11.2: Ester Formation and Waxes

Learning Objectives

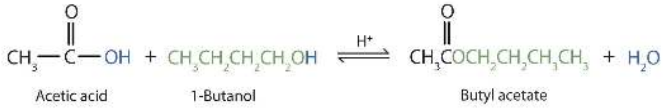
- Understand how esters are made.
- Identify the molecules used to make waxes.
- Describe the primary biological function of waxes.

Esterification

Esters can be prepared by esterification, a reaction in which a carboxylic acid and an alcohol, heated in the presence of an acid catalyst, form an ester and water through a dehydration followed by condensation reaction.



The reaction is reversible. As a specific example of an esterification reaction, butyl acetate can be made from acetic acid and 1-butanol.



Waxes

Lipid waxes are esters of one long-chain fatty acid with one unbranched long-chain alcohol containing 14 to 30 C's per chain. For example, myricyl palmitate found in beeswax is the ester of palmitic acid and myricyl alcohol.

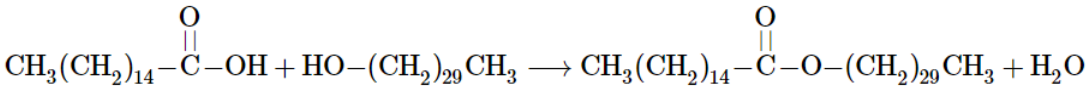


Figure **11.2.1**: palmitic acid and myricyl alcohol gives myricyl palmitate in beeswax

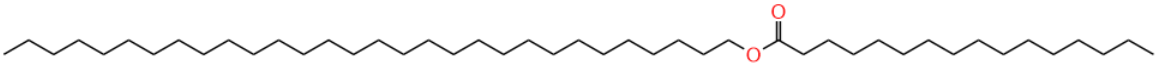
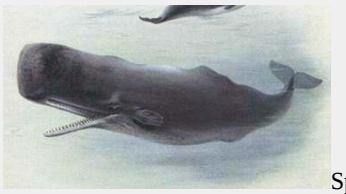


Figure **11.2.2**: A wax ester - Myricyl palmitate

Waxes form a protective and waterproof coating on leaves, fruits, and stems of plants and on the skin, fur, and feathers of animals. Plant waxes on the surfaces of leaves, stems, flowers, and fruits protect the plant from dehydration and invasion by harmful microorganisms. In animals the waxes serve as protective coatings, keeping the surfaces of feathers, skin, and hair pliable, water repellent and buoyant. In fact, if the waxy coating on the feathers of a water bird is dissolved as a result of the bird swimming in an oil slick, the feathers become wet and heavy, and the bird, unable to maintain its buoyancy, drowns. Some microorganisms use waxes as energy storage molecules.

Some of the most common lipid waxes are listed in Table 11.2.1. These include beeswax obtained from the honeycomb, carnauba wax from palm trees, jojoba wax from jojoba bushes, and spermaceti obtained from the head and cavity of sperm whales. Beeswax and carnauba wax are used as protective coating on cars, furniture, floors, etc. Jojoba wax and spermaceti are used to make candles and cosmetics. A mixture of waxes obtained from wool is used in lotions for face and skin softening.

Table 11.2.1: Common waxes

Wax name	Condensed formula	Source
Beeswax	$CH_3(CH_2)_{14}-\overset{\overset{O}{\|}}{C}-O-(CH_2)_{29}CH_3$	 Honeycomb
Carnauba wax	$CH_3(CH_2)_{24}-\overset{\overset{O}{\|}}{C}-O-(CH_2)_{29}CH_3$	 Palm tree
Jojoba wax	$CH_3(CH_2)_{18}-\overset{\overset{O}{\|}}{C}-O-(CH_2)_{19}CH_3$	 Jojoba bush
Spermaceti	$CH_3(CH_2)_{14}-\overset{\overset{O}{\|}}{C}-O-(CH_2)_{15}CH_3$	 Sperm whale

Petroleum Waxes and Earwax

Petroleum-derived waxes or paraffin are also called waxes but they are not esters, they are mixtures of long-chain hydrocarbons. Earwax is a mixture of esters of glycerol, esters of cholesterol, phospholipids, etc. from dead skin and from secretions of cerumen glands. Like other lipid waxes, it protects the ear canal against bacteria, fungi, and water. Excess earwax can cause blockage in the ear canal and hearing loss.

Summary

Esters are made by the reaction of a carboxylic acid with an alcohol, a process that is called esterification. Waxes are formed from long chain fatty acids and long chain alcohols.

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11.3: Triglycerides- Fats and Oils

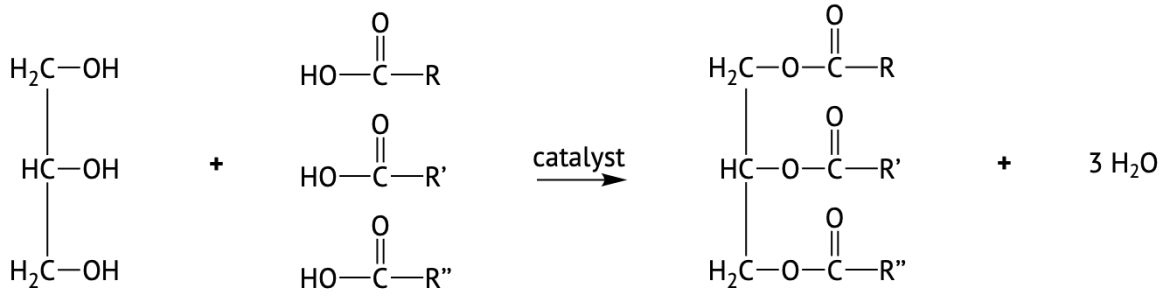
Learning Objectives

- Explain why fats and oils are referred to as triglycerides.
- Explain how the fatty acid composition of the triglycerides determines whether a substance is a fat or oil.

Fats and oils are the most abundant lipids in nature. They provide energy for living organisms, insulate body organs, and transport fat-soluble vitamins through the blood.

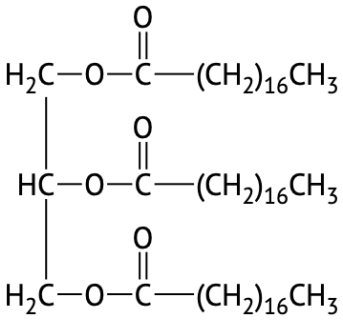
Structures of Triglycerides

Fats and oils are called triglycerides (or *triacylglycerols*) because they are esters composed of three fatty acid units joined to *glycerol*, a trihydroxy alcohol:

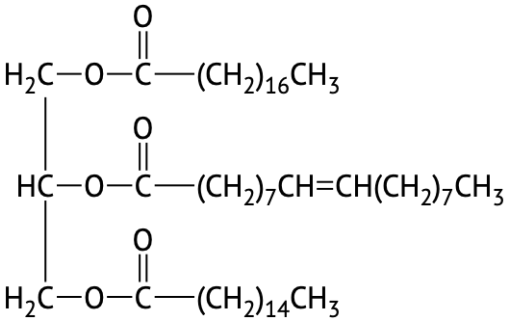


glycerol **3 fatty acids** **triacylglycerol**

If all three OH groups on the glycerol molecule are esterified with the same fatty acid, the resulting ester is called a ***simple triglyceride***. Although simple triglycerides have been synthesized in the laboratory, they rarely occur in nature. Instead, a typical triglyceride obtained from naturally occurring fats and oils contains two or three different fatty acid components and is thus termed a ***mixed triglyceride***.



Tristearin
a simple triglyceride



a mixed triglyceride

A triglyceride is called a **fat** if it is a solid at 25°C; it is called an **oil** if it is a liquid at that temperature. These differences in melting points reflect differences in the number of double bonds and number of carbon atoms in the constituent fatty acids. Triglycerides obtained from animal sources are usually fats, while those of plant origin are generally oils. Therefore, we commonly speak of animal fats and vegetable oils.

No single formula can be written to represent the naturally occurring fats and oils because they are highly complex mixtures of triglycerides in which many different fatty acids are represented. Table **11.3.1** shows the fatty acid compositions of some common fats and oils. The composition of any given fat or oil can vary depending on the plant or animal species it comes from, as well as on diet and climatic factors. To cite just one example, lard from corn-fed hogs is more highly saturated than lard from peanut-fed hogs. Palmitic acid is the most abundant of the saturated fatty acids, while oleic acid is the most abundant unsaturated fatty acid.

Table **11.3.1**: Average Fatty Acid Composition of Some Common Fats and Oils (%)*

	Lauric	Myristic	Palmitic	Stearic	Oleic	Linoleic	Linolenic
Fats							
butter (cow)	3	11	27	12	29	2	1
tallow		3	24	19	43	3	1
lard		2	26	14	44	10	
Oils							
canola oil			4	2	62	22	10
coconut oil†	47	18	9	3	6	2	
corn oil			11	2	28	58	1
olive oil			13	3	71	10	1
peanut oil			11	2	48	32	
soybean oil			11	4	24	54	7

*Totals less than 100% indicate the presence of fatty acids with fewer than 12 carbon atoms or more than 18 carbon atoms.

†Coconut oil is highly saturated. It contains an unusually high percentage of the low-melting C_8, C_{10}, and C_{12} saturated fatty acids.

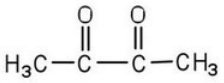
Terms such as *saturated fat* or *unsaturated oil* are often used to describe the fats or oils obtained from foods. Saturated fats contain a high proportion of saturated fatty acids, while unsaturated oils contain a high proportion of unsaturated fatty acids. The high consumption of saturated fats is a factor, along with the high consumption of cholesterol, in increased risks of heart disease.

Biological Roles of Triglycerides

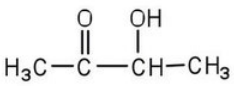
One of the primary biological roles of triglycerides is to provide energy. Gram for gram, they pack more than twice the caloric content of carbohydrates and proteins. The oxidation of fats and oils supplies about 9 kcal of energy for every gram oxidized, whereas the oxidation of carbohydrates or proteins supplies 4 kcal/g. Although the high caloric content of fats may be bad news for the dieter, it says something about the efficiency of nature's designs. Our bodies use carbohydrates, primarily in the form of glucose, for our *immediate* energy needs. Our capacity for storing carbohydrates for later use is limited to tucking away a bit of glycogen in the liver or in muscle tissue. We store our *reserve* energy in lipid form. Triglycerides have other biological functions besides energy storage. They serve as protective padding and insulation for vital organs. The layers of fats in seals, penguins, and other aquatic animals provide thermal insulation and bouyancy in water. Furthermore, without lipids in our diets, we would be deficient in the fat-soluble vitamins A, D, E, and K since they transport fat-soluble vitamins through the blood..

Physical Properties of Fats and Oils

Contrary to what you might expect, *pure* fats and oils are colorless, odorless, and tasteless. The characteristic colors, odors, and flavors that we associate with some of them are imparted by foreign substances that are lipid soluble and have been absorbed by these lipids. For example, the yellow color of butter is due to the presence of the pigment carotene; the taste of butter comes from two compounds—diacetyl and 3-hydroxy-2-butanone—produced by bacteria in the ripening cream from which the butter is made.



Diacetyl



3-hydroxy-2-butanone

Fats and oils are lighter than water, having densities of about 0.8 g/cm^3. They are poor conductors of heat and electricity and therefore serve as excellent insulators for the body, slowing the loss of heat through the skin. Since they contain 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 composed of alkanes such as hexane ($CH_3CH_2CH_2CH_2CH_2CH_3$) or octane ($CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$).

Summary

Fats and oils are molecules known as triglycerides, which are esters of three fatty acids linked to glycerol. An increase in the percentage of shorter-chain fatty acids and/or unsaturated fatty acids lowers the melting point of a triglyceride.

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11.4: Oxidation and Reduction Reactions of Triglycerides

Learning Objectives

- Describe the importance of key reactions of triglycerides such as oxidation and reduction of alkenes.
- Understand the origin and consequences of trans fats in our diet.

Triglycerides contain alkene and ester functional groups. In this section we will study two reactions based on the alkene functional group.

- Oxidation of alkenes
- Reduction of alkenes

Oxidation of Alkenes

When an akene is exposed to air, a slow oxidation reaction takes place. Atmospheric O_2 is responsible for oxidation and breaking of the double bond. The breakdown products are aldehydes and carboxylic acids as shown in figure 11.4.1.

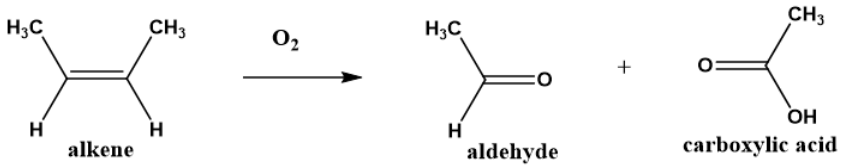


Figure **11.4.1**: Oxidation of an alkene double bond to form aldehydes and carboxylic acids.

Triglycerides (fats and oils) that are in contact with moist air undergo oxidation of the alkene functional group and hydrolysis of esters. The breakdown products are small organic molecules that are aldehydes and carboxylic acids, that have unpleasant odors. When this happens, the triglyceride has spoiled or has gone rancid. Vegetable oils have a greater number of double bonds and so spoil more quickly compared to fats. See figure 11.4.2.

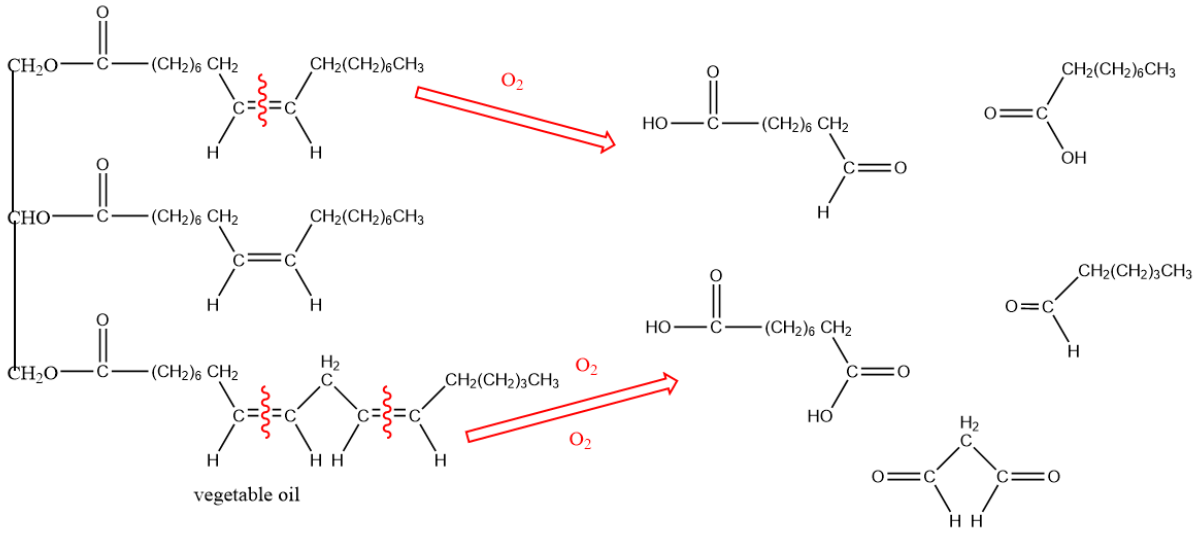


Figure **11.4.2**: Oxygen in air can cause the oxidation of the alkenes in a triglyceride, producing small organic molecules that have unpleasant odors.

To make the oils last longer you can keep them refrigerated. The rate of the oxidation reaction is slower at lower temperatures. The oils may be stored in tightly capped bottles to lower the exposure to O_2. Antioxidant molecules slow the oxidation of alkenes may

also be added to the oils to increase shelf life. The oils are often stored in dark bottles to limit exposure to light. Removing some of the double bonds by partial hydrogenation also slows rancidity or spoilage.

Reduction of Alkenes or Catalytic Hydrogenation

Alkenes react with H_2 in presence of a catalyst to alkanes as shown in figure 11.4.3.

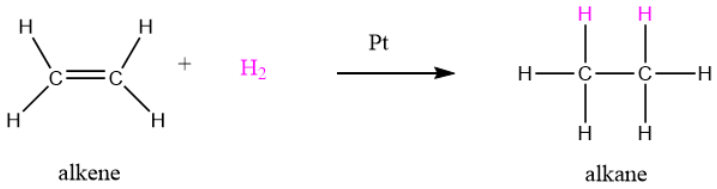
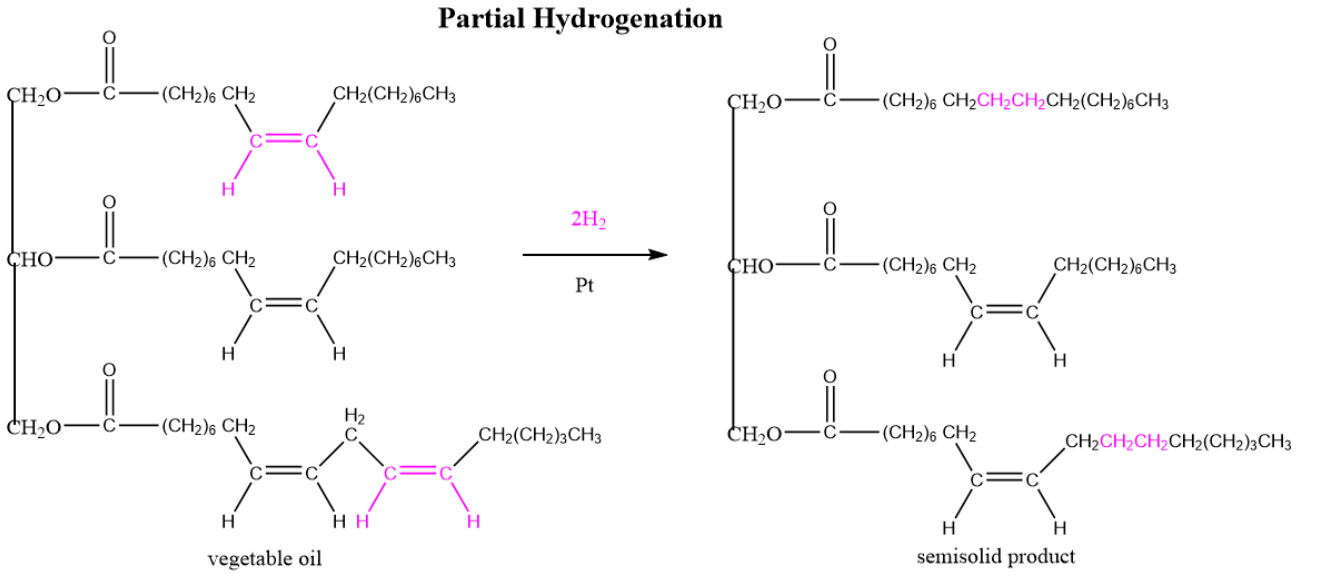


Figure **11.4.3**: Reduction of an alkene to an alkane.

The alkene functional groups in unsaturated triglycerides can undergo the same catalytic reduction reaction of figure 11.4.3. Commercially during partial hydrogenation inexpensive and abundant vegetable oils (canola, corn, soybean) is treated with H_2 and Pt, However the reaction is halted before all the double bonds are removed. Figure 11.4.4 shows the two double bonds in purple are hydrogenated. Partial hydrogenation converts vegetable oils into a semisolid product which has a good consistency for cooking but does not contain the cholesterol that is present in fats from animal sources. With fewer double bonds the partially hydrogenated vegetable oils do not oxidize (spoil) as rapidly.



x

Figure **11.4.4**: Partial Reduction of a triglyceride.

Trans fats

During the partial hydrogenation of vegetable oils an isomerization reaction occurs which converts a cis-double bond of a triglyceride to a trans-double bond as shown in figure 11.4.5.

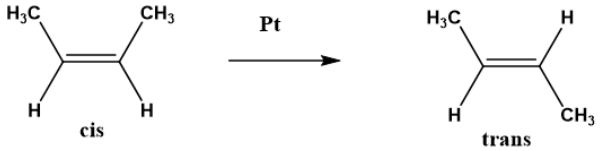


Figure 11.4.5: Cis to trans isomer formation during partial hydrogenation.

A cis-double bond in a fatty acid introduces a kink in the structure. This kink or bend in the hydrocarbon part does not let the molecules pack with a large contact area between neighboring molecules. The London Forces are weaker in a triglyceride with cis-double bonds. Trans-fatty acids do not have a bend in the structure, and thus pack closely together in the same way that the saturated fats do as shown in figure 11.4.6.

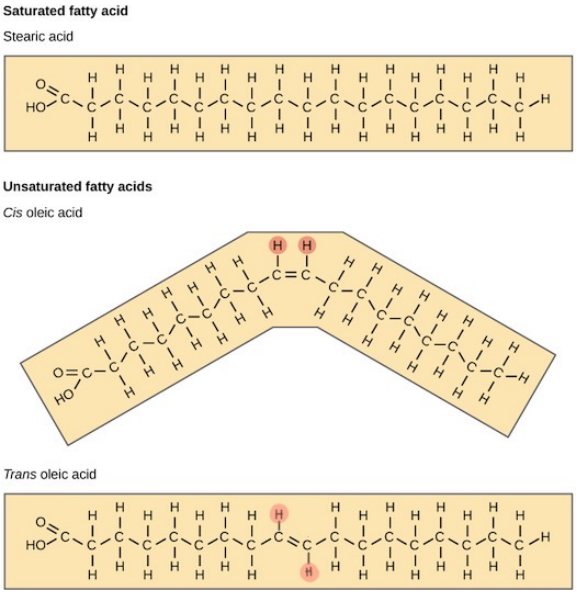


Figure 11.4.6: Comparison of a trans fatty acid with a saturated and a cis fatty acid.

Studies have shown that trans fats increase the incidence of heart disease. They raise your LDL (bad cholesterol) and lower your HDL (good cholesterol). Too much trans fats in your diet increases the risk of developing heart disease and stroke. It's also associated with a higher risk of developing type 2 diabetes.

Naturally-occurring trans fats are produced in the gut of some animals (milk and meat products) may contain small quantities of these fats.

Summary

The oxidation of the alkene functional group in a triglyceride can form compounds with a disagreeable odor. Double bonds present in unsaturated triglycerides can be hydrogenated to convert oils to a semisolid product like margarine. During the process of partial hydrogenation some of the cis double bonds may be isomerized to trans.

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11.5: Neutralization of Fatty Acids and Hydrolysis of Triglycerides

Learning Objectives

- Describe how carboxylic acids react with basic compounds.
- Identify the products of an acidic or basic hydrolysis of an ester.
- Understand the interactions of soap and detergent with hard water.

Neutralization Reaction of a Carboxylic Acid with NaOH

Carboxylic acids react with aqueous solutions of sodium hydroxide (NaOH) to form salts and water in a double replacement reaction. This reaction is an acid-base reaction known as neutralization reaction.

$$CH_3COOH + NaOH(aq) \rightarrow CH_3COO^-Na^+(aq) + H_2O$$

The salt is the conjugate base of the carboxylic acid is called the carboxylate ion ($CH_3COO^-Na^+$).

Carboxylate ions ($RCOO^-Na^+$) are more soluble in water than carboxylic acids (RCOOH).

Neutralization Reaction of a Fatty Acid with NaOH

Fatty acids are long chain carboxylic acids. They are insoluble in water and hence hydrophobic. When treated with NaOH, fatty acids are converted into carboxylate ion salts (ionic compounds) containing Na^+ and amphipathic fatty acid anion.

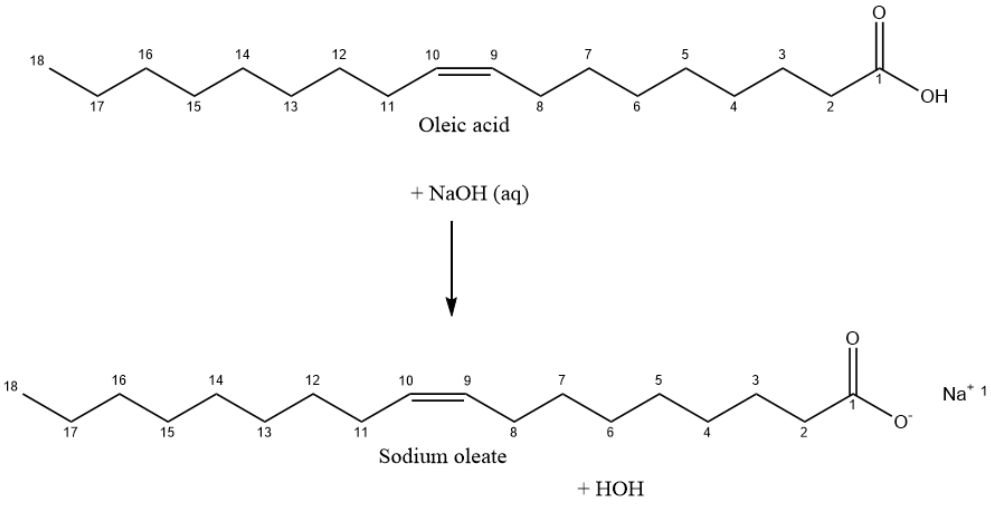


Figure 11.5.1: Oleic acid (hydrophobic), when reacted with aq NaOH, it is converted to an ionic compound sodium oleate (amphipathic)

Like all fatty acids, oleic acid is hydrophobic. In figure 11.5.1 oleic acid is converted into sodium oleate. Sodium oleate is soap. In aqueous solutions, soaps will spontaneously form **micelles**, a spherical structure that allows the hydrophobic tails to avoid contact with water and simultaneously form cluster in the center via London dispersion forces. The hydrophilic heads make up the outer surface and dissolves in water.

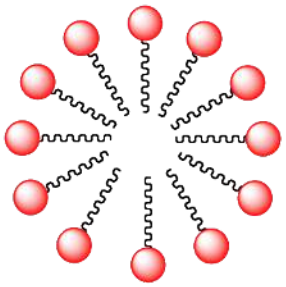


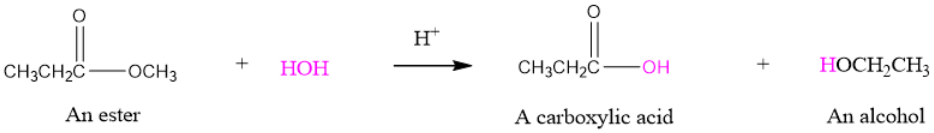
Figure **11.5.2**: Micelle

Dirt and grime usually adhere to skin and clothing by combining with body oils, cooking fats, and lubricating greases. These substances are nonpolar and so they are not miscible in water. Washing with water alone will not remove them. These nonpolar oils and grease along with dirt will be trapped in the hydrophobic interior of the micelle of soap, while the micelle is easily rinsed away with water.

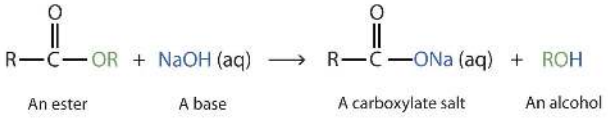
Hydrolysis of Esters

Hydrolysis is "splitting with water." Hydrolysis is the most important reaction of esters. The hydrolysis of esters is catalyzed by either an acid or a base. During acid catalysed hydrolysis the ester is heated with water containing a small amount of a strong-acid catalyst such as sulfuric acid. The sulfuric acid is a source of the hydrogen ion (H^+), the catalyst in the reaction.

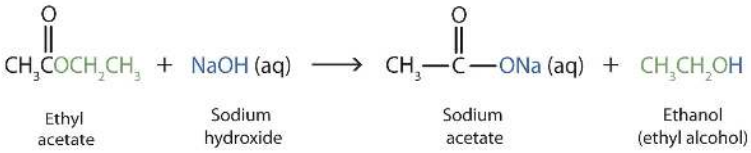
Acidic hydrolysis of an ester gives a carboxylic acid and an alcohol. Note that the -OH of the carboxylic acid and the H of the alcohol comes from water (shown in purple).



When a base (such as sodium hydroxide [NaOH] or potassium hydroxide [KOH]) is used to hydrolyze an ester, the products are a carboxylate salt and an alcohol.



As a specific example, ethyl acetate and NaOH react to form sodium acetate and ethanol:



Because soaps are prepared by the alkaline hydrolysis of fats and oils, alkaline hydrolysis of esters is called saponification (Latin *sapon*, meaning "soap," and *facere*, meaning "to make"). In a saponification reaction, the base is a reactant and a catalyst.

Saponification of Fats and Oils

Fats and oils are esters, they can be hydrolyzed in the presence of an acid, a base, or specific enzymes known as lipases. The hydrolysis of fats and oils in the presence of a base is used to make soap and is called saponification. Soaps are prepared through the hydrolysis of triglycerides (often from tallow, coconut oil, or both) using water and sodium hydroxide.

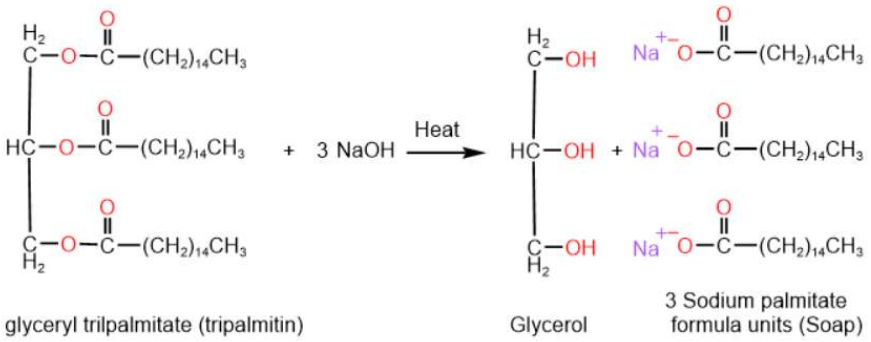


Figure 11.5.3: Saponification of a triglyceride.

Sodium salts of fatty acids make hard soaps while potassium salts make soft soaps. Three of the most common soaps are sodium stearate, sodium oleate, and sodium linoleate. Salts of saturated fatty acids make hard soap and those of polyunsaturated fatty acids make soft soap. Perfumes are added for scent and dyes for color, sand is added in scouring soap, and the air is blown into the soap to make it float on water.

Looking Closer: Soaps

Ordinary soap is a mixture of the sodium salts of various fatty acids, produced in one of the oldest organic syntheses practiced by humans (second only to the fermentation of sugars to produce ethyl alcohol). Both the Phoenicians (600 BCE) and the Romans made soap from animal fat and wood ash. Even so, the widespread production of soap did not begin until the 1700s. Soap was traditionally made by treating molten lard or tallow with a slight excess of alkali in large open vats. The mixture was heated, and steam was bubbled through it. After saponification was completed, the soap was precipitated from the mixture by the addition of sodium chloride (NaCl), removed by filtration, and washed several times with water. It was then dissolved in water and reprecipitated by the addition of more NaCl. The glycerol produced in the reaction was also recovered from the aqueous wash solutions.

Pumice or sand is added to produce scouring soap, while ingredients such as perfumes or dyes are added to produce fragrant, colored soaps. Blowing air through molten soap produces a floating soap. Soft soaps, made with potassium salts, are more expensive but produce a finer lather and are more soluble. They are used in liquid soaps, shampoos, and shaving creams.

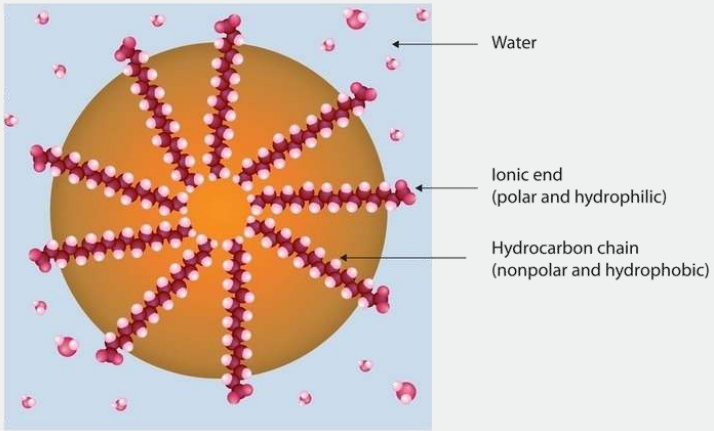


Figure 11.5.4: Micelle of Soap.

Dirt and grime usually adhere to skin, clothing, and other surfaces by combining with body oils, cooking fats, lubricating greases, and similar substances that act like glues. Because these substances are not miscible in water, washing with water alone does little to remove them. Soap removes them, because soap molecules have a dual nature. One end, called the *head*, carries an ionic charge (a carboxylate anion) and therefore dissolves in water; the other end, the *tail*, has a hydrocarbon structure and dissolves in oils. The hydrocarbon tails dissolve in the soil; the ionic heads remain in the aqueous phase. With the oil no longer "gluing" the dirt to the soiled surface (skin, cloth, dish), the soap-enclosed dirt can easily be rinsed away.

Hard water

Water is considered hard water, if it contains soluble Ca^{2+}, Mg^{2+} and Fe^{3+} ions. These ions are present in water as soluble bicarbonate compounds. If these ions are removed or absent, the water is considered soft water. Hard water can be purified by using water softeners or ion-exchange resins that replace Ca^{2+} and Mg^{2+} with Na^{+} ions.

Soaps

A soap is the sodium or potassium salt of a fatty acid. The tail part of the soap molecule is nonpolar, but the polar "head" at the carboxyl end permits a small portion of the molecule to be water soluble. Hence soaps are amphipathic. The cleansing power of soap is dependent on this water solubility. However, hard water containing Ca^{2+} and Mg^{2+} reacts with soap and forms insoluble scales (scum) with soap. Figure 11.5.5 shows the formation of insoluble precipitate of soap with Ca^{2+} ion in hard water. This reduces the cleansing properties of soap since less soap molecules are available to form the micelle.

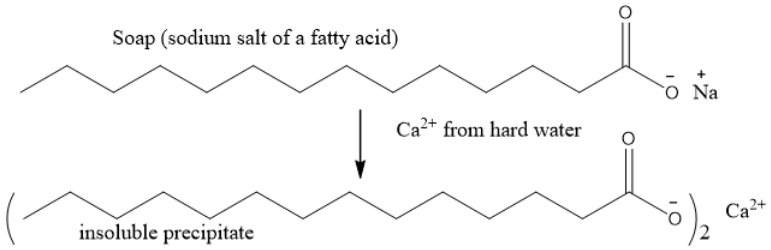


Figure 11.5.5: Formation of insoluble scales of soap with calcium ions in hard water.

Detergents

A detergent is the sodium salt of a sulfonic acid. It has similar cleansing properties to soap since detergents can form micelles just like soap. However, detergents differ from soaps in its interactions with hard water. The calcium and magnesium ions from hard water interact with a detergent and from compounds that are soluble in water. Figure 11.5.6 shows the formation of a soluble compound of detergent with Ca^{2+} ion in hard water. So, calcium and magnesium ions from hard water have no effect on the cleansing properties of a detergent.

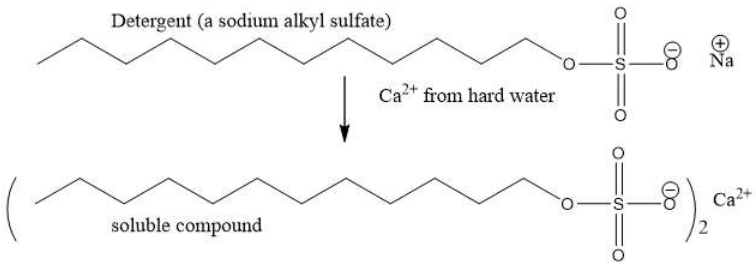


Figure 11.5.6: Formation of soluble compound of detergent with calcium ions in hard water.

Summary

Fatty acids can react with bases to form an amphipathic compound called soap. Soaps are made by the hydrolysis reaction of a triglyceride with an aqueous solution of a base.

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11.6: Steroids

Learning Objectives

- To identify the functions of steroids produced in mammals.

Triglycerides are *saponifiable*. That means triglycerides can react with aqueous base (alkali) to yield simpler components, such as glycerol and fatty acid salts (soap). Lipid samples extracted from cellular material, however, also contain a small but important fraction that does not react with alkali. Such a lipid is non saponifiable. The most important nonsaponifiable lipids are the steroids. These compounds include the *bile salts*, *cholesterol*, and certain *hormones* (such as cortisone and the sex hormones).

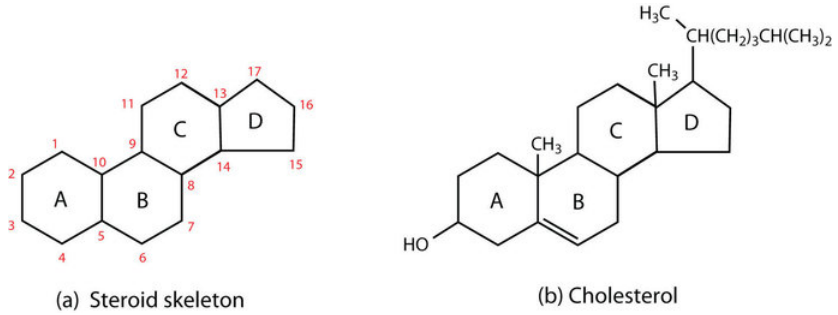


Figure **11.6.1** Steroids. (a) The four-fused-ring steroid skeleton uses letter designations for each ring and the numbering of the carbon atoms. (b) The cholesterol molecule follows this pattern.

Steroids occur in plants, animals, yeasts, and molds but not in bacteria. They may exist in free form or combined with fatty acids or carbohydrates. All steroids have a characteristic structural component consisting of four fused rings. Chemists identify the rings by capital letters and number the carbon atoms as shown in Figure **11.6.1*a***. Slight variations in this structure or in the atoms or groups attached to it produce profound differences in biological activity.

Cholesterol

Cholesterol (Figure **11.6.1*b***) does not occur in plants, but it is the most abundant steroid in the human body (240 g is a typical amount). Excess cholesterol is believed to be a primary factor in the development of atherosclerosis and heart disease, which are major health problems in the United States today. About half of the body's cholesterol is interspersed in the lipid bilayer of cell membranes. Much of the rest is converted to cholic acid, which is used in the formation of bile salts. Cholesterol is also a precursor in the synthesis of sex hormones, adrenal hormones, and vitamin D.

Excess cholesterol not metabolized by the body is released from the liver and transported by the blood to the gallbladder. Normally, it stays in solution there until being secreted into the intestine (as a component of bile) to be eliminated. Sometimes, however, cholesterol in the gallbladder precipitates in the form of gallstones (Figure **11.6.2**). Indeed, the name *cholesterol* is derived from the Greek *chole*, meaning "bile," and *stereos*, meaning "solid."



Figure **11.6.2**: Numerous small gallstones made up largely of cholesterol, all removed in one patient. Grid scale 1 mm

To Your Health: Cholesterol and Heart Disease

Heart disease is the leading cause of death in the United States for both men and women. The Centers for Disease Control and Prevention reported that heart disease claimed 631,636 lives in the United States (26% of all reported deaths) in 2006.

Scientists agree that elevated cholesterol levels in the blood, as well as high blood pressure, obesity, diabetes, and cigarette smoking, are associated with an increased risk of heart disease. A long-term investigation by the National Institutes of Health showed that among men ages 30 to 49, the incidence of heart disease was five times greater for those whose cholesterol levels were above 260 mg/100 mL of serum than for those with cholesterol levels of 200 mg/100 mL or less. The cholesterol content of blood varies considerably with age, diet, and sex. Young adults average about 170 mg of cholesterol per 100 mL of blood, whereas males at age 55 may have cholesterol levels at 250 mg/100 mL or higher because the rate of cholesterol breakdown decreases with age. Females tend to have lower blood cholesterol levels than males.

To understand the link between heart disease and cholesterol levels, it is important to understand how cholesterol and other lipids are transported in the body. Lipids, such as cholesterol, are not soluble in water and therefore cannot be transported in the blood (an aqueous medium) unless they are complexed with proteins that are soluble in water, forming assemblages called *lipoproteins*. Lipoproteins are classified according to their density, which is dependent on the relative amounts of protein and lipid they contain. Lipids are less dense than proteins, so lipoproteins containing a greater proportion of lipid are less dense than those containing a greater proportion of protein.

Research on cholesterol and its role in heart disease has focused on serum levels of low-density lipoproteins (LDLs) and high-density lipoproteins (HDLs). One of the most fascinating discoveries is that high levels of HDLs reduce a person's risk of developing heart disease, whereas high levels of LDLs increase that risk. Thus the serum LDL:HDL ratio is a better predictor of heart disease risk than the overall level of serum cholesterol. Persons who, because of hereditary or dietary factors, have high LDL:HDL ratios in their blood have a higher incidence of heart disease.

How do HDLs reduce the risk of developing heart disease? No one knows for sure, but one role of HDLs appears to be the transport of excess cholesterol to the liver, where it can be metabolized. Therefore, HDLs aid in removing cholesterol from blood and from the smooth muscle cells of the arterial wall.

LDL's transport cholesterol from the liver to the cells where they are incorporated into membrane or transformed to other steroids. Therfore high LDL levels are warning signs of atherosclerosis, the build up of cholesterol containing deposits in the arteries.

Dietary modifications and increased physical activity can help lower total cholesterol and improve the LDL:HDL ratio. The average American consumes about 600 mg of cholesterol from animal products each day and also synthesizes approximately 1 g of cholesterol each day, mostly in the liver. The amount of cholesterol synthesized is controlled by the cholesterol level in the blood; when the blood cholesterol level exceeds 150 mg/100 mL, the rate of cholesterol biosynthesis is halved. Hence, if cholesterol is present in the diet, a feedback mechanism suppresses its synthesis in the liver. However, the ratio of suppression is not a 1:1 ratio; the reduction in biosynthesis does not equal the amount of cholesterol ingested. Thus, dietary substitutions of unsaturated fat for saturated fat, as well as a reduction in consumption of *trans* fatty acids, is recommended to help lower serum cholesterol and the risk of heart disease.

Steroid Hormones

Hormones are chemical messengers that are released in one tissue and transported through the circulatory system to one or more other tissues. They are molecules that regulate the function of organs and tissues. One group of hormones is known as steroid hormones because these hormones are synthesized from cholesterol, which is also a steroid. There are two main groups of steroid hormones: adrenocortical hormones and sex hormones.

Adrenocorticoid Hormones

The adrenocortical hormones, such as cortisone and cortisol are produced by the adrenal gland, which is located adjacent to each kidney. They are produced from cholesterol through progesterone. These compounds are anti-inflammatory agents. Cortisol and its analogs are therefore used pharmacologically as immunosuppressants after transplant operations and in the treatment of severe skin allergies and autoimmune diseases, such as rheumatoid arthritis.

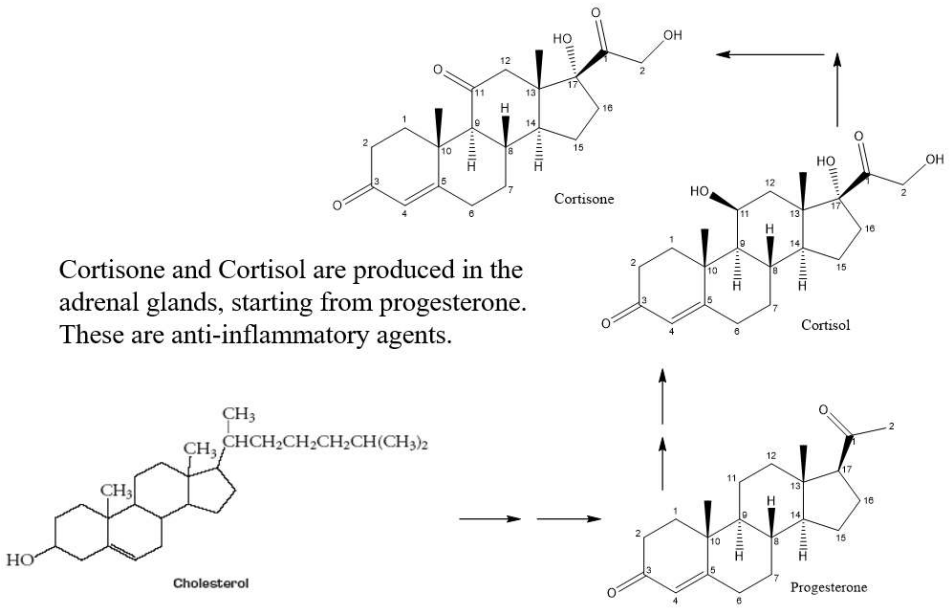


Figure **11.6.3** Manufacture of cortisol and cortisone starts from cholesterol.

Sex Hormones

The sex hormones are a class of steroid hormones secreted by the gonads (ovaries or testes), the placenta, and the adrenal glands. They are produced from cholesterol by the shortening of the hydrocarbon chain and alterations to the ring producing progesterone. Progesterone is then converted through several steps to testosterone and estradiol.Testosterone is a male sex hormones, controlling the primary sexual characteristics of males. Two kinds of sex hormones are of particular importance in females: progesterone, which prepares the uterus for pregnancy and prevents the further release of eggs from the ovaries during pregnancy, and the estradiol, which are mainly responsible for the development of female secondary sexual characteristics. Both males and females produce these hormones, differing in the amounts of secreted hormones rather than in the presence or absence of one or the other. Cholesterol is the starting material for the production of these hormones.

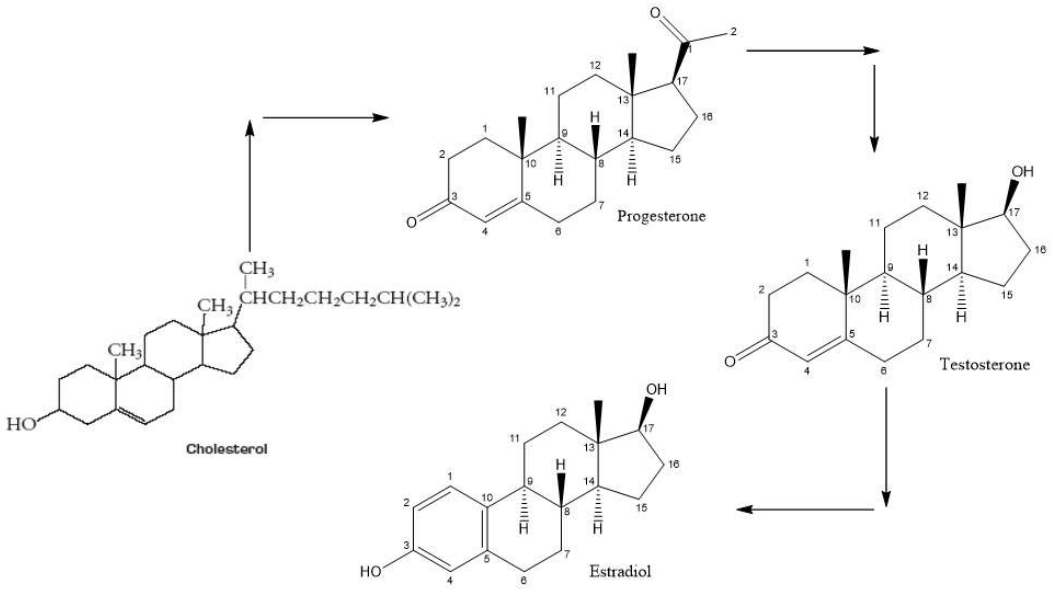
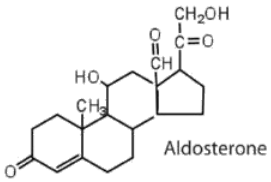
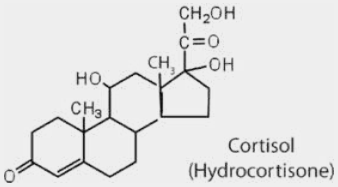
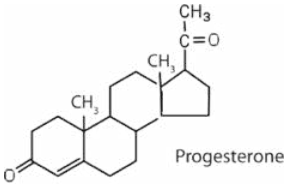
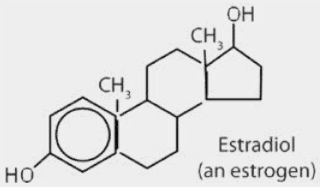
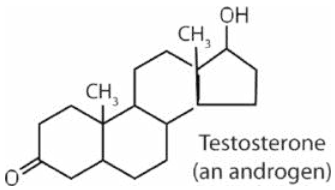


Figure **11.6.4** Cholesterol is the starting material for the synthesis of the sex hormpnes.

Table 11.6.1: Representative Steroid Hormones and Their Physiological Effects

Hormone	Effect
 Aldosterone	regulates salt metabolism; stimulates kidneys to retain sodium and excrete potassium
 Cortisol (Hydrocortisone)	stimulates the conversion of proteins to carbohydrates
 Progesterone	regulates the menstrual cycle; maintains pregnancy
 Estradiol (an estrogen)	stimulates female sex characteristics; regulates changes during the menstrual cycle
 Testosterone (an androgen)	stimulates and maintains male sex characteristics

Sex hormones, both natural and synthetic, are sometimes used therapeutically. For example, a woman who has had her ovaries removed may be given female hormones to compensate. Some of the earliest chemical compounds employed in cancer chemotherapy were sex hormones. For example, estrogens are one treatment option for prostate cancer because they block the release and activity of testosterone. Testosterone enhances prostate cancer growth. Sex hormones are also administered in preparation for sex-change operations, to promote the development of the proper secondary sexual characteristics. Oral contraceptives are synthetic derivatives of the female sex hormones; they work by preventing ovulation.

Bile Salts

Bile is a yellowish green liquid (pH 7.8–8.6) produced in the liver. The most important constituents of bile are bile salts, which are sodium salts of glycocholate and taurocholate. They are synthesized from cholesterol in the liver, stored in the gallbladder, and then secreted in bile into the small intestine.

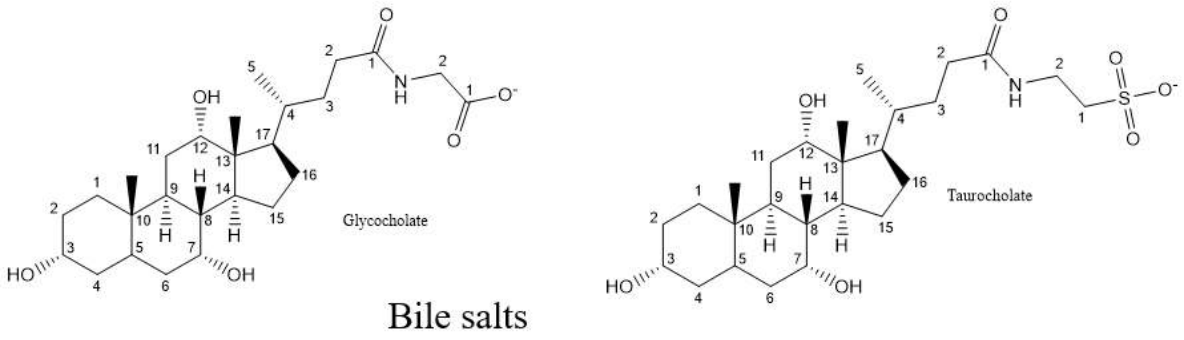


Figure **11.6.5** Glycocholate and Taurocholate are examples of bile salts.

Because they contain both hydrophobic and hydrophilic groups, bile salts are amphipathic and highly effective emulsifying agents; they break down large fat globules into smaller ones and keep those smaller globules suspended in the aqueous digestive environment. Enzymes can then hydrolyze fat molecules more efficiently. Thus, the major function of bile salts is to aid in the digestion of dietary lipids.

> Surgical removal is often advised for a gallbladder that becomes infected, inflamed, or perforated. This surgery does not seriously affect digestion because bile is still produced by the liver, but the liver's bile is more dilute and its secretion into the small intestine is not as closely tied to the arrival of food.

Summary

Steroids have a four-fused-ring structure and have a variety of functions. Cholesterol is a steroid found in mammals that is needed for the rigidity of cell membranes, synthesis of bile acids, and several hormones. Bile salts are secreted into the small intestine to aid in the digestion of fats.

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11.7: Amphipathic Lipids and Membranes

Learning Objectives

- Describe the structure of membrane lipids
- Describe membrane components and how they are arranged.

Cell Membrane

All living cells are surrounded by a cell membrane. Plant cells and animal cells contain a cell nucleus that is also surrounded by a membrane and holds the genetic information for the cell. Everything between the cell membrane and the nuclear membrane, including intracellular fluids and various subcellular components such as the mitochondria and ribosomes is called the cytoplasm.

Lipids found in Membranes

The lipids in cell membranes are amphipathic. They have dual characteristics: part of the lipid is ionic and therefore dissolves in water, whereas the rest has a hydrocarbon structure and therefore dissolves in nonpolar substances. The ionic part is referred to as hydrophilic, meaning “water loving,” and the nonpolar part as hydrophobic, meaning “water fearing” (repelled by water). The amphipathic lipids commonly found in membranes are phospholipids and glycolipids.

Phospholipids are lipids containing phosphorus. The phosphate ion (PO_4^{3-}) is one of the components used in their formation. It is a major component of cell membranes. A phospholipid consists of a hydrophilic (water-loving) head and hydrophobic (water-fearing) tail (see figure 11.7.1). The phospholipid is similar to a triglyceride in which a fatty acid has been replaced by a phosphate group and an amine alcohol.

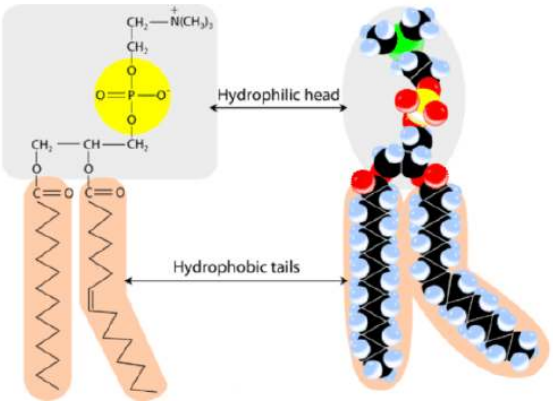


Figure 11.7.1: A phospholipid consists of a head and a tail. The "head" of the molecule contains the phosphate group and is hydrophilic, meaning that it will dissolve in water. The "tail" of the molecule is made up of two fatty acids, which are hydrophobic and do not dissolve in water.

The phospholipid drawn above is a **lecithin**. Lecithins occur in all living organisms. Egg yolks are especially rich in lecithins. Commercial-grade lecithins isolated from soybeans are widely used in foods as emulsifying agents. An **emulsifying agent** is used to stabilize an emulsion—a dispersion of two liquids that do not normally mix, such as oil and water. Many foods are emulsions. Milk is an emulsion of butterfat in water. The emulsifying agent in milk is a protein called *casein*. Mayonnaise is an emulsion of oil in water, stabilized by lecithins present in egg yolk.

Glycolipids are sugar-containing lipids. These lipids contain sphingosine rather than a glycerol and hence may be addressed as sphingolipids. The structure of sphingosine is shown in figure 11.7.2 to the left.The glycolipids are found on the outer surface of the cell membrane, acting as distinguishing surface markers for the cell and thus serving in cellular recognition and cell-to-cell communication. Shown in figure 11.7.2 is a glycolipid composed of sphingosine, a fatty acid, and galactose. Sphingolipids are important constituents of the membranes of nerve and brain cells.

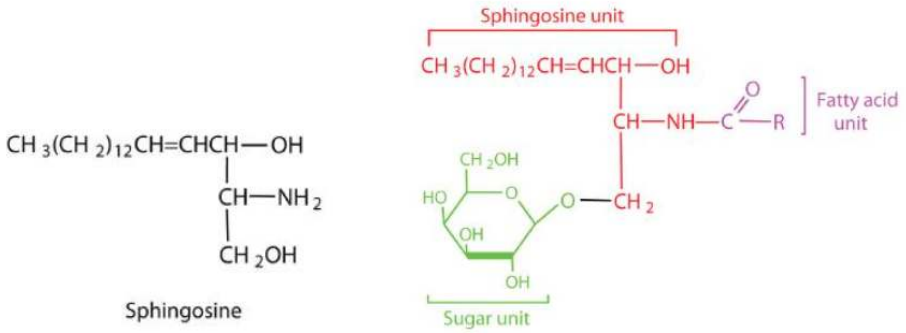


Figure **11.7.2**: Sphingosine is an amino alcohol found in all sphingolipids, Glycolipids contain a sugar unit.

Sphingolipids may be phospholipids or glycolipids that contain the unsaturated amino alcohol sphingosine rather than glycerol. Shown below in figure 11.7.3 is the structure of a phospholipid that is also a sphingolipid.

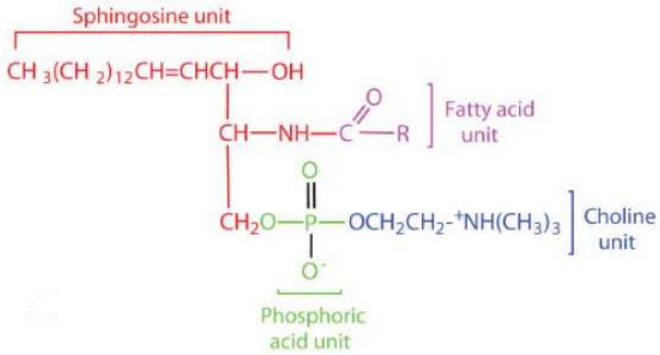


Figure **11.7.3**: A sphingolipid may also be a phospholipid, as evidenced by the phosphoric acid unit in its structure.

Lipid Bilayers

The membranes of all cells have a fundamentally similar structure. Bilayers shown in figure 11.7.4 are double layers of lipids arranged so that the hydrophobic tails are sandwiched between an inner surface and an outer surface consisting of hydrophilic heads. The hydrophilic heads are in contact with water on either side of the bilayer, whereas the tails, are inside the bilayer, are prevented from having contact with the water.

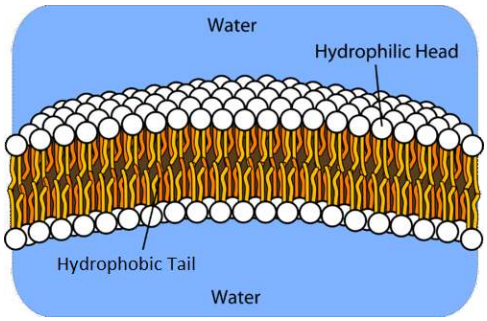


Figure **11.7.4**: In a water solution, phospholipids form a bilayer where the hydrophobic tails point towards each other on the interior and only the hydrophilic heads are exposed to the water.

In the bilayer interior, the hydrophobic tails (that is, the fatty acid portions of lipid molecules) interact by means of London's forces. The interactions are weakened by the presence of unsaturated fatty acids. As a result, the membrane components are free to move about to some extent, and the membrane is described as fluid. Cholesterol deposits in the bilayers adds rigidity to the membrane structure. The fused ring portion of cholesterol is rigid and inflexible. Bilayers like this make up every cell membrane.

Very few ions or polar molecules could pass through their hydrophobic “sandwich filling” to enter or leave any cell. Ions and polar species do cross the membrane, aided by proteins that are present in the lipid bilayer. Some proteins span the hydrophobic interior of the bilayer. Small ions and molecules soluble in water enter and leave the cell by way of channels through these proteins. Other proteins are more loosely associated with the surface of the lipid bilayer like tiles in a mosaic. This is known as the **fluid mosaic model** of the membrane structure.

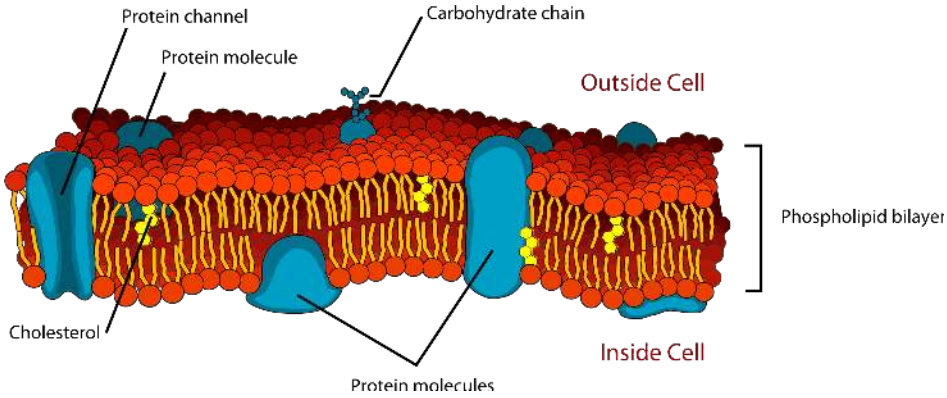


Figure **11.7.5**: The phospholipid bilayer of a cell membrane contains embedded protein molecules which allow for selective passage of ions and molecules through the membrane.

Summary

Lipids are important components of biological membranes. The membrane lipids are amphipathic: part of the molecule is hydrophilic, and part of the molecule is hydrophobic. Membrane lipids may be classified as phospholipids, glycolipids, and/or sphingolipids. Proteins are another important component of biological membranes. Deposits of cholesterol adds rigidity to the membrane.

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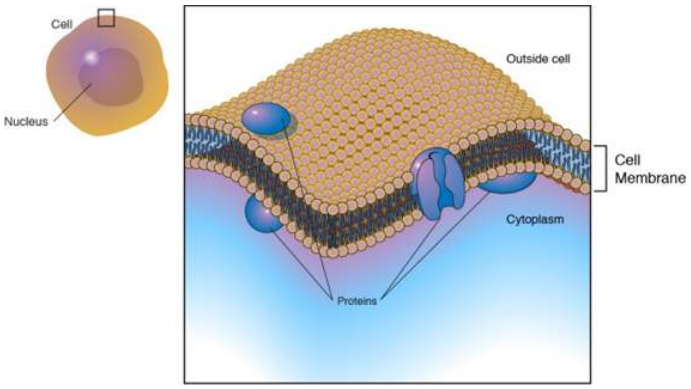
11.8: Transport Across Cell Membrane

Learning Objectives

- Explain how various compounds cross the membrane.
- Relate to electron transport chain as examples of facilitated diffusion and active transport of H^+ ions.

What is a cell membrane?

The cell membrane, also known as the plasma membrane is a lipid bilayer that separates the cell interior from the extracellular space.



Composition of the cell membrane

A cell membrane is a complex structure with several components as shown in Figure **11.8.1**. and described here.

- Phospholipid bilayer that has polar heads on the outside in contact with water and nonpolar tails inside the bilayer. Unsaturated fatty acids in the lipid hydrophobic tails increase the membrane fluidity. The more the proportion of unsaturated fatty acids the higher the fluidity.
- Cholesterol interspersed between phospholipids gives rigidity to the membrane. The more the proportion of cholesterol the more rigid the membrane.
- Two types of proteins: integral proteins that span the membrane and serve as transporters of species and peripheral proteins that are loosely attached to the outer side of the membrane that act as enzymes to facilitate the interaction with the cell's environment.
- Glycoproteins and glycolipids have carbohydrate parts attached to the outer lipid layer and serve the purpose of cell-to-cell recognition.

The cell membrane controls the movement of substances in and out of the cells and organelles. It is selectively permeable to ions and organic molecules. It plays a role in cell adhesion, cell signaling, and attachment surface for the cytoskeleton to give the shape to the cells and attached to the extracellular matrix to hold the cells together in tissues.

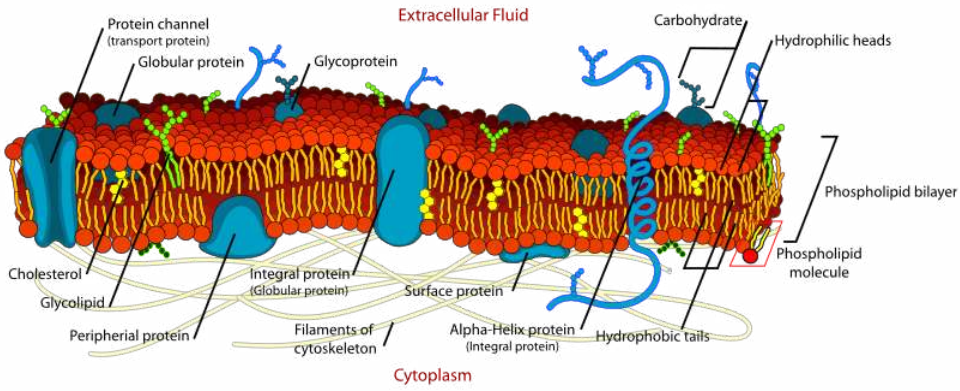


Figure **11.8.1**: A detailed structure of cell membrane. (Copyright; LadyofHats, Public domain, via Wikimedia Commons)

Transport through the cell membrane

The cell membrane is a partition between intracellular and extracellular spaces, but some substances needed by the cell need to enter and some products or wastes need to exit the cell. The cell membrane allows a selective movement of substances in and out of the cell in several ways.

Diffusion

Nonpolar molecules, such as O_2, CO_2, and N_2, can move across the membrane from a higher concentration region to a lower concentration region through the process of diffusion, as illustrated in Figure **11.8.2**. Diffusion does not require energy and hence this is called passive transport.

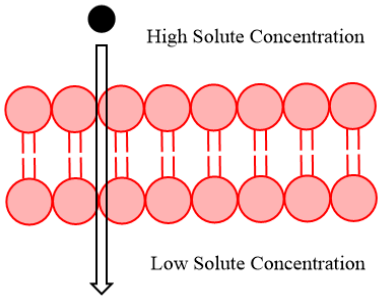


Figure **11.8.2**: Illustration of diffusion across the cell membrane.

Facilitated Diffusion

Proteins that span the membrane form channels through which polar molecules and ions can diffuse more rapidly than by simple diffusion. The proteins have the channel size that match the size of the substance or they change the shape to adjust to the size of the substance that needs to be selectively transported through the facilitated transport, as illustrated in Figure **11.8.3** right. Ions such as chloride ion (Cl^-), bicarbonate ion (HCO_3^-), and polar molecules like glucose molecules do not move fast enough through simple diffusion and are transported by the facilitated diffusion process to meet the need of the cells. Movement of water involves facilitated diffusion with the help of protein channels called aquaporins. Movement of water across cell membranes gives rise to osmosis. Facilitated diffusion does not require energy and hence this is also an example of passive transport.

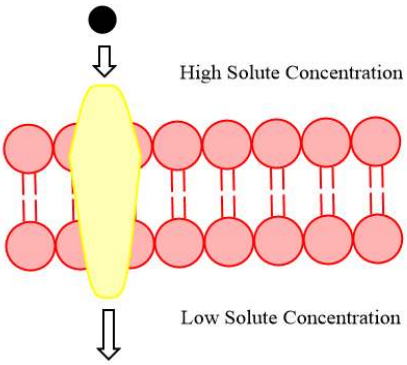


Figure **11.8.3**: Illustration of facilitated diffusion transport across the cell membrane.

Active transport

Sometimes substance need to be moved against the concentration gradient, from lower to higher concentration. This process requires the input of energy. Polar molecules and ions are transported across membranes through a protein channel in the direction of lower to higher concentration. For example, K^+ concentration is greater inside the cell and Na^+ is greater outside the cell. In the conduction of nerve impulses and contraction of muscles, K^+ moves into the cell, and Na^+ moves out of the cell by active transport. The active transport process, is illustrated in figure 11.8.4.

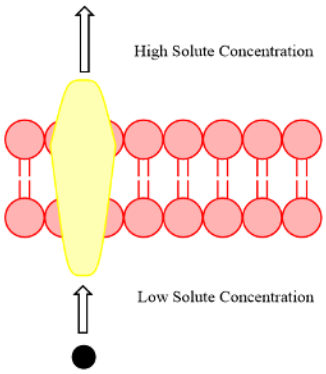


Figure **11.8.4**: Illustration of active transport across the cell membrane. Active transport requires energy.

The Electron Transport Chain

Facilitated diffusion and active transport of the hydrogen ion, (H^+) is encountered in a process called the electron transport chain which happens in the inner mitochondrial membrane.

At the end of a turn of the citric acid cycle, energy-rich harvest molecules such as GTP, $FADH_2$ and 3 NADH molecules are produced. The primary task of the last stage of cellular respiration, the **electron transport chain**, is to transfer the potential energy from NADH and $FADH_2$ to ATP molecules. The ATP molecule is the "battery" which power work within the cell.

The electron transport chain is a group of protein embedded in the inner mitochondrial membrane (see figure 11.8.5). These proteins are energy carrier molecules and are arranged in sequence within the membrane so that energy-carrying electrons pass from one to another, losing a little energy in each step. Complex I accepts electrons from NADH. As a result NADH is oxidized to NAD^+. Complex II accepts electrons from $FADH_2$. $FADH_2$ is oxidized to FAD. Electrons supplied by NADH and $FADH_2$ move from one protein complex to another in the inner mitochondrial membrane. The energy is harnessed to pump hydrogen ions, from the matrix (low H^+ concentration) into the intermembrane space (high H^+ concentration) in a process called active transport. As a result the concentration of concentration of H^+ ions become higher in the intermembrane space creating an electrochemical gradient.

Hydrogen ions then flow down the gradient - from high concentration region to low concentration. The ion channel/enzyme ATP synthase allows the H^+ to flow from intermembrane space to the matrix in a facilitated diffusion process. The energy released during the facilitated diffusion of hydrogen ions converts a ADP and a phosphate ion to an ATP in a process known as **oxidative phosphorylation.**

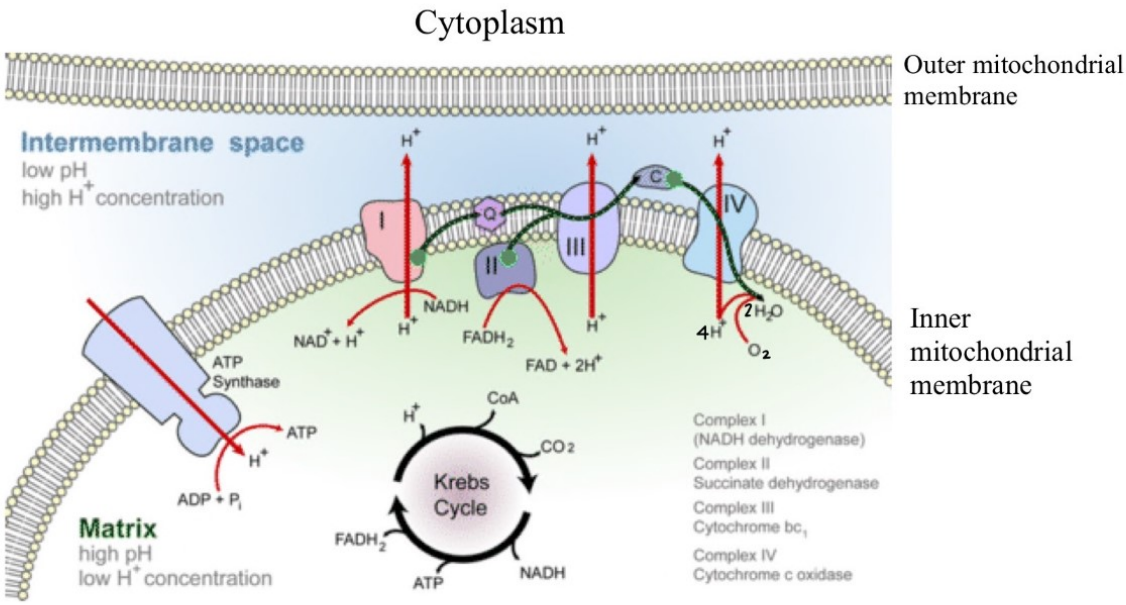


Figure 11.8.5: The electron transport chain embedded in the mitochondrial inner membrane capture high-energy electrons from the carrier molecules and use them to concentrate hydrogen ions in the intermembrane space. Hydrogen ions flow down the gradient back into the matrix through ATP synthase channel which capture the energy to convert ADP to ATP. (CC BY-NC 3.0; Mariana Ruiz Villarreal (LadyofHats) for the CK-12 Foundation).

Oxygen is the final electron acceptor which combines with hydrogen ions to form water. One glucose molecule is converted to approximately 32 ATP's with glycolysis, aerobic conversion of pyruvate to acetyl-CoA, the citric acid cycle, and the electron transport chain. Cellular respiration transfers the energy from one molecule of glucose to molecules of ATP, releasing carbon dioxide and water as waste. Food energy has become energy which can be used for work within the cell such as active transport of molecules across membranes and building large organic and biological molecules.

Summary

There are three different ways molecules and ions move across a cell membrane. They are diffusion, facilitated diffusion, and active transport. Active transport requires energy, while diffusion and facilitated diffusion do not. Most polar molecules and ions require a protein channel during transport.

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11.9: Digestion of Triglycerides

Learning Objectives

- Describe the process of digestion of triglycerides.

Triglycerides are esters and esters can undergo hydrolysis. When this hydrolysis happens in the digestive system the process is called digestion. Lipid digestion begins in the stomach. Lingual lipases present in the saliva are activated when they reach the acidic conditions of the stomach. Gastric lipases secreted in the stomach start the hydrolysis of the ester functional groups in triglycerides. Most of the digestion of triglycerides happen in small intestine. A hormone secreted in this region stimulates the gallbladder to discharge bile. The principal constituents of **bile** are the bile salts, which emulsify large, water-insoluble lipid droplets, disrupting some of the hydrophobic interactions holding the lipid molecules together and suspending the resulting smaller globules (micelles) in the aqueous digestive medium. These changes greatly increase the surface area of the lipid particles, allowing for contact with the enzyme lipases and thus rapid digestion of the fats.

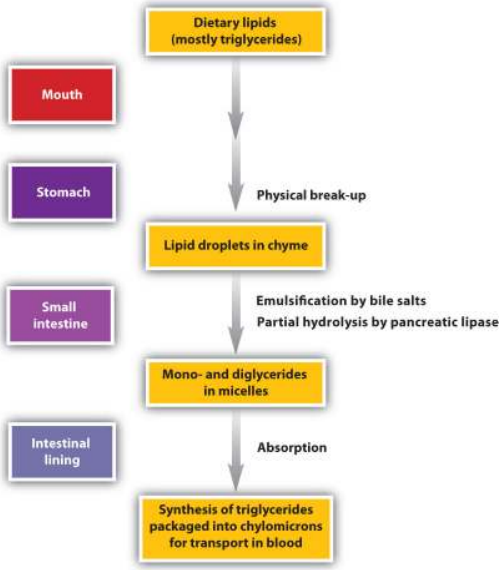


Figure **11.9.1**: The Principal Events and Sites of Triglyceride Digestion

The lipases catalyze the digestion of a triglyceride first to diglyceride and a fatty acid. The diglyceride is then hydrolyzed to a monoglyceride and a fatty acid as shown in figure 11.9.2.

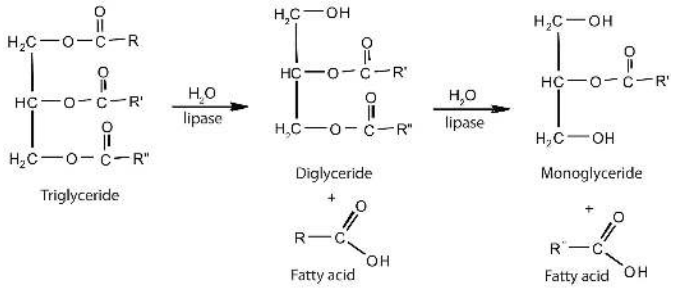


Figure **11.9.2**: Digestion of triglycerides.

The large and hydrophobic long-chain fatty acids and monoglycerides are not so easily suspended in the watery intestinal chyme. However, bile salts resolve this issue by enclosing them in a **micelle**, which is a tiny sphere with polar (hydrophilic) ends facing the watery environment and hydrophobic tails turned to the interior, creating a receptive environment for the long-chain fatty acids. The core also includes cholesterol and fat-soluble vitamins. Without micelles, lipids would sit on the surface of chyme and never come in contact with the absorptive surfaces of the cells.

The free fatty acids and monoglycerides then enter the cells that line the small intestine and are converted back into triglycerides. The triglycerides are eventually carried through blood by lipoproteins.

Summary

Bile salts help to emulsify the hydrophobic triglycerides in the aqueous gastric solution. The enzyme lipase catalyze the hydrolysis (digestion) of the triglycerides to fatty acids and monoglycerides. These products of digestion then diffuse into the cells that line the small intestine and then converted back to triglycerides.

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11.10: Fatty Acid Catabolism

Learning Objectives

- To describe the reactions needed to completely oxidize a fatty acid to carbon dioxide and water.

Like glucose, the fatty acids released in the digestion of triglycerides and other lipids are broken down in a series of sequential reactions accompanied by the gradual release of usable energy. Some of these reactions are oxidative and require nicotinamide adenine dinucleotide (NAD^+) and flavin adenine dinucleotide (FAD). The enzymes that participate in fatty acid catabolism are located in the mitochondria, along with the enzymes of the citric acid cycle, the electron transport chain, and oxidative phosphorylation. This localization of enzymes in the mitochondria is of the utmost importance because it facilitates efficient utilization of energy stored in fatty acids and other molecules.

Fatty acid catabolism is initiated in the cytoplasm. There the fatty acid is activated by conversion to an energy-rich fatty acid derivative of coenzyme A called *fatty acyl-coenzyme A* (CoA). The activation is catalyzed by *acyl-CoA synthetase*. For each molecule of fatty acid activated, one molecule of coenzyme A and one molecule of adenosine triphosphate (ATP) are used, equaling a net utilization of the two high-energy bonds in one ATP molecule (which is therefore converted to adenosine monophosphate [AMP] rather than adenosine diphosphate [ADP]). For this reason the activation of one fatty acid is considered to require the energy equivalent of 2 ATP's.

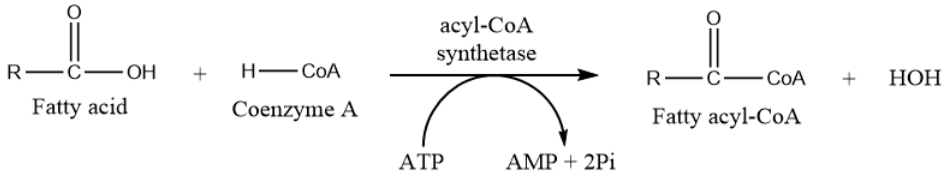


Figure **11.10.1**: Activation of the fatty acid by CoA for catabolism.

The fatty acyl-CoA diffuses to the inner mitochondrial membrane for the catabolism.

Catabolism of Fatty Acids

Catabolism of the fatty acyl-CoA occurs in the mitochondrial matrix via a sequence of four reactions known collectively as β-oxidation because the β-carbon undergoes successive oxidations in the progressive removal of two carbon atoms (Figure **11.10.2**). The fatty acyl-CoA formed in the final step becomes the substrate for the first step in the next round of the spiral.

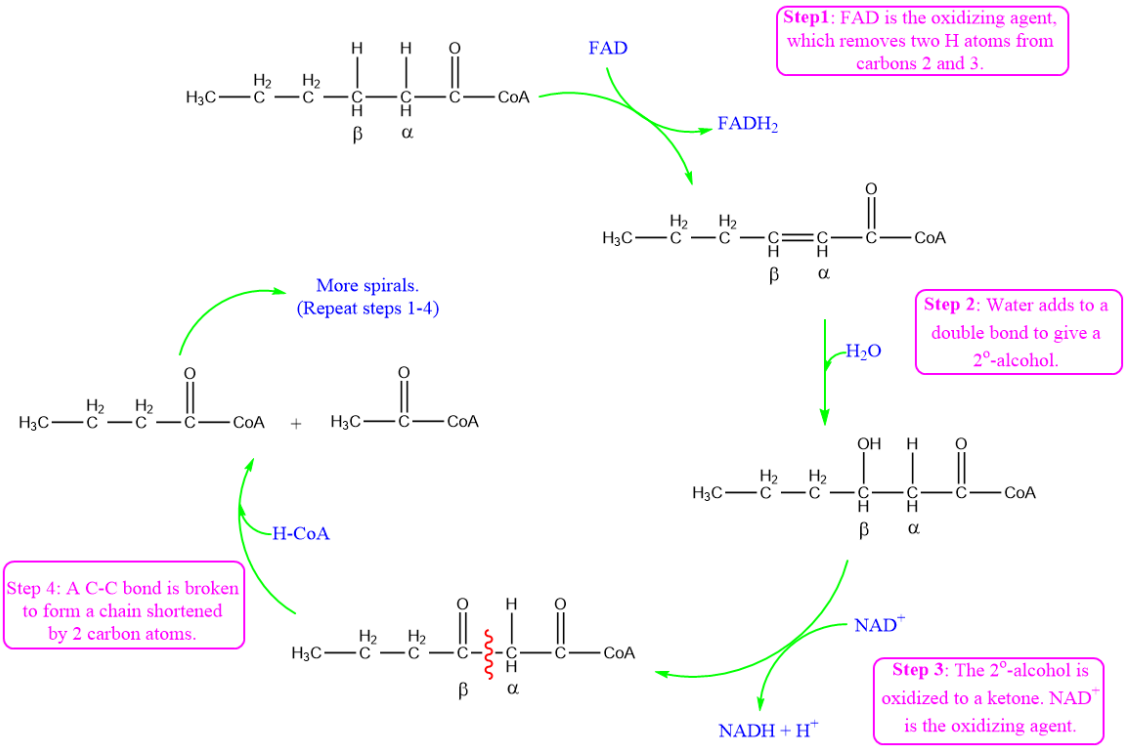


Figure **11.10.2**: A spiral in fatty acid catabolism. Each spiral shortens the fatty acid chain by two carbons.

STEP 1: First oxidation of an alkane (dehydrogenation)

A fatty acyl-CoA is oxidized to yield a trans double bond between the α- and β- carbons (the second and third carbons). FAD is oxidizing agent and is converted to $FADH_2$, which moves into the electron transport chain.

STEP 2: Hydration (addition of water across the double bond)

The trans alkene is then hydrated to form a secondary alcohol. The hydroxyl group is placed on the β-carbon.

STEP 3: Second oxidation (oxidation of an alcohol)

The secondary alcohol is then oxidized to a ketone by NAD^+ acting as the oxidizing agent.

STEP 4: Cleavage

Acetyl-CoA cleaves off to yield a fatty acid that is two carbons shorter than before. The chain shortened fatty acid will start back at step 1 and go through another spiral.

Note

In each spiral, 1 molecule of acetyl-CoA, 1 molecule of NADH, and 1 molecule of $FADH_2$ are produced. The final spiral yields two acetyl-CoA molecules.

> *Because each shortened fatty acyl-CoA cycles back to the beginning of the pathway, β-oxidation is referred to as the fatty acid spiral.*

The fate of the acetyl-CoA obtained from fatty acid oxidation depends on the needs of an organism. It may enter the citric acid cycle and be oxidized to produce energy, it may be used for the formation of water-soluble derivatives known as ketone bodies, or it may serve as the starting material for the synthesis of fatty acids.

Looking Closer: Ketone Bodies

In the liver, most of the acetyl-CoA obtained from fatty acid oxidation is oxidized by the citric acid cycle. However, some of the acetyl-CoA is used to synthesize a group of compounds known as *ketone bodies*: acetoacetate, β-hydroxybutyrate, and acetone shown in figure 11.10.3.

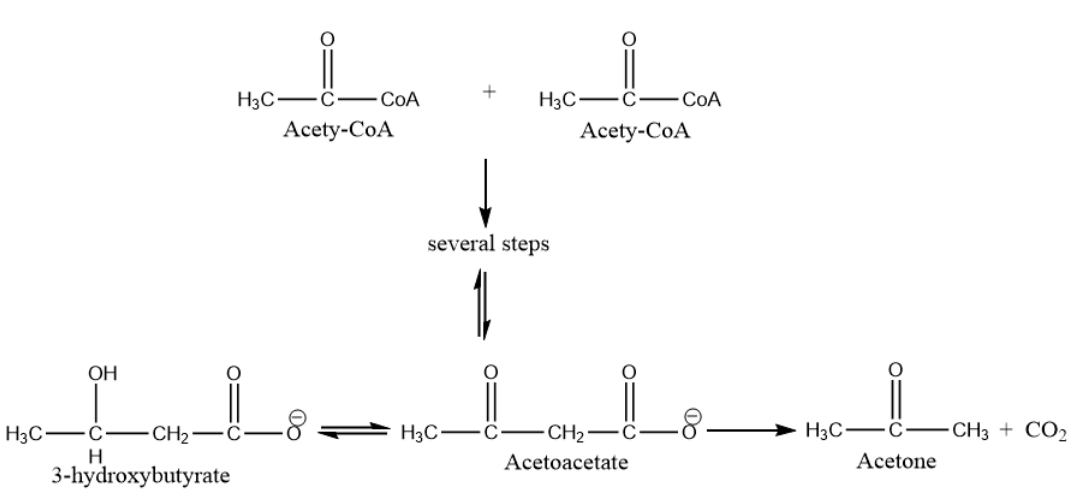


Figure **11.10.3**: Ketone bodies are produced from acetyl-CoA not consumed by the citric acid cycle.

The acetoacetate and β-hydroxybutyrate synthesized by the liver are released into the blood for use as a metabolic fuel (to be converted back to acetyl-CoA) by other tissues, particularly the kidney and the heart. Thus, during prolonged starvation, ketone bodies provide about 70% of the energy requirements of the brain. Under normal conditions, the kidneys excrete about 20 mg of ketone bodies each day, and the blood levels are maintained at about 1 mg of ketone bodies per 100 mL of blood.

In starvation, diabetes mellitus, and certain other physiological conditions in which cells do not receive sufficient amounts of carbohydrate, the rate of fatty acid oxidation increases to provide energy. This leads to an increase in the concentration of acetyl-CoA. The increased acetyl-CoA cannot be oxidized by the citric acid cycle because of a decrease in the concentration of oxaloacetate, which is diverted to glucose synthesis. In response, the rate of ketone body formation in the liver increases further, to a level much higher than can be used by other tissues. The excess ketone bodies accumulate in the blood and the urine, a condition referred to as *ketosis*. When the acetone in the blood reaches the lungs, its volatility causes it to be expelled in the breath. The sweet smell of acetone, a characteristic of ketosis, is frequently noticed on the breath of severely diabetic patients.

Because two of the three kinds of ketone bodies are weak acids, their presence in the blood in excessive amounts overwhelms the blood buffers and causes a marked decrease in blood pH (to 6.9 from a normal value of 7.4). This decrease in pH leads to a serious condition known as *acidosis*. One of the effects of acidosis is a decrease in the ability of hemoglobin to transport oxygen in the blood. In moderate to severe acidosis, breathing becomes labored and very painful. The body also loses fluids and becomes dehydrated as the kidneys attempt to get rid of the acids by eliminating large quantities of water. The lowered oxygen supply and dehydration lead to depression; even mild acidosis leads to lethargy, loss of appetite, and a generally run-down feeling. Untreated patients may go into a coma. At that point, prompt treatment is necessary if the person's life is to be saved.

Note

The number of times a spiral is repeated for a fatty acid containing n carbon atoms is (n/2 – 1) because the final spiral yields two acetyl-CoA molecules.

12-Carbon Fatty Acid	$\xrightarrow{5\text{ spirals}}$	6 Acetyl-CoA
14-Carbon Fatty Acid	$\xrightarrow{6\text{ spirals}}$	7 Acetyl-CoA
18-Carbon Fatty Acid	$\xrightarrow{8\text{ spirals}}$	9 Acetyl-CoA

ATP Yield from Fatty Acid Oxidation

The amount of ATP obtained from fatty acid oxidation depends on the size of the fatty acid being oxidized. As an example, we'll study octanoic acid, a saturated long chain carboxylic acid acid with 8 carbon atoms. Calculating its ATP yield provides a model for determining the ATP yield of all other fatty acids.

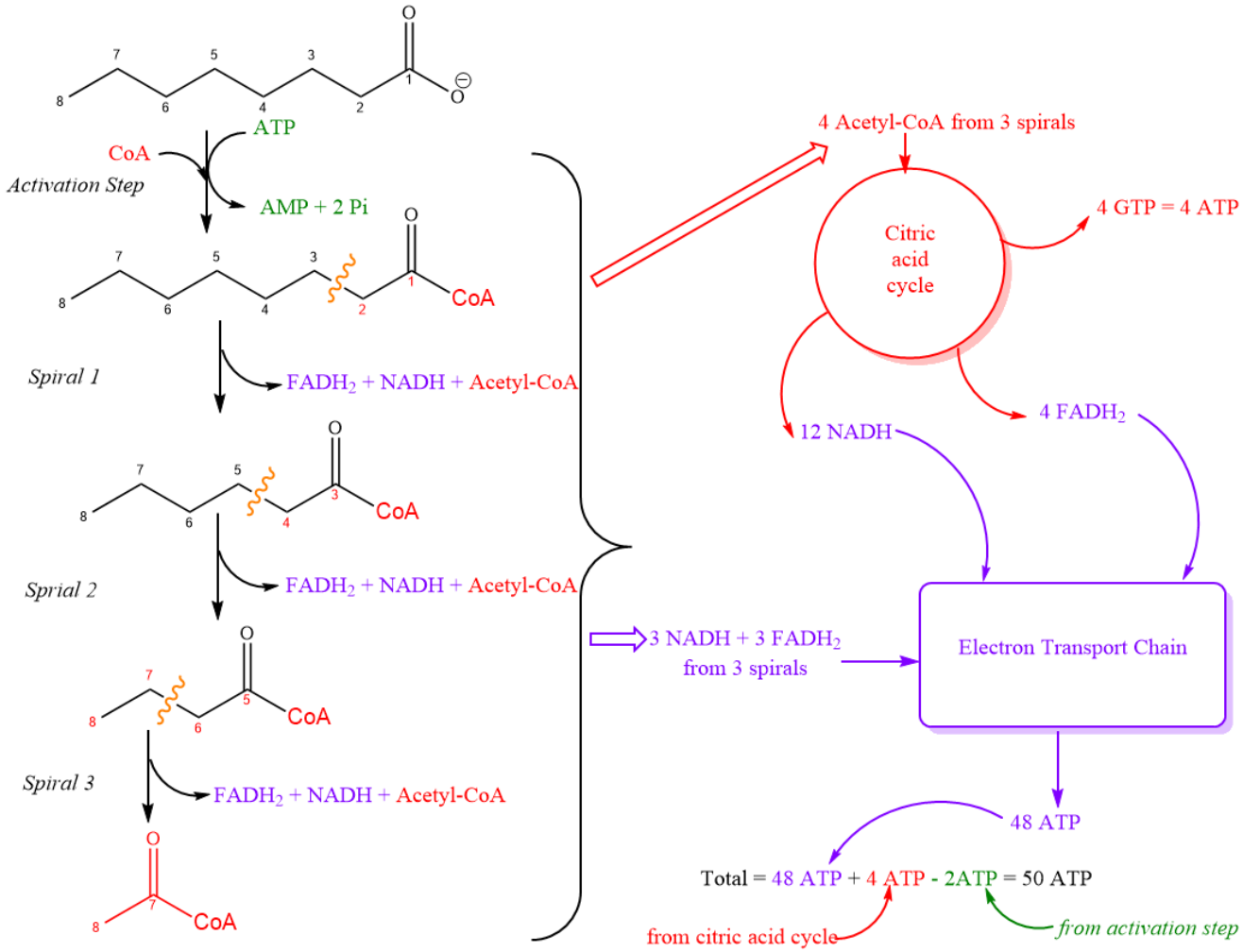


Figure **11.10.4**: Catabolism of octanoic acid.

Activation and Spirals

The catabolism of 1 molecule of octanoic acid requires 2 molecules of ATP (for activation) and forms 4 molecules of acetyl-CoA, 3 molecules of NADH, and 3 molecules of $FADH_2$. This complete conversion requires 3 spirals to be repeated.

Citric Acid Cycle

The 4 molecules of acetyl-CoA are then processed by the citric acid cycle. Recall from earlier that each molecule of acetyl-CoA metabolized by the citric acid cycle yields 3 molecules of NADH, 1 molecule of $FADH_2$, and 1 molecule of GTP. The 4 molecules of acetyl-CoA require 4 citric acid cycles and produce 12 molecules of NADH, 4 molecules of $FADH_2$, and 4 molecules of GTP.

Electron Transport Chain

The reduced coenzymes NADH and $FADH_2$ from the spirals and the citric acid cycle then enter the electron transport chain and are converted to ATP. Each NADH molecule is responsible for the formation of 2.5 ATP and each $FADH_2$ forms 1.5 ATP. There are 12 + 3 = 15 NADH and 4 + 3 = 7 $FADH_2$ molecules from the 3 spirals and 4 Citric Acid Cycles that enter the ETC to produce 48 ATP (15 x 2.5 + 7 x 1.5).

Total ATP count

The 4 molecules of GTP from the citric acid cycle are equivalent in energy to ATP and so are added to 48 ATP from the ETC. Two ATP molecules are required for the activation step and hence subtracted. The total ATP yield from octanoic acid is 48+4-2= 50 ATP.

> *The oxidation of fatty acids produces large quantities of water. This water, sustains migratory birds and animals (such as the camel) for long periods of time.*

Summary

- Fatty acids, obtained from the breakdown of triglycerides and other lipids, are oxidized through a series of four reactions known in a spiral.
- In each round of the spiral, 1 molecule of acetyl-CoA, 1 molecule of NADH, and 1 molecule of $FADH_2$ are produced.
- The acetyl-CoA, NADH, and $FADH_2$ are used in the citric acid cycle, the electron transport chain, and oxidative phosphorylation to produce ATP.

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11.11: End of Chapter Problems

Fatty Acids

1. In terms of structure what distinguishes:

a. fatty acids from carboxylic acids

b. saturated fatty acid from unsaturated fatty acid.

2. Which has lower melting point? Explain.

a. lauric acid or myristic acid.

b. linoleic acid or oleic acid

Waxes

1. List the biological functions of waxes.

2. What property of waxes allows them to fulfill their primary biological functions?

Triglycerides

1. Construct a triglyceride from myristic acid, palmitic acid, oleic acid, and glycerol. Would you expect the triglyceride to be a liquid or a solid at room temperature? Explain.

2. List biological functions of triglycerides.

3. Why are triglycerides not found in cell membranes?

4. Explain with IMF's why butter is a fat (solid) at room temperature?

5. Explain with IMF's why canola oil is a liquid at room temperature?

6. True or False

a. Triglycerides are carboxylic acids

b. Saturated fats contain only saturated fatty acids.

c. Unsaturated fats contain only unsaturated fatty acids.

Reactions of Lipids

1. Vegetable oils tend to become rancid more quickly than animal fats. Why?

2. Vegetable oils are easier to digest than animal fats. Why?

3. List four different ways to slow the spoilage (rancidity) of vegetable oil.

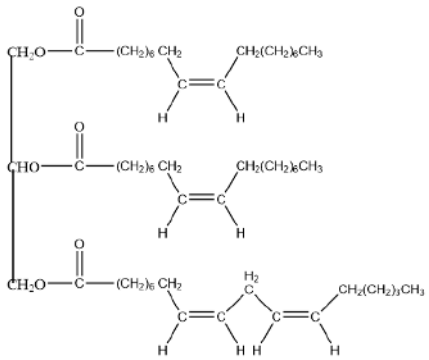
4. What does it mean to partially hydrogenate vegetable oil?

5. What are trans fats? How are they produced? Why do restaurants ban the use of *trans* fats in Ca?

6. Why do food companies partially hydrogenate vegetable oils? Why do they not fully hydrogenate vegetable oils?

7. Which is likely to have most trans-fats: vegetable oil, margarine, or butter? Explain.

8. Draw the products formed when the triglyceride below is saponified. Label each compound in the reaction based on the solubility classification.



9. Draw a picture of a soap micelle and clearly label the parts that interact with water and grease. Explain how soap works to clean grease in your hands.

10. What is hard water? Explain what happens when a soap interacts with a hard water ion.

Steroids

1. How is cholesterol used in the human body?

2. What are the biological functions of HDL and LDL?

3. What are the signs of heart disease in terms of HDL and LDL?

4. A person whose diet contains no cholesterol can still have high cholesterol levels. Explain.

5. What is the role of bile salts in digestion?

Membranes & Transport

1. Draw a sketch of the lipid bilayer. Name the kinds of lipids that make up membrane structure? Why are these lipids used to make up membrane structure?

2. What is the role of cholesterol in membranes?

3. Plant membranes contain a higher proportion of unsaturated fatty acids than do animal membranes. Which type of membrane would you expect to be more fluid (flexible), plant or animal?

4. The phospholipid known as Lecithin is an emulsifying agent but triglycerides are not. Account for the difference.

5. The lipids present in the cell membranes of fish, especially fish that live in cold water, contain more unsaturated fatty acid residues than do lipids found in the membranes of land animals. Explain why.

6. To function properly, membranes must flexible or fluid. In light of this fact, propose an explanation of why cell membranes in the feet and legs of reindeer contain a higher percentage of unsaturated fatty acids than do the cell membranes in the interior of its body.

7. Compare diffusion and facilitated diffusion. How are they similar and different?

8. Compare facilitated diffusion and active transport. How are they similar and different?

Digestion of Triglycerides

1. What part of the digestive system is involved in the digestion of triglycerides?

2. Name the functional group hydrolyzed when a triglyceride undergoes digestion.

3. Name an enzyme that catalyzes the digestion of triglycerides.

4. What products are formed when a triglyceride undergoes digestion?

5. Explain the role of bile salts in lipid digestion?

Fatty Acid Catabolism

1. Answer the following question based on a spiral of fatty acid catabolism:

a. What is the function of FAD in step 1 of the spiral?

b. Identify the type of reaction in step 2 of the spiral?

c. What is the function of NAD^+ in step 3 of the spiral?

d. What is the harvest of each spiral? (note: last spiral is different)

e. What is different about the harvest in the last spiral?

2. Calculate the **net number of ATP's** produced when a myristic acid is activated, undergoes oxidation via adequate number of spirals, acetyl-CoA's pass through the correct number of citric acid cycles, and all the reduced coenzymes (NADH & $FADH_2$) pass through the electron transport chain.

3. Calculate the **net number of ATP's** produced when one 10-carbon fatty acid is activated, undergoes oxidation via adequate number of spiral, acetyl-CoA's pass through the correct number of citric acid cycles, and all the reduced coenzymes (NADH & $FADH_2$) pass through the electron transport chain.

4. Where in a cell does fatty acid oxidation take place?

5. a. What are ketone bodies?

b. Why is acidosis (low blood pH) associated with the overproduction of ketone bodies?

c. In 2004 a low carb diet craze swept the US. People on the diet had very bad breath. What molecule is the reason for the bad breath?

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CHAPTER OVERVIEW

12: Gases

12.1: Gases and Pressure
12.2: Vapor Pressure
12.3: The Pressure-Temperature Law
12.4: The Volume-Temperature Law
12.5: The Pressure-Volume Law
12.6: The Combined Gas Law
12.7: Henry's Law
12.8: Partial Pressure and Dalton's Law
12.9: End of Chapter Problems

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12.1: Gases and Pressure

Learning Objectives

- Understand properties of gases, gas pressure and atmospheric pressure.
- List common pressure units, understand the workings of a barometer.
- Describe the relationship between atmospheric pressure and altitude.
- Understand systolic and diastolic blood pressures.
- Know the elements that are monoatomic and diatomic gases.
- Recognize STP conditions.

The gas phase is unique among the three states of matter. There are some simple models we can use to predict the physical behavior of all gases—independent of their identities. We cannot do this for the solid and liquid states. In fact, the development of this understanding of the behavior of gases represents the historical dividing point between alchemy and modern chemistry. Initial advances in the understanding of gas behavior were made in the mid 1600s by Robert Boyle, an English scientist who founded the Royal Society (one of the world's oldest scientific organizations).

How is it that we can model all gases independent of their chemical identity? The answer is in a group of statements called the kinetic theory of gases:

- Gases are composed of tiny particles that are separated by large distances.
- Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container.
- The kinetic energy of gas particles is related to the temperature of a gas.
- Gas particles do not experience any force of attraction or repulsion with each other.

Did you notice that none of these statements relates to the identity of the gas? This means that all gases should behave similarly. A gas that follows these statements perfectly is called an *ideal gas*. Most gases show slight deviations from these statements and are called *real gases*. The kinetic theory also states that there is no interaction between individual gas particles. There are, in fact, intermolecular interactions in real gases, however the kinetic theory assumes that gas particles are so far apart that the individual particles don't "feel" (interact with) each other.

Gas Pressure

One of the statements of the kinetic theory mentions collisions. As gas particles are constantly moving, they are also constantly colliding with each other and with the walls of their container. There are forces involved as gas particles bounce off the container walls (Figure **12.1.1**). The force generated by gas particles divided by the area of the container walls yields pressure.

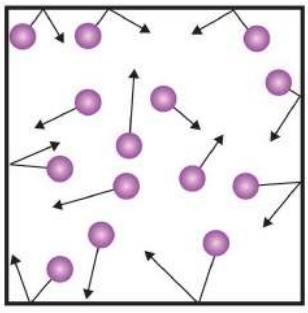


Figure **12.1.1**: Gas Pressure. Pressure is what results when gas particles collide with the walls of their container.

Atmospheric Pressure

Just as we exert pressure on a surface because of gravity, so does our atmosphere. We live at the bottom of an ocean of gases that becomes progressively less dense with increasing altitude. Approximately 99% of the mass of the atmosphere lies within 30 km of Earth's surface, and half of it is within the first 5.5 km. Every point on Earth's surface experiences a net pressure called *atmospheric pressure*.

Barometer

Atmospheric pressure can be measured using a barometer. This device invented in 1643 by one of Galileo's students, Evangelista Torricelli (1608–1647). A barometer may be constructed from a long glass tube that is closed at one end. It is filled with mercury and placed upside down in a dish of mercury without allowing any air to enter the tube. Some of the mercury will run out of the tube, but a relatively tall column remains inside (Figure **12.1.2**) . Why doesn't all the mercury run out? Gravity is exerting a downward force on the mercury in the tube, but it is opposed by the pressure of the atmosphere pushing down on the surface of the mercury in the dish, which has the net effect of pushing the mercury up into the tube. Because there is no air above the mercury inside the tube in a properly filled barometer (it contains a *vacuum*), there is no pressure pushing down on the column. Thus the mercury runs out of the tube until the pressure exerted by the mercury column itself exactly balances the pressure of the atmosphere. Under normal weather conditions at sea level, the two forces are balanced when the top of the mercury column is approximately 760 mm above the level of the mercury in the dish, as shown in (Figure **12.1.2**) . This value varies with meteorological conditions and altitude. In Denver, Colorado, for example, at an elevation of about 1 mile, or 1609 m (5280 ft), the height of the mercury column is 630 mm rather than 760 mm.

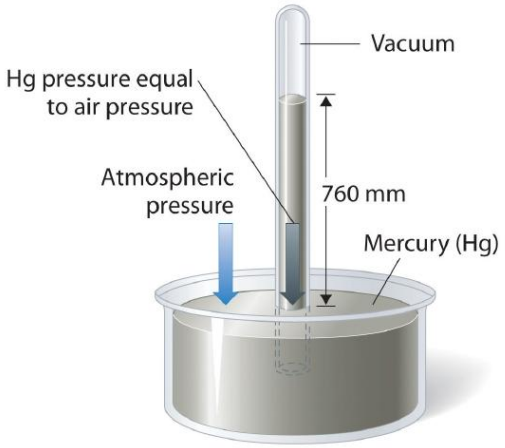


Figure **12.1.2**: The pressure exerted by the atmosphere on the surface of the pool of mercury supports a column of mercury in the tube that is about 760 mm tall.

While mercury barometers were the workhorses for pressure measurement into the last quarter of the 20th century, they have been replaced by electronic gauges. One motivation was safety. Mercury and mercury vapor are heavy metal poisons.

Units of Atmospheric Pressure

Common units of pressure are the atmosphere (atm), which was originally defined as the average pressure of Earth's atmosphere at sea level; and mmHg (millimeters of mercury), which is the pressure generated by a column of mercury 1 mm high. The unit millimeters of mercury is also called a torr, named after the Italian scientist Evangelista Torricelli, who invented the barometer in the mid-1600s. A more precise definition of atmosphere, in terms of torr, is that there are exactly 760 torr in 1 atm.

1 atm = 14.7 psi = 760 torr = 760 mm of Hg

Given all the relationships between these pressure units, the ability to convert from one pressure unit to another is a useful skill.

✔ Example **12.1.1**: Converting Pressures

Write a conversion factor to determine how many atmospheres are in 1,547 mm of Hg.

Solution

Because there are 760.0 mm of Hg in 1 atm, we can use this conversion factor to do the mathematical conversion:

$$1,547\ \text{mmHg} \times \frac{1\ \text{atm}}{760.0\ \text{mmHg}} = 2.036\ \text{atm}$$

Note how the mm of Hg units cancel algebraically.

? Exercise 12.1.1: Converting Pressures

Write a conversion factor to determine how many millimeters of mercury are in 9.65 atm.

Answer

$$9.65\ \text{atm} \times \frac{760\ \text{mmHg}}{1\ \text{atm}} = 7,330\ \text{mmHg}.$$

Atmospheric Pressure and Altitude

As altitude increases, the air becomes thinner, the atmospheric pressure decreases (Figure **12.1.3**) . The oxygen levels in the air decreases as the altitude increases. That is why mountain climbers experience an inadequate supply of oxygen to the body at high altitudes -a medical condition called Acute Mountain Sickness.

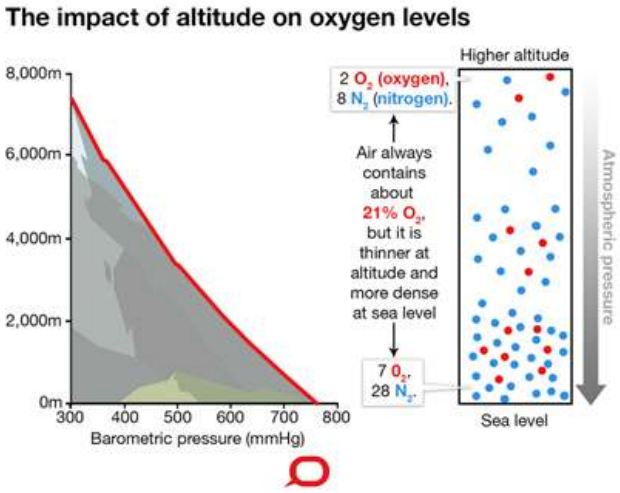


Figure **12.1.3**: Air pressure decreases and the air thins and oxygen levels in it decreases as the altitude increases. Source: Brendan Scott/The Conversation, CC BY-ND https://images.theconversation.com/f....png?ixlib=rb-

Blood pressure

The heart pumps the blood into the circulatory system. When the heart contracts, it applies the pressure on the blood in it, and the blood pumps out of the heart into the circulatory system, as illustrated in (Figure **12.1.4**). The blood pressure in the circulatory system is highest at this point, and it is called **systolic** pressure. It can be in the range of 100 to 200 mm Hg. A desirable systolic pressure range is 100 to 120 mm Hg. When heart muscles relax, the cavity in the heart expands and more blood fills in the heart. The blood pressure in the circulatory system is minimum at this point, and it is called **diastolic** pressure. The diastolic pressure may vary in the range of 60 to 110 mm Hg. A desirable diastolic pressure is less than 80 mm Hg.

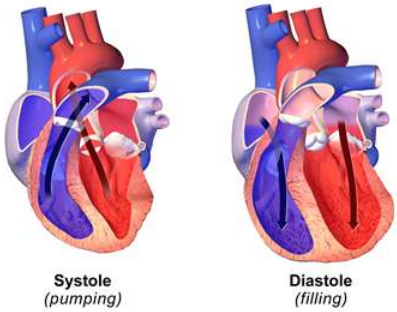


Figure **12.1.4**: Diagram of heart illustrating systole, i.e., blood pumping out of heart vs Diastole i.e., blood filling into the heart. Source: BruceBlaus / CC BY-SA (https://creativecommons.org/licenses/by-sa/4.0)

The device used to measure blood pressure is called a **sphygmometer or sphygmomanometer**. It consists of a cuff that wraps around the upper arm, a pump to inflate air in the cuff, and a stethoscope to hear the sound of blood flow, as illustrated in (Figure

12.1.5) . The cuff is inflated with air to a pressure above the systolic pressure and it results in cutting off the blood flow through the brachial artery in the upper arm. A stethoscope is used to hear the sound of blood flow. No sound is heard at this point. The pressure in the cuff is slowly reduced. When the pressure in the cuff is equal to systolic pressure, the blood begins to spurt into the artery, and a tapping sound is heard through the stethoscope. At the point when the pressure is equal to the diastolic pressure, the blood flows freely through the artery, and the tapping sound disappears. The blood pressure reading is reported as a set of two numbers, e.g., 120/80, where the higher number is systolic, and the lower number is the diastolic pressure.

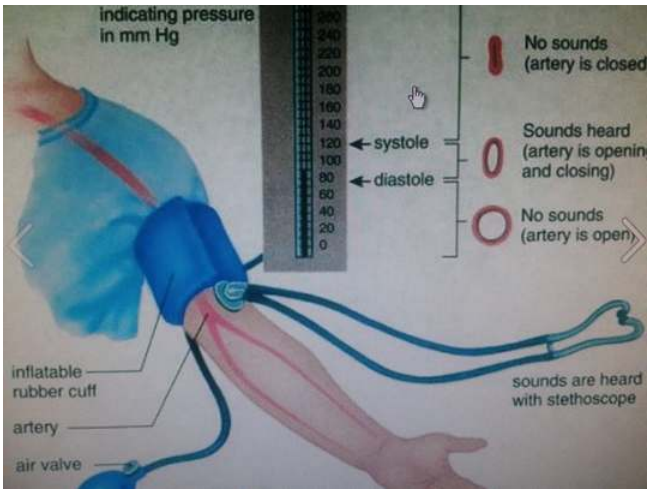


Figure 12.1.5: A healthcare professional performing blood pressure monitoring on a patient. Source: Roa'a najim / CC BY-SA (https://creativecommons.org/licenses/by-sa/3.0)

Monoatomic and Diatomic gases

Figure 12.1.6 shows the locations in the periodic table of elements that are commonly found in the gaseous, liquid, and solid states. Except for hydrogen, the elements that occur naturally as gases are on the right side of the periodic table. Of these, all the noble gases He, Ne, Ar, Kr, Xe, and Rn are monatomic gases, whereas the other gaseous elements are diatomic molecules (H_2, N_2, O_2, F_2, and Cl_2). Bromine (Br_2) and mercury (Hg) are liquids under normal conditions (25°C and 1.0 atm, commonly referred to as “room temperature and pressure”). The rest of the elements are all solids under normal conditions.

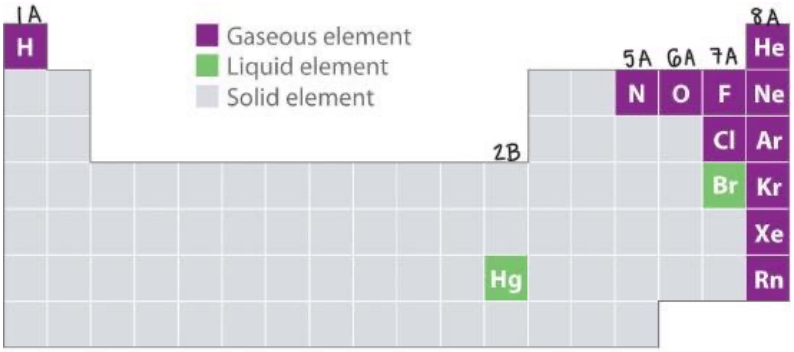


Figure 12.1.6: Elements That Occur Naturally as Gases, Liquids, and Solids at 25°C and 1 atm

Standard Temperature and Pressure (STP)

A temperature of 0.0°C (273.15 Kelvin) and under 1 atmosphere of pressure (760 mm of Hg) is referred to as STP.

Key Takeaway

The gas phase has certain general properties characteristic of that phase. The atmosphere applies a pressure on the surface of the earth known as the atmospheric pressure. The barometer is used to measure the atmospheric pressure. The atmospheric pressure changes with altitude. Blood pressure is the pressure applied by the blood on the walls of the arteries and is measured by a device called the sphygmometer.

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12.2: Vapor Pressure

Learning Objectives

- Describe the relationship between atmospheric pressure, vapor pressure and the boiling point of a liquid.
- Understand the principles behind the workings of an autoclave.

Vapor Pressure

When a partially filled container of liquid is sealed with a stopper, some liquid molecules at the surface evaporate into the **vapor phase**. However, the vapor molecules cannot escape from the container. So, after a certain amount of time, the space above the liquid reaches a point where it cannot hold any more vapor molecules. Now, some of the vapor molecules condense back into a liquid. The system reaches the point where the rate of evaporation is equal to the rate of condensation (see figure 12.2.1). This is considered an equilibrium between the liquid and vapor phase.

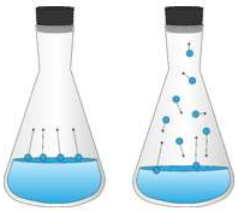


Figure **12.2.1**: Equilibrium between liquid phase and vapor phase.

An equilibrium can be illustrated by an equation with a double arrow, meaning that the process is occurring in both directions and at the same rate.

$$H_2O(l) \rightleftharpoons H_2O(g)$$

The forward direction represents the evaporation of water, while the reverse direction represents the condensation of water vapor.

Because the vapor molecules cannot escape the container, the vapor molecules above the surface of the liquid exert a pressure on the walls of the container. The **vapor pressure (P_{vap})** is a measure of the pressure (force per unit area) exerted by a gas above a liquid in a sealed container. Vapor pressure is a property of a liquid based on the strength of its intermolecular forces. A liquid with weak intermolecular forces evaporates more easily and has a higher vapor pressure. A liquid with stronger intermolecular forces does not evaporate easily, and thus has a lower vapor pressure.

Vapor Pressure and Temperature

The temperature of water also plays a role in the magnitude of vapor pressure (P_{vap}). The vapor pressure of liquids increases with temperature. This is due to increased kinetic energy of molecules at higher temperature. More molecules have the minimum kinetic energy to be bounced of the surface of water. Table 12.2.1 shows the variations of vapor pressure as temperature increases.

Table **12.2.1**: Vapor Pressure of Water at Various Temperatures

Temperature	Vapor Pressure
20.0 °C	17.5 mm Hg
40.0 °C	55.3 mm Hg
60.0 °C	149.4 mm Hg
80.0 °C	355.1 mm Hg
100.0 °C	760.0 mm Hg
125.0 °C	1740.9 mm Hg

Boiling and Boiling Point

Boiling is a special form of evaporation where conversion from a liquid state to the gaseous state occurs in a liquid through bubble formation. For liquids in open containers the vapor pressure (P_{vap}) steadily increases as the liquid is heated. The P_{vap} reaches a value that is equal to the atmospheric pressure (P_{atm}). When these two pressures are equal the vapor bubbles quickly rise to the surface because they are less dense than the liquid itself and the liquid is said to be boiling. The *normal boiling point* of water is 100.0 ^{o}C under an atmospheric pressure of 760.0 mm of Hg (1.00 atm).

$$P_{atm} = P_{vap} \quad \Longrightarrow \quad \text{Water Boils}$$

Note

The **boiling point** of a liquid is the temperature at which the vapor pressure of the liquid equals the atmospheric pressure.

Altitude and Boiling Point

The *actual boiling point* depends on the atmospheric pressure. At a pressure greater than 1.00 atm, water boils at a temperature greater than 100.0 °C because the increased pressure forces vapor molecules above the surface to condense. Hence the molecules must have greater kinetic energy to escape from the surface. Conversely, at pressures less than 1.00 atm, water boils below 100.0°C.

At high altitudes, the dependence of the boiling point of water on pressure is significant. Table 12.2.2 lists the boiling points of water at several locations with different altitudes. The lower boiling point of water has major consequences for cooking. It will take longer to cook a pot of noodles in the Rockies and even longer in the Himalayas. Cake mixes are often sold with separate high-altitude instructions.

Table **12.2.2**: The Boiling Points of Water at Various Locations on Earth

Place	Altitude above Sea Level (ft)	Atmospheric Pressure (mmHg)	Boiling Point of Water (°C)
Mt. Everest, Nepal/Tibet	29,028	240	70
Bogota, Colombia	11,490	495	88
Denver, Colorado	5280	633	95
Washington, DC	25	759	100
Dead Sea, Israel/Jordan	−1312	799	101.4

Pressure Cookers and Autoclave

Pressure cookers, which have a seal that allows the pressure inside them to exceed 1 atm, are used to cook food more rapidly by raising the boiling point of water and thus the temperature at which the food is being cooked. At twice the atmospheric pressure of 1520 torr, water boils at 120.0 ^{o}C. The high pressures reached in autoclaves, allows water to boil at temperatures high enough to kill most infectious agents. Therefore the autoclave such as in figure 12.2.2 is used for sterilization of surgical equipment in a medical facility and media and equipment in a microbiology laboratory.



Figure 12.2.2: **Autoclave**: Large autoclave used for moist sterilization of media and equipment.

Summary

The **vapor pressure** is a measure of the pressure exerted by a gas above a liquid in a closed container. The **boiling point** of a liquid is the temperature at which the vapor pressure of the liquid equals the atmospheric pressure. Autoclaves, allows water to boil at high temperatures under high pressures. The high temperatures kill most infectious agents.

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12.3: The Pressure-Temperature Law

Learning Objective

- Understand the relationship between pressure of a gas and temperature.

The properties of a gas can be related to each other under certain conditions. The properties are pressure (*P*), volume (*V*), temperature (*T*, in kelvins), and amount of material expressed in moles (*n*). A sample of gas cannot have any random values for these properties. Instead, only certain values, dictated by some simple mathematical relationships, will occur.

The first simple relationship, referred to as a gas law, is between the pressure of a gas and its temperature. If the amount of gas in a sample and its volume are kept constant, then as the temperature of a gas is increased, the pressure of the gas increases proportionately. Consider a gas in a cylinder with a piston in Figure **12.3.1**. Increasing temperature increases the average kinetic energy (KE) and the average velocity of the gas molecules resulting in more frequent and more forceful collisions which result in increased gas pressure applied on the piston or the walls of the gas container.

Gay-Lussac's law

Gay-Lussac's law states that the pressure of a gas is directly proportional to the absolute temperature provided the volume and amount of gas are not changed.

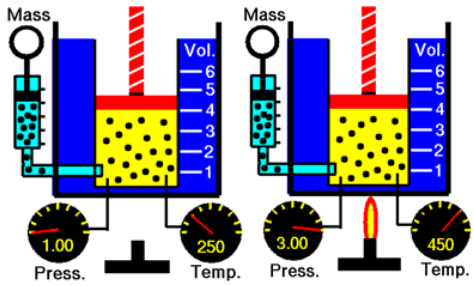


Figure **12.3.1**: Increasing temperature increases pressure, i.e., $\frac{P_1}{T_1} = \frac{P_2}{T_2}$. Source: NASA's Glenn Research Center / Public domain.

The mathematical forms of Gay-Lussac's law are the following.

$$P \propto T$$

, or

$$P = \mathrm{k}T$$

, or

$$\frac{P}{T} = \mathrm{k},$$

where k is a constant, P is pressure, and T is the temperature (in kelvin scale) of the gas. Since Since $\frac{P}{T}$ is a constant, it implies that

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} = \mathrm{k},$$

where P_1 is the initial pressure, T_1 is the initial temperature in Kelvin, P_2 is the final pressure, and T_2 is the final temperature in Kelvin, provided the amount of gas and volume do not change.

Example 12.3.1

The pressure of an oxygen tank containing 15.0 L oxygen is 965 torr at 55.0 ^{o}C. What will be the pressure when the tank is cooled to 16.0 ^{o}C.

Solution

First, convert the temperatures to the Kelvin scale before applying gas laws.

Given: T_1 = 55.0 °C + 273.15 = 328.2 K, T_2 = 16.0 °C + 273.15 = 289.2 K, P_1 = 965 torr, P_2 = ?

Formula:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2},$$

rearrange the formula to isolate the desired variable:

$$P_2 = \frac{P_1 T_2}{T_1}$$

Plug in the values in the rearranged formula and calculate:

$$P_2 = \frac{965 \text{ torr} \times 289.2 \text{ K}}{328.2 \text{ K}} = 850. \text{ torr}$$

Applications

Oxygen Cylinders

The relationship between temperature and pressure must be kept in mind while using pressurized O_2 (g) cylinders in the hospital. When oxygen is administered to a patient the pressure of the gas in the tank decreases. As per Gay-Lussac's law the temperature of the tank will decrease too. The surrounding will be cooler too. Be sure to provide an extra blanket to the your patients to keep them comfortable.

Car Tires

It is important to keep your car tires inflated to the correct pressure to maximize gas mileage. The recommended inflation pressure is based on the measurements made when the tire is cold. As you drive your car, your tire heats up. For every 10 °F increase in temperature, the tire pressure increases by 1 psi.

Summary

The pressure of a gas is directly proportional to the temperature in the Kelvin scale provided the volume and amount of gas is not changed.

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12.4: The Volume-Temperature Law

Learning Objectives

- Understand the relationship between volume of a gas and temperature.

Consider a gas under constant pressure in a cylinder with a piston in Fig. 7.3.1. Increasing temperature increases the average kinetic energy (*KE*) of the gas molecules. This results in more forceful collisions inside the chamber. The gas volume starts to increase since the pressure in the container does not change. In other words, increasing temperature increases the volume of the gas if the pressure and amount of gas are not changed.

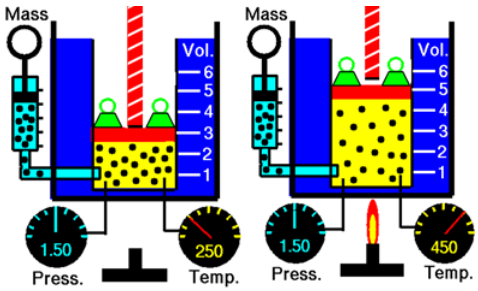


Figure **12.4.1**: Increasing temperature increases the volume of the gas i.e., $\frac{V_1}{T_1} = \frac{V_2}{T_2} = \mathbf{k}$. Source: NASA's Glenn Research Center / Public domain.

If two related parameters increase or decrease together, they are directly proportional to each other.

Charles's law

Charles's law states that the volume of a given amount of gas is directly proportional to the temperature in the Kelvin scale at constant pressure.

Fig. 12.4.2 demonstrates that the volume of a gas decreases when the gas is cooled down.



Figure **12.4.2**: Air balloon (left) shrinks when its temperature is decreased by submerging in liquid nitrogen (middle) and re-expands when returned to room temperature condition (right). Source: Ryan Poling aka expictura on Flickr.https://en.Wikipedia.org/wiki/File:Nitrogen.ogv#file, CC BY 2.0

The mathematical forms of Charles's law are the following.

$$V \propto T$$

, or

$$V = \mathrm{k}T$$

, or

$$\frac{V}{T} = \mathrm{k}$$

, where k is a constant, V is volume, and T is the temperature (in kelvin scale) of the gas. Since $\frac{V}{T}$ is a constant, it implies that

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} = \mathrm{k}$$

where V_1 is the initial volume, T_1 is the initial temperature in Kelvin, V_2 is the final volume, and T_2 is the final temperature in Kelvin, provided the amount of gas and pressure do not change. Note that the kelvin scale is used in Charles's law because the kelvin scale does not have negative numbers. If the given temperature is not in the kelvin scale, first convert the temperature to the Kelvin scale and then use the gas laws for the calculations.

Example 12.4.1

A sample of CO_2 occupies 3.23 L volume at 25.0 ^oC. Calculate the volume of the gas at 50.0 ^oC if pressure and amount of gas do not change?

Solution

Given: T_1 = 25.0 ^oC + 273.15 = 298.2 K, T_2 = 50.0^oC + 273.15 = 323.2 K, V_1 = 3.23 L, V_2 = ?

Formula: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, rearrange the formula to isolate the desired variable: $\boldsymbol{V_2 = \frac{V_1 T_2}{T_1}}$

Plug in the values in the rearranged formula and calculate: $\boldsymbol{V_2 = \frac{3.23\ \text{L} \times 323.2\ \text{K}}{298.2\ \text{K}} = 3.50\ \text{L}}$

Applications

Car Tires

As winter approaches, the temperature drops rapidly. The volume of the tire decreases as temperature decreases. For every 10°F decrease in temperature, the tire pressure decreases by 1 psi. It is important to keep your car tires inflated to the correct pressure to maximize gas mileage.

Hot Air Balloons

Charles's law explains the drifting of warm air upward in the atmosphere. As the gas is warmed, its volume increases and its density decreases which makes the gas drift upward. A hot air balloon, shown in Fig. 12.4.3 operates using hot air.



Figure **12.4.3**: A hot air balloon seen from a view directly below. The burner, or flame, is firing into the envelope above. The warm air is less dense than the atmospheric air which makes the balloon rise in the air. Source: Arpingstone at English Wikipedia / Public domain.

Summary

The volume of a gas is directly proportional to the temperature in the Kelvin scale provided the pressure and amount of gas is not changed.

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12.5: The Pressure-Volume Law

Learning Objectives

- Understand the relationship between pressure and volume of a gas.

Pressure-Volume Relationship

Consider a gas in a cylinder with a piston pushing it down as illustrated in Fig. 12.5.1. The temperature and amount of gas is kept constant. Objects are placed on the piston to change the pressure or volume of the gas in the cylinder. Increasing the pressure on the piston, e.g., by adding more weight to it, causes the piston to move down, reducing the gas volume. The gas molecules have less distance to travel before striking the piston surface which increases the collision frequency and causes the gas pressure to increase.

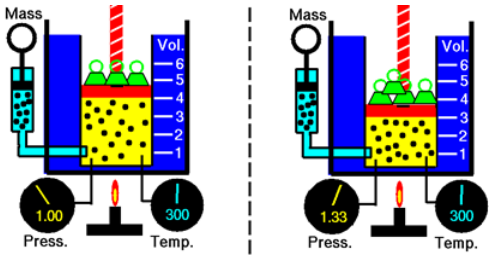


Figure **12.5.1**: Illustration of Boyle's law: for a given amount of gas at a constant temperature, the volume is inversely proportional to the pressure. Source: NASA's Glenn Research Center / Public domain.

When a change in one parameter causes a change in another, the parameters are related. When an increase in one parameter causes a decrease in another, the two are inversely proportional to each other. Robert Boyle studied the quantitative relationship between the volume and pressure of the gas, keeping the quantity of gas and the temperature constant. The research concluded in a law called Boyle's law, which states that:

Boyle's law

The volume of a gas is inversely proportional to the pressure of the gas provided the temperature and the amount of the gas remain constant.

The mathematical form of Boyle's law is:

$$V \propto \frac{1}{P}$$

, or

$$V = \frac{\mathrm{k}}{P}$$

, or

$$PV = \mathrm{k}$$

, where k is a constant. Since the product PV is a constant, it implies that:

$$P_1V_1 = P_2V_2 = k$$

, i.e., a product of initial pressure (P_1) and initial volume (V_1) is equal to the product of final pressure (P_2) and final volume (V_2) of gas provided the quantity of the gas and temperature does not change.

Example 12.5.1

The pressure of a 1.32 L sample of SO_2 gas at 0.532 atm is increased to 1.231 atm. Calculate the new volume of the gas if the temperature and the quantity of the gas remain the same?

Solution

Given: P_1 = 0.532 atm, P_2 = 1.231 atm, V_1 = 1.32 L V_2 = ?

Formula: $P_1V_1 = P_2V_2 = k$, rearrange to isolate the desired variable: $V_2 = \frac{P_1V_1}{P_2}$.

Plug in the values in the rearranged formula and calculate: $V_2 = \frac{0.532\ \text{atm} \times 1.32\ \text{L}}{1.231\ \text{atm}} = 0.570\ \text{L}$

Example 12.5.2

An oxygen tank holds 20.0 L of oxygen at a pressure of 10.0 atm. What is the final pressure when the gas is released and occupies a volume of 200. L?

Solution

Given: V_1 = 20.0 L, V_2 = 200. L, P_1 = 10.0 atm P_2 = ?

Formula: $P_1V_1 = P_2V_2 = k$, rearrange to isolate the desired variable: $P_2 = \frac{P_1V_1}{V_2}$.

Plug in the values in the rearranged formula and calculate: $P_2 = \frac{10.0\ \text{atm} \times 20.0\ \text{L}}{200.\ \text{L}} = 1.00\ \text{atm}$.

Applications

Respiration: Inhalation and Exhalation

Respiration is important for survival. Boyle's law explains the mechanism of respiration. Lungs are elastic structures like balloons placed in the thoracic cavity, as illustrated in Fig. 12.5.2. The diaphragm muscle makes a flexible floor and ribs surround the cavity.

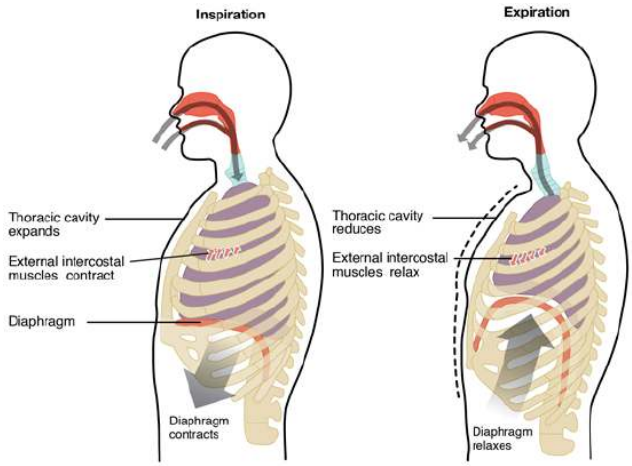


Figure 12.5.2: Illustration of breathing process –inspiration and expiration occur due to the expansion and contraction of the thoracic cavity, respectively. Source: OpenStax College / CC BY (https://creativecommons.org/licenses/by/3.0)

The inhalation or the inspiration process starts when the diaphragm moves down, expanding the thoracic cavity. Volume increases, the air pressure decreases inside the thoracic cavity and the atmospheric air flows into the lungs until the pressure in the lungs is equal to the outside pressure.

The exhalation or the expiration process starts when the diaphragm moves upwards, contracting the thoracic cavity. Volume decreases and the air pressure increases inside the thoracic cavity that pumps the air out of the lungs into the atmosphere.

Diving

When a person dives into the water, at 10 meters in depth, the pressure increases two-fold. Recall the relationship between altitude and atmospheric pressure. As altitude decreases, atmospheric pressure increases. The lungs behave like elastic balloons contracts by a factor of two. To avoid stress on the lungs, the diver requires decompression stops, to allow the lungs to adjust to the increase in atmospheric pressure.

Summary

The pressure of a gas is inversely proportional to the volume provided the temperature and amount of gas are constant.

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12.6: The Combined Gas Law

Learning Objectives

- Understand mathematical relationship between pressure, volume, and temperature give the amount of gas is constant.

Combined gas law

The laws relating to pressure P, volume V, and temperature T for a constant amount of a gas are the following:

If amount of gas is constant:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This is a combination of Boyle's law, Charles law and the Gay-Lussac's law.

1. If $T_1 = T_2$: then $P_1V_1 = P_2V_2$, this is the Boyle's law.
2. If $P_1 = P_2$: then $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, this is the Charles's law.
3. If $V_1 = V_2$: then $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, this is the Gay Lussac's law.

The combined gas law allows calculating the effect of varying two parameters on the third.

Example 12.6.1

A weather balloon contains $212\ \text{L}$ of helium at $25.0°\text{C}$ and $750.\ \text{mmHg}$. What is the volume of the balloon when it ascends to an altitude where the temperature is $-40.0°\text{C}$ and $540.\ \text{mmHg}$, assuming the quantity of gas remains the same?

Solution

Given and desired parameters (temperatures must be converted to Kelvin scale):

$$\begin{aligned} &\text{P}_1 = 750.\ \text{mmHg}, \quad \text{V}_1 = 212\ \text{L}, \quad \text{T}_1 = 25.0°\text{C} + 273.15 = 298.2\ \text{K} \\ &\text{P}_2 = 540.\ \text{mmHg}, \quad \text{V}_2 = ? \qquad\quad \text{T}_2 = -40.0°\text{C} + 273.15 = 233.2\ \text{K} \end{aligned}$$

Formula:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2},$$

rearrange the formula to isolate the desired parameter:

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}.$$

Calculations:

$$V_2 = \frac{750.\ \cancel{\text{mmHg}} \times 212\ \text{L} \times 233.2\ \cancel{\text{K}}}{298.2\ \cancel{\text{K}} \times 540.\ \cancel{\text{mmHg}}} = 230.\ \text{L}.$$

Summary

The combined gas law is a combination of Boyle's, Charles', and the Gay-Lussac's law.

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12.7: Henry's Law

Learning Objectives

- Describe the effect of pressure on the solubility of gases in water and apply this concept in health-related conditions.

Solubility of a Gas in Water

Earlier you were introduced to the effect of intermolecular attractive forces on solution formation. The chemical structures of the solute and solvent dictate the types of forces possible and, consequently, are important factors in determining solubility.

For example, under similar conditions, the water solubility of oxygen is approximately three times greater than that of helium.

The solubility of oxygen (nonpolar) in the liquid hydrocarbon hexane, C_6H_{14}(nonpolar), is approximately 20 times greater than it is in water (polar).

Solubility of a Gas and Temperature

Temperature is another factor that affects gas solubility. Gas solubility decreases as temperature increases. This is one of the major impacts resulting from the thermal pollution of natural bodies of water. When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills.

Henry's Law (Solubility of a Gas and Pressure)

The solubility of a gaseous solute is also affected by pressure. **Gas solubility increases as the pressure of the gas increases.** Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at high pressure. When the beverage container is opened, a familiar hiss is heard as the carbon dioxide gas pressure is released, and some of the dissolved carbon dioxide is typically seen leaving solution in the form of small bubbles (Figure **12.7.1**). With time, the dissolved carbon dioxide concentration will decrease to its equilibrium value and the beverage will become "flat."



Figure **12.7.1**: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (Credit: modification of work by Derrick Coetzee)

Decompression Sickness ("The Bends")

Decompression sickness (DCS), or "the bends," is an effect of the increased pressure of the air inhaled by scuba divers when swimming underwater at considerable depths. In addition to the pressure exerted by the atmosphere, divers are subjected to additional pressure due to the water above them, experiencing approximately 2 atm at 10 m of depth. Therefore, the air inhaled by a diver while submerged contains gases at the corresponding higher pressure, and the concentrations of the gases dissolved in the diver's blood are proportionally higher per Henry's law.

As the diver ascends to the surface of the water, the atmospheric pressure decreases and the dissolved gases becomes less soluble. If the ascent is too rapid, the gases escaping from the diver's blood may form bubbles that eventually bursts capillaries and cause a variety of symptoms ranging from rashes and joint pain to paralysis and death. To avoid DCS, divers must ascend from depths at relatively slow speeds (10 or 20 m/min) or otherwise make several decompression stops, pausing for several minutes at given

depths during the ascent. This allows excess dissolved gases in the blood to be exhaled. If stops or slow speeds are not an option, then the diver should exhale rapidly during the ascent.

Divers may also use a special mixture of helium and oxygen. Nitrogen in air is replaced by helium. He gas is not as soluble in the blood serum, so the effects of DCS is minimized. Under similar conditions, the water solubility of nitrogen is greater than that of helium.

When these preventive measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers.

Hyperbaric Chamber

Hyperbaric (high pressure) oxygen treatment in another application of Henry's Law. This is an air chamber that is at two to three times the atmospheric pressure. The solubility of oxygen increases with an increase in pressure. A patient placed in a hyperbaric chamber has a higher concentration of oxygen dissolved in blood. The higher concentration of oxygen is toxic to many strains of bacteria. Therefore the hyperbaric chambers are used to treat burn patients, gangrene, tetenus, in surgeries, and to treat some cancers.

The hyperbaric chambers are also used to treat carbon monoxide (CO) poisoning because the higher concentration of oxygen in the chamber can displace the CO bound with hemoglobin faster than atmospheric oxygen does.

Another use of hyperbaric chambers is to treat scuba divers suffering from the bends. If a diver ascends too quickly, the nitrogen dissolved in blood makes bubbles in the vessels that block the blood flow, a condition called bends. The divers suffering from the bends are placed in a hyperbaric chamber at high pressure, and then the pressure is slowly decreased to atmospheric pressure. The nitrogen dissolved in blood under higher pressure and slowly diffuses out through the lungs as the pressure is gradually decreased.



(a) (b)

Figure **12.7.2**: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.

Summary

The solubility of a gas in a liquid is proportional to the pressure of the gas over the liquid. The higher the pressure of a gas above a liquid, the greater its solubility.

Contributions & Attributions

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12.8: Partial Pressure and Dalton's Law

Learning Objectives

- Explain Dalton's law of partial pressure and apply the partial pressure concept in respiration.

Dalton's law of Partial Pressures

Dalton's law applies to *a* mixture of two or more gases. This law states that in a mixture of gases, the total pressure is the sum of the partial pressures of all the components. If there are three different gases in a container then the law is given by the following mathematical equation.

$$P_{total} = P_1 + P_2 + P_3$$

The **partial pressure** of a gas is the pressure that gas would exert if it occupied the container by itself. Partial pressure is represented by the letter *P*.

Figure 12.8.1 below demonstrates the concept of partial pressure in more concrete terms, showing the pressure of each gas alone in a container and then showing the gases combined pressure once mixed.



Figure **12.8.1**: Alone, Hydrogen has a pressure of 0.4 atm, while Nitrogen has a pressure of 0.6 atm. When placed in the same container, the total pressure is 1 atm, with each gas contributing its partial pressure. Image credits: OpenStax College

Air is a mixture of 78.0 % nitrogen, 21.1 % oxygen, and 0.9 % of other gases such as argon, helium, carbon dioxide, water vapors, methane, and hydrogen. The atmospheric pressure is the sum of the partial pressures of components of air. At sea level the atmospheric pressure is 760 torr.

$$\text{At sea level: } P_{total} = P_{atm} = 760 \text{ torr} = P_{N2} + P_{O2} + P_{others} = 593 \text{ torr} + 160. \text{ torr} + 7 \text{ torr}$$

At the top of Mount Everest the atmospheric pressure is 200 torr.

$$\text{Mount Everest: } P_{total} = P_{atm} = 200 \text{ torr} = P_{N2} + P_{O2} + P_{others} = 156 \text{ torr} + 42.2. \text{ torr} + 2 \text{ torr}$$

Therefore as altitude increases, and atmospheric pressure decreases, the partial pressure of oxygen decreases from 160 torr at sea level to 42.2 torr at Mount Everest. So climbers have to carry oxygen supplies. Otherwise climbers face the risk of **Acute Mountain Sickness**. Some of the symptoms are headaches, nausea, and shortness of breath.

Some runners take advantage of high altitude training. Why? P_{O2} is low at high altitude and so training at high altitudes increases aerobic capacity. The person tends to make more red blood cells to increase the intake of oxygen.

The concept of partial pressures also explains why some people with respiratory problems are uncomfortable in low pressure air that is very moist. When the atmospheric pressure is low and the air is moist the $P_{water\ vapor}$ is high and so the P_{O2} is low.

Basic Principles of Gas Exchange during Respiration

Gas exchange during respiration occurs primarily through diffusion. Diffusion is a process in which transport is driven by a concentration gradient. Gas molecules move from a region of high concentration to a region of low concentration. Blood that is low in oxygen concentration and high in carbon dioxide concentration undergoes gas exchange with air in the lungs. The air in the lungs has a higher concentration of oxygen and lower in concentration of carbon dioxide than that of oxygen-depleted blood. This concentration gradient allows for gas exchange during respiration.

Partial pressure is a measure of the concentration of the individual components in a mixture of gases. The partial pressure of a gas above a liquid is directly related to the amount of gas dissolved in the liquid. The rate of diffusion of a gas is proportional to its partial pressure within the total gas mixture.

Partial pressure of oxygen in ambient air is 160 mm Hg. In the arterial blood, the P_{O2} = 100 mm Hg. The venous blood has a P_{O2} = 40 mm Hg. The P_{O2} in the remote tissues is less than 40 torr. This pressure gradient drives the oxygen into the tissue cells via diffusion. See figure 12.8.2 for visual connection.

At the same time, the venous blood has a P_{CO2} = 46 mm Hg. The arterial blood has a P_{CO2} = 40 mm Hg and the ambient air has a P_{CO2} = 0.2 mm Hg A net movement of CO_2 takes place. The pressure gradient drives CO_2 out of tissue cells. CO_2(g) is a bi-product of metabolism.

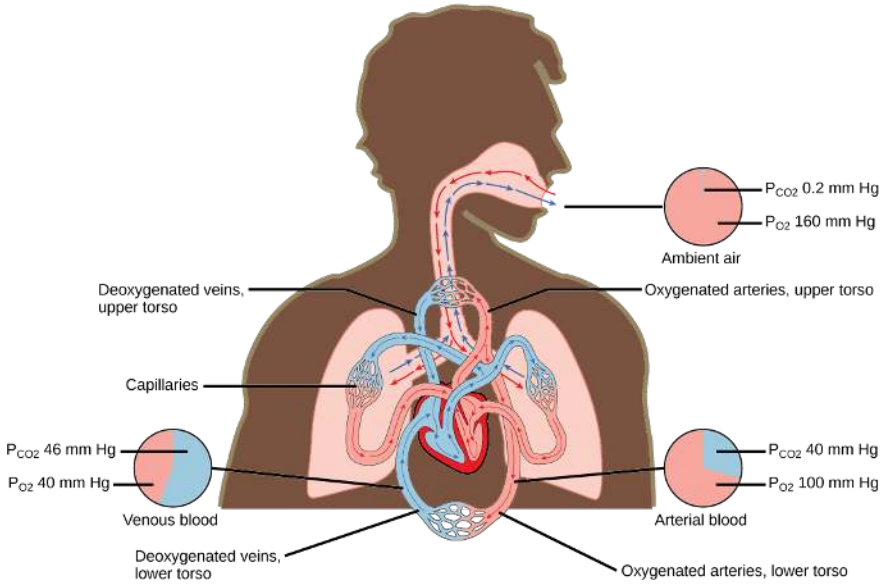


Figure 12.8.2: The partial pressures of oxygen and carbon dioxide change as blood moves through the body.

In short, the changes in partial pressure of gases drives the oxygen into the tissues and the carbon dioxide into the blood from the tissues. The blood is then transported to the lungs where differences in partial pressures result in the movement of carbon dioxide out of the blood into the lungs, and oxygen into the blood.

Summary

The partial pressure of a gas is the pressure that gas would exert if it occupied the container by itself. Dalton's law applies to *a* mixture of two or more gases. This law states that in a mixture of gases, the total pressure is the sum of the partial pressures of all the components. Gas exchange during respiration occurs primarily through diffusion. Gas molecules move from a region of high concentration to a region of low concentration.

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12.9: End of Chapter Problems

Gases and Pressure

1. How many atmospheres are in 1,547 mm of Hg?

2. How many millimeters of mercury are in 9.65 atm?

3. Covert 9.2 psi to:

a. atm.

b. torr.

4. Explain how a barometer works?

5. What is systolic pressure and diastolic pressure?

6. What is STP?

7. If you go camping in the mountains, why does it take longer to cook rice than it does at sea level? Explain with the help of the vapor pressure concept.

8. Explain the working of the autoclave with the help of the vapor pressure concept.

Gas Law Applications

Identify the gas laws involved in the following problems:

1. What is the relationship between temperature and pressure that must be kept in mind when oxygen cylinders are in use at the **hospital**?

2. Aerosol cans should be stored in a cool place. Why? Explain with the help of a gas law.

3. Early winter, your tires on your car look deflated. Why? Explain with the help of a gas law.

4. Rapid ascent in an elevator may cause the volume of intestinal gas to increase and cause intestinal cramps. Why? Explain with the help of a gas law.

5. In respiration (exhaling), explain how changes in pressure in chest cavity changes the volume of the lungs. What gas law applies?

6. In respiration (inhaling), explain how changes in pressure in chest cavity changes the volume of the lungs. What gas law applies?

7. Suppose you were diving rather deep in the ocean and had to surface quickly while leaving your tank at the bottom. What precaution would you take so as not to put pressure on your **lungs**? Gas law?

8. During severe bleeding, ADH (a hormone released by the hypothalamus) causes vasoconstriction (shrinking of the blood vessels) to take place. What effect does a decrease in blood vessel volume have on blood pressure?

9. What happens to the solubility of carbon dioxide gas in water when:

   a. the pressure of CO_2 over the solution is increased.

   b. the temperature is increased.

10. Why is breathing 100% O_2 in a pressurized hyperbaric chamber more effective at treating gangrene than breathing 100% O_2 at atmospheric pressure?

11. What is Bends? Explain with the help of a solubility law.

12. Two unopened bottles of carbonated water are at the same temperature. If one is opened at the top of a mountain and the other at sea level, which will produce more bubbles? Explain

13. Some people with respiratory problems is very uncomfortable in low pressure air that is very moist. Why

Gas Laws Calculations

Identify the gas laws involved in the following problems:

1. The pressure of an oxygen tank containing 15.0 L oxygen is 965 torr at 55.0 ^{o}C. What will be the pressure when the tank is cooled to 16.0 ^{o}C.

2. A sample of CO_2 occupies 3.23 L volume at 25.0 ^{o}C. Calculate the volume of the gas at 50.0 ^{o}C if pressure and amount of gas do not change?

3. The pressure of a 1.32 L sample of SO_2 gas at 0.532 atm is increased to 1.231 atm. Calculate the new volume of the gas if the temperature and the quantity of the gas remain the same?

4. An oxygen tank holds 20.0 L of oxygen at a pressure of 10.0 atm. What is the final pressure when the gas is released and occupies a volume of 200. L?

5. A weather balloon contains 212 L of helium at 25.0∘C and 750. mmHg. What is the volume of the balloon when it ascends to an altitude where the temperature is −40.0∘C and 540. mmHg, assuming the quantity of gas remains the same?

6. A sample of a gas at a temperature of 37.0^{o}C has a pressure of 725 torr and occupies a volume of 3.44 L.

a. Calculate the new pressure is the temperature is held constant and the volume is increased to 6.17 L.

b. Calculate the new pressure if the volume is held constant and the temperature is decreased to 10.0^{o}C.

c. Calculate the new volume if the pressure is held constant and the temperature is increased to 150.0^{o}C.

d. Calculate the new pressure if the temperature is decreased to 25.0^{o}C and the volume is increased to 15.9 L.

Respiration and Gases

1. Explain with the concept of partial pressures how oxygen gas (O_2) moves from the lungs to the tissues.

2. Explain with the concept of partial pressures how carbon dioxide gas (CO_2) moves from the tissues to the lungs.

3. Some runners train at high altitudes. Why?

4. Why does O_2 have greater solubility in blood than in water?

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Glossary

Acetyl-CoA | A molecule that participates in many biochemical reactions in protein, carbohydrate and lipid metabolism. Its main function is to deliver the acetyl group to the citric acid cycle (Krebs cycle) to be oxidized for energy production.

addition polymerization | A reaction in which monomers add to one another to produce a polymeric product that contains all the atoms of the starting monomers.

addition reactions | A reaction in which substituent groups join to hydrocarbon molecules at points of unsaturation—the double or triple bonds.

alcohol | An organic compound with an OH functional group on an aliphatic carbon atom.

aldehyde | An organic compound with a carbonyl functional group that has an hydrogen atom attached and either a hydrocarbon group or a second hydrogen atom.

alkaloid | A nitrogen-containing organic compound obtained from plants that has physiological properties.

alkanes (or saturated hydrocarbons) | A hydrocarbon with only carbon-to-carbon single bonds and existing as a continuous chain of carbon atoms also bonded to hydrogen atoms

alkenes | A hydrocarbon with one or more carbon–carbon double bonds.

alkyl group | A hydrocarbon group derived from an alkane by removal of a hydrogen atom.

alkyl halide (or haloalkane) | A compound resulting from the replacement of a hydrogen atom of an alkane with a halogen atom.

Alkynes | A hydrocarbon with a carbon–carbon triple bond.

alloy | A solid solution of a metal with other substances dissolved in it.

alpha particle | A type of radioactive emission that is equivalent to a helium atom nucleus.

amide | An organic compound with a carbonyl group joined to a nitrogen atom from ammonia or an amine.

amine | An organic compound derived from ammonia by the replacement of one, two, or three of the hydrogens atoms by alkyl or aryl groups.

amino group | An NH_2 unit.

amorphous | A solid with no regular structure.

amphiprotic | A substance that can either donate or accept a proton, depending on the circumstances.

amylopectin | A branched polymer of glucose units found in starch.

anabolism | Metabolic reactions in which molecules are synthesized.

anaerobic metabolism | A biochemical process that takes place in the absence of oxygen.

anions | A negatively charged ion.

anomeric carbon | The carbon atom that was the carbonyl carbon atom in the straight-chain form of a monosaccharide.

anticodon | A set of three nucleotides on the tRNA that is complementary to, and pairs with, the codon on the mRNA.

antioxidants | A substance in foods that acts as a reducing agent.

aromatic compound | Any compound that contains a benzene ring or has certain benzene-like properties.

aromatic hydrocarbons | A hydrocarbon with a benzene-like structure.

aryl group | A group derived from an aromatic hydrocarbon by the removal of a hydrogen atom.

atomic bomb | A weapon that depends on a nuclear chain reaction to generate immense forces.

atomic mass | A weighted average of the masses of all the element's naturally occurring isotopes.

atomic mass unit | One-twelfth the mass of a ^{12}C atom.

atomic radius | The approximate size of an atom.

autoionization of water | The process by which water ionizes into hydronium ions and hydroxide ions as it acts as an acid and a base.

Avogadro's number | The value 6.022×10^{23}.

balanced | A property of a chemical equation when there are the same number of atoms of each element in the reactants and products.

base | A compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.

Base (or basic) units | A fundamental unit of SI.

beta particle | A type of radioactive emission that is equivalent to an electron.

Bilayers | A double layer of lipids arranged so that nonpolar tails are found between an inner surface and outer surface consisting of hydrophilic heads.

Bile | The yellowish green liquid produced in the liver.

biochemistry | The chemistry of molecules found in living organisms.

boiling point | The temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid).

boiling point elevation | The raising of the boiling point of a solution versus the pure solvent.

bond length | The distance between two nuclei in a covalent bond.

Boyle's law | The gas law that relates pressure and volume.

Brønsted-Lowry base | A compound that accepts a hydrogen ion (H^+) in a reaction; a proton acceptor.

buffer | A solution that resists dramatic changes in pH.

calorie | A unit of energy widely used in the health professions and everyday life.

capacity | The amount of strong acid or base a buffer can counteract.

carbohydrates | A compound composed of carbon, hydrogen, and oxygen atoms that is a polyhydroxy aldehyde or ketone or a compound that can be broken down to form such a compound. It is one of the three main components of the human diet.

carboxyl group | A functional group that contains a carbon–oxygen bond and an OH group also attached to the same carbon atom.

carboxylic acids | An organic compound that has a carboxyl functional group.

cations | A positively charged ion.

cerebrosides | A sphingolipid that contains a fatty acid unit, a sphingosine unit, and galactose or glucose.

chain reaction | An exponential growth in a process.

Charles's law | The gas law that relates volume and absolute temperature.

chemical bond | A very strong attraction between two atoms.

chemical equilibrium (or equilibrium) | The condition in which the extent of a chemical reaction does not change any further.

chemical formula | A concise list of the elements in a compound and the ratios of these elements.

Chemical properties | A characteristic that describes how matter changes its chemical structure or composition.

chemical reaction | A representation of a chemical change.

chemical symbol | A one- or two-letter abbreviation for an element.

chiral carbon | A carbon atom that has four different groups attached to it.

Cholesterol | A steroid that is found in mammals.

cis-trans isomers (or geometric isomers) | Isomers that have different configurations because of the presence of a rigid structure such as a double bond or ring.

citric acid cycle | A cyclic sequence of reactions that brings about the oxidation of a two-C unit to carbon dioxide and water.

codon | A set of three nucleotides on the mRNA that specifies a particular amino acid.

coefficient | A number that gives the number of molecules of a substance in a balanced chemical equation.

coenzymes | A cofactor that is an organic molecule.

colligative properties | A characteristic of solutions that depends only on the number of dissolved particles.

combination (composition) reaction | A chemical reaction that makes a single substance from two or more reactants.

combined gas law | The gas law that relates pressure, volume, and absolute temperature.

combustion reaction | A chemical reaction in which a substance combines with molecular oxygen to make oxygen-containing compounds of other elements in the reaction.

competitive inhibitor | A compound that resembles a particular substrate and competes with the substrate for binding at the active site of an enzyme to slow the rate of the reaction.

complementary bases | Specific base pairings in the DNA double helix.

compound | A substance that can be broken down into chemically simpler components.

concentration | How much solute is dissolved in a certain amount of solvent.

condensed structural formulas | An organic chemical formula that shows the hydrogen atoms (or other atoms or groups) right next to the carbon atoms to which they are attached.

conversion factor | A fraction that has equivalent quantities in the numerator and the denominator but expressed in different units.

core electrons | An electron in a lower-numbered shell of an atom.

covalent network bonding | A type of interaction in which all the atoms in a sample are covalently bonded to other atoms.

curie (Ci) | A unit of radioactivity equal to 3.7×10^{10} decays per second.

cyclic hydrocarbons | A hydrocarbon with a ring of carbon atoms.

cycloalkanes | A cyclic hydrocarbon with only single bonds.

cytochromes | A protein that contains an iron porphyrin in which iron can alternate between Fe(II) and Fe(III).

cytoplasm | Everything between the cell membrane and the nuclear membrane.

decomposition reaction | A chemical reaction in which a single substance is converted into two or more products.

Denaturation | Any change in the three-dimensional structure of a macromolecule that renders it incapable of performing its assigned function.

Density | The mass of an object divided by its volume.

Derived units | A combinations of the SI base units.

diatomic molecules | A two-atom grouping that behaves as a single chemical entity.

digestion | The breakdown of food molecules by hydrolysis reactions into the individual monomer units in the mouth, stomach, and small intestine.

dipole-dipole interaction | An attraction between polar molecules.

Dispersion forces | A force caused by the instantaneous imbalance of electrons about a molecule.

dispersion forces (or London forces) | A force caused by the instantaneous imbalance of electrons about a molecule.

dissociation | The process of cations and anions of an ionic solute separating when the solute dissolves.

Disulfide linkages | A covalent bond that forms by the oxidation and linkage of two sulfur atoms from the side chains of two cysteine residues.

double bond | Two pairs of electrons being shared by two atoms in a molecule.

electron | A subatomic particle with a negative electric charge.

electron configuration | A shorthand description of the arrangement of electrons in an atom.

electron transport chain (or respiratory chain) | An organized sequence of oxidation-reduction reactions that ultimately transports electrons to oxygen, reducing it to water.

electronegativity | A relative measure of how strongly an atom attracts electrons when it forms a covalent bond.

emulsion | A dispersion of two liquids that do not normally mix.

enantiomers | Stereoisomers that are nonsuperimposable mirror images of each other.

endothermic | A process that absorbs energy.

Energy | The ability to do work.

equivalents (Eq) | One mole of charge (either positive or negative).

essential fatty acids | A fatty acid that must be obtained from the diet because it cannot be synthesized by the human body.

ester | An organic compound derived from a carboxylic acid and an alcohol in which the OH of the acid is replaced by an OR group.

esterification | The formation of an ester from a carboxylic acid and an alcohol.

ether | An organic compound that has an oxygen atom between two hydrocarbon groups.

Exact numbers | A number that is defined or counted.

exothermic | A process that gives off energy.

fats | A compound, composed largely of hydrocarbon chains, that supplies energy for the body.

Feedback inhibition | A normal biochemical process that makes use of noncompetitive inhibitors to control some enzymatic activity.

fibrous proteins | A protein that is elongated or fiberlike and insoluble in water.

formula mass | The sum of the masses of the elements in the formula of an ionic compound.

formula unit | A set of oppositely charged ions that compose an ionic compound.

freezing point depression | The lowering of the freezing point of a solution versus the pure solvent.

functional group | A structural arrangement of atoms and/or bonds that imparts a wide range of important properties to organic compounds.

Fusion | A nuclear process in which small nuclei are combined into larger nuclei, releasing energy.

galactosemia | A genetic disease caused by the absence of one of the enzymes needed to convert galactose to glucose.

gamma rays | A type of radioactive emission that is a very energetic form of electromagnetic radiation.

gangliosides | A sphingolipid that contains a fatty acid unit, a sphingosine unit, and a complex oligosaccharide.

gas law | A simple mathematical formula that relates two or more properties of a gas.

Gastric juice | A mixture of water, inorganic ions, hydrochloric acid, and various enzymes and proteins found in the stomach.

Geiger counter | An electrical device that detects radioactivity.

genes | The basic unit of heredity.

genetic code | The identification of each group of three nucleotides and its particular amino acid.

genetic diseases | A hereditary condition caused by an altered DNA sequence.

Globular proteins | A protein that is generally spherical in structure and soluble in water.

glycols | An alcohol with two OH functional groups.

glycolysis | The metabolic pathway in which glucose is broken down to two molecules of pyruvate with the corresponding production of ATP.

glycosidic linkage | The carbon–oxygen-carbon linkage between monosaccharide units in more complex carbohydrates, such as disaccharides or polysaccharides.

groups (or families) | A column of elements on the periodic table.

half reactions | A chemical reaction that shows only oxidation or reduction.

half-life | The amount of time it takes for one-half of a radioactive isotope to decay.

halogenated hydrocarbons | A hydrocarbon in which one or more hydrogen atoms has been replaced by a halogen atom.

halogenation | A reaction in which a halogen reacts at a carbon-to-carbon double or triple bond to add halogen atoms to carbon atoms.

heat | The transfer of energy from one part of the universe to another due to temperature differences.

heat of vaporization | The amount of heat per gram or per mole required for a phase change that occurs at the boiling point.

heterocyclic compounds | A cyclic compound in which one or more atoms in the ring is an element other than a carbon atom.

homogeneous mixtures | A mixture that acts as a single substance so that it is not obvious that two or more substances are present.

homologous series | Any family of compounds in which adjacent members differ from each other by a definite factor.

hydration | Solvation by water molecules.

Hydrogen bonding | Bonding between a highly electronegative oxygen atom or nitrogen atom and a hydrogen atom attached to another oxygen atom or nitrogen atom.

hydrogenation | A reaction in which hydrogen gas reacts at a carbon-to-carbon double or triple bond or a carbon-to-oxygen double bond to add hydrogen atoms to carbon atoms.

hydrolysis | The reaction of a substance with water.

ideal gas law | The gas law that relates volume, pressure, temperature, and amount of a gas.

ideal gas law constant | The constant the appears in the ideal gas law.

immiscible | Liquids that do not dissolve in each other.

induced-fit model | A model that says an enzyme can undergo a conformational change when it binds substrate molecules.

inner transition metals | An element in the two rows beneath the main body on the periodic table. Such metals are also called the lanthanide and actinide elements.

inorganic chemistry | The study of the chemistry of all other elements.

intermolecular interactions | A force of attraction between different molecules.

Ionic bonding | Bonding that results from electrostatic attractions between positively and negatively charged groups.

ionic compounds | A compound formed with an ionic bond.

ionic interactions | An attraction due to ions of opposite charges.

irreversible inhibitor | A substance that inactivates an enzyme by bonding covalently to a specific group at the active site.

isoelectric point | The pH at which a given amino acid exists in solution as a zwitterion.

isomers | Compounds having the same molecular formula but different structural formulas and properties.

isothermal | A process that occurs at constant temperature.

isotopes | Atoms of the same element that have different numbers of neutrons.

IUPAC System of Nomenclature | A systematic way of naming chemical substances so that each has a unique name.

joule | The SI unit of energy, work, and heat.

ketogenic amino acids | An amino acid that is converted to acetoacetyl-CoA or acetyl-CoA, which can be used for the synthesis of ketone bodies but not glucose.

ketone | An organic compound whose molecules have a carbonyl functional group between two hydrocarbon groups.

ketoses | A monosaccharide that contains a ketone functional group on the second carbon atom.

kinetic theory of gases | The fundamental theory of the behavior of gases.

L sugars | A sugar whose Fischer projection terminates in the same configuration as L-glyceraldehyde.

lattice energy | The strength of interactions between atoms that make ionic bonds.

law | A general statement that explains a large number of observations.

law of conservation of matter | In any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant.

Lewis diagrams | A representation that shows valence electrons as dots around the chemical symbol of an atom (also called Lewis electron dot diagrams).

line-angle formula | An organic chemical formula in which carbon atoms are implied at the corners and ends of lines. Each carbon atom is understood to be attached to enough hydrogen atoms to give each carbon atom four bonds.

lipids | A compound isolated from body tissues that is more soluble in organic solvents than in water.

lock-and-key model | A model that portrays an enzyme as conformationally rigid and able to bond only to a substrate or substrates that exactly fit the active site.

mass number | The sum of the numbers of protons and neutrons in a nucleus of an atom.

mass-mass calculations | A stoichiometry calculation converting between the mass of one substance and the mass of a different substance in a chemical reaction.

mass/mass percent | A concentration unit that relates the mass of the solute to the mass of the solution.

mass/volume percent | A concentration unit that relates the mass of the solute to the volume of the solution.

Matter | Anything that has mass and takes up space.

metabolic pathway | A series of biochemical reactions by which an organism converts a given reactant to a specific end product.

microscopic | A view of the universe in which one is working with a few atoms or molecules at a time.

mitochondria | Small, oval organelles with double membranes; the “power plants” of a cell.

modern atomic theory | The fundamental concept that all elements are composed of atoms.

molar mass | The mass of 1 mol of atoms or molecules.

Molarity | Number of moles of solute per liter of solution.

mole-mass calculations | A stoichiometry calculation converting between masses and moles of different substances in a chemical reaction.

mole-mass conversion | The conversion from moles of material to the mass of that same material.

molecular formulas | A chemical formula for a covalent compound.

molecular mass | The mass of a molecule, which is the sum of the masses of its atoms.

molecule | A discrete group of atoms connected by covalent bonds.

mutagens | A chemical or physical agent that cause mutations.

mutarotation | The ongoing interconversion between anomeric forms of a monosaccharide to form an equilibrium mixture.

mutation | Any chemical or physical change that alters the nucleotide sequence in DNA.

neutralization | The reaction of acid and base to make water and a salt.

neutron | A subatomic particle with no electric charge.

nomenclature | The systematic naming of chemical compounds.

nonbonding pairs (or lone pairs) | Electron pair that does not participate in covalent bonds.

noncompetitive inhibitor | A compound that can combine with either the free enzyme or the enzyme-substrate complex at a site distinct from the active site to slow the rate of the reaction.

nonelectrolytes | A compound that does not ionize at all when it dissolves.

nonpolar covalent bond | A covalent bond with a balanced electron distribution across the bond.

Nuclear energy | The controlled harvesting of energy from fission reactions.

nuclear reactor | An apparatus designed to carefully control the progress of a nuclear reaction and extract the resulting energy for useful purposes.

nucleotides | A monomer unit that is linked together to form nucleic acids.

nucleus | The central part of an atom that contains protons and neutrons.

octet rule | The idea that atoms tend to have eight electrons in their valence shell.

oil | A triglyceride that is a liquid at room temperature.

optimum pH | The pH at which a particular enzyme exhibits maximum activity.

Organic chemistry | The study of the chemistry of carbon compounds.

organic compound | A compound containing carbon atoms.

Osmolarity | A way of reporting the total number of particles in a solution to determine the osmotic pressure.

osmotic pressure | The tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal.

oxidative deamination | A reaction in which glutamate loses its amino group as an ammonium ion and is oxidized back to α-ketoglutarate.

oxidative phosphorylation | The process that links ATP synthesis to the operation of the electron transport chain.

oxidizing agent | A species that causes oxidation, which is itself reduced.

parts per billion (ppb) | The mass of a solute compared to the mass of a solution times 1,000,000,000.

peptide bond | The amide bond joining two amino acid units in a peptide or protein.

period | A row of elements on the periodic table.

periodic table | A chart of elements that groups the elements by some of their properties.

peripheral proteins | A protein that is more loosely associated with the membrane surface.

pH scale | A logarithmic scale that relates the concentration of the hydrogen ion in solution.

phase | A form of matter that has the same physical properties throughout.

phase change | A physical process in which a substance goes from one phase to another.

phases | A certain form of matter that includes a specific set of physical properties.

phenols | An aromatic compound with an OH group attached directly to a benzene ring.

photosynthesis | The process by which plants use solar energy to convert carbon dioxide and water to glucose.

point mutations | A change in which one nucleotide is substituted, added, or deleted.

polar | A molecule with a net unequal distribution of electrons in its covalent bonds.

polyamide | A condensation polymer in which the monomer units are joined by an amide linkage.

polyatomic ions | An ion with more than one atom.

polycyclic aromatic hydrocarbons (PAHs) | An aromatic hydrocarbon consisting of fused benzene rings sharing a common side.

polymers | A giant molecule formed by the combination of monomers in a repeating manner.

polypeptides | A chain of about 50 or more amino acids.

polysaccharides | A carbohydrate containing many monosaccharide units.

polyunsaturated fatty acids | A fatty acid that has two or more carbon-to-carbon double bonds.

power | The exponent in a number expressed in scientific notation.

pressure | Force divided by area.

primary (1°) alcohol | A compound with an OH group on a carbonatom that is attached to only one other carbon atom.

primary structure | The sequence of amino acids in a polypeptide chain or protein.

products | A substance on the right side of the arrow in a chemical equation.

Proteins | A compound of high molar mass consisting largely or entirely of amino acids linked together.

proton | A subatomic particle with a positive charge.

purines | A heterocyclic amine consisting of a pyrimidine ring fused to a five-member ring with two nitrogen atoms.

quantized | Having a fixed value.

quantum mechanics | The modern theory of electron behavior.

quaternary structure | The arrangement of multiple subunits in a protein.

rad | A unit of radioactive exposure equal to 0.01 J/g of tissue.

radioactivity | Emanations of particles and radiation from atomic nuclei.

reducing sugar | Any carbohydrate capable of reducing a mild oxidizing agent, such as Tollens' or Benedict's reagents, without first undergoing hydrolysis.

rem | A unit of radioactive exposure that includes a factor to account for the type of radioactivity.

replication | The process in which the DNA in a dividing cell is copied.

respiration | The process by which cells oxidize organic molecules in the presence of gaseous oxygen to produce carbon dioxide, water, and energy in the form of ATP.

retroviruses | An RNA virus that directs the synthesis of a DNA copy in the host cell.

ribonucleic acid | The nucleic acid responsible for using the genetic information encoded in DNA.

ribosomes | A cellular substructure where proteins are synthesized.

round | The process of assessing the final significant figure of a quantity to determine if it should be kept or moved higher.

saponification | The hydrolysis of fats and oils in the presence of a base to make soap.

science | The process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations.

scientific method | An organized procedure for learning answers to questions.

secondary (2°) alcohol | A compound with an OH group on a carbon atom that is attached to two other carbon atoms.

secondary structure | The fixed arrangement of the polypeptide backbone.

semimetals | An element whose properties are intermediate between metals and nonmetals.

shells | A grouping of electrons within an atom.

significant figures | All the digits of a measured quantity known with certainty and the first uncertain, or estimated, digit.

single bond | A covalent bond formed by a single pair of electrons.

solubility | The limit of how much solute can be dissolved in a given amount of solvent.

solute | The minor component of a solution.

solution | Another name for a homogeneous mixture.

specific heat | A proportionality constant that relates heat to a temperature change.

Sphingolipids | A lipid that contains the unsaturated amino alcohol sphingosine.

Sphingomyelins | A sphingolipid that contains a fatty acid unit, a phosphoric acid unit, a sphingosine unit, and a choline unit.

spontaneous fission | The breaking apart of an atomic nucleus into smaller nuclei.

standard temperature and pressure | 273 K (0°C) and 1.00 atm pressure.

steroids | A lipid with a four-fused-ring structure.

Stock system | The system of indicating a cation's charge with roman numerals.

stoichiometry | The study of the numerical relationships between the reactants and the products in a balanced chemical equation.

strong acid | An acid that is 100% ionized in aqueous solution.

strong base | A base that is 100% ionized in aqueous solution.

structural formula | A chemical formula that shows how the atoms of a molecule are attached to one another.

subshells | A grouping of electrons within a shell.

substrate-level phosphorylation | The synthesis of ATP by the direct transfer of a phosphate group from a metabolite to ADP.

substrates | A compound on which an enzyme acts.

tertiary (3°) alcohol | A compound with an OH group on a carbon atom that is attached to three other carbon atoms.

tertiary (3°) amine | A compound that has three alkyl or aryl groups on the nitrogen atom.

tertiary structure | The unique three-dimensional shape of a polypeptide chain as a whole.

theory | A general statement that describes a large set of observations and data.

Thiols | A compound with an SH functional group.

torr | Another name for millimeters of mercury.

tracer | A substance that can be used to follow the pathway of that substance through some structure.

transamination | An exchange of functional groups between any amino acid and an α-keto acid.

transcription | The process in which RNA is synthesized from a DNA template.

translation | The process in which the information encoded in mRNA is used to direct the sequencing of amino acids to synthesize a protein.

triglycerides | An ester composed of three fatty acid units linked to glycerol and found in fats and oils.

triple bonds | Three pairs of electrons being shared by two atoms in a molecule.

unit | The scale of measurement for a quantity.

unsaturated | A solution whose solute is less than its solubility limit.

unsaturated hydrocarbons | An alkene or alkyne having one or more multiple (double or triple) bonds between carbon atoms.

valence shell electron pair repulsion | The general concept that estimates the shape of a simple molecule.

vapor pressure | The pressure of a vapor that is in equilibrium with its liquid phase.

vapor pressure depression | The lowering of the vapor pressure of a solution versus the pure solvent.

Viruses | An infectious agent that is much smaller and simpler than bacteria.

vitamins | An organic compound that is essential in very small amounts for the maintenance of normal metabolism.

volume | The amount of space that a given substance occupies.

volume/volume percent | A concentration unit that relates the volume of the solute to the volume of the solution.

weak base | A base that is less than 100% ionized in aqueous solution.

zwitterion | An electrically neutral compound that contains both negatively and positively charged groups.

β-oxidation | A sequence of four reactions in which fatty acyl-CoA molecules are oxidized, leading to the removal of acetyl-CoA molecules.

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