MPC CHEM 30A INTRODUCTION TO CHEMISTRY FOR HEALTH SCIENCES



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MPC CHEM 30A Introduction to Chemistry for Health Sciences

Compiled and edited by Jessica Baum, Ph.D

for CHEM 30A: Introductory Chemistry for Health Sciences at Monterey Peninsula College

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1: INTRODUCTION TO CHEMISTRY, THE PERIODIC TABLE, AND IONIC COMPOUNDS

1.1: Using the Scientific Method

1.2: Defining Chemistry

1.3: Classification of Matter

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1.1: Using the Scientific Method

Learning Objectives

- To identify the components of the scientific method

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments, which in turn lead to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.1.1).

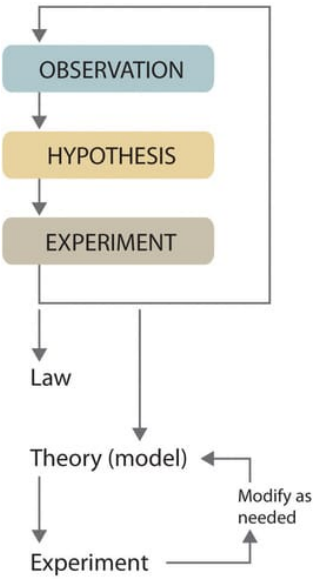


Figure 1.1.1: The Steps in the Scientific Method. Image used with permission (CC BY-SA-NC; anonymous).

Step 1: Make observations

Observations can be qualitative or quantitative. **Qualitative observations** describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: the outside air temperature is cooler during the winter season, table salt is a crystalline solid, sulfur crystals are yellow, and dissolving a penny in dilute nitric acid forms a blue solution and a brown gas. **Quantitative observations** are measurements, which by definition consist of both a number and a unit. Examples of quantitative observations include the following: the melting point of crystalline sulfur is 115.21° Celsius, and 35.9 grams of table salt—whose chemical name is sodium chloride—dissolve in 100 grams of water at 20° Celsius. For the question of the dinosaurs' extinction, the initial observation was quantitative: iridium concentrations in sediments dating to 66 million years ago were 20–160 times higher than normal.

Step 2: Formulate a hypothesis

After deciding to learn more about an observation or a set of observations, scientists generally begin an investigation by forming a hypothesis, a tentative explanation for the observation(s). The hypothesis may not be correct, but it puts the scientist's understanding of the system being studied into a form that can be tested. For example, the observation that we experience alternating periods of light and darkness corresponding to observed movements of the sun, moon, clouds, and shadows is consistent with either of two hypotheses:

a. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun, or
b. the sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. For the disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this hypothesis does not lend itself to direct testing by any obvious experiment, but scientists can collect additional data that either support or refute it.

Step 3: Design and perform experiments

After a hypothesis has been formed, scientists conduct experiments to test its validity. Experiments are systematic observations or measurements, preferably made under controlled conditions—that is, under conditions in which a single variable changes.

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether the original hypothesis is valid. In which case he can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified thus requiring further experimentation.

Step 5: Development into a law and/or theory

More experimental data are then collected and analyzed, at which point a scientist may begin to think that the results are sufficiently reproducible (i.e., dependable) to merit being summarized in a law, a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply says what happens; it does not address the question of why.

One example of a law, the law of definite proportions, which was discovered by the French scientist Joseph Proust (1754–1826), states that a chemical substance always contains the same proportions of elements by mass. Thus, sodium chloride (table salt) always contains the same proportion by mass of sodium to chlorine, in this case 39.34% sodium and 60.66% chlorine by mass, and sucrose (table sugar) is always 42.11% carbon, 6.48% hydrogen, and 51.41% oxygen by mass.

Whereas a law states only what happens, a theory attempts to explain why nature behaves as it does. Laws are unlikely to change greatly over time unless a major experimental error is discovered. In contrast, a theory, by definition, is incomplete and imperfect, evolving with time to explain new facts as they are discovered.

Because scientists can enter the cycle shown in Figure 1.1.1 at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

A Real World Application of the Scientific Method

Observation: You enter a room and flip the light switch, but the light does not go on.

Hypothesis: Determine more than one possible hypothesis to explain your observation.

Experiments: What experiments could you design and perform to test your hypothesis?

Exercise 1.1.1

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Ice always floats on liquid water.
b. Birds evolved from dinosaurs.
c. When 10 g of ice were added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
d. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution:

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is an educated guess regarding the origin of birds, so it is a hypothesis.
c. The temperature is measured before and after a change is made in a system, so these are quantitative observations.
d. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

Exercise 1.1.2

Classify each statement as a law, a theory, an experiment, a hypothesis, a qualitative observation, or a quantitative observation.

a. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
b. Heat always flows from hot objects to cooler ones, not in the opposite direction.
c. The universe was formed by a massive explosion that propelled matter into a vacuum.

Answer a

experiment

Answer b

law

Answer c

theory

Summary

The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. The key steps in the scientific method include the following:

- Step 1: Make observations.
- Step 2: Formulate a hypothesis.
- Step 3: Test the hypothesis through experimentation.
- Step 4: Accept or modify the hypothesis .
- Step 5: Development into a law and/or a theory

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1.2: Defining Chemistry

Learning Objectives

- To define and describe the science of chemistry
- To distinguish between the terms macroscopic and microscopic



Figure 1.2.1 (Credit: User:Phanton/Wikipedia; Source: https://commons.wikimedia.org/wiki/File:Decorative_Soaps.jpg; License: Public Domain)

How does soap relate to chemistry?

Chemistry affects every area of our lives. Here is just one example of chemistry in action – the making of soaps. Soap was once made by boiling animal fat in ashes – the product was hard on the skin and not very pleasant to use. Today, soap manufacture involves complicated chemical processes to provide a wide variety of soaps for different skin types. Colors and odors can be custom-made for that individual experience.

"What is this made of? How can we produce this material quickly and at a low cost? Will this product harm us or help us?"—these are all questions that can be answered using the science of chemistry.

What is Chemistry?

Science is a general term used to describe the principled, rigorous study of the natural world. Many interconnected disciplines fall under this broader concept. For example, physics is the study of motion and forces. Biology is the study of living things. Geology is the study of the Earth and the rocks and minerals of which it is comprised. Chemistry is the study of the composition of matter and the changes that matter undergoes. Matter is anything that has mass and takes up space. Virtually everything around us is matter, including both living and nonliving things. Chemistry affects nearly everything we see and every action we take. Chemistry explains why milk that is left in the refrigerator for too long turns sour. Chemistry explains why certain pollutants called chlorofluorocarbons have done lasting damage to the ozone layer of our planet. Chemistry explains why the leaves of deciduous trees turn from green in the summer to various shades of red and yellow in the autumn (Figure below).



Figure 1.2.2: Chemical reactions in the leaves of deciduous trees cause them to change color from green to red, orange, or yellow before dropping to the ground. (Credit: User:Jongleur100/Wikimedia Commons; Source: https://commons.wikimedia.org/wiki/File:Country_lane.jpg; License: Public Domain)

Chemistry touches every area of our lives. The medicines we take, the food we eat, the clothes we wear—all of these materials and more are, in some way or another, products of chemistry.

YouTube Video

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Chemists look at the world in two ways, often simultaneously. The two worlds of the chemist are the **macroscopic** world and the **microscopic** world. Macroscopic refers to substances and objects that can be seen, touched, and measured directly. Microscopic refers to the small particles that make up all matter. Chemists must observe matter and do experiments macroscopically; then make generalizations and propose explanations that are microscopic in nature. For example, anyone can observe the physical change in appearance that occurs as an iron object, such as a tractor, is left out in the elements and gradually turns to rust. However, a chemist looking at the rusting tractor considers the individual atoms that make up the iron, and how they are changing as a result of exposure to oxygen in the air, and water from rain. Throughout the study of chemistry, there is often a switch back and forth between the macroscopic and microscopic worlds.



Figure 1.2.3: *Rusting artillery shells.* (Courtesy of the US Marine Corps, Source: http://commons.wikimedia.org/wiki/File:USMC-060507-M-7799R-003.jpg; License: Public Domain)

Summary

- Chemistry is the study of matter and the changes it undergoes.
- Chemistry considers both macroscopic and microscopic information.

Review

1. Give two examples of chemistry in your everyday life.
2. What is the macroscopic world?
3. What is the microscopic world?

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1.3: Classification of Matter

Learning Objectives

- Describe the solid, liquid and gas phases.
- 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.

The States of Matter

Matter typically exists in one of three states: **solid**, **liquid**, or **gas**. The state a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states.



Figure 1.3.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From top to bottom quartz (solid), water (liquid), nitrogen dioxide (gas),

Note

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

Solid

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

Liquid

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

Mercury

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

YouTube Video

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Video $1.3.1$*: Mercury boiling to become a gas.*

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

Gas

Gases have the following characteristics:

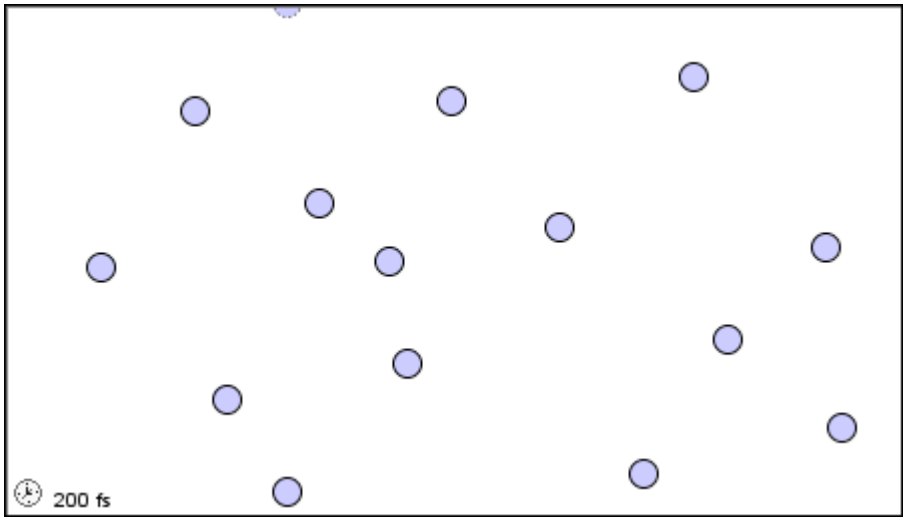
- No definite shape (takes the shape of its container)
- No definite volume
- Particles move in random motion with little or no attraction to each other
- Highly compressible

The characteristics of the three states of matter are listed in Table $1.3.1$ and an animation of the movement and position of the individual particles is shown in Figure $1.3.2$

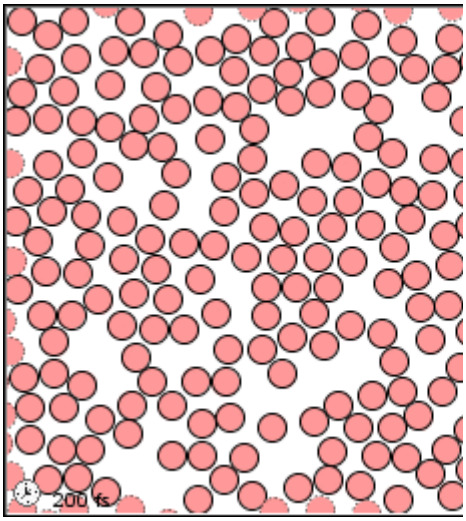
Table $1.3.1$ *Characteristics of the Three States of Matter*

Characteristics	Solids	Liquids	Gases
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak

Characteristics	Solids	Liquids	Gases
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions



The gaseous state



The liqui

Figure 1.3.2 *A microscopic model showing particles (atoms or molecules) in the gaseous, liquid, and solid states.*

Example 1.3.1

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 1.3.1

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: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Substances and Mixtures

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure 1.3.1). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning it's the same everywhere) and properties that are constant throughout the sample (meaning 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**.

Ordinary table salt is called sodium chloride. It is considered a pure 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 pure substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

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

Elements and Compounds

Elements and compounds are both examples of pure substances. A pure substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A pure 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.

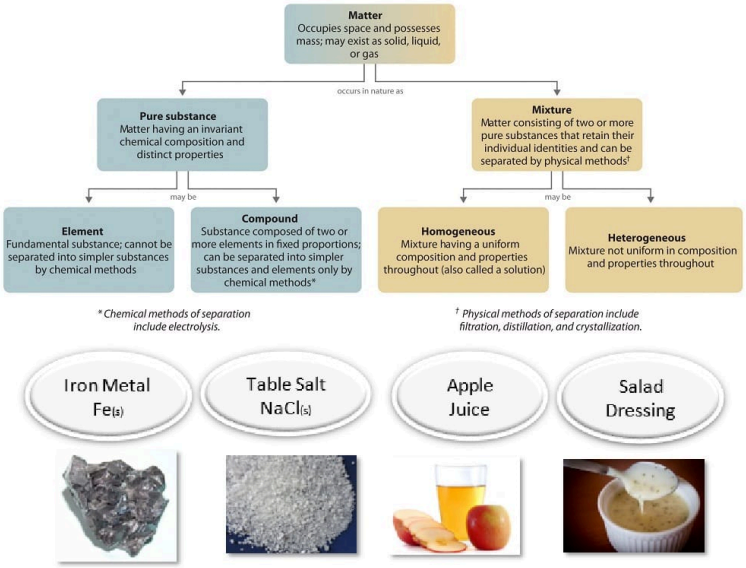


Figure 1.3.2: Relationships between the types of matter and the methods used to separate mixtures.

Atoms, Elements, and Compounds

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Video 1.3.2 *Atoms, elements, and compounds (first five minutes of video)*

Example 1.3.2

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

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

Given: a chemical substance

Asked for: its classification

Strategy:

A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution:

a. **A)** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
B) Because the composition of the solution is uniform throughout, it is ***a homogeneous mixture***.
b. **A)** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
B) Because its composition is not uniform throughout, orange juice is ***a heterogeneous mixture***.
c. **A)** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence a compact disc is not chemically pure.
B) The regions of different composition indicate that a compact disc is ***a heterogeneous mixture.***
d. **A)** Aluminum oxide is a single, chemically ***pure compound***.
e. **A)** Selenium is one of the known ***elements***.

Exercise 1.3.2

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: available at the end of this section.

Example 1.3.3

How would a chemist categorize each example of matter?

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

Solution

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

Exercise 1.3.3

How would a chemist categorize each example of matter?

a. coffee
b. hydrogen
c. an egg

Answer a: available at the end of this section.

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
- 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.

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Answers

Exercise 1.3.1
solid

Exercise 1.3.1
gas

Exercise 1.3.1
solid

Exercise 1.3.2
homogeneous mixture (solution)
element
heterogeneous mixture
compound

Exercise 1.3.3
a homogeneous mixture (solution), assume it's filtered coffee
element
heterogeneous mixture.

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1.4: Physical and Chemical Changes

Learning Outcomes

- Distinguish between physical and chemical changes.
- Give examples of physical and chemical changes.

Physical Changes

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. **Melting** is an example of a **physical change**. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.



Figure 10.4.1: *Melting ice in the Beaufort Sea.*

When a piece of wood is ground into sawdust, that change is irreversible since the sawdust could not be reconstituted into the same piece of wood that it was before. Cutting the grass or pulverizing a rock would be other irreversible physical changes. Firewood also represents an irreversible physical change since the pieces cannot be put back together to form the tree.



Figure 10.4.2: *Firewood being cut is a physical change because the composition doesn't change when being cut.*

Chemical Changes

When exposed to air, an object made of iron will eventually begin to rust (see figure below).



Figure 10.4.3: *Rust (iron oxide) forms on an unprotected iron surface.*

As the rust forms on the surface of the iron, it flakes off to expose more iron, which will continue to rust. Rust is clearly a substance that is different from iron. Rusting is an example of a chemical change.

A **chemical property** *describes the ability of a substance to undergo a specific chemical change*. A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust. A more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are *burn, rot, explode,* and *ferment*. Chemical properties are very useful as a way of identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

A chemical change is also called a chemical reaction. A **chemical reaction** *is a process that occurs when one or more substances are changed into one or more new substances*. Zinc (Zn) is a silver-gray element that can be ground into a powder. If zinc is mixed at room temperature with powdered sulfur (S), a bright yellow element, the result will simply be a mixture of zinc and sulfur. No chemical reaction occurs. However, if energy is provided to the mixture in the form of heat, the zinc will chemically react with the sulfur to form the compound zinc sulfide (ZnS). Pictured below are the substances involved in this reaction.

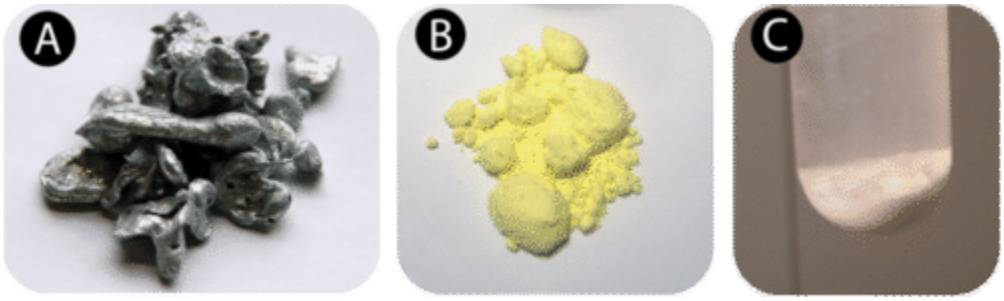


Figure 10.4.4: *Zinc (A) and sulfur (B) are two elements that undergo a chemical reaction when heated to form the the compound zinc sulfide (C).*

The reaction between zinc and sulfur can be depicted in something called a chemical equation. In words, we could write the

reaction as:

$$\text{zinc} + \text{sulfur} \rightarrow \text{zinc sulfide} \tag{1.4.1}$$

A more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved:

$$\text{Zn} + \text{S} \rightarrow \text{ZnS} \tag{1.4.2}$$

The substance(s) to the left of the arrow in a chemical equation are called reactants. A **reactant** *is a substance that is present at the start of a chemical reaction.* The substance(s) to the right of the arrow are called products. A **product** *is a substance that is present at the end of a chemical reaction*. In the equation above, zinc and sulfur are the reactants that chemically combine to form zinc sulfide as a product.

Example 1.4.1

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 contains a yellow solid.

Solution

1. Physical: When water boils, no bonds are broken or formed. The liquid water changes into gaseous water. Both forms are H_2O.
2. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes tend to indicate chemical changes.
3. Physical: None of the properties changed. The green mixture is still green and the colorless solution is still colorless. They have just been mixed together. No color *change* occurred or other evidence of chemical change.
4. Chemical: The formation of a solid when two solutions are mixed and the color change from colorless to yellow indicate a chemical change.

Recognizing Chemical Reactions

How can you tell if a chemical reaction is taking place? Certain visual clues indicate that a chemical reaction is likely (but not necessarily) occurring, including the following examples:

1. A change of color occurs during the reaction.
2. A gas is produced during the reaction.
3. A solid product, called a precipitate, is produced in the reaction.
4. A visible transfer of energy occurs in the form of light as a result of the reaction.

When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced (see figure below). The production of a gas is also an indication that a chemical reaction may be occurring.



Figure 10.4.5: *Zinc reacts with hydrochloric acid to produce bubbles of hydrogen gas.*

When a colorless solution of lead (II) nitrate is added to a colorless solution of potassium iodide, a yellow solid called a precipitate is instantly produced (see figure below). A **precipitate** *is a solid product that forms from a reaction and settles out of a liquid mixture*. The formation of a precipitate may also indicate the occurrence of a chemical reaction.

$$\mathrm{Pb(NO_3)_2}\,(aq) + 2\mathrm{KI}\,(aq) \rightarrow \mathrm{PbI_2}\,(s) + 2\mathrm{KNO_3}\,(aq) \tag{1.4.3}$$



Figure 10.4.6: *A yellow precipitate of solid lead (II) iodide forms immediately when solutions of lead (II) nitrate and potassium iodide are mixed.*

Exercise 1.4.1

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

1. A mirror breaks.
2. An iron nail rusts in moist air
3. Copper metal melts at high temperature.
4. In a car, a catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer 1 available at the end of this section.

Summary

- In a physical change, the identity of the substance does not change, e.g. freezing water to produce solid water (ice).
- In a chemical change the chemical structure of the substance changes and the product(s) are represented by different chemical formulas, e.g. the electrolysis of water which splits it into hydrogen and oxygen.

Contributors and Attributions

-

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Answers

Exercise 1.4.1

Physical change because the pieces are still mirror, just smaller.

Chemical change because the iron and oxygen (from the air) are forming a new substance with a different chemical composition: rust

Physical change because the substance is still copper, just in a liquid form instead of solid. All phase changes (e.g. melting, boiling, etc.) are physical changes.

Chemical change because the nitrogen dioxide no longer exists. It has become separate nitrogen and oxygen.

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1.5: Atomic Theory

Learning Objectives

- State the modern atomic theory.
- Learn how atoms are constructed.

The smallest unit of an element that maintains the identity of that element is called an **atom**. Individual atoms are extremely small. It would take about fifty million atoms in a row to make a line that is 1 cm long. The period at the end of a printed sentence has several million atoms in it. Atoms are so small that it is difficult to believe that all matter is made from atoms—but it is.

The concept that atoms play a fundamental role in chemistry is formalized by the **modern atomic theory**, first stated by John Dalton, an English scientist, in 1808. It consists of three parts:

1. All matter is composed of atoms.
2. Atoms of the same element are the same; atoms of different elements are different.
3. Atoms combine in whole-number ratios to form compounds.

These concepts form the basis of chemistry. Although the word *atom* comes from a Greek word that means "indivisible," we understand now that atoms themselves are composed of smaller parts called *subatomic particles*. The first part to be discovered was the **electron**, a tiny subatomic particle with a negative charge. It is often represented as e^-, with the right superscript showing the negative charge. Later, two larger particles were discovered. The **proton**, a subatomic particle with a positive charge, is a more massive (but still tiny) subatomic particle with a positive charge, represented as p^+. The neutron is a subatomic particle with about the same mass as a proton but no charge. It is represented as either n or n^0. We now know that all atoms of all elements are composed of electrons, protons, and (with one exception) neutrons. Table 1.5.1 summarizes the properties of these three subatomic particles.

Table 1.5.1***:*** *Properties of the Three Subatomic Particles*

Name	Symbol	Mass (approx.; kg)	Charge
Proton	p^+	1.6×10^{-27}	1+
Neutron	n, n^0	1.6×10^{-27}	none
Electron	e^-	9.1×10^{-31}	1–

How are these particles arranged in atoms? They are not arranged at random. Experiments by Ernest Rutherford in England in the 1910s pointed to a nuclear model with atoms that has the protons and neutrons in a central nucleus with the electrons in orbit about the nucleus. The relatively massive protons and neutrons are collected in the center of an atom, in a region called the nucleus of the atom (plural *nuclei*). The electrons are outside the nucleus and spend their time orbiting in space about the nucleus. (Figure 1.5.1).

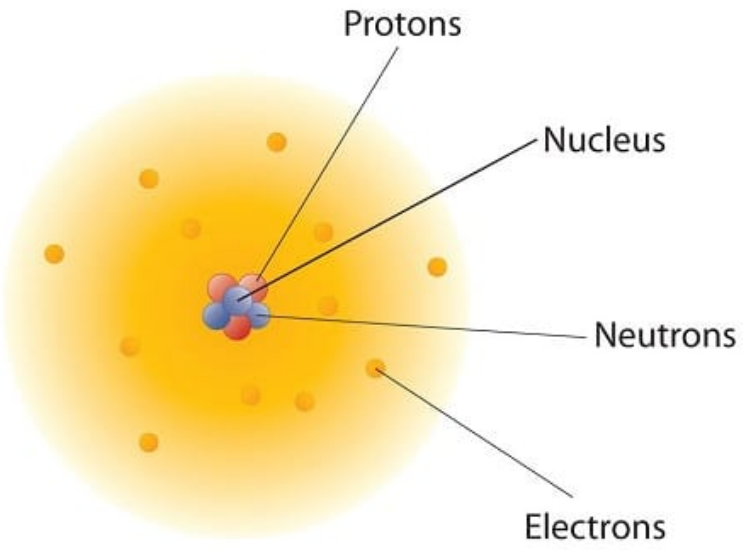


Figure 1.5.1***:*** *The Structure of the Atom.Atoms have protons and neutrons in the center, making the nucleus, while the electrons orbit the nucleus.*

The modern atomic theory states that atoms of one element are the same, while atoms of different elements are different. What makes atoms of different elements different? **The fundamental characteristic that all atoms of the same element share is the *number of protons***. All atoms of hydrogen have one and only one proton in the nucleus; all atoms of iron have 26 protons in the nucleus. This number of protons is so important to the identity of an atom that it is called the **atomic number**, the number of protons in an atom of the element. Thus, hydrogen has an atomic number of 1, while iron has an atomic number of 26. Each element has its own characteristic atomic number.

Atoms of the same element can have different numbers of neutrons, however. Atoms of the same element (i.e., atoms with the same number of protons) with different numbers of neutrons are called **isotopes**. Most naturally occurring elements exist as isotopes. For example, most hydrogen atoms have a single proton in their nucleus. However, a small number (about one in a million) of hydrogen atoms have a proton and a neutron in their nuclei. This particular isotope of hydrogen is called deuterium. A very rare form of hydrogen has one proton and two neutrons in the nucleus; this isotope of hydrogen is called tritium. The sum of the number of protons and neutrons in the nucleus is called the **mass number** of the isotope.

Neutral atoms have the same number of electrons as they have protons, so their overall charge is zero. However, as we shall see later, this will not always be the case.

Example 1.5.1:

a. The most common carbon atoms have six protons and six neutrons in their nuclei. What are the atomic number and the mass number of these carbon atoms?
b. An isotope of uranium has an atomic number of 92 and a mass number of 235. What are the number of protons and neutrons in the nucleus of this atom?

Solution

a. If a carbon atom has six protons in its nucleus, its atomic number is 6. If it also has six neutrons in the nucleus, then the mass number is 6 + 6, or 12.
b. If the atomic number of uranium is 92, then that is the number of protons in the nucleus. Because the mass number is 235, then the number of neutrons in the nucleus is 235 – 92, or 143.

Exercise 1.5.1

The number of protons in the nucleus of a tin atom is 50, while the number of neutrons in the nucleus is 68. What are the atomic number and the mass number of this isotope?

Answer

Atomic number = 50, mass number = 118

Key Takeaways

- Chemistry is based on the modern atomic theory, which states that all matter is composed of atoms.
- Atoms themselves are composed of protons, neutrons, and electrons.
- Each element has its own atomic number, which is equal to the number of protons in its nucleus.

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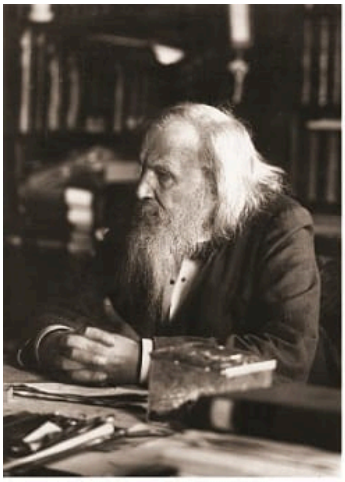
1.6: The Periodic Table

Skills to Develop

- State the periodic law and explain the organization of elements in the periodic table
- Predict the general properties of elements based on their location within the periodic table
- Identify metals, nonmetals, and metalloids by their properties and/or location on the periodic table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K): These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive than are Ca, Sr, and Ba; Li, Na, and K form compounds with oxygen in a ratio of two of their atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one of their atoms to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of any of the elements above.

Dimitri Mendeleev in Russia (1869) and Lothar Meyer in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time. Both published tables with the elements arranged according to increasing atomic mass. But Mendeleev went one step further than Meyer: He used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were yet unknown. The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work. Although Mendeleev and Meyer had a long dispute over priority, Mendeleev's contributions to the development of the periodic table are now more widely recognized (Figure 1.6.1).



(a)

Reihen	Gruppe I. — R^1O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

(b)

Figure 1.6.1*: (a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by "Den fjättrade ankan"/Wikimedia Commons)*

By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses. The modern statement of this relationship, the periodic law, is as follows: *the properties of the elements are periodic functions of their atomic numbers*. A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column (Figure 1.6.2). Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. The elements are arranged in seven horizontal rows, called periods or series, and 18 vertical columns, called groups. Groups are labeled at the top of each column. In the United States, the labels traditionally were numerals with capital letters. However, IUPAC recommends that the numbers 1 through 18 be used, and these labels are more common. For the table to fit on a single page, parts of two of the rows, a total of 14 columns, are usually written

below the main body of the table.

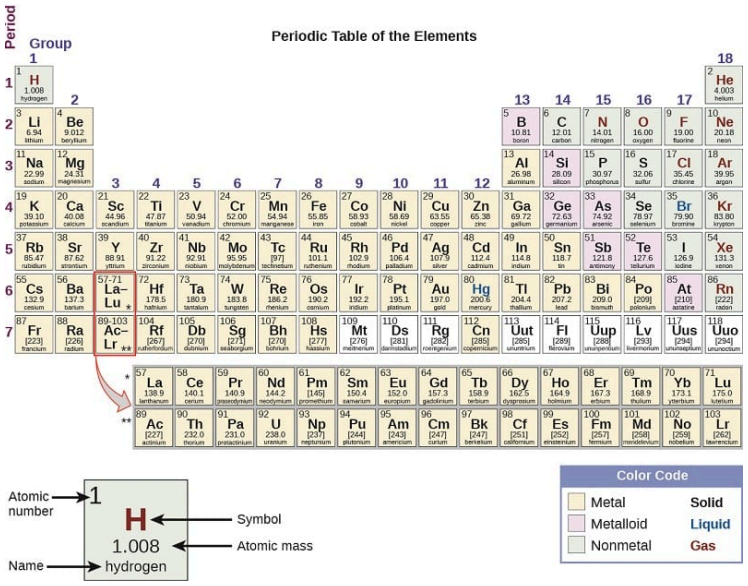


Figure 1.6.2*: Elements in the periodic table are organized according to their properties.*

Many elements differ dramatically in their chemical and physical properties, but some elements are similar in their behaviors. For example, many elements appear shiny, are malleable (able to be deformed without breaking) and ductile (can be drawn into wires), and conduct heat and electricity well. Other elements are not shiny, malleable, or ductile, and are poor conductors of heat and electricity. We can sort the elements into large classes with common properties: metals (elements that are shiny, malleable, good conductors of heat and electricity—shaded yellow); nonmetals (elements that appear dull, poor conductors of heat and electricity—shaded green); and metalloids (elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals—shaded purple).

The elements can also be classified into the main group elements (or representative elements) in the columns labeled 1, 2, and 13–18; the transition metals in the columns labeled 3–12; and inner transition metals in the two rows at the bottom of the table (the top-row elements are called lanthanides and the bottom-row elements are actinides; Figure 1.6.3). The elements can be subdivided further by more specific properties, such as the composition of the compounds they form. For example, the elements in group 1 (the first column) form compounds that consist of one atom of the element and one atom of hydrogen. These elements (except hydrogen) are known as alkali metals, and they all have similar chemical properties. The elements in group 2 (the second column) form compounds consisting of one atom of the element and two atoms of hydrogen: These are called alkaline earth metals, with similar properties among members of that group. Other groups with specific names are the pnictogens (group 15), chalcogens (group 16), halogens (group 17), and the noble gases (group 18, also known as inert gases). The groups can also be referred to by the first element of the group: For example, the chalcogens can be called the oxygen group or oxygen family. Hydrogen is a unique, nonmetallic element with properties similar to both group 1 and group 17 elements. For that reason, hydrogen may be shown at the top of both groups, or by itself.

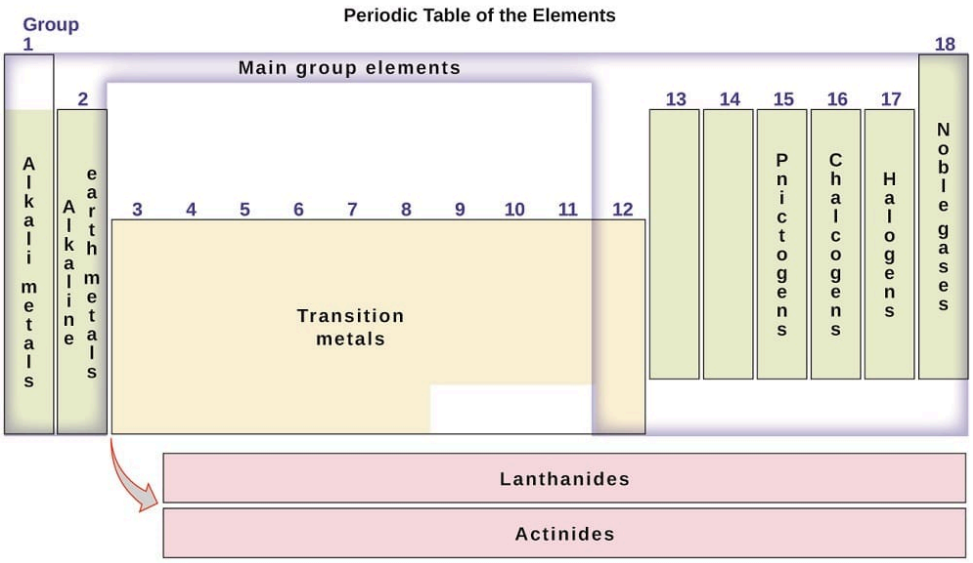


Figure 1.6.3***:*** *The periodic table organizes elements with similar properties into groups.*

Example 1.6.1: Naming Groups of Elements

Atoms of each of the following elements are essential for life. Give the group name for the following elements:

a. chlorine
b. calcium
c. sodium

Solution

The family names are as follows:

a. halogen
b. alkaline earth metal
c. alkali metal

Exercise 1.6.1

Give the group name for each of the following elements:

a. krypton
b. barium
c. lithium

Answer a

noble gas

Answer b

alkaline earth metal

Answer c

alkali metal

As described above, elements in the same column of the periodic table tend to have similar properties. This is because they have the same number of valence electrons. Remember that protons and neutrons are found in the nucleus of an atom, with electrons in the space around the nucleus. However, the electrons are not arranged randomly! They are organized into layers called shells or levels. In this class we will be most interested in the outermost shell, called the valence shell, because those electrons can be gained, lost, or shared with other atoms. Studying valence electrons gives us a lot of insight into how an atom (and therefore an element) will behave.

The arrangement of the periodic table tells us how many valence electrons an atom of an element has. Group 1 elements, hydrogen and the alkali metals, each have one valence electron. Group 2 elements, the alkaline earth metals, each have two valence electrons. We won't concern ourselves with the number of valence electrons for the transition metals or inner transition metals, but the pattern continues for the remaining main group elements. Elements in groups 13-18 have three to eight valence electrons, respectively.

In studying the periodic table, you might have noticed something about the atomic masses of some of the elements. Element 43 (technetium), element 61 (promethium), and most of the elements with atomic number 84 (polonium) and higher have their atomic mass given in square brackets. This is done for elements that consist entirely of unstable, radioactive isotopes (you will learn more about radioactivity in the nuclear chemistry chapter). An average atomic weight cannot be determined for these elements because their radioisotopes may vary significantly in relative abundance, depending on the source, or may not even exist in nature. The number in square brackets is the atomic mass number (and approximate atomic mass) of the most stable isotope of that element.

Summary

The discovery of the periodic recurrence of similar properties among the elements led to the formulation of the periodic table, in which the elements are arranged in order of increasing atomic number in rows known as periods and columns known as groups. Elements in the same group of the periodic table have similar chemical properties. Elements can be classified as metals, metalloids, and nonmetals, or as a main-group elements, transition metals, and inner transition metals. Groups are numbered 1–18 from left to right. The elements in group 1 are known as the alkali metals; those in group 2 are the alkaline earth metals; those in 17 are the halogens; and those in 18 are the noble gases.

Glossary

actinide
: inner transition metal in the bottom of the bottom two rows of the periodic table

alkali metal
: element in group 1

alkaline earth metal
: element in group 2

group
: vertical column of the periodic table

halogen
: element in group 17

inert gas
: (also, noble gas) element in group 18

inner transition metal
: (also, lanthanide or actinide) element in the bottom two rows; if in the first row, also called lanthanide, or if in the second row, also called actinide

lanthanide
inner transition metal in the top of the bottom two rows of the periodic table

main-group element
(also, representative element) element in columns 1, 2, and 12–18

metal
element that is shiny, malleable, good conductor of heat and electricity

metalloid
element that conducts heat and electricity moderately well, and possesses some properties of metals and some properties of nonmetals

noble gas
(also, inert gas) element in group 18

nonmetal
element that appears dull, poor conductor of heat and electricity

period
(also, series) horizontal row of the periodic table

periodic law
properties of the elements are periodic function of their atomic numbers.

periodic table
table of the elements that places elements with similar chemical properties close together**representative element**

(also, main-group element) element in columns 1, 2, and 12–18

transition metal
element in columns 3–11

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1.7: The Importance of Ions

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Learning Objectives

- To distinguish between cations and anions
- To identify the charges of monatomic ions when given the name

Chemists are frequently concerned with the movement of electrons. When an atom gains or loses electrons, it becomes a charged particle called an ion. When this occurs, the number of protons and electrons in the nucleus does not change. For atoms that lose electrons, an overall positive charge will result (**#protons > #electrons**) and the particle is called a cation. Metal atoms (located on the left side of the periodic table) lose electrons to become cations.

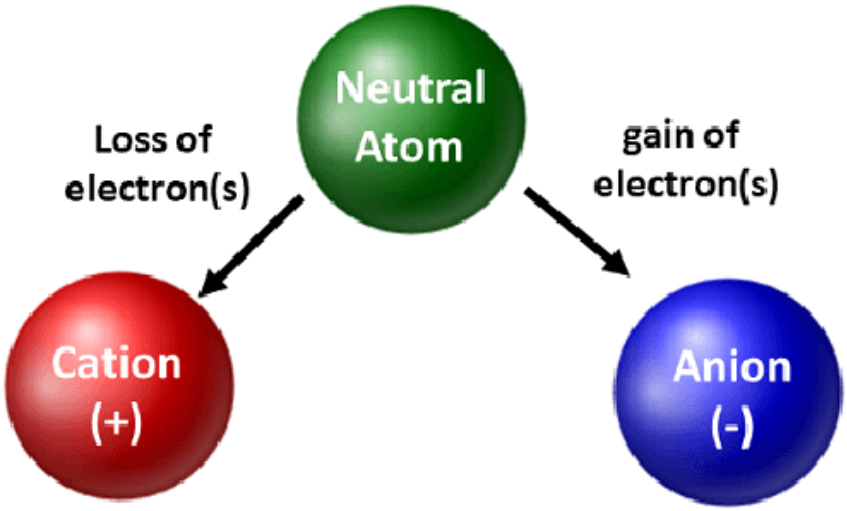


Figure 1.7.1: Neutral atoms become ions by either losing an electron (to form positively charges cations) or gaining electrons (to form negatively charged anions). (Public Domain; Delmar Larsen)

Unlike metal atoms, nonmetals can gain electrons to become anions. Anions have an overall negative charge (**#electrons > #protons**). With the exception of the noble gases, all atoms on the periodic table can lose or gain electrons to achieve electronic stability. Different types of bonding occur when atoms lose, gain, or share electrons. These types of atomic connections will be further discussed in the next section and again later in the course.

Interactive: Building an Atom

Build an atom out of protons, neutrons, and electrons, and see how the element, charge, and mass change. Then play a game to test your ideas!

At this point, you should be able to calculate the number of protons and electrons in a particle when given its identity and charge. In addition, you should be able to classify ions as cations or anions. For example, how many protons and electrons are in a magnesium ion with a charge of +2? Is it a cation or an anion?

Magnesium has the symbol Mg on the periodic table and its atomic number is 12. That means that it has 12 protons. Magnesium, a metal, can exist as either an atom with an overall charge of zero or as a cation. When it has a charge of +2 that means that there are two fewer electrons than protons. A magnesium ion has 12 protons (which are positive) and 10 electrons (which are negative)

resulting in an overall charge of +2.

The ions of main group elements tend to have charges that follow patterns as shown in the figure below. This is based on the number of valence electrons an atom has. Remember that atoms of hydrogen and the alkali metals each have one valence electron. That valence electron can be lost resulting in a cation with a charge of +1. Alkaline earth metals each have two valence electrons and can form cations with a +2 charge. Nonmetals in groups 15-17 have five to seven valence electrons and can gain the 1-3 valence electrons that they need to achieve eight valence electrons, a particularly stable configuration. Remember that because electrons are negative, an atom that gains extra electrons becomes negative overall.

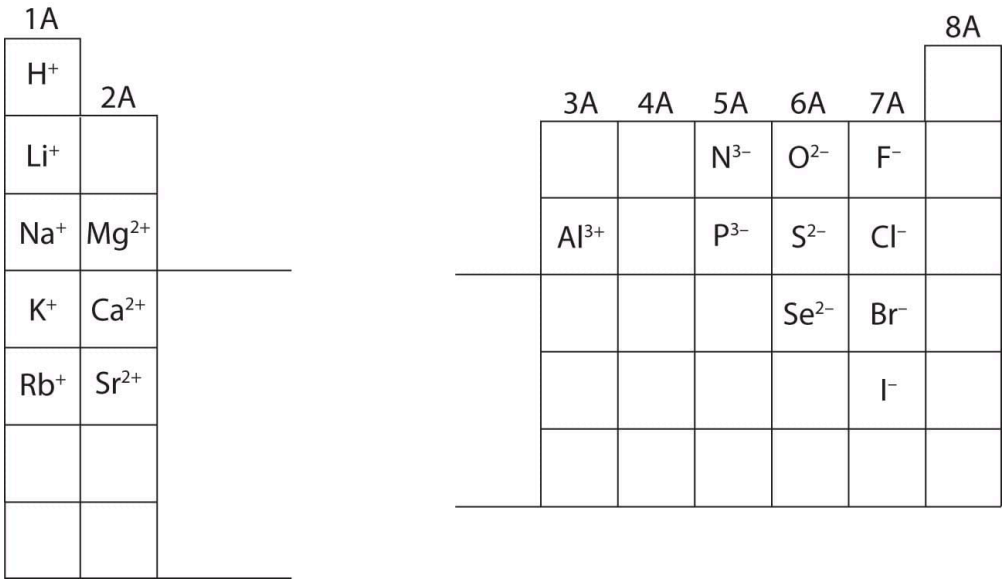


Figure 1.7.2: Periodic Table with ion charges of common main group elements. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

Most transition metals, as well as the metals at the bottom of groups 13-15, can form more than one ion. For example, iron can be found as ions with a +2 charge or a +3 charge. These are called iron(II) ions and iron(III) ions, respectively. Other elements have different possible charges such as nickel(I) and nickel(II) or lead(II) and lead(IV). You don't need to memorize all of these possibilities! Instead, notice that the name of the ion gives the charge as a Roman numeral.

The noble gases in group 8 already have eight valence electrons, so they do not readily form ions. Note: Helium is an exception. It has only two electrons, both of which are valence electrons. However, like the rest of the noble gases, helium is stable does not form ions.

Example 1.7.1: Ion Classification

Calculate the number of protons and neutrons for each particle below using the information in Figure 1.7.2. Label each as being an atom, cation, or anion. Refer to the periodic table for masses, atomic numbers, and specific ion charges.

a. Aluminum ion
b. Zirconium atom
c. Sulfur ion

Solutions

a. There are 13 protons and 10 electrons in the aluminum ion. This ion is positively charged which means it has lost electrons and forms a cation.
b. There are 40 protons and 40 electrons in the zirconium atom. This is an atom which has no overall charge.
c. There are 16 protons and 18 electrons in the sulfur ion. This ion is negatively charged which means it has gained electrons and forms and anion.

Summary

- Cations are positive ions that result when a metal loses one or more electrons. Anions are negative ions that form when a nonmetal gains one or more electrons.
- Some elements have ions whose charges can be predicted based on their location on the periodic table. For example, Group 1 elements form ions with a +1 charge and Group 17 elements form ions with a -1 charge.
- Many metals can form ions with different charges such as iron(II), Fe^{2+}, and iron(III), Fe^{3+}. The Roman numeral in the name of the ion indicates the positive charge of the cation.

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1.8: Ionic and Covalent Bonds

Learning Objectives

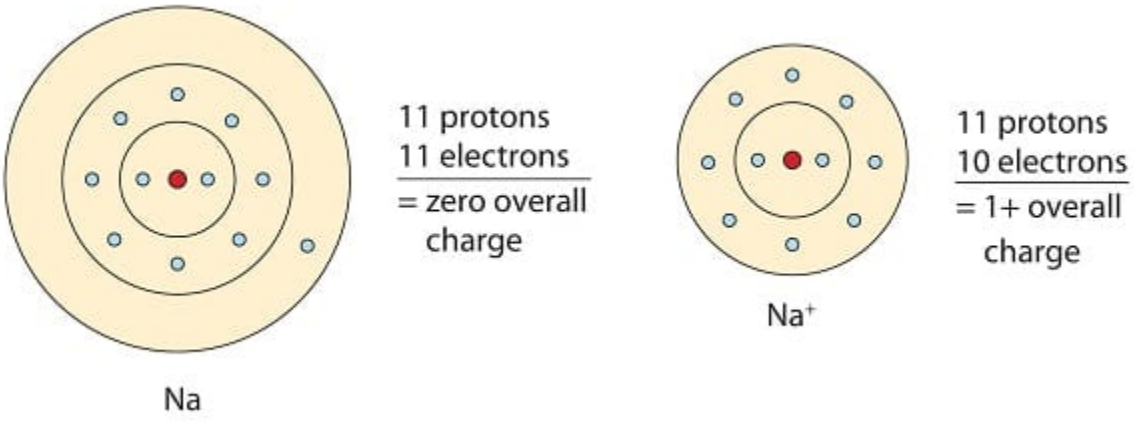
- To distinguish between ionic and covalent bonds
- To describe how the octet rule can be used to predict an ion's charge

There are many types of chemical bonds and forces that hold atoms or ions together in compounds. The two most basic types of bonds are characterized as either ionic or covalent. In ionic bonding, electrons are transferred from one atom to another. Ionic bonds require at least atom that easily loses electrons to form a cation (usually a metal) and at least one atom that easily gains electrons to form an anion (usually a nonmetal). The cation and anion are attracted to one another and form a bond. In contrast, atoms that are more similar to one another (e.g. two nonmetals) can share electrons in covalent bonds which will be covered later in the course.

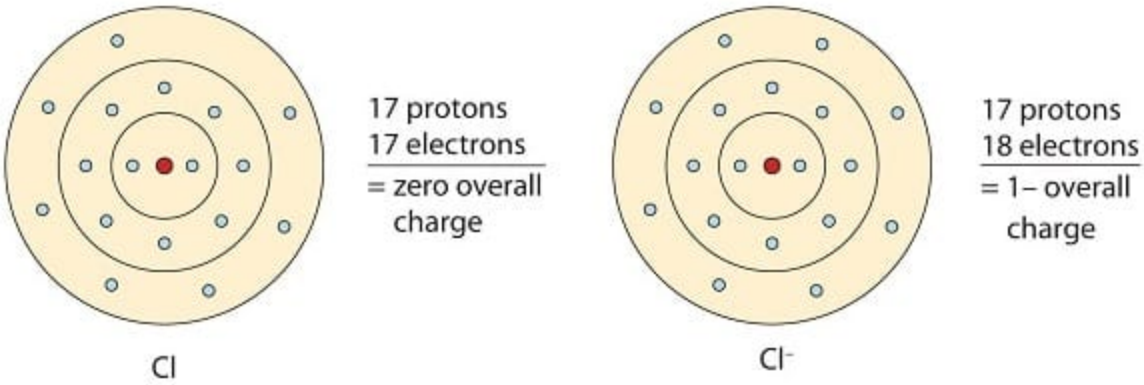
Introduction

Ionic bonding is the complete transfer of valence electron(s) between atoms. In ionic bonds, a metal can lose electrons to become a positively charged cation, whereas a nonmetal accepts those electrons to become a negatively charged anion. Ionic bonds require an electron donor, often a metal, and an electron acceptor, a nonmetal.

Ionic bonding is observed because metals have few electrons in their outer-most shell. You can think of an atom like an onion with many layers. If the outermost layer of the onion is torn and incomplete, that is similar to an incomplete octet. If you remove that outermost, incomplete layer of the onion, there is a complete layer underneath that is now the new outermost layer. By losing their valence electrons, metals can satisfy the octet rule. This sodium atom donates the lone electron in its valence orbital in order to achieve octet configuration. This creates a positively charged cation due to the loss of electron.



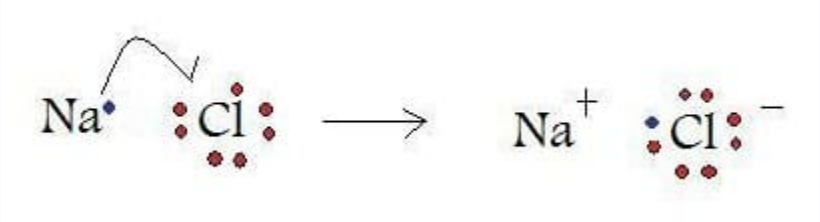
On the other hand, nonmetals generally have close to 8 electrons in their valence shells and can more easily accept electrons to achieve an octet. This chlorine atom receives one electron to achieve its octet, resulting in a negatively charged anion.



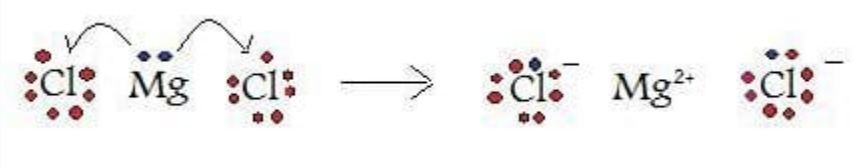
Because sodium ions have a 1+ charge and chloride ions have a 1- charge, the ratio of sodium ions to chloride ions in sodium chloride is 1:1. Its formula is NaCl. In the example below you will also see $MgCl_2$ in which the ratio of cations to anions is 1:2 because magnesium ions have a 2+ charge whereas chloride ions only have a 1- charge.

In ionic bonding, more than one electron can be donated or received to satisfy the octet rule. The charges on the anion and cation correspond to the number of electrons donated or received. In ionic bonds, the net charge of the compound must be zero. This is important for understanding what happens in nature and for predicting the formulas of ionic compounds! The ratio of cations and anions in an ionic compound always results in an overall charge of zero.

Example 1.8.1: Chloride Salts



In this example, the sodium atom is donating its 1 valence electron to the chlorine atom. This creates a sodium cation and a chlorine anion. Notice that the net charge of the resulting compound is 0.



In this example, the magnesium atom is donating both of its valence electrons to chlorine atoms. Each chlorine atom can only accept 1 electron before it can achieve its noble gas configuration; therefore, 2 atoms of chlorine are required to accept the 2 electrons donated by the magnesium. Notice that the net charge of the compound is 0.

Exercise 1.8.1

Do these compounds contain ionic or covalent bonds?

1. CH_4
2. Fe_2O_3
3. KNO_3

4. H_2O
5. $BeCl_2$

Answer 1 available at the end of this section.

Summary

1. Metals and nonmetals tend to form ionic bonds whereas nonmetals tend to bond with one another via covalent bonding.
2. The charge of an ion is based on the number of valence electrons it gains or loses to acheive an octet.

References

1. Petrucci, Ralph H. General Chemistry: Principles and Modern Applications. Upper Saddle River, NJ: Pearson Education, 2007.
2. Brown, Theodore L., Eugene H. Lemay, and Bruce E. Bursten. Chemistry: The Central Science. 6th ed. Englewood Cliffs, NJ: Prentice Hall, 1994.

Answers

Exercise 1.8.1

Covalent because carbon and hydrogen are both nonmetals

Ionic because iron is a metal and oxygen is a nonmetal

Ionic

Covalent

Ionic

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1.9: Formulas for Ionic Compounds

Learning Objectives

- Write the chemical formula for a simple ionic compound.
- Recognize polyatomic ions in chemical formulas.

We have already encountered some chemical formulas for simple ionic compounds. A chemical formula is a list of the elements in a compound and the ratio of the number of particles of each element. 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 patterns called crystals (Figure 1.9.1). There are no individual NaCl “particles” in the crystal; instead, there is a continuous pattern of alternating sodium and chloride ions. In order to identify this compound using a chemical formula we use the ratio of sodium ions to chloride ions, expressed in the lowest possible whole numbers. 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. The term molecule is not applicable to an ionic compound; instead the lowest ratio is called a formula unit. Although it is convenient to think that NaCl crystals are composed of individual NaCl units, Figure 1.9.1 shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by several ions of opposite charge.

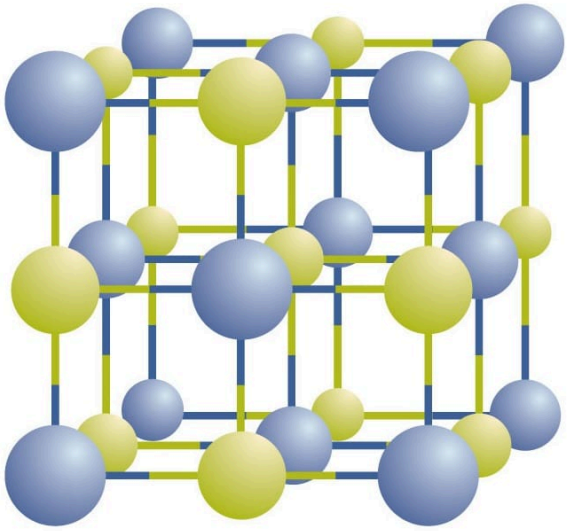


Figure 1.9.1 A Sodium Chloride Crystal. A crystal contains a three-dimensional pattern of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions.

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 component species are ions, not neutral atoms, even though the formula does not contain charges. Finally, the proper formula for an ionic compound always obeys the following rule: 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.

By convention, assume that there is only one atom if a subscript is not present. We do not use 1 as a subscript.

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 LiBr.

When an ionic compound is formed from magnesium and oxygen, the magnesium ion has a 2+ charge, and the oxygen atom 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 ionic compound is MgO.

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

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

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. Instead of writing the chemical formula for this ionic compound as MgClCl, we use a numerical subscript when there is more than one ion of a given type: $MgCl_2$. This chemical formula says that there are one magnesium ion and *two* chloride ions in this formula. 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+}

Example 1.9.1

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

1. The sodium ion and the sulfur ion
2. The aluminum ion and the fluoride ion
3. The 3+ iron ion and the oxygen ion

Solution

1. To obtain a valence shell octet, sodium forms an ion with a 1+ charge, while the sulfur ion has a 2– charge. Two sodium 1+ ions are needed to balance the 2– charge on the sulfur ion. **The formula is Na_2S** which indicates two sodium ions and one sulfide ion.
2. The aluminum ion has a 3+ charge, while the fluoride ion formed by fluorine has a 1– charge. Three fluorine 1– ions are needed to balance the 3+ charge on the aluminum ion. **This combination is written as AlF_3.**
3. Iron can form two possible ions, but the ion with a 3+ charge is specified here. The oxygen atom 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 2– oxygen ions will give 6–, thereby balancing the overall positive and negative charges. Thus, **the formula for this ionic compound is Fe_2O_3.**

Exercise 1.9.1

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

1. The calcium ion and the oxygen ion
2. The 2+ copper ion and the sulfur ion
3. The 1+ copper ion and the sulfur ion

Answer available at the end of this section.

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1– charge. Table 1.9.1 lists the most common polyatomic ions.

Table 1.9.1: Some Common Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$
carbonate ion	CO_3^{2-}
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
nitrate ion	NO_3^-
sulfite ion	SO_3^{2-}
sulfate ion	SO_4^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Example 1.9.2

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

1. The potassium ion and the sulfate ion
2. The calcium ion and the nitrate ion

Solution

1. 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**.
2. 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 1.9.2

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 available at the end of this section.

Interpreting Names of Ionic Compounds

Each ionic compound has a name consisting of the name of the cation followed by the name of the anion.

Cation Names

Many cation names are the same as the name of the corresponding element on the periodic table. For example, sodium can form sodium ions and magnesium can form magnesium ions.

However, most transition metals (Groups 3 - 12) and post-transition metals (at the bottom of Groups 13 - 16) can form ions with different charges in different environments, so we need a way of distinguishing these possibilities. For example, iron can form cations with a +2 charge and with a +3 charge. It would be confusing if we called both of those iron ions.

Instead, when multiple different charges are possible we use Roman numerals as part of the name of the ion to indicate which one we are referring to. The two types of iron ions are called iron(II) ions and iron(III) ions. The Roman numeral tells the charge of the ion. The most common Roman numerals used to represent charges are: 1 = I, 2 = II, 3 = III, and 4 = IV. The charges represented by Roman numerals are always positive (they are only used for cations).

It is possible to have polyatomic cations. A common one is the ammonium ion, NH_4^+. Unlike most cations, ammonium is not made from a metal.

Anion Names

Unlike cations, element names are modified when a substance becomes an anion. For monatomic anions (monatomic means "one atom"), the ending of the element name is removed and replaced by -ide. For example, toothpaste contains fluoride ions, not fluorine atoms.

Polyatomic anions are much more common than polyatomic cations. Luckily you can generally recognize them by the last three letters of their names: -ate and -ite. Unfortunately there are some exceptions like hydroxide ions and cyanide ions which are polyatomic despite having an -ide ending.

Example 1.9.3

Write the formula for each compound.

1. Sodium oxide
2. Ammonium chloride
3. Iron(III) sulfide

Solution

1. Sodium ions have a charge of +1 ; oxide ions have a charge of -2. Therefore, twice as many sodium ions as oxide ions are needed and **the formula is Na_2O**.

2. Ammonium is a polyatomic cation with the formula NH_4^+ where the charge is +1 (represented by the + sign written as a superscript). The subscript 4 is not part of the charge, it tells the number of hydrogen atoms in each ammonium ion. Chloride ions have a -1 charge. Since these ions have charges of +1 and -1 they can combine at a 1:1 ratio and t**he formula is NH_4Cl**.
3. The Roman numeral tells us that in this case the iron has a charge of +3. The sulfide ion has a charge of -2. Finding the ratio on this one is tricky! **The formula is Fe_2S_3**. Two iron(III) ions add up to a total positive charge of +6 and three sulfide ions add up to a total negative charge of -6. This makes a charge of zero overall.

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:

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 (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 hydrogen phosphate ions—HPO_4^{2-} and $H_2PO_4^-$—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 (SO_4^{2-}), but this ion is present in seawater.

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.

Contributors and Attributions

- Anonymous

Answers

Exercise 1.9.1

1. CaO

2. CuS
3. Cu_2S

Exercise 1.9.2

1. $MgCO_3$
2. $Al(C_2H_3O_2)_3$

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2: NUCLEAR CHEMISTRY

Nuclear reactions are quite different from the chemical reactions that we will focus on throughout most of the book. The main difference is that an atom of one element can turn into an atom of another element! Nuclear chemistry has many applications in medicine. Benefits include use in diagnostics and treatment, but radiation can also cause illness and cancer. In this chapter you will read about radiation, the decay of radioisotopes, and nuclear equations.

2.1: Isotopes

2.2: Radiation

2.3: Nuclear Reactions and Nuclear Equations

2.4: Half-Life

2.5: Radiation and the Body

2.6: Uses of Radioactive Isotopes in Medicine

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2.1: Isotopes

Learning Objectives

- Understand how isotopes differ in particles and mass.
- Write symbols for isotopes.
- Identify the most abundant isotope when given specific values.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Isotopes are atoms of the same element that contain different numbers of neutrons. For these species, the number of electrons and protons remain constant. This difference in neutron amount affects the atomic mass (**A**) but not the atomic number (**Z**). Isotopes of an element have the same physical and chemical properties (except for mass). For this reason, it is difficult to distinguish between an element's isotopes. Nuclear scientists can identify and separate different types of atomic nuclei (Figure 2.1.1), but the technology required for this process is more sophisticated that what would be found in a typical chemical laboratory.

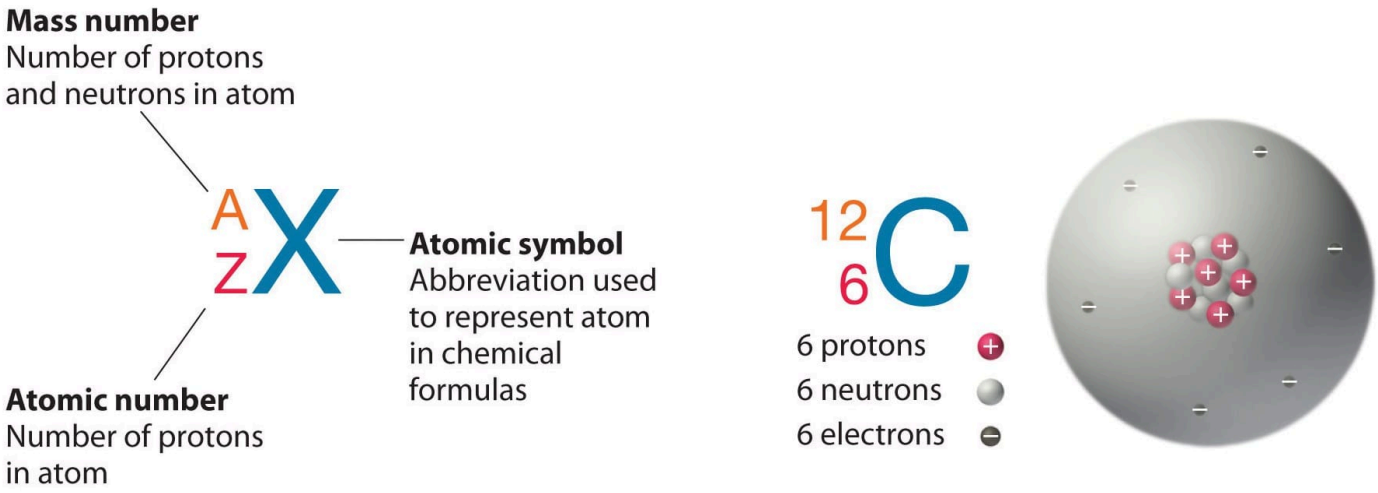


Figure 2.1.1: Unlike protons, the number of neutrons is not absolutely fixed for most elements. Atoms that have the same number of protons, and hence the same atomic number, but different numbers of neutrons are called isotopes. All isotopes of an element have the same number of protons and electrons, which means they exhibit the same chemistry. The isotopes of an element differ only in their atomic mass, which is given by the mass number (A), the sum of the numbers of protons and neutrons. (CC BY-NS-SA; anonymous by request)

The element carbon (C) has an atomic number of 6, which means that all neutral carbon atoms contain 6 protons and 6 electrons. In a typical sample of carbon-containing material, 98.89% of the carbon atoms also contain 6 neutrons, so each has a mass number of 12 because the mass number is the sum of protons and neutrons. An isotope of any element can be uniquely represented as ${}^{A}_{Z}X$where X is the atomic symbol of the element. The isotope of carbon that has 6 neutrons is therefore ${}^{12}_{6}\mathrm{C}$. The subscript indicating the atomic number is actually redundant because the atomic symbol already uniquely specifies *Z*. Consequently, it is more often written as ${}^{12}\mathrm{C}$, which is read as “carbon-12.” Nevertheless, the value of Z is commonly included in the notation for *nuclear* reactions because these reactions involve changes in Z.

Most elements on the periodic table have at least two stable isotopes. For example, in addition to ${}^{12}\mathrm{C}$, a typical sample of carbon contains 1.11% ${}^{13}_{6}\mathrm{C}$, with 7 neutrons and 6 protons, and a trace of ${}^{14}_{6}\mathrm{C}$, with 8 neutrons and 6 protons. The nucleus of ${}^{14}_{6}\mathrm{C}$ is not stable, however, but undergoes a slow radioactive decay that is the basis of the carbon-14 dating technique used in archaeology. Many elements other than carbon have more than one stable isotope; tin, for example, has 10 isotopes. There are about twenty elements that exist in only one isotopic form (sodium and fluorine are examples of these). Most scientists cannot tell you how many isotopic forms exist unless they consult an isotopic table.

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 2.1.2 compares the three isotopes of hydrogen.

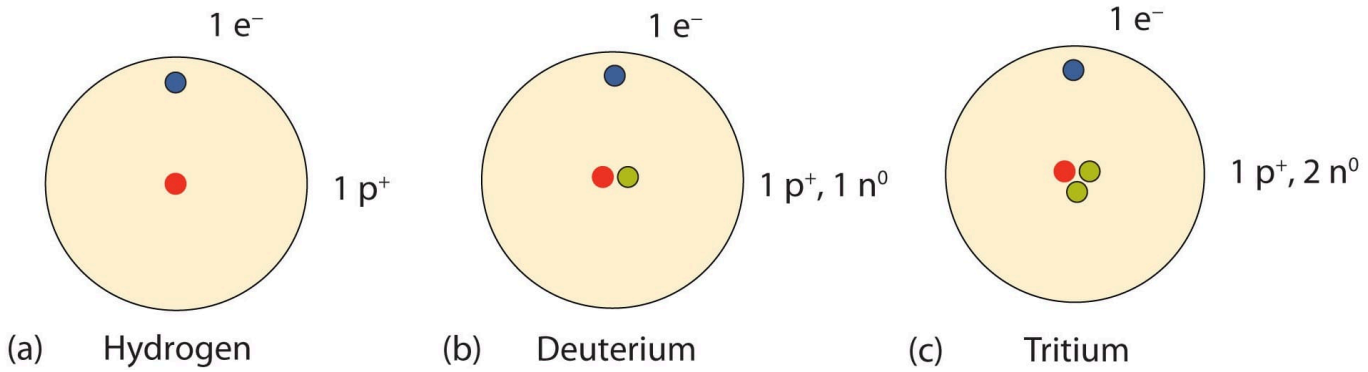


Figure 2.1.2: a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

There are currently over 3,500 isotopes known for all the elements. It would be very inefficient to give different names to each isotope like we do for hydrogen! When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. A/Z and symbol-mass formats can be used to display periodic table information. When viewing either of these two notations, isotopic differences can be obtained.

> 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.

Look at the **A/Z formats** for the three isotopes of hydrogen in Table 2.1.1. Note how the atomic number (bottom value) remains the same while the atomic masses (top number) are varied. All isotopes of a particular element will vary in neutrons and mass. This variance in mass will be visible in the **symbol-mass format** of same isotopes as well.

Table 2.1.1

Common Name	A/Z formats	symbol-mass format	Expanded Name
Hydrogen	${}^{1}_{1}\mathrm{H}$	H-1	hydrogen-1
Deuterium	${}^{2}_{1}\mathrm{H}$	H-2	hydrogen-2
Tritium	${}^{3}_{1}\mathrm{H}$	H-3	hydrogen 3

Using a periodic table, A/Z or symbol-mass formats can be utilized to determine the amount of subatomic particles (protons, neutrons, and electrons) contained inside an isotope. Be sure to use the mass number (A) instead of the atomic mass to calculate the number of neutrons in the nucleus!

Example 2.1.1

How many neutrons are in each atom?

1. ${}^{36}_{17}\mathrm{Cl}$
2. ${}^{58}_{26}\mathrm{Fe}$
3. carbon-14

Solution

1. To find the number of neutrons subtract mass number (top) - number of protons (bottom). 36 particles total – 17 protons = **19 neutrons.**
2. Subtract mass number - atomic number = number of neutrons. 58 particles total - 26 protons = **32 neutrons**.
3. The notation carbon-14 means ${}^{6}_{14}\mathrm{C}$ where the reader is expected to get the atomic number from the periodic table. 14 particles total - 6 protons = **8 neutrons**.

Exercise 2.1.2

For the species below, translate the A/Z format to symbol-mass. Then, calculate the number of subatomic particles for each atom.

a. $^{197}_{79}\mathrm{Au}$
b. $^{23}_{11}\mathrm{Na}$
c. $^{239}_{94}\mathrm{Pu}$

Answer a

Au-197 contains 79 protons, 79 electrons, and 118 neutrons.

Answer b

Na-23 contains 11 protons, 11 electrons, and 12 neutrons.

Answer c

Pu-239 contains 94 protons, 94 electrons, and 145 neutrons.

Calculating Atomic Mass

The **atomic mass** of an element is the weighted mass of all the naturally occurring isotopes. On the periodic table, it is assumed that this mass has units of amu (atomic mass unit) which can be abbreviated by using the letter u. To determine the most abundant isotopic form of an element, compare given isotopes to the weighted average on the periodic table. For example, the three hydrogen isotopes in Table 2.1.1 are hydrogen-1, hydrogen-2, and hydrogen-3. The atomic mass or weighted average of hydrogen is around 1.008 amu (see the periodic table). Of the three hydrogen isotopes, hydrogen-1 is closest in mass to the weighted average; therefore, it is the most abundant. The other two isotopes of hydrogen are quite rare, but are very exciting in the world of nuclear science.

Example 2.1.2

Identify the true statements:

a. Aluminum-27 is more abundant than aluminum-25.
b. An possible isotope of bromine could be bromine-35.
c. A chemist could easily distinguish between cesium-132 and cesium-133 by noting chemical and physical properties.
d. Most scientists know that calcium has 24 isotopes.

Answer a

This statement is true. Aluminum with a mass of 27 is closest to the mass on the periodic table. It would be more abundant than aluminum-25.

Answer b

This statement is false. An appropriate mass number of an isotope of bromine would be in around 80 amu (atomic mass units), not 35 (which is the atomic number). Both the protons and neutrons contribute to the mass of the atom but the atomic number is just the number of protons.

Answer c

This statement is false. Chemists cannot distinguish between isotopes by looking at various properties. Most isotopes have similar solubilities, densities, and colors.

Answer d

This statement is false. Unless a scientist works heavily with a particular atom, he or she is not aware of the many forms that cannot exist. They would need a reference guide to know how many natural and artificial isotopes exist for a particular element.

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

$$\text{Atomic mass} = (\%_1)\,(\text{mass}_1) + (\%_2)\,(\text{mass}_2) + \cdots \tag{2.1.1}$$

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

An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all the abundances are not provided, it is safe to assume that all numbers should add up to 100%.

Example 2.1.3: Atomic mass of boron

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

Solution

Boron has two isotopes. We will use the Equation 2.1.1:

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

Substitute these into the equation, and we get:

$$\begin{aligned}\text{Atomic mass} &= (0.20)\,(10) + (0.80)\,(11)\\ &= 10.8\,\text{amu}\end{aligned}$$

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

Example 2.1.4: Atomic mass of neon

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

Solution

Neon has three isotopes. We will use the equation:

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

Substitute these into the equation, and we get:

$$\begin{aligned}\text{Atomic mass} &= (0.9092)\,(19.99) + (0.003)\,(20.99) + (0.0885)\,(21.99)\\ &= 20.17\,\text{amu}\end{aligned}$$

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

The Commission on Isotopic Abundances and Weights publishes isotopic abundances that you can use to practice calculating atomic masses. Then you can use the periodic table to check your answers! Note: These are listed as fractional abundances rather than percentages so you don't need to divide them by 100.

Healthcare Connection

Some isotopes of an element emit radiation. While it is prudent to be cautious around radiation, it also serves many useful purposes. One example is sterilization of medical equipment. Because radiation can kill microorganisms, it can be used to ensure that medical equipment will not transfer bacteria from one patient to another.

One common source of radiation for sterilization is cobalt-60 which emits gamma radiation. It is especially used to prepare single use medical products, but it can be used on other products as well. Demand for sterilized single use medical products increased in 2020 due to the Covid-19 pandemic and the production of test kits and vaccines for the virus.



Source: National Academies of Sciences, Engineering, and Medicine; Division on Earth and Life Studies; Nuclear and Radiation Studies Board; Committee on Radioactive Sources: Applications and Alternative Technologies. Radioactive Sources: Applications and Alternative Technologies. Washington (DC): National Academies Press (US); 2021 Jun 14. 5, Radioactive Sources and Alternative Technologies in Sterilization. Available from: https://www.ncbi.nlm.nih.gov/books/NBK573888/

Summary

- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.
- An element's atomic mass is the weighted average of the masses of the isotopes of an element
- **An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all the abundances are not provided, it is safe to assume that all numbers should add up to 100%.**

Contributors

-
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2.2: Radiation

Learning Objectives

- To define and give examples of the major types of radioactivity.

Atoms are composed of subatomic particles—protons, neutrons, and electrons. Protons and neutrons are located in the nucleus and provide most of the mass of the atom, while electrons circle the nucleus in shells and account for an atom's size. The following notation succinctly identifies an isotope of a particular atom:

$$^{37}_{17}\mathrm{Cl} \tag{2.2.1}$$

The element in this example, represented by the symbol *Cl*, is chlorine. Its atomic number, 17, is the lower left subscript on the symbol and is the number of protons in the atom. The mass number, the superscript to the upper left of the symbol, is the sum of the number of protons and neutrons in the nucleus of this particular isotope. In this case, the mass number is 37, which means that the number of neutrons in the atom is 37 – 17 = 20 (that is, the mass number of the atom minus the number of protons in the nucleus equals the number of neutrons). Occasionally, the atomic number is omitted in this notation because the symbol of the element itself conveys its characteristic atomic number. Another way of expressing a particular isotope is to list the mass number after the element name: chlorine-37.

Atomic theory in the 19th century presumed that nuclei had fixed compositions. But in 1896, the French scientist Henri Becquerel found that a uranium compound placed near a photographic plate made an image on the plate, even if the compound was wrapped in black cloth. He reasoned that the uranium compound was emitting some kind of radiation that passed through the cloth to expose the photographic plate. Further investigations showed that the radiation was a combination of particles and electromagnetic rays, with its ultimate source as the atomic nucleus. These emanations were ultimately called, collectively, **radioactivity**.

There are three main forms of radioactive emissions. The first is called an alpha particle, which is symbolized by the Greek letter α. An alpha particle is composed of two protons and two neutrons, and so it is the same as a helium nucleus. (We often use $^{4}_{2}\mathrm{He}$ to represent an alpha particle.) It has a 2+ charge.

The second major type of radioactive emission is called a beta particle, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus, not from the shells of electrons about the nucleus! Because a beta particle is an electron, it has a 1– charge and we can represent it as $^{0}_{-1}\mathrm{e}$ or β^{-}.

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called gamma rays, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed *ionizing radiation*).

Alpha, beta, and gamma emissions have different abilities to penetrate matter. The relatively large alpha particle is easily stopped by matter (although it may impart a significant amount of energy to the matter it contacts). Beta particles penetrate slightly into matter, perhaps a few centimeters at most. Gamma rays can penetrate deeply into matter and can impart a large amount of energy into the surrounding matter. Table $2.2.1$ summarizes the properties of the three main types of radioactive emissions.

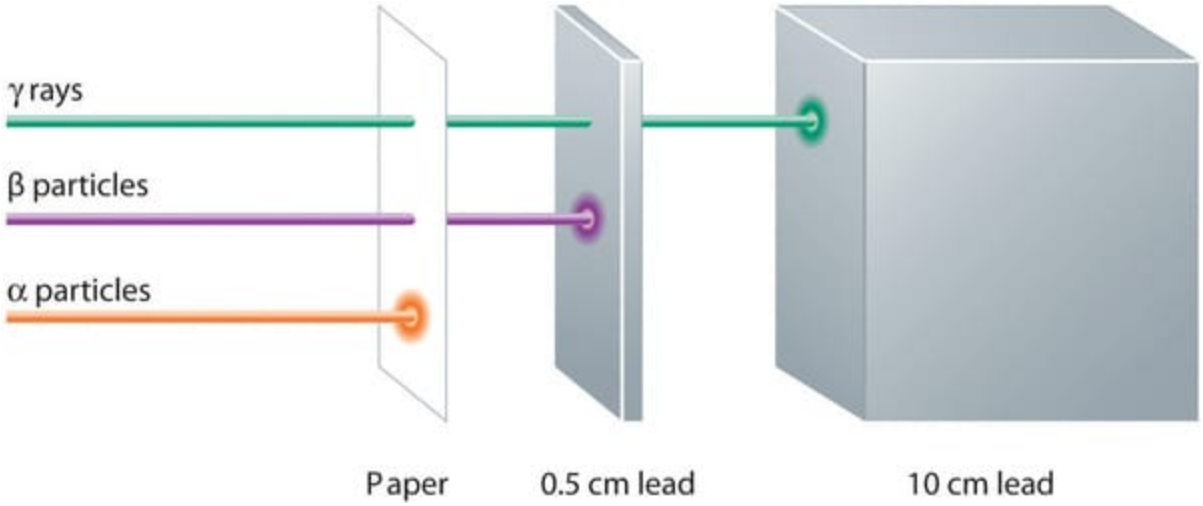


Figure 2.2.2: Different emissions exhibit different pentration powers. (CC BY-NC-SA 3.0; anonymous)

Table 2.2.1: The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, ${}_{2}^{4}\mathrm{He}$	β, ${}_{-1}^{0}\mathrm{e}$	γ
identity	helium nucleus	electron	electromagnetic radiation
charge	2+	1–	none
mass number	4	0	0
penetrating power	minimal (will not penetrate skin)	short (will penetrate skin and some tissues slightly)	deep (will penetrate tissues deeply)

Healthcare Connection

For many people, one of the largest sources of exposure to radiation is from radon gas (Rn-222).

Uranium-238 exists in small amounts in rocks. It decays into radium226 which then decays into radon-222. While the uranium and radium are solids and therefore tend to stay in place in the ground, radon is a gas. The radon slowly escapes from the ground and gradually seeps into homes and other structures above. Since it is about eight times more dense than air, radon gas accumulates in basements and lower floors, and slowly diffuses throughout buildings (Figure 2.2.1).

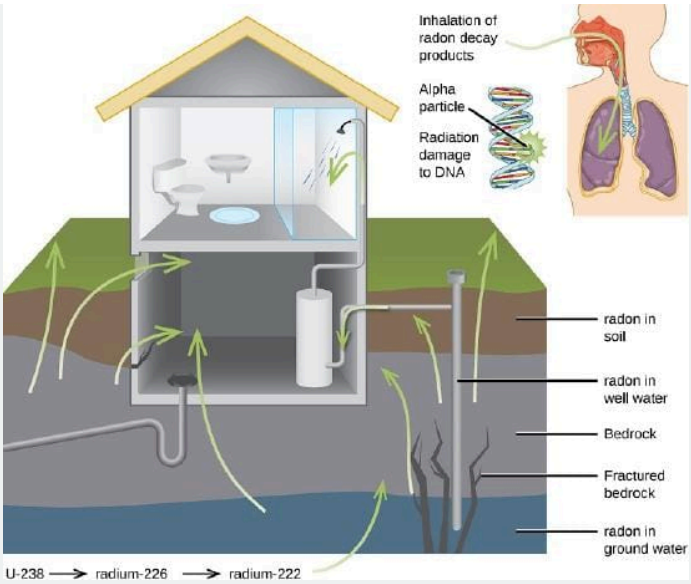


Figure 2.2.1: Radon-222 seeps into houses and other buildings through cracks in concrete foundations and basement floors, stone or porous cinder block foundations, and openings for water and gas pipes.

Radon is found in buildings across the country, with amounts dependent on location. The average concentration of radon inside houses in the US (1.25 pCi/L) is about three times the level found in outside air, and about one in six houses have radon levels high enough that remediation efforts to reduce the radon concentration are recommended. Radon-222 is an α emitter and when it is inhaled the ionizing radiation can damage DNA. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

Summary

- The major types of radioactivity include alpha particles, beta particles, and gamma rays.
- Alpha particles are the largest and the least penetrating, but the most ionizing.

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2.3: Nuclear Reactions and Nuclear Equations

Learning Objectives

- To write nuclear equations to represent nuclear reactions.
- To identify a missing particle when provided with the rest of a nuclear equation.

Reaction Formats

As radioisotopes decay, they tend to follow predictable formats where the original atom gives off, or emits, radiation and at the same time changes into a different element. Some of the common formats are described here.

Alpha Emission

When a radioactive atom emits an alpha particle, the original atom's atomic number decreases by two (because of the loss of two protons), and its mass number decreases by four (because of the loss of four nuclear particles). We can represent the emission of an alpha particle with a nuclear equation—for example, the alpha-particle emission of uranium-235 is as follows:

$$^{235}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\alpha^{+} \quad {}^{231}_{90}\mathrm{Th} \tag{2.3.1}$$

Like mathematical equations, nuclear equations relate things that are equal. However, nuclear equations use an arrow instead of an equal sign to show the passage of time. Substances written on the left are present at the beginning of the reaction and substances written on the right are present at the end. How do we know that a product of the reaction is $^{231}_{90}\mathrm{Th}$? We use the law of conservation of matter, which says that matter cannot be created or destroyed. This means we must have the same number of protons and neutrons on both sides of the equation. If our uranium nucleus loses 2 protons, there are 90 protons remaining, identifying the element as thorium. Moreover, if we lose 4 nuclear particles of the original 235, there are 231 remaining. Thus, we use subtraction to identify the isotope of the thorium atom—in this case, $^{231}_{90}\mathrm{Th}$.

Chemists often use the names *parent isotope* and *daughter isotope* to represent the original atom and the product other than the alpha particle. In the previous example, $^{235}_{92}\mathrm{U}$ is the parent isotope, and $^{231}_{90}\mathrm{Th}$ is the daughter isotope. When one element changes into another in this manner, it undergoes *radioactive decay*.

Example 2.3.1: Radon-222

Write the nuclear equation that represents the radioactive decay of radon-222 by alpha particle emission and identify the daughter isotope.

Solution

Radon has an atomic number of 86, so the parent isotope is represented as $^{222}_{86}\mathrm{Rn}$. We represent the alpha particle as $^{4}_{2}\alpha$ and use subtraction (222 – 4 = 218 and 86 – 2 = 84) to identify the daughter isotope as an isotope of polonium, $^{218}_{84}\mathrm{Po}$:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\alpha^{+} \quad {}^{218}_{84}\mathrm{Po}$$

Exercise 2.3.1: Polonium-209

Write the nuclear equation that represents the radioactive decay of polonium-209 by alpha particle emission and identify the daughter isotope.

Answer available at the end of this section.

Beta Emission

The net effect of beta particle emission on a nucleus is that a neutron is converted to a proton. The overall mass number stays the same, but because the number of protons increases by one, the atomic number goes up by one. Carbon-14 decays by emitting a beta particle:

$$^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\beta \tag{2.3.2}$$

Again, the sum of the atomic numbers is the same on both sides of the equation, as is the sum of the mass numbers. (Note that the beta particle is assigned an "atomic number" of 1–, equal to its charge. This helps us keep track of the atomic numbers as a proton changes into a neutron.)

Gamma Emission

For example, in the radioactive decay of radon-222, both alpha and gamma radiation are emitted, with the latter having an energy of 8.2×10^{-14} J per nucleus decayed:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{218}_{84}\mathrm{Po} + {}^{4}_{2}\alpha^{+}\ \gamma \tag{2.3.3}$$

This may not seem like much energy, but if 1 mole of radon atoms were to decay, the gamma ray energy would be 49 million kJ!

Electron Capture

The format is different than the ones described above. Here an electron is absorbed into the nucleus of an atom causing a proton to change into a neutron. Because the number of protons changes, the atom changes into a different element. For example, aluminum-26 undergoes electron capture to become magnesium-26.

$$^{26}_{13}\mathrm{Al} + {}^{0}_{-1}\mathrm{e} \rightarrow {}^{26}_{12}\mathrm{Mg} \tag{2.3.4}$$

Using the symbol $^{0}_{-1}e$ helps us keep track of the changing atomic numbers as a proton changes into a neutron.

Table 2.3.1: Summary of nuclear reaction formats

Description	Format	Examples
A parent atom decays and emits alpha, beta, or positron radiation	parent → daughter + radiation	alpha decay, beta emission, etc.
An atom absorbs an electron into the nucleus (and a proton changes into a neutron)	parent + e^- → daughter	electron capture

Example 2.3.2: Boron-12

Write the nuclear equation that represents the radioactive decay of boron-12 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is $^{12}_{5}\mathrm{B}$ while one of the products is a beta particle, $^{0}_{-1}\beta$. So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 12, and its atomic number must be 6. The element having an atomic number of 6 is carbon. Thus, the complete nuclear equation is as follows:

$$^{12}_{5}\mathrm{B} \rightarrow {}^{12}_{6}\mathrm{C} + {}^{0}_{-1}\beta^{+}\ \gamma$$

The daughter isotope is $^{12}_{6}\mathrm{C}$.

Exercise 2.3.2: Iodine-131

Write the nuclear equation that represents the radioactive decay of iodine-131 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Answer available at the end of this section.

Bombardment

Occasionally, an atomic nucleus breaks apart into smaller pieces in a radioactive process called *spontaneous fission* (or fission). Typically, the daughter isotopes produced by fission are a varied mix of products, rather than a specific isotope as with alpha and beta particle emission. Often, fission produces excess neutrons that will sometimes be captured by other nuclei, possibly inducing additional radioactive events. Uranium-235 undergoes spontaneous fission to a small extent. One typical reaction is

$$^{235}_{92}\mathrm{U} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 2{}^{1}_{0}\mathrm{n} \tag{2.3.5}$$

where ${}^{1}_{0}\mathrm{n}$ is a neutron. As with any nuclear process, the sums of the atomic numbers and the mass numbers must be the same on both sides of the equation. Spontaneous fission is found only in large nuclei. The smallest nucleus that exhibits spontaneous fission is lead-208.

Fission is the radioactive process used in nuclear power plants and one type of nuclear bomb.

To write the nuclear equation for a reaction:

1. Write the symbols for each particle on the correct side of the arrow. Starting materials that are present before the reaction go on the left. Products that are made during the reaction go on the right. If there is an unknown particle, i.e. one that is mentioned but not identified specifically, leave a space for it.
2. Find the sum of the mass numbers on both sides of the equation. If they are not equal, that helps you identify the value of the mass number for the unknown particle. Repeat this step for the atomic numbers.
3. Identify the symbol for the unknown particle. If it is an element, get the 1 - 2 letter symbol for the element from the periodic table by looking up the atomic number. Important: Do not use the mass number to identify an element! Many elements can have isotopes with the same mass number, but an atomic number uniquely identifies an element. The unknown particle may also be a subatomic particle (${}^{1}_{1}p$, ${}^{1}_{0}n$, or ${}^{0}_{-1}e$) or radiation (${}^{4}_{2}\alpha$, ${}^{0}_{-1}\beta$, or ${}^{0}_{+1}\beta$), so consider whether the numbers you solved for match one of those particles.

Example 2.3.3

The bombardment of curium-245 with calcium-48 produces livermorium-290 and some number of neutrons. Write the nuclear equation, including the proper number of neutrons.

Solution

1. First determine whether each particle should be before or after the arrow and fill in the known numbers based on the mass numbers provided in the problem and the atomic numbers found on the periodic table. The phrase "bombardment of ___ with ___" is indicating how the reaction starts. These particles go on the left. The word "produces" indicates what will be after the arrow. ${}^{245}_{96}\mathrm{Cm} + {}^{48}_{20}\mathrm{Ca} \rightarrow {}^{290}_{116}\mathrm{Lv} + {}^{1}_{0}\mathrm{n}$
2. Next, solve for the number of neutrons that are produced. The sum of the mass numbers on each side of the arrow should be equal. On the left, 245 + 48 = 293. On the right, livermorium has a mass number of 290 and the neutrons make up the difference. There must be three neutrons in order to reach a total of 293 on the right.
3. The symbol for a neutron is ${}^{1}_{0}n$. Reactions like this produce multiple separate neutrons. They are not grouped together into one particle. Therefore, we use a coefficient in front of the neutron symbol to indicate the number of copies produced.

$$^{245}_{96}\mathrm{Cm} + {}^{48}_{20}\mathrm{Ca} \rightarrow {}^{290}_{116}\mathrm{Lv} + 3{}^{1}_{0}\mathrm{n}$$

Example 2.3.4

Uranium-238 can be hit by a small atom to produce neptunium-238 and two neutrons. Write the nuclear equation for this reaction, including the symbol for the unidentified atom.

Solution

1. First write the symbols for the particles that you know on the correct side of the arrow. Leave a space for the unknown particle, if relevant. In this case a collision initiates the reaction so the uranium and "small atom" are on the left. The neptunium and neutrons are produced, so they are written after the arrow. $^{238}_{92}\mathrm{U} + ? \rightarrow {}^{238}_{93}\mathrm{Np} + 2{}^{1}_{0}\mathrm{n}$
2. Next, determine the missing mass number and atomic number. Remember that the sums need to be the same on both sides of the arrow. For the mass numbers 238 + ? = 238 + (2)1 where the mass number for the neutron is multiplied by two because there are two neutrons. The missing mass number is a 2. For the atomic numbers, 92 + ? = 93 + (2)0 so the unknown atom has an atomic number of 1.
3. Finally, write the missing symbol. The atomic number of 1 tells us that the unknown element is hydrogen. Remember not to use the mass number to identify an element.

$$^{238}_{92}\mathrm{U} + {}^{2}_{1}\mathrm{H} \rightarrow {}^{238}_{93}\mathrm{Np} + 2{}^{1}_{0}\mathrm{n}$$

Exercise 2.3.3

A lead-208 atom collides with another atom to form darmstadtium-269 and a neutron. Write the nuclear equation for this reaction, including the symbol for the unidentified atom.

Answer available at the end of this section.

Summary

- Particles present before the reaction start are written before/to the left of the arrow. Particles that result from a reaction are written after/to the right of the arrow. Note: The arrow always points from left to right.
- Nuclear equations should account for all of the protons and neutrons involved in a nuclear reaction.
  - The sums of the atomic numbers should be the same on both sides of the equation
  - The sums of the mass numbers should be the same on both sides of the equation
- If an unknown particle needs to be identified, figure out its mass number and atomic number first, then identify the 1-2 letter symbol that represents it.
  - Use the atomic number, not the mass number to identify an element.
  - Use both numbers to identify a subatomic particle or radiation.

Concept Review Exercise

1. What are the major types of radioactivity? Write chemical equations demonstrating each type.

Answer

1. The major types of radioactivity are alpha decay, beta decay, and gamma ray emission; alpha decay with gamma emission: $^{222}_{86}\mathrm{Rn} \rightarrow {}^{218}_{84}\mathrm{Po} + {}^{4}_{2}\mathrm{He} + \gamma$; beta decay: $^{14}_{6}\mathrm{C} \rightarrow {}_{7}{}^{14}\mathrm{N} + {}^{0}_{-1}\mathrm{e}$ (answers will vary)

Answers

Exercise 2.3.1: Polonium-209

$^{209}_{84}\mathrm{Po} \rightarrow {}^{4}_{2}\alpha^{+} {}^{205}_{82}\mathrm{Pb}$

Exercise 2.3.2: Iodine-131

$^{131}_{53}\mathrm{I} \rightarrow {}^{131}_{54}\mathrm{Xe} + {}^{0}_{-1}\beta^{+} \gamma$

Exercise 2.3.3

$^{208}_{82}\mathrm{Pb} + ^{62}_{28}\mathrm{Ni} \rightarrow {}^{269}_{110}\mathrm{Ds} + {}^{1}_{0}\mathrm{n}$

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2.4: Half-Life

Learning Objectives

- To define *half-life*.
- To solve word problems involving half-life, time, and mass of a radioactive substance.

Some isotopes are stable indefinitely, while others are radioactive and decay over time. As time passes, less and less of the radioactive isotope will be present, and the level of radioactivity decreases. An interesting and useful aspect of radioactive decay is the **half-life**. The half-life of a radioactive isotope is the amount of time it takes for one-half of the radioactive isotope to decay. The half-life of a specific radioactive isotope is constant; it is not affected by conditions and is independent of the initial amount of that isotope. This can be represented by

$$\text{total time} = \text{number of half} - \text{lives} \times \text{length of half} - \text{life} \tag{2.4.1}$$

where the number of half-lives is an integer, the length of half-life is the duration of each half-life, and the total time how long the decay of the isotope was observed. If the number of half-lives is not an integer a different equation is needed. This book will only use examples in which the number of half-lives is a whole number.

Consider the following example. Suppose we have 100.0 g of ^{3}H (tritium, a radioactive isotope of hydrogen). It has a half-life of 12.3 years. After 12.3 y, half of the sample will have decayed to ^{3}He by emitting a beta particle, so that only 50.0 g of the original ^{3}H remains. After another 12.3 y—making a total of 24.6 y—another half of the remaining ^{3}H will have decayed, leaving 25.0 g of ^{3}H. After another 12.3 y—now a total of 36.9 y—another half of the remaining ^{3}H will have decayed, leaving 12.5 g of ^{3}H. This sequence of events is illustrated in Figure 2.4.1.

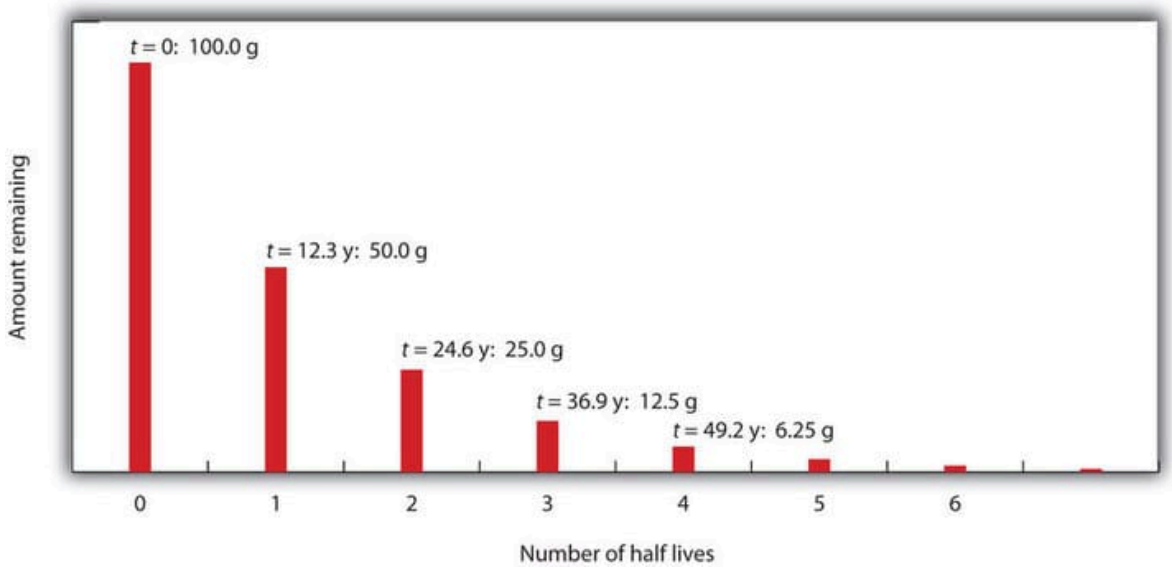


Figure 2.4.1: Radioactive Decay. During each successive half-life, half of the initial amount will radioactively decay.

Please note that "amount remaining" is referring specifically to the amount of ^{3}H that is remaining. As it undergoes beta decay it turns into helium according to this equation: ${}^{3}_{1}H \rightarrow {}^{0}_{-1}\beta + {}^{3}_{2}He$. Atoms still exist, but they are helium instead of hydrogen. Throughout this section we will focus on the parent and generally omit information about the daughter.

We can keep track of the mass of parent that remains by using a table. After each half-life the mass of the parent isotope has decreased by a factor of two. Notice that the initial mass of the parent isotope corresponds to zero half-lives. The initial mass is the mass before the first half-life has occurred.

Table 2.4.1:Mass of hydrogen-3 remaining after each half-life in its beta decay reaction

Number of Half-Lives	Mass of ^{3}H remaining
0	100.0 g
1	50.0 g
2	25.0 g
3	12.5 g
4	6.25 g
5	3.13 g
6	1.56 g

Alternatively, we can determine the amount of a radioactive isotope remaining after a given number half-lives by using the following expression:

$$\text{amount remaining} = \text{initial amount} \times \left(\frac{1}{2}\right)^{n} \tag{2.4.2}$$

where *n* is the number of half-lives. This expression works even if the number of half-lives is not a whole number whereas the table (dividing by 2 repeatedly) is only accurate when the number of half-lives is an integer.

To solve half-life problems involving the initial mass, final mass, half-life, and total time where you are given three of the values and asked to solve for the fourth, the general procedure is:

1. Calculate the number of half-lives. How you do this depends on whether you were provided with the initial mass and final mass or whether you were provided with the half-life and total time.
2. Use the number of half-lives that you calculated from step 1 along with the additional information that was provided in the problem (and not yet used) to solve for the requested value.

Example 2.4.1: Fluorine-20

The half-life of ^{20}F is 11.0 seconds. If a sample initially contains 5.00 g of ^{20}F, how much ^{20}F remains after 44.0 s?

Solution

Step 1: Since two times were provided, but only one mass, use the times to calculate the number of half-lives that pass during this problem. The wording "the half-life is" tells us the length of each half-life; in this case it is 11.0 s. The 44.0 s is the total time. Rearrange 2.4.1 to solve for the number of half-lives which will be a small whole number.

total time = number of half-lives x length of half-life

44.0 s = number of half-lives x 11.0 s

number of half-lives = 4 half-lives

Step 2: Now we know that the fluorine-20 decays to half its mass four times (the number of half-lives). We know the original mass and are solving for the final mass.

- Option 1: When the initial mass and number of half-lives are known you can start with the initial mass and divide it by 2 once for each half-life as seen in Table 2.4.1. In this case the initial mass is 5.00 g and there are four half-lives, so divide by 2 four times: 5.0 g ÷ 2 ÷ 2 ÷ 2 ÷ 2 = 0.313 g. If organizing your work into a table, be sure that the first row matches 5.00 g (the initial mass) with 0 half-lives.
- Option 2: We can use Equation 2.4.2 with $n = 4$. Substituting and solving results in the following:

$$\begin{aligned}\text{amount remaining} &= 5.00\ g \times \left(\frac{1}{2}\right)^4 \\ &= 5.00\ g \times \frac{1}{16} \\ &= 0.313\ g\end{aligned}$$

Using either option we find that **0.313 g ^{20}F remains**.

Exercise 2.4.1: Titanium-44

The half-life of ^{44}Ti is 60.0 y. A sample initially contains 0.600 g of ^{44}Ti. How much ^{44}Ti remains after 180.0 y?

Answer available at the end of this section.

Half-lives of isotopes range from fractions of a microsecond to billions of years. Table 2.4.1 lists the half-lives of some isotopes.

Table 2.4.1: Half-Lives of Various Isotopes

Isotope	Half-Life
^{3}H	12.3 y
^{14}C	5,730 y
^{40}K	1.26×10^{9} y
^{51}Cr	27.70 d
^{90}Sr	29.1 y
^{131}I	8.04 d
^{222}Rn	3.823 d
^{235}U	7.04×10^{8} y
^{238}U	4.47×10^{9} y
^{241}Am	432.7 y
^{248}Bk	23.7 h
^{260}Sg	4 ms

Example 2.4.2: Iodine-131

The isotope I-131 is used in treatment for thyroid cancer and has a half-life of 8.05 days. If a clinic initially has 400 mg, how long will it take for the sample to decay to only 25 mg of I-131?

Solution

We begin by determining how many half-lives it would take for the sample to decay from 400 mg to 25 mg of I-131. There are several ways to do this! We know that for each half-life the mass of I-131 is reduced to one half of the previous value so we could solve this conceptually:

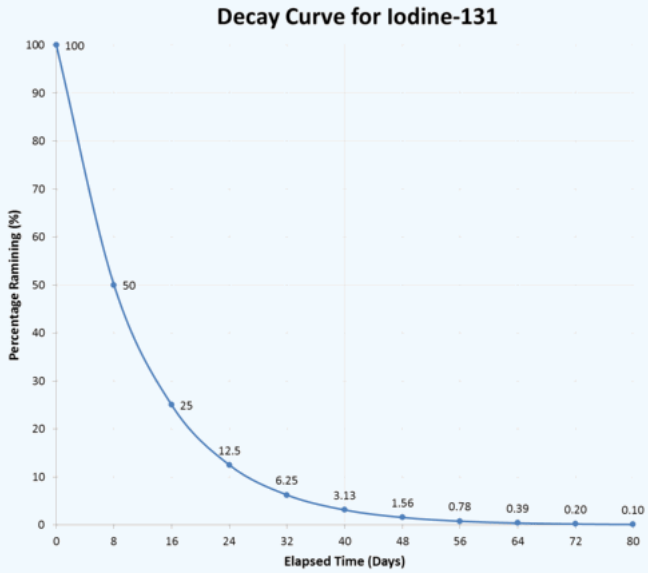
$$\frac{400\ \text{mg}}{2} = 200\ \text{mg}$$
$$\frac{200\ \text{mg}}{2} = 100\ \text{mg}$$
$$\frac{100\ \text{mg}}{2} = 50\ \text{mg}$$
$$\frac{50\ \text{mg}}{2} = 25\ \text{mg}$$

We had to divide the original mass by 2 four times to get the final mass. Therefore, it took four half-lives to get from 400 mg to 25 mg of I-131.

Once the number of half-lives has been determined, we can calculate the total time.

$$\begin{aligned}\text{total time} &= \text{half-life} \times \text{number of half-lives} \\ &= 8.05\ \text{days} \times 4\ \text{half-lives} = 32.2\ \text{days}\end{aligned}$$

A more sophisticated way to solve the problem is to make use of percentages or fractions. The graph below shows the percentage of I-131 remaining after each half-life. These percentages are always the same, regardless of the initial mass of I -131.



Typical radioactive decay curve.

The graph above illustrates a typical decay curve. Since the half-life for $\mathrm{I}-131$ is 8.05 days, the x-axis has units of days. The y-axis indicates the percentage of sample remaining. Initially, (at elapsed time = 0), 100 % of the sample remains. After one half-life (8.05 days for $\mathrm{I}-131$), 50 % remains. After two half-lives, 25 % of the initial sample remains. We can use this pattern as an alternate way of calculating the number of half-lives when the initial and final masses are known.

Alternate method of solving:

$$\frac{\text{final mass}}{\text{initial mass}} \times 100 = \text{percentage remaining} \tag{2.4.3}$$

$$\frac{25\ \text{mg}}{400\ \text{mg}} \times 100 = 6.25$$

This result, 6.25 % remaining, corresponds to four half-lives. Careful! Don't count 100 % as the first half-life. At time = 0, nothing has happened yet. 50 % remaining = 1 half-life, 25 % remaining = 2 half-lives, 12.5 % remaining = 3 half-lives, and 6.25 % remaining = 4 half-lives. These percentages are the same for all isotopes so you could calculate them once and then use them for multiple problems.

To finish this problem, multiply the number of half-lives by the length of each half-life to get the total time.

Exercise 2.4.2

A sample of Ac-225 originally contained 80 grams and after 50 days only 2.5 grams of the original Ac-225 remain. What is the half life of Ac-225?

Answer available at the end of this section.

Looking Closer: Half-Lives of Radioactive Elements

Many people think that the half-life of a radioactive element represents the amount of time an element is radioactive. In fact, it is the time required for half—not all—of the element to decay radioactively. Occasionally, however, the daughter element is also radioactive, so its radioactivity must also be considered.

The expected working life of an ionization-type smoke detector is about 10 years. In that time, americium-241, which has a half-life of about 432 y, loses less than 4% of its radioactivity. A half-life of 432 y may seem long to us, but it is not very long as half-lives go. Uranium-238, the most common isotope of uranium, has a half-life of about 4.5×10^9 y, while thorium-232 has a half-life of 14×10^9 y.

On the other hand, some nuclei have extremely short half-lives, presenting challenges to the scientists who study them. The longest-lived isotope of lawrencium, ^{262}Lr, has a half-life of 3.6 h, while the shortest-lived isotope of lawrencium, ^{252}Lr, has a half-life of 0.36 s. As of this writing, the largest atom ever detected has atomic number 118, mass number 293, and a half-life of 120 ns (nanoseconds). Can you imagine how quickly an experiment must be done to determine the properties of elements that exist for so short a time?

Key Takeaways

- Natural radioactive processes are characterized by a half-life, the time it takes for half of the material to decay radioactively.
- The amount of material left over after a certain number of half-lives can be easily calculated.

Concept Review Exercises

1. Define *half-life*.
2. Describe a way to determine the amount of radioactive isotope remaining after a given number of half-lives.

Answers

1. Half-life is the amount of time needed for half of a radioactive material to decay.
2. take half of the initial amount for each half-life of time elapsed

Answers

Exercise 2.4.1: Titanium-44

0.075 g.

Exercise 2.4.2

10 days. Notice that the phrase "what is the half-life?" means "how long is the half-life?".

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2.5: Radiation and the Body

Learning Objectives

- To list sources of radiation in daily life

What is a person's annual exposure to radioactivity and radiation? Table 2.5.1 lists the sources and annual amounts of radiation exposure in units of millirems, mrem. It may surprise you to learn that about 82% of the radioactivity and radiation exposure we receive is from natural sources—sources we cannot avoid. Ten percent of the exposure comes from our own bodies—largely from ^{14}C and ^{40}K.

Table 2.5.1: Average Annual Radiation Exposure (Approximate)

Source	Amount (mrem)
radon gas	200
medical sources	53
radioactive atoms in the body naturally	39
terrestrial sources	28
cosmic sources	28
consumer products	10
nuclear energy	0.05
Total	358

Flying from New York City to San Francisco adds 5 mrem to your overall radiation exposure because the plane flies above much of the atmosphere, which protects us from most cosmic radiation.

The actual effects of radioactivity and radiation exposure on a person's health depend on the type of radioactivity, the length of exposure, and the tissues exposed. Table 2.5.2 lists the potential threats to health at various amounts of exposure over short periods of time (hours or days). The first column lists amounts of radiation; as the dose of radiation increases the effect increases as well.

Table 2.5.2: Effects of Short-Term Exposure to Radioactivity and Radiation

Exposure (rem)	Effect
1 (over a full year)	no detectable effect
~20	increased risk of some cancers
~100	damage to bone marrow and other tissues; possible internal bleeding; decrease in white blood cell count
200–300	visible "burns" on skin, nausea, vomiting, and fatigue
>300	loss of white blood cells; hair loss
~600	death

One of the simplest ways of detecting radioactivity is by using a piece of photographic film embedded in a badge or a pen. On a regular basis, the film is developed and checked for exposure. A comparison of the exposure level of the film with a set of standard exposures indicates the amount of radiation a person was exposed to.

Key Takeaway

- Radiation is around us all the time, but in small doses
- Larger amounts of radiation lead to greater physiological effects

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2.6: Uses of Radioactive Isotopes in Medicine

Learning Objectives

- To learn some applications of radioactivity in medicine.

Radioactive isotopes have a variety of medical applications. Generally, however, they are useful either because we can detect their radioactivity or we can use the energy they release.

Radioactivity and Imaging/Diagnostics

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A tracer is a substance that can be used to follow the pathway of that substance through some structure such as the human body.

One example of a diagnostic application of radiation is using radioactive iodine-131 to test for thyroid activity (Figure 2.6.2). The thyroid gland in the neck is one of the few places in the body with a significant concentration of iodine. To evaluate thyroid activity, a measured dose of iodine-131 is administered to a patient, and the next day a scanner is used to measure the amount of radioactivity in the thyroid gland. The amount of radioactive iodine that collects there is directly related to the activity of the thyroid, allowing trained physicians to diagnose both hyperthyroidism and hypothyroidism. Iodine-131 has a half-life of only 8 d, so the potential for damage due to exposure is minimal. Technetium-99 can also be used to test thyroid function. Bones, the heart, the brain, the liver, the lungs, and many other organs can be imaged in similar ways by using the appropriate radioactive isotope.

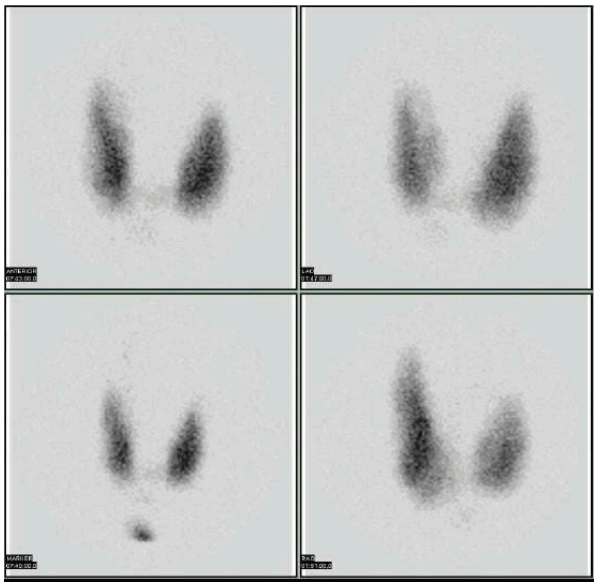


Figure 2.6.2: Medical Diagnostics. Radioactive iodine can be used to image the thyroid gland for diagnostic purposes. Source: Scan courtesy of Myo Han, Wikipedia, Thyroid Scan(opens in new window) [en.Wikipedia.org].

Very little radioactive material is needed in these diagnostic techniques because the radiation emitted is so easy to detect.

Radiation and Medical Treatment

Therapeutic applications of radiation usually require much larger doses than diagnostic procedures because their purpose is to preferentially kill diseased tissues. For example, if a thyroid tumor is detected, a much larger infusion (thousands of rem, as opposed to a diagnostic dose of less then 40 rem) of iodine-131 could help destroy the tumor cells. Similarly, radioactive strontium is used to not only detect but also ease the pain of bone cancers. Table 2.6.1 lists several radioactive isotopes and their medical

uses.

Table 2.6.1: Some Radioactive Isotopes That Have Medical Applications

Isotope	Use
^{32}P	cancer detection and treatment, especially in eyes and skin
^{59}Fe	anemia diagnosis
^{60}Co	gamma ray irradiation of tumors
^{99m}Tc	brain, thyroid, liver, bone marrow, lung, heart, and intestinal scanning; blood volume determination
^{131}I	diagnosis and treatment of thyroid function
^{133}Xe	lung imaging
^{198}Au	liver disease diagnosis

In addition to the direct application of radioactive isotopes to diseased tissue, the gamma ray emissions of some isotopes can be directed toward the tissue to be destroyed. Cobalt-60 is a useful isotope for this kind of procedure.

To Your Health: Positron Emission Tomography Scans

One relatively rare form of radioactivity is called *positron emission*. It is similar to beta particle emission, except that instead of emitting an electron, a nucleus emits a positively charged electron, called a *positron*. A positron is actually a piece of antimatter; therefore, when a positron encounters an electron, both particles are converted into high-energy gamma radiation.

Isotopes that emit positrons can be employed in a medical imaging technique called *positron emission tomography* (*PET*). A patient receives a compound containing a positron-emitting isotope, either intravenously or by ingestion. The radioactive compound travels throughout the body, and the patient is then pushed slowly through a ring of sensors that detect the gamma radiation given off by the annihilation of positrons and electrons. A computer connected to the sensors constructs a three-dimensional image of the interior of part or all of the patient's body, allowing doctors to see organs or tumors or regulate the function of various organs (such as the brain or the heart) to diagnose the medical condition of the patient.



Figure 2.6.3: (left) Combined apparatus for positron emission tomography (PET) and X-ray computer tomography (CT), Siemens Biograph (right) Whole-body PET scan using 18F-FDG. Images used with permission from Wikipedia.

Two isotopes that undergo positron emission are carbon-11 and fluorine-18, with half-lives of 20.4 and 110 min, respectively. Both isotopes can be incorporated into sugar molecules and introduced into the body. Doctors can use the intensity of gamma ray emission to find tissues that metabolize the sugar faster than other tissues; fast-metabolizing tissue is one sign of a malignant (i.e., cancerous) tumor. Researchers use similar techniques to map areas of the brain that are most active during specific tasks, such as reading or speaking.

PET is one of many diagnostic and treatment methods that physicians use to improve the quality of our lives. It is one of the many positive uses of radioactivity in society.

Key Takeaway

- Radioactivity has several practical applications in medicine

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3: MEASUREMENTS, UNIT CONVERSIONS, AND DENSITY

3.1: Expressing Numbers - Scientific Notation
3.2: Significant Figures
3.3: Units for Measurements
3.4: Converting Units
3.5: Solving Multistep Conversion Problems
3.6: Density

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3.1: Expressing Numbers - Scientific Notation

Learning Objectives

- Express a large number or a 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 times 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 next 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. × 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 3.1.1: Scientific Notation

Express each number in scientific notation.

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

Answer a available at the end of this section.

Exercise 3.1.1

Express each number in scientific notation.

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

Answer a available at the end of this section.

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 3.1.2

Express each number in standard, or conventional notation.

a. 5.27×10^4
b. 1.0008×10^6

Answer a available at the end of this section.

Exercise 3.1.2

Express each number in standard, or conventional notation.

a. 6.98×10^8
b. 1.005×10^2

Answer a available at the end of this section.

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 times 10^{-3}, or 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:

$$0.\underbrace{006} = 6 \times 10^{-3}$$

Note:

In writing scientific notations, 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 n places, the power (n) of 10 is positive.

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

- Numbers that are less than one have a negative power in scientific notation. If the decimal point is moved to the right n places, the power (n) of 10 is negative. $0.\underbrace{006} = 6 \times 10^{-3}$

Example 3.1.3

Express each number in scientific notation.

a. 0.000006567
b. −0.0004004
c. 0.000000000000123

Answer a available at the end of this section.

Exercise 3.1.3

Express each number in scientific notation.

a. 0.000355
b. 0.314159
c. −0.051204

Answer a available at the end of this section.

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 3.1.4

Change the number in scientific notation to standard form.

a. 6.22×10^{-2}
b. 9.9×10^{-9}

Answer a available at the end of this section.

Exercise 3.1.4

Change the number in scientific notation to standard form.

a. 9.98×10^{-5}
b. 5.109×10^{-8}

Answer a available at the end of this section.

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.

Exercise 3.1.1

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

Answer available at the end of this section.

Exercise 3.1.2

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?

Answer available at the end of this section.

Key Takeaway

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

Answers

Example 3.1.1: Scientific Notation

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

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}.

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

Exercise 3.1.1

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

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}.

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

Example 3.1.2

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

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

Exercise 3.1.2

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

Moving the decimal point two places to the right gives 100.5

Example 3.1.3

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

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.

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

Exercise 3.1.3

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

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

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

Example 3.1.4

0.0622

0.0000000099

Exercise 3.1.4

0.0000998

0.000000005109

Exercise 3.1.1

Scientific notation is more convenient than listing a large number of zeros.

Exercise 3.1.2

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.

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3.2: Significant Figures

Learning Objective

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

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

$$337 \div 217 = 1.5529953917$$

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

Consider using a ruler to measure the width of an object, as shown in Figure 3.2.1. The object is definitely more than 1 cm long, so we know that the first digit in our measurement is 1. We see by counting the tick marks on the ruler that the object is at least three ticks after the 1. If each tick represents 0.1 cm, then we know the object is at least 1.3 cm wide. But our ruler does not have any more ticks between the 0.3 and the 0.4 marks, so we can't know exactly how much the next decimal place is. But with a practiced eye we can estimate it. Let us estimate it as about six-tenths of the way between the third and fourth tick marks, which estimates our hundredths place as 6, so we identify a measurement of 1.36 cm for the width of the object.

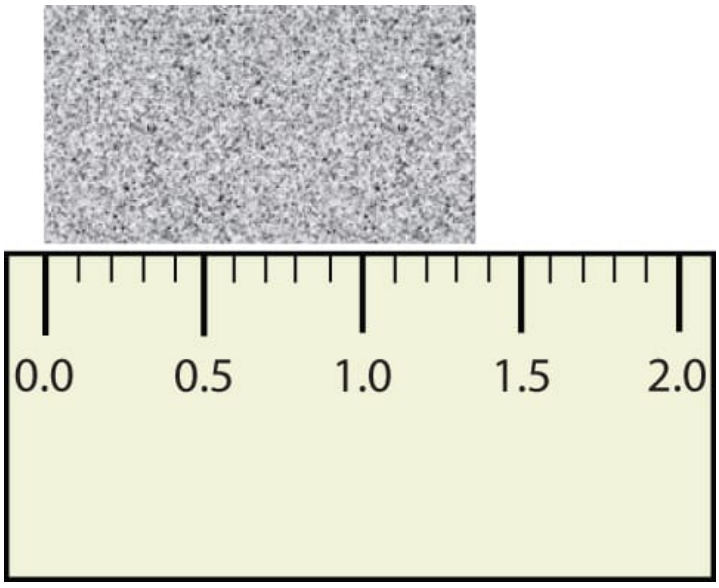


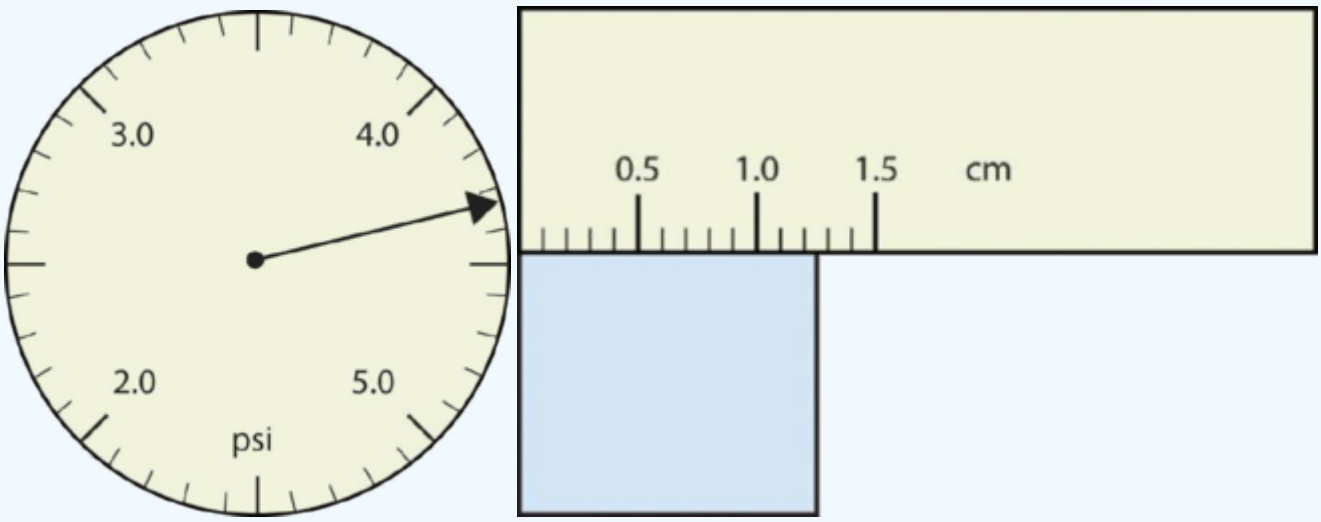
Figure 3.2.1: Expressing Width. What is the proper way to express the width of this object?

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

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

Example 3.2.1

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

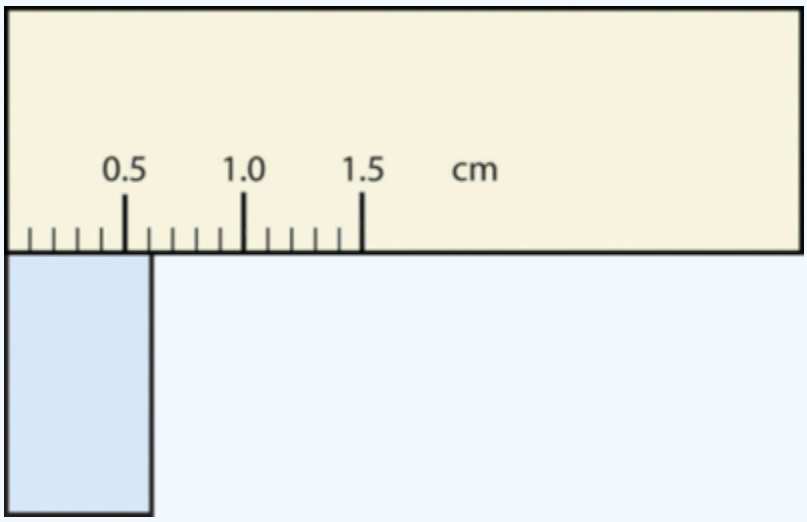


Solution

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

Exercise 3.2.1

What would be the reported width of this rectangle?



Answer available at the end of this section.

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

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

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

So, by these rules, the population figure of the United States has only three significant figures: the 3, the 6, and the zero between them. The remaining six zeros simply put the 306 in the millions position. (See Figure 3.2.2 for another example.)



Figure 3.2.2: Significant Figures. How many of the numbers in this display are actually significant? (Public Domain.)

Example 3.2.2

Give the number of significant figures in each measurement.

a. 36.7 m
b. 0.000606 s
c. 2,002 kg
d. 306,490,000 people

Solution

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

Exercise 3.2.2

Give the number of significant figures in each measurement.

a. 0.000601 m
b. 65.080 kg

Answer a available at the end of this section.

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

$$\begin{array}{r} 1.2 \\ \underline{4.41} \\ 5.61 \end{array}$$

↑ limit final answer to the tenths column: 5.6

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

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

$$\begin{array}{r} 77.2 \\ \underline{10.46} \\ 87.66 \end{array}$$

↑ limit final answer to the tenths column and round up: 87.7

Example 3.2.3

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

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

Solution

a. If we use a calculator to add these two numbers, we would get 119.902. However, most calculators do not understand significant figures, and we need to limit the final answer to the tenths place. Thus, we drop the 02 and report a final answer of 119.9 (rounding down).

b. A calculator would answer 201.867. However, we have to limit our final answer to the hundredths place. Because the first number being dropped is 7, which is greater than 5, we round up and report a final answer of 201.87.

Exercise 3.2.3

Express the answer for

$$3.445 + 90.83 - 72.4$$

to the proper number of significant figures.

Answer available at the end of this section.

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

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

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

Example 3.2.4: Significant Figures

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

a. 76.4 × 180.4 = ?
b. 934.9 ÷ 0.00455 = ?

Solution

a. The first number has three significant figures, while the second number has four significant figures. Therefore, we limit our final answer to three significant figures: 76.4 × 180.4 = 13,782.56 = 13,800.
b. The first number has four significant figures, while the second number has three significant figures. Therefore, we limit our final answer to three significant figures: 934.9 ÷ 0.00455 = 205,472.5275... = 205,000.

Exercise 3.2.4

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

a. 22.4 × 8.314 = ?
b. 1.381 ÷ 6.02 = ?

Answer a available at the end of this section.

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

Summary

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

Answers

Exercise 3.2.1

0.63 cm

Exercise 3.2.2

three significant figures

five significant figures

Exercise 3.2.3

21.9

Exercise 3.2.4

186

0.229

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3.3: Units for Measurements

Learning Objectives

- Identify the appropriate metric unit for measuring items of various dimensions.
- Use metric and English units in unit analysis problems.

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

The number in the measurement can be represented in different ways, including decimal form and scientific notation. For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as 2.98×10^5 kg. The mass of the average mosquito is about 0.00000025 kilograms, which can be written as 2.5×10^{-6} kg.

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

We usually report the results of scientific measurements in SI units, an updated version of the metric system, using the units listed in Table 3.3.1. Other units can be derived from these base units. The standards for these units are fixed by international agreement, and they are called the International System of Units or SI Units (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964.

***Table** 3.3.1**:** Base Units of the SI System*

Property Measured	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K
amount of substance	mole	mol

Metric System

Sometimes we use units that are fractions or multiples of a base unit. In the metric system these fractions or multiples are always powers of 10. Fractional or multiple SI units are named using a prefix and the name of the base unit. For example, a length of 1000 meters is also called a kilometer because the prefix *kilo* means "one thousand," which in scientific notation is 10^3 (1 kilometer = 1000 m = 10^3 m).

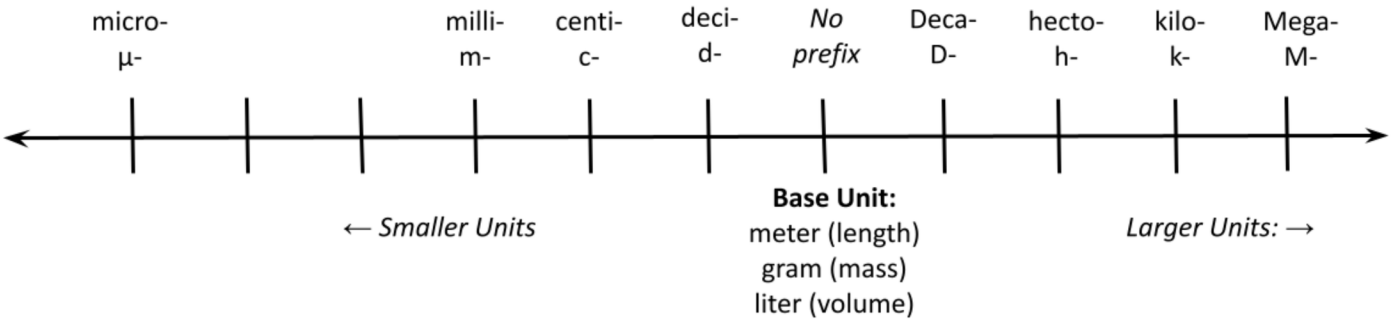


Figure 3.3.1: Prefixes used with SI and metric units. Each step on the number line represents a factor of 10.

The number line in Figure 3.3.1 can be used to determine the relationship between units with the same base unit. Meters is a base unit; it is abbreviated m with no prefix. Kilometers is abbreviated km. Meters and kilometers are three steps apart on the number line. This corresponds to three factors of ten or 10 x 10 x 10 = 1000. The smaller units are written to the left on the number line in Figure 3.3.1. It takes more of the smaller unit to equal a larger unit. Thus, it takes 1000 m to equal the length of 1 km.

Example 3.3.1

What is the relationship between microliters (μL and milliliters (mL)?

Solution

1. Check that the two units have the same base unit. The number line cannot be used for units with different base units. In this case both units end with L, the base unit liters, so we can proceed.
2. Find the prefixes on the number line and count how many steps it take to get from one to the other. It takes three steps to get from micro to milli.
3. Each step is worth a factor of ten so three steps is 10 x 10 x 10 = 1000. In scientific notation this can be written as 1 x 10^3 where the exponent on the ten is the number of steps.
4. There are many of the smaller unit in one of the larger unit, so the relationship is 1000 μL = 1 mL.

Exercise 3.3.1

What is the relationship between centimeters and kilometers?

Answer available at the end of this section.

Exercise 3.3.2

What is the relationship between centiliters and deciliters?

Answer available at the end of this section.

SI Base Units

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

Length

The standard unit of length in both the SI and original metric systems is the meter (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. A meter is about 3 inches longer than a yard (Figure 3.3.2); one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 1 x 10^{3} m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 1 x 10^{-2} m) or millimeters (1 mm = 0.001 m = 1 x 10^{-3} m).

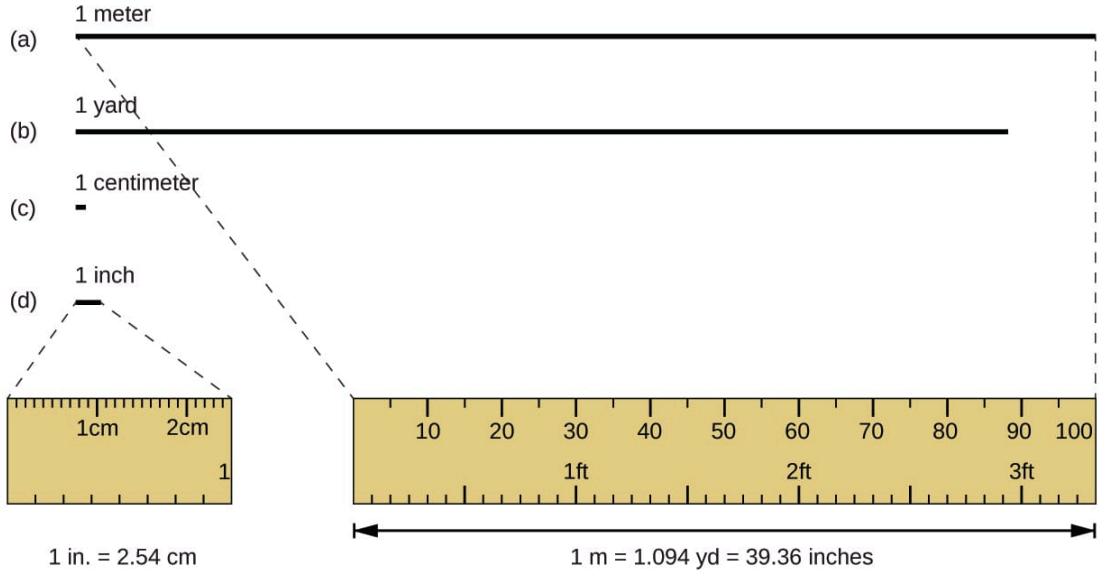


Figure 3.3.2: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

Mass

The standard unit of mass in the SI system is the kilogram (kg). A kilogram was originally defined as the mass of a liter of water (a cube of water with an edge length of exactly 0.1 meter). In 1889, it was redefined by a certain cylinder of platinum-iridium alloy, which was kept in France (Figure 3.3.3). Any object with the same mass as this cylinder was said to have a mass of 1 kilogram. (This led to uncertainties unacceptable to the precision of modern instrumentation. Since May 20, 2019, a new definition has be used for the kilogram, based on the unchanging Planck's constant.).

One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (1 x 10^{-3} kg).



Figure 3.3.3: This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology).

Temperature

Temperature is an intensive property. The SI unit of temperature is the kelvin (K). The IUPAC convention is to use kelvin (all lowercase) for the word, K (uppercase) for the unit symbol, and neither the word “degree” nor the degree symbol (°). The degree Celsius (°C) is also allowed in the SI system, with both the word “degree” and the degree symbol used for Celsius measurements. Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places. Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C) by definition, and normal human body temperature is approximately 310 K (37 °C). The conversion between these two units and the Fahrenheit scale will be discussed later in this chapter.

Time

The SI base unit of time is the second (s). Small and large time intervals can be expressed with the appropriate prefixes; for example, 3 microseconds = 0.000003 s = 3×10^{-6} and 5 megaseconds = 5,000,000 s = 5×10^{6} s. Alternatively, hours, days, and years can be used.

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (Figure 3.3.4). The standard volume is a cubic meter (m^3), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

A more commonly used unit of volume is derived from the decimeter (0.1 m, or 10 cm). A cube with edge lengths of exactly one decimeter contains a volume of one cubic decimeter (dm^3). A liter (L) is the more common name for the cubic decimeter. One liter is about 1.06 quarts. A cubic centimeter (cm^3) is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for **c**ubic **c**entimeter) is often used by health professionals. A cubic centimeter is also called a milliliter (mL) and is 1/1000 of a liter.

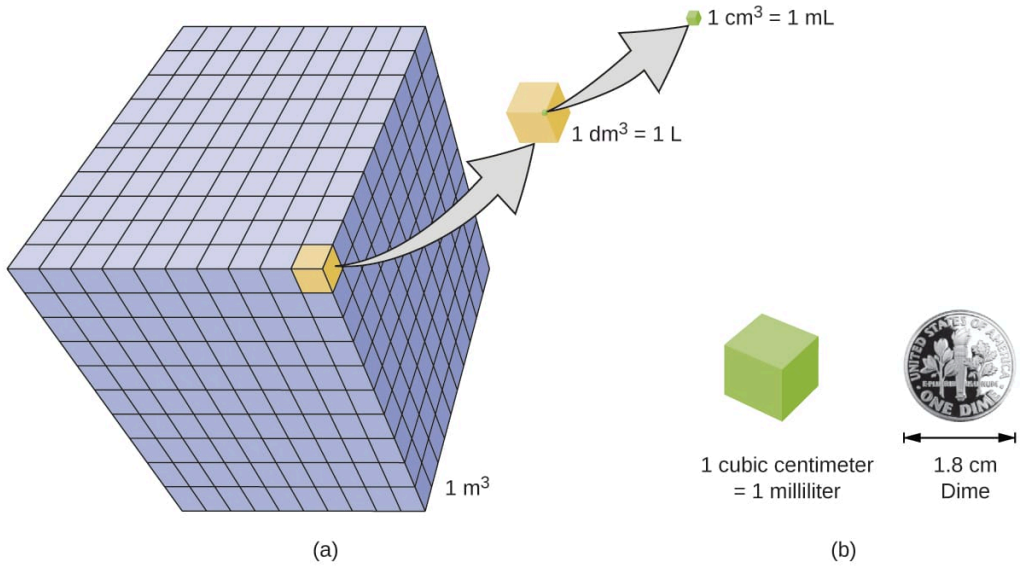


Figure 3.3.4: (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

Exercise 3.3.3

What type of measurement does each of the following correspond to? Answer options are: length, mass, temperature, time, volume.

- 4.5 g
- 1500 mL
- 225 ms

Answer available at the end of this section.

Summary

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

Glossary

Celsius (°C)
unit of temperature; water freezes at 0 °C and boils at 100 °C on this scale

cubic centimeter (cm^3 or cc)
volume of a cube with an edge length of exactly 1 cm

cubic meter (m^3)
SI unit of volume

kelvin (K)
SI unit of temperature; 273.15 K = 0 °C

kilogram (kg)
standard SI unit of mass; 1 kg = approximately 2.2 pounds

length
measure of one dimension of an object

liter (L)
(also, cubic decimeter) unit of volume; 1 L = 1,000 cm^3

meter (m)
standard metric and SI unit of length; 1 m = approximately 1.094 yards

milliliter (mL)
1/1,000 of a liter; equal to 1 cm^3

second (s)
SI unit of time

SI units (International System of Units)
standards fixed by international agreement in the International System of Units (*Le Système International d'Unités*)

unit
standard of comparison for measurements

volume
amount of space occupied by an object

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- Christy VanRooyen, Oregon Tech
- Crash Course Chemistry: Crash Course is a division of Complexly and videos are free to stream for educational purposes.
- Sci Show is a division of Complexly and videos are free to stream for educational purposes.

Answers

Exercise 3.3.1

1 x 10^5 cm = 1 km (or 100,000 cm = 1 km)

Exercise 3.3.2

10 cL = 1 dL

Exercise 3.3.3

As you go through the course, associate units with the type of measurement they correspond to. This will help you interpret word problems.

- 4.5 g = **mass**. The unit g is grams.
- 1500 mL = **volume**. The unit mL is milliliters.
- 225 ms = **time**. The unit ms is milliseconds.

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3.4: Converting Units

Learning Objective

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

You may already know how to convert common units that you use in your everyday life. Continue to use that process when it works for you! However, in this course you may encounter unfamiliar problems that are more difficult to figure out. In that case, I recommend using the unit analysis method which will be demonstrated here.

Consider a simple example: how many feet are there in 4 yards? Many people can answer that there are 12 feet in 4 yards. If there are 3 feet in 1 yard, and there are 4 yards, then there are 4 × 3 = 12 feet in 4 yards.

To solve a problem like this one using unit analysis, we use equalities which are pairs of values that are equal to one another. In this case, 1 yard (yd) equals 3 feet (ft):

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

Conversion Factors

In the unit analysis problem solving method we will use equalities as conversion factors by rewriting them as fractions equal to one. For example:

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

or

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

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

Unit Analysis

To use the unit analysis method, start with the original quantity. This will generally be a measurement with a simple (not compound) unit that is not already part of a relationship or equality. In the problem above our original quantity was 4 yd.

Next, multiply this quantity by a conversion factor that is equal to 1 (so that it does not change the value of the quantity). Remember that conversion factors are fractions made up of two values that are equal. Eventually we want the unit of yards to divide out leaving only a unit of feet for the answer, so we set up the conversion factor so that yards is on the bottom where it will cancel the original unit and feet is on the top. That will become the unit for the answer.

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

You'll follow this pattern when solving other unit analysis problems in the future. The unit for the original value will go on the bottom of the conversion factor and the desired unit will go on the top of the conversion factor.

The 4 yd term can be thought of as 4yd/1; that is, it can be thought of as a fraction with 1 in the denominator. We are essentially multiplying fractions. If the same thing appears in the numerator and denominator of a fraction, they cancel. (You may have seen this with variables in algebra where an x that appears on both the top and bottom of a fraction divides out. It is the same with units). In this case, what cancels is the unit *yard*:

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

That is all that we can cancel. Now, multiply the fractions to get the final answer. To multiply fractions start with the number at the far left. Multiply by any values on the top of fractions and divide by any values on the bottom of fractions:

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

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

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

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

Now we construct a conversion factor by dividing one quantity into both sides. Can you figureout which version of the conversion factor is needed for this problem?

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

or

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

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

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

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

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

Example 3.4.1

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

Solution

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

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

b. We will use the fact that 1,000 mm = 1 m. Of the two possible conversion factors, the appropriate one has the mm unit in the denominator:

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

Applying this conversion factor, we get:

$$555\,mm \times \frac{1m}{1000mm} = 0.555\,m$$

Exercise 3.4.1

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

Answer a available at the end of this section.

Answer b available at the end of this section.

Exercise 3.4.2

How many milliliters are in 607.8 kL?

Answer available at the end of this section.

Chemistry is Everywhere: The Gimli Glider

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



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

The 767 took off from Montreal on its way to Ottawa, ultimately heading for Edmonton, Canada. About halfway through the flight, all the engines on the plane began to shut down because of a lack of fuel. When the final engine cut off, all electricity (which was generated by the engines) was lost; the plane became, essentially, a powerless glider. Captain Robert Pearson was an experienced glider pilot, although he had never flown a glider the size of a 767. First Officer Maurice Quintal quickly determined that the aircraft would not be able make it to Winnipeg, the next large airport. He suggested his old Royal Air

Force base at Gimli Station, one of whose runways was still being used as a community airport. Between the efforts of the pilots and the flight crew, they managed to get the airplane safely on the ground (although with buckled landing gear) and all passengers off safely.

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

Key Takeaways

- Units can be converted to other units using the proper conversion factors.
- Conversion factors are constructed from equalities that relate two different units.
- Conversions can be a single step or multi-step.
- Unit conversion is a powerful mathematical technique in chemistry that must be mastered.
- Exact numbers do not affect the determination of significant figures.

Answers

Exercise 3.4.1

6.708×10^{-5} L

Exercise 3.4.1

5.68×10^{-2} km

Exercise 3.4.2

6.078×10^{8} mL

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3.5: Solving Multistep Conversion Problems

Learning Objective

- Convert a measurement into a different unit in cases where an equality between the two units is not available; i.e. perform conversions that take two or more steps.

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 inches. 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 miles which is in the same measurement system as inches. We know that 0.6214 mi =1 km.

Then we convert miles to feet, remembering that $1\ \text{mi} = 5280\ \text{ft}$. Finally, we convert feet to inches using 1 ft. = 12 in.

Concept Map

$$\text{kilometers} \rightarrow \text{miles} \rightarrow \text{feet} \rightarrow \text{inches}$$

Calculation

$$54.7\ \cancel{\text{km}} \times \frac{0.6214\ \cancel{\text{mi}}}{1\ \cancel{\text{km}}} \times \frac{5280\ \cancel{\text{ft.}}}{1\cancel{\text{mi}}} \times \frac{12\ \cancel{\text{in.}}}{1\cancel{\text{ft.}}} = 2,153,643\ \text{in.} \tag{3.5.1}$$

$$= 2.15 \times 10^6\ \text{in.} \tag{3.5.2}$$

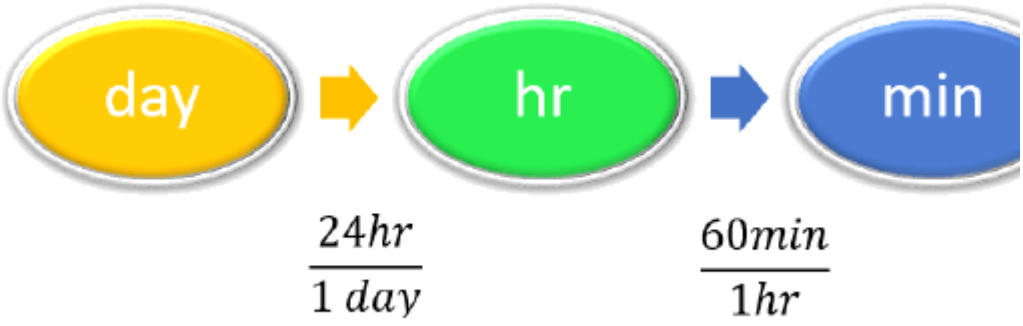
In each step, the previous unit is canceled and the next unit in the sequence is produced, each successive unit canceling out until only the unit needed in the answer is left.

Example $3.5.1$: Unit Conversion

How many seconds are in a day?

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given"information and what the problem is asking you to "find."	Given: 1 day Find: s

Steps for Problem Solving	Unit Conversion
List other known quantities	1 day = 24 hours 1 hour = 60 minutes 1 minute = 60 seconds
Prepare a concept map	
Calculate	$1\text{ d} \times \frac{24\text{ hr}}{1\text{ d}} \times \frac{60\text{ min}}{1\text{ hr}} \times \frac{60\text{ s}}{1\text{ min}} = 86,400\text{ s}$

Example 3.5.1

An IV bag contains 250 mL of solution. How many quarts (qt.) of solution are in the bag? Note: 1 L = 1.06 qt.

Solution

When using the unit analysis method, if you are solving for a simple unit such as quarts you should start with the measured or given number that has a simple unit. A simple unit has only one part such as mL. A compound unit has two parts such as g/mL (grams per milliliter). Do not start with an equality like 1 L = 1.06 qt. Equalities will be used to make conversions later in the problem.

$$250\ mL \tag{3.5.3}$$

When you make a conversion factor put the unit from the top of the previous step on the bottom of the fraction.

$$250\ mL \times \frac{}{mL} \tag{3.5.4}$$

Consider whether you have the information needed to convert directly to the unit for the answer (an equality that relates the unit that you have to the unit that you want). In this case we do not have that information. However, we can use the metric system to convert from mL to L and then use the note provided in the problem to convert from L to qt.

$$250\ mL \times \frac{L}{mL} \times \frac{qt.}{L} \tag{3.5.5}$$

It is helpful to plan the units out first (either one step at a time or for all of them at once) and then fill in the numbers. Making sure that the units are in the right place will help you to know whether to multiply or divide by the numbers in the equalities.

Fill in the numbers. Make sure that each conversion factor (fraction) is made out of two values that are equal. Using the metric system, 1000 mL = 1 L. Using the information provided in the problem, 1 L = 1.06 qt. Match the numbers up with the correct units.

$$250\ mL \times \frac{1\ L}{1000\ mL} \times \frac{1.06\ qt.}{1\ L} \tag{3.5.6}$$

To solve, start with the given number, multiply it by values in numerators (tops of fractions) and divide by values in denominators (bottoms of fractions). It is not necessary to multiply or divide by 1 in the calculator since that does not change the value of a number. For this problem you could type: 250 ÷ 1000 x 1.06

$$250\ mL \times \frac{1\ L}{1000\ mL} \times \frac{1.06\ qt.}{1\ L} = 0.265\ qt. \tag{3.5.7}$$

The given value, 250 mL has only two significant figures. The metric system equality, 1000 mL = 1 L, is exact because those values are in the same measurement system. The other equality, 1 L = 1.06 qt., is not exact because L and qt. are from two different systems. Based on 1.06 that relationship has three significant figures. The value with the least significant figures was the 250 mL with two significant figures, so the answer should be rounded to two significant figures: **0.27 qt.**

Exercise 3.5.1

Perform each conversion in a multistep calculation.

a. 65.0 hours to milliseconds (note: the metric system can be used to convert seconds to milliseconds)
b. 15 inches to millimeters
c. 1200 feet to kilometers

Answer a:

$$65.0\ h \times \frac{60\ min}{1\ h} \times \frac{60\ s}{1\ min} \times \frac{1000\ ms}{1\ s} = 2.34 \times 10^8\ ms$$

Answer b:

380 mm

Answer c:

0.37 km

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. This image was released by the National Cancer Institute, an agency part of the National Institutes of Health. Image used with permission (Public Domain; Rhoda Baer (Photographer) via NIH).

Summary

In multistep 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.

Contributors

- Anonymous

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3.6: Density

Learning Objectives

- Calculate density, mass, or volume when given 2 of these three variables.
- Identify what units are required for the density equation.
- Review metric conversions.
- Compare densities of different chemical substances.
- Compare any chemical substance's density to the density of water.

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

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{3.6.1}$$

$$d = \frac{m}{V} \tag{3.6.2}$$

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. That means that density has a compound unit; it always has multiple parts. 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 3.6.1. The density units used in the table are equal to one another because 1 mL = 1 cm^3. Usually mL is used for liquids and cm^3 is used for solids, but this can vary based on how the volume was measured. Note: Densities change with temperature so the values in the table will vary at high or low temperatures.

Table 3.6.1: Densities of Some Common Substances

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



Figure 3.6.1: Separatory Funnel containing oil and water (with green dye) to display density differences. The less dense substance floats. (CC BY-SA 3.0; PRHaney via Wikipedia).

Density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3. How can you determine what mass of aluminum you have without a scale or balance? You can use the volume and density to calculate it. If you multiply the given volume by the known density from (Table 3.6.1), the volume units will divide out and leave you with mass units, telling you the mass of the sample.

There are two ways of knowing that multiplication is needed in this case:

If using unit analysis, start with the value with a simple unit. In this case we will consider the volume unit, cm^3, to be simpler than the density unit, g/mL or g/cm^3.

$$7.88\ cm^3 \times \frac{\quad}{cm^3} \tag{3.6.3}$$

To use the density as a conversion factor, split it into two parts. For aluminum the density tells us that 2.7 g = 1 mL or 2.7 g = 1 cm^3. These relationships are not always true, they are specific to aluminum. Density is needed to relate mass and volume for a substance.

$$7.88\ cm^3 \times \frac{7.88\ g}{1\ cm^3} = 21\ g$$

If using algebra, start with Equation 3.6.1

$$\text{density} = \frac{m}{V}$$

and insert the relevant numbers

$$\frac{2.7g}{cm^3} = \frac{m}{7.88\ cm^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

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

Example 3.6.1: Mercury

What is the mass of 44.6 mL of mercury? The density of mercury is 13.6 g/mL.

Solution

Option 1: Unit Analysis

Start with the given number. Set up a conversion factor and put the unit from the given number on the bottom. (Leave room to put a number in front of the unit.)

$44.6\ mL \times \frac{}{mL}$

Consider what unit will go on the top of the conversion factor. We are asked to solve for mass and the density provides a relationship between volume and mass. To use the density, which has a compound unit, we split the unit into two lines or, in other words, we make the compound unit into a conversion factor. For mercury 13.6 g = 1 mL (this is only true for mercury. You cannot use 13.6 g = 1 mL in other problems).

$44.6\ mL \times \frac{13.6\ g}{1\ mL}$

The diagonal division sign in 13.6 g/mL becomes the horizontal division sign in $\frac{13.6\ g}{1\ mL}$. The number, 13.6, goes with g because 13.6 g/mL means 13.6 grams per milliliter or 13.6 grams per 1 mL.

Calculate the answer:

$44.6\ mL \times \frac{13.6\ g}{1\ mL} = 606.56\ g$ or **607 g** with three significant figures.

Option 2: Algebra

If you need to review how to isolate variables, please see the appendix of this textbook.

Use the definition of density (Equation 3.6.1)

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

$$\frac{13.6g}{mL} = \frac{m}{44.6\,mL}$$

Remember to cross multiply here in order to isolate variable. Then, report answer with correct units.

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

The mass of the mercury is **607 g**.

Exercise 3.6.1

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

Answer available at the end of this section.

Interactive Element

Explore how different masses and volumes affect density.

- Does changing an object's mass or volume affect its density?
- How does density relate to whether an object sinks or floats in water?

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

Example 3.6.2: Wine Cork

A cork stopper from a bottle of wine has a mass of 0.133 oz. If the density of cork is 0.22 g/mL, what is the volume of the cork in milliliters? Note: 1 lb. = 16 oz. and 1 lb. = 453.6 g. *Regardless of the method that is used, you should still be able to obtain the same (and correct) answer.*

Solution

If you plan to use unit analysis to solve the problem, start by identifying the number with the simple unit.

$$0.133\ oz. \times \frac{}{oz.}$$

Consider whether the starting unit matches half of the density unit. It does not, so we need to do a conversion before using the density. Since the starting value is a mass we should convert it to the same mass unit found in the density, grams.

$$0.133\ oz. \times \frac{1\ lb.}{16\ oz.} \times \frac{453.6\ g}{1\ lb.}$$

Now we can use the density to convert from mass to volume. Use the density as a conversion factor by splitting it into two parts.

$$0.133\ oz. \times \frac{1\ lb.}{16\ oz.} \times \frac{453.6\ g}{1\ lb.} \times \frac{1\ mL}{0.22\ g}$$

Keep following the pattern with the units (where the unit from the top of the previous step is written on the bottom of the current step). This ensures that everything except the unit that is needed for the answer divides out. A density of 0.22 g/mL means that 1 mL of cork weighs 0.22 g. The number goes with the unit that it is next to; the other unit has a value of 1 unless something else is specified.

$$0.133\ oz. \times \frac{1\ lb.}{16\ oz.} \times \frac{453.6\ g}{1\ lb.} \times \frac{1\ mL}{0.22\ g} = \mathbf{17\ mL}$$

Because the density only had 2 significant figures, the answer can only have 2 significant figures. Densities are measured, not exact.

Exercise 3.6.2

What is the volume in liters of 3780 g of gold? The density of gold is 19.3 g/cm^3. Note: 1 mL = 1 cm^3

Answer available at the end of this section.

Exercise 3.6.3

What is the volume in liters of 5.0 lbs. of iron? Iron has a density of 7.87 g/mL or g/cm^3, 1 lb. = 453.6 g, and 1 cm^3 = 1 mL.

Answer available at the end of this section.

Summary

- Density is the relationship between mass and volume for a substance
- Densities are different for different substances. To convert between mass and volume for a sample its density must be known.
- Density problems can be solved using unit analysis, algebra, or cross multiplying.

Contributors and Attributions

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Answers

Exercise 3.6.1

Use the density value for iron from Table 3.6.1

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

$$\frac{7.87g}{cm^3} = \frac{m}{25.0\,cm^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

$$25.0\,\cancel{cm^3} \times \frac{7.87\,g}{\cancel{cm^3}} = 197g \text{ of iron}$$

Exercise 3.6.2

0.196 L

Exercise 3.6.3

0.29 L

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4: STRUCTURE OF MOLECULAR COMPOUNDS

You often encounter compounds in your daily life. Water, carbon dioxide, and sugar are all compounds. You already learned about ionic compounds in Chapter 1. In this chapter you will learn about the other main category of compounds: molecular compounds. We'll focus on how the atoms are bonded together, the shape of the molecules, and whether or not electrons are shared equally between atoms.

4.1: Introduction to Molecular Compounds

4.2: Drawing Lewis Structures

4.3: Molecular Shapes- The VSEPR Theory

4.4: Unequal Sharing- Polar Covalent Bonds

4.5: Shapes and Properties- Polar and Nonpolar Molecules

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4.1: Introduction to Molecular Compounds

Learning Objectives

- To contrast ionic compounds and molecular compounds.

Molecular Compounds vs Ionic Compounds

In general, molecular compounds are composed of only nonmetals and ionic compounds are composed of both metals and nonmetals, but there are exceptions such as ammonium chloride, NH_4Cl, which contains only nonmetals but is an ionic compound. Therefore, it is helpful to have a deeper understanding of the differences between these two types of compounds. Table 4.1.1 contrasts these categories.

Table 4.1.1: Ionic vs Molecular Compounds

Property	Ionic Compounds	Molecular Compounds
Typical elements	Both metals and nonmetals (with some exceptions)	Only nonmetals
Type of bond	Ionic bonds; may also contain covalent bonds	Covalent bonds
Smallest unit	Formula unit (lowest ratio of ions)	Molecule

The octet rule states that atoms tend to gain, lose, or share valence electrons in order to get a stable arrangement of eight valence electrons. Ionic bonds and covalent bonds represent two methods for following this rule. When an atom gains or loses valence electrons it then becomes an ion and has the ability to participate in ionic bonds. When an atom shares valence electrons with another atom, a covalent bond results.

Ionic compounds form a lattice structure because the charged particles are attracted to all nearby ions with an opposite charge. Figure 4.1.1(a) shows a sample lattice structure where the smaller green spheres represent sodium ions and the larger red spheres represent chloride ions. Each ion is surrounded by ions of opposite charge on all sides. We cannot identify certain pairs of ions that belong together as partners, but there is a 1:1 ratio of Na^+ and Cl^- ions resulting in a formula unit of NaCl.

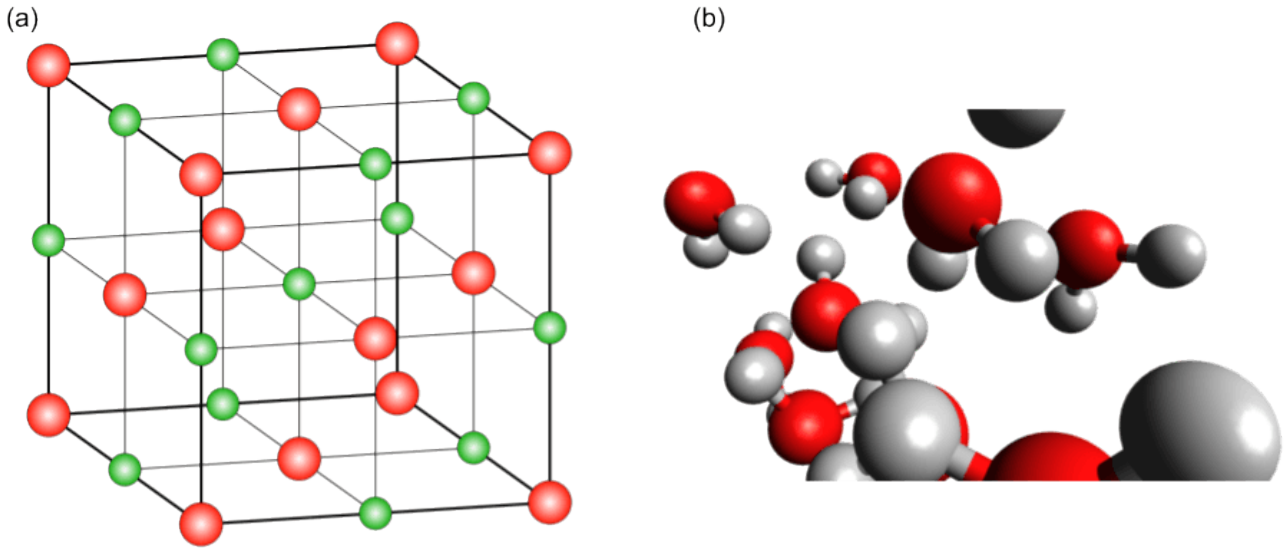


Figure 4.1.1: (a) A lattice structure as found in ionic compounds such as NaCl. (b) A collection of water molecules. Prolineserver, CC BY-SA 3.0, via Wikimedia Commons and Lizard125, CC BY-SA 4.0, via Wikimedia Commons

On the other hand, Figure $4.1.1$(b) shows a collection of water molecules in the gas phase (water vapor). In molecules each atom is bonded to one or more other specific atoms. The lines in the figure represent covalent bonds which are shared electrons. If we look closely at the image we can identify specific H_2O molecules; each H atom "belongs to" a certain O atom. The shared valence electrons are being shared between specific atoms.

Summary

- Molecular compounds contain covalent bonds where atoms share valence electrons in order to satisfy the octet rule.
- Covalent bonds connect specific atoms into molecules.

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4.2: Drawing Lewis Structures

Learning Objectives

- To draw Lewis structures for molecular compounds

Representing Valence Electrons as Bonds and Lone Pairs

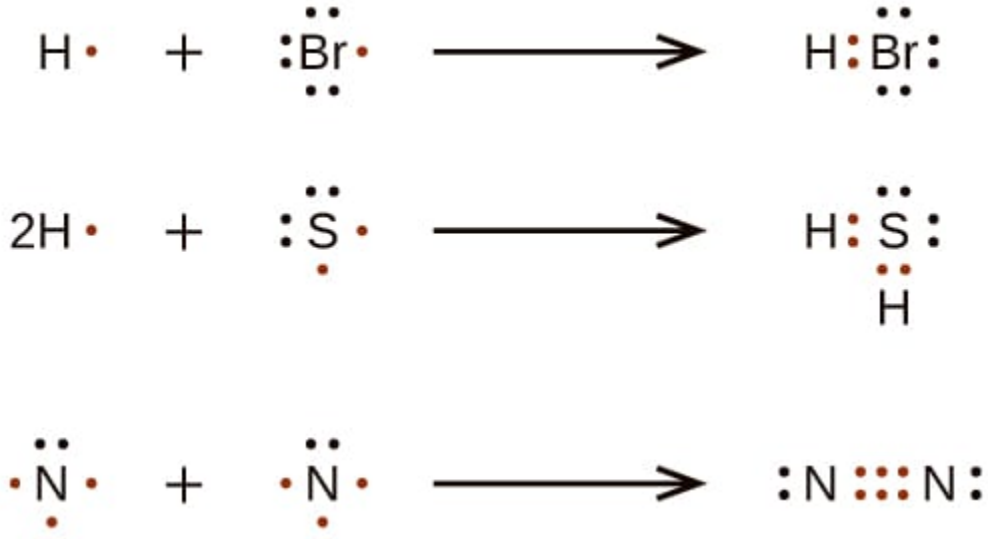
Lewis structures are used to show the arrangement of valence electrons in a molecule.

Dots are used to show valence electrons that belong to only one atom. Lines are used to represent a pair of shared electrons. Core electrons (any electrons that are not valence electrons) are not shown. Electrons are almost always arranged in pairs, either as a pair of dots or as a line which is worth two electrons.

Lewis structures are used to clearly show whether valence electrons are shared or unshared. They are optimized for a 2D paper or screen and do not clearly show the correct angles between atoms or the overall 3D shape of a molecule; 3D arrangements will be covered in a later section.

Drawing Lewis Structures

For very simple molecules and molecular ions, we can write the Lewis structures by merely pairing up the unpaired electrons on the constituent atoms. In the examples below individual atoms are shown with the correct number of valence electrons. Put one valence electron on each side of the symbol before putting any of the dots in pairs. The unpaired electrons from each atom can be shared to create a bond.



For larger molecules the method above is tedious and it can be complicated to determine where multiple bonds (double or triple bonds) should be. For organic compounds an easy method to determine the Lewis structure is to start at the edges of the molecule and fill in the number of bonds and lone pairs for each atom so that they follow the patterns shown in

Typical number of bonds and lone pairs for nonmetals

Location of nonmetal on the Periodic Table	# of Covalent Bonds	# of Lone Pairs
Hydrogen	1	0
Group 14	4	0

Location of nonmetal on the Periodic Table	# of Covalent Bonds	# of Lone Pairs
Group 15	3	1
Group 16	2	2
Group 17	1	3

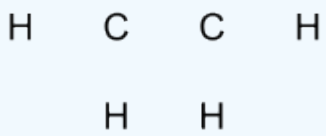
The nonmetals in Group 18 already have 8 valence electrons, so they do not usually form any bonds. The patterns for the Group 14 - 18 nonmetals result in an octet of electrons because each covalent bond consists of two electrons and each lone pair is also two electrons. Hydrogen is an exception; it is satisfied with only two electrons because it is so small. There are other exceptions to the octet rule, but they are not covered here.

When given the molecular formula, which just states the number of atoms of each element such as H_2O or C_2H_4, then the element listed first goes in the middle, unless it is hydrogen. Hydrogen is never found in the middle of a molecule because it is not able to form more than one bond. Lay out the atoms. For small molecules they tend to be arranged symmetrically.

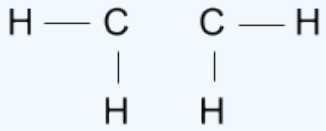
Example 4.2.1

Draw the Lewis structure for C_2H_4.

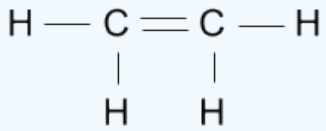
Solution



1. Lay out the atoms. Carbon goes in the middle because it is listed first (and because hydrogen can never go in the middle). Atoms tend to be arranged symmetrically for small molecules, unless you are told otherwise. Therefore we give two hydrogen atoms to each carbon atom. All atoms can be arranged at 90 or 180 degree angles. It does not matter for Lewis structures if you chose to put the H atoms above, below, or to the sides of the carbon atoms.



2. Work from the outside in. Hydrogen is satisfied with one bond and zero lone pairs.



3. Each carbon should have four bond (and zero lone pairs). The hydrogen atoms are full and bonds cannot point towards the empty space above the carbon atoms because the lines represent electrons that are shared by two atoms. Therefore, two bonds are needed between the two carbon atoms.

The bonds represent shared electrons, so each atom in the Lewis structure has the correct bonding pattern: one bond and zero lone pairs for hydrogen atoms and four bonds and zero lone pairs for carbon atoms.

Exercise 4.2.1

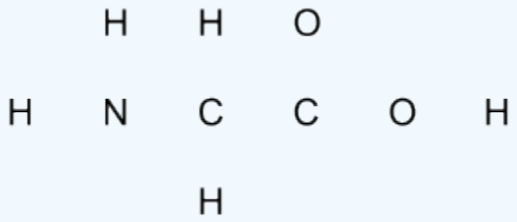
Draw the Lewis structure for CH_2O.

Answer available at the end of this section.

For larger structures you may be provided with the layout of the atoms. Follow the same procedure: start from the outside filling in bonds and lone pairs according to the expected pattern for each element.

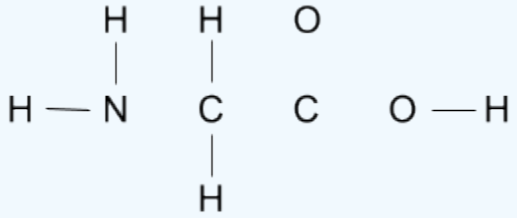
Example 4.2.2

Fill in the bonds and lone pairs to complete the Lewis structure. Note: The upper O atom is bonded only to C, not to H.



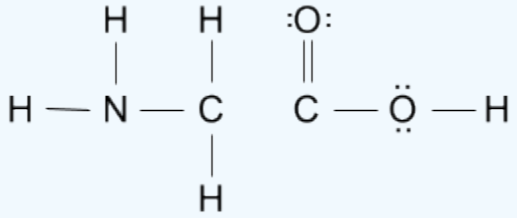
Solution

Each H atom can only have one bond, so you can start by filling those in.

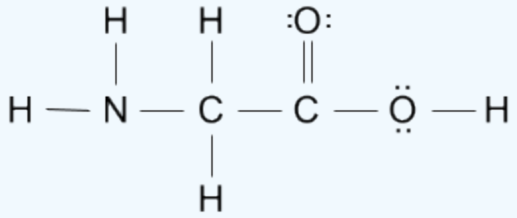


Next do the N or either of the O atoms. Leave the C atoms for later since they are in the middle and more information about their neighboring atoms is needed before completing them.

Nitrogen should have three bonds and one lone pair. Each oxygen atom should have two bonds and two lone pairs. The upper oxygen atom is only bonded to one other atom so it needs a double bond. The oxygen atom on the right is between two atoms so it needs two single bonds. Lone pairs must go on an empty side; they should never be drawn between two atoms.



Each carbon atom should have four bonds and zero lone pairs.



For polyatomic ions and compounds that have exceptions to the octet rule, it is helpful to follow the step-by-step procedure outlined here:

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

Let us determine the Lewis structures of OF_2 and HCN as examples in following this procedure:

1. Determine the total number of valence (outer shell) electrons in the molecule or ion. For a molecule, we add the number of valence electrons (use the main group number) on each atom in the molecule. This is the total number of electrons that must be used in the Lewis structure.

$$O + 2\,(F) = OF_2$$

$$6e^- + (2 \times 7e^-) = 20e^-$$

$$H + C + N = HCN$$

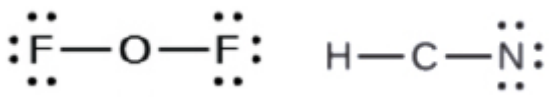
$$1e^- + 4e^- + 5e^- = 10e^-$$

2. Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom and connecting each atom to the central atom with a single (one electron pair) bond. Note that H and F can only form one bond, and are always on the periphery rather than the central atom.



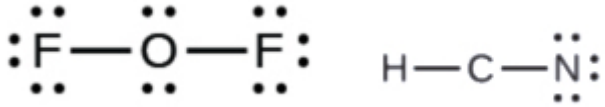
3. Distribute the remaining electrons as lone pairs on the terminal atoms (except hydrogen) to complete their valence shells with an octet of electrons.

- In OF_2, six electrons are placed on each F.
- In HCN, six electrons placed on N



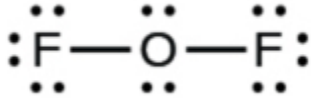
4. Place all remaining electrons on the central atom.

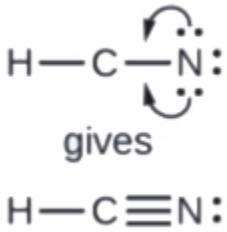
- In OF_2, 4 electrons are placed on O.
- In HCN: no electrons remain (the total valence of $10e^-$ is reached) so nothing changes.



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

- In OF_2, each atom has an octet as drawn, so nothing changes.
- In HCN, form two more C–N bonds



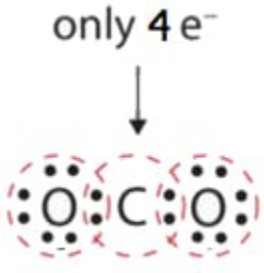


H dash C dash N with 6 dots around the N. 4 of the dots are pointing to the dash connecting the C and the N. This turns the formula into H dash C and three lines connecting the C and N together, with two dots on the outside of the N.

Finally, check to see if the total number of valence electrons are present in the Lewis structure. And then, inspect if the H atom has 2 electrons surrounding it and if each of the main group atoms is surrounded by 8 electrons.

MULTIPLE BONDS

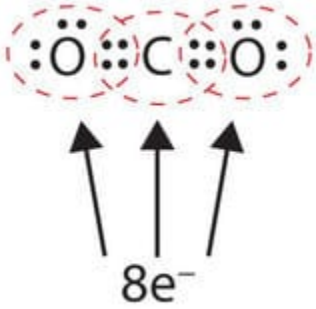
In many molecules, the octet rule would not be satisfied if each pair of bonded atoms shares only two electrons. Review HCN in Step 5 above. Another example is carbon dioxide (CO_2). CO_2 has a total valence of $4e^- + (2 \times 6e^-) = 16e^-$. Following steps 1 to 4, we draw the following:



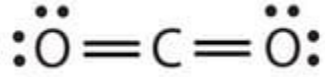
This does not give the carbon atom a complete octet; only four electrons are in its valence shell. This arrangement of shared electrons is far from satisfactory.



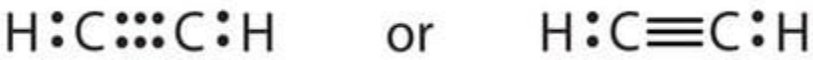
In this case, more than one pair of electrons must be shared between two atoms for both atoms to have an octet. A second electron pair from each oxygen atom must be shared with the central carbon atom shown by the arrows above. A lone pair from each O must be converted into a bonding pair of electrons.



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:



Some molecules contain triple bonds (like HCN, shown above). Triple bonds are covalent bonds in which *three* pairs of electrons are shared by two atoms. Another compound that has a triple bond is acetylene (C_2H_2), whose Lewis diagram is as follows:



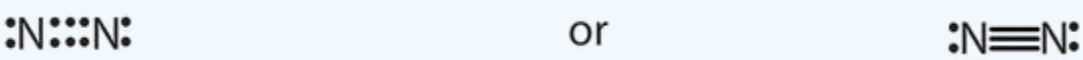
Example 4.2.1

Draw the Lewis diagram for each molecule.

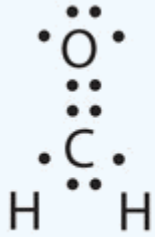
1. N_2
2. CH_2O (The carbon atom is the central atom.) One application of CH_2O, also called formaldehyde, is the preservation of biological specimens. Aqueous solutions of CH_2O are called formalin and have a sharp, characteristic (pungent) odor.

Solution

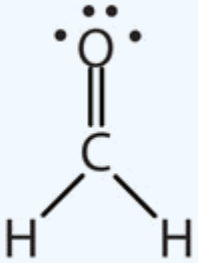
1. The total number of electrons is 2 x 5 = 10 electrons. The bond between the two nitrogen atoms is a triple bond. The Lewis diagram for $N_2$is as follows:



2. The total number of electrons is 4 x 2(1) + 6 = 12 electrons. In CH_2O, the central atom is surrounded by two different types of atoms. The Lewis diagram that fills each atom's valence electron shell is as follows:



or



Exercise 4.2.1

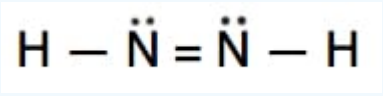
Draw the Lewis diagram for each molecule.

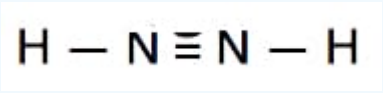
a. O_2
b. C_2H_4

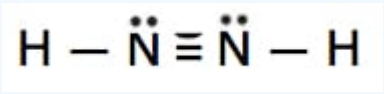
Answer a: available at the end of this section.

Example 4.2.2

Which is the correct Lewis structure for N_2H_2?

A. 

B. 

C. 

Solution

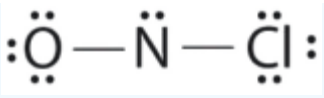
Lewis structure A is the correct answer. It has a total of (2 x $5e^-$) + (2 x $1e^-$) = $12e^-$. Each of the N atoms satisfy the octet requirement and the H atoms follow the duet rule.

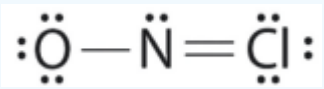
Structure B is electron deficient. It has only $10e^-$ instead of 12.

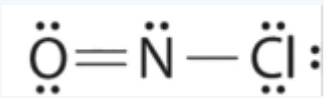
Structure C has 14 (2 extra) electrons. The N atoms do not satisfy the octet.

4.2.2Exerc

Which is the correct Lewis structure for NOCl?

A. 

B. 

C. 

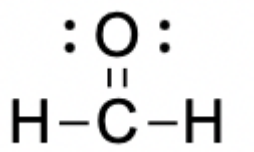
Answer available at the end of this section.

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.

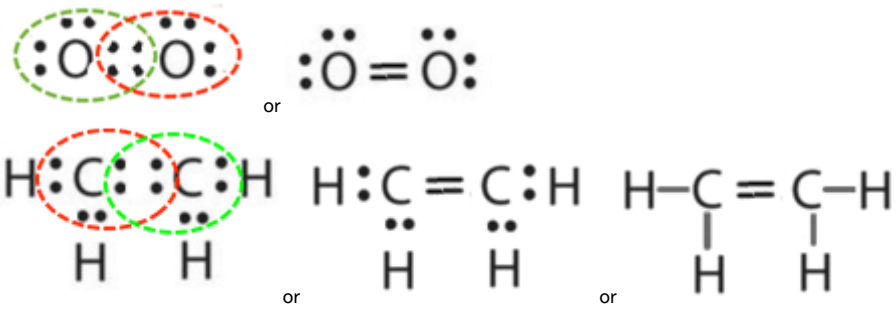
Answers

Exercise 4.2.1



The carbon atom goes in the middle because it is listed first. Each hydrogen atom gets one bond. The oxygen atom gets two bonds. They both have to go to the carbon atom (as a double bond) because the oxygen atom is only bonded to one other atom. The oxygen atom also has two lone pairs.

Exercise 4.2.1



Structure A violates the octet rule; N is surrounded by only $6e^-$.

Structure B violates the octet rule; Cl has $10e^-$ around it. Furthermore, there are a total of $20e^-$ instead of $18e^-$.

Structure C is the correct structure. It has a total of $6e^- + 5e^- + 7e^- = 18e^-$. Each atom is surrounded by 8 electrons (octet rule).

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4.3: Molecular Shapes- The VSEPR Theory

Learning Objective

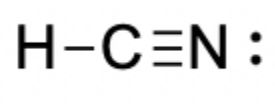
- Determine the shape of simple molecules and polyatomic ions.

Molecules have shapes. There is an abundance of experimental evidence for that from their physical properties to their chemical reactivity. Small molecules, molecules with a single central atom, have shapes that can be easily predicted. Valence shell electron pair repulsion (VSEPR) is the theory of molecular shapes. It says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from each other as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond (single, double, or triple) and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Example 4.3.1

1. How many electron groups are there around the carbon atom in HCN?
2. How many of the groups are bonding groups and how many are nonbonding groups?

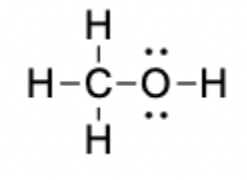


Solution

1. Look at the electrons immediately next to the C atom. It has a single bond and a triple bond. This is **two electron groups**. Note 1: Carbon always has four bonds, but it does not always have four electron groups. Double and triple bonds each count as one electron group. Note 2: The lone pair on the N atom does not count because the question asked about the carbon atom.
2. In HCN carbon has **two bonding groups**. Any bond (single, double, triple) counts as a bonding group. Lone pairs are nonbonding groups.

Exercise 4.3.1

1. How many electron groups are there around the carbon atom in ethanol? How many of the groups are bonding groups and how many are nonbonding groups?
2. How many electron groups are there around the oxygen atom in ethanol? How many of the groups are bonding groups and how many are nonbonding groups?



Answer available at the end of this section.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible: 180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. An example is carbon dioxide, CO_2:

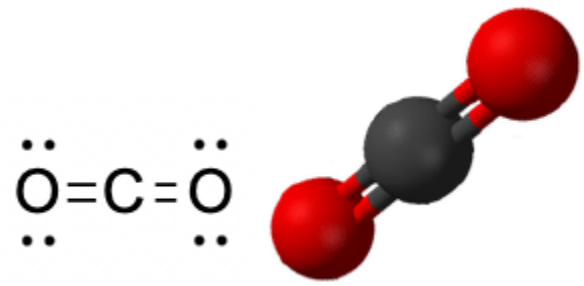


Figure 4.3.1 Carbon dioxide Lewis Structure and 3D ball-and-stick model.

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle: 120° apart and in a plane. The shape of such molecules is trigonal planar. An example is formaldehyde, CH_2O, in Figure 4.3.2. The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

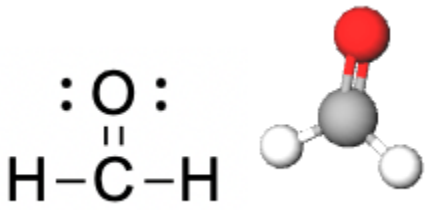


Figure 4.3.2 Formaldehyde, CH_2O, Lewis Structure and 3D ball-and-stick model.

A formaldehyde molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 4.3.3 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

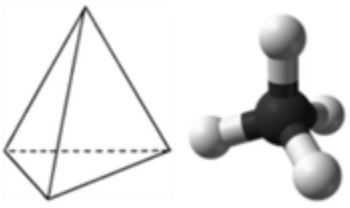


Figure 4.3.3 Tetrahedral structure of methane, CH_4.

You don't need to be an artist to communicate the 3D shapes of molecules. Figure 4.3.4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

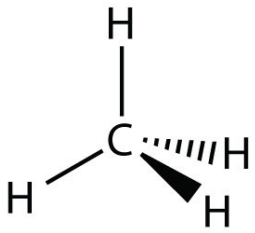


Figure 4.3.4 Using wedges and dashes to depict a 3D structure.

NH_3 is an example of a molecule whose central atom has four electron groups but only three of them are bonded to surrounding atoms.

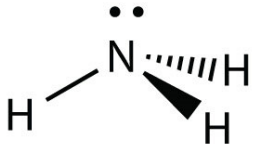


Figure 4.3.5 Ammonia, NH_3, has a trigonal pyrimidal shape.

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is trigonal pyramidal.

H_2O is an example of a molecule whose central atom has four electron groups but only two of them are bonded to surrounding atoms.

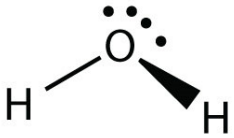


Figure 4.3.6 Lewis structure of water with a bent shape. A wedge shows one bond coming out of the plane and one of the lone pairs is going into the plane of the paper/screen.

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom but only one electron group bonded to another atom is linear because there are only two atoms in the molecule.

Figure 4.3.7 illustrates several representations of the water, ammonia, and methane molecules. The dash-wedge structures clearly show the bonds and chemical symbols. The ball-and-stick models look similar to plastic model kits that are commonly used in chemistry classes. The space-filling models in which atoms are not perfect spheres and touch one another rather than being attached by sticks are the most realistic.

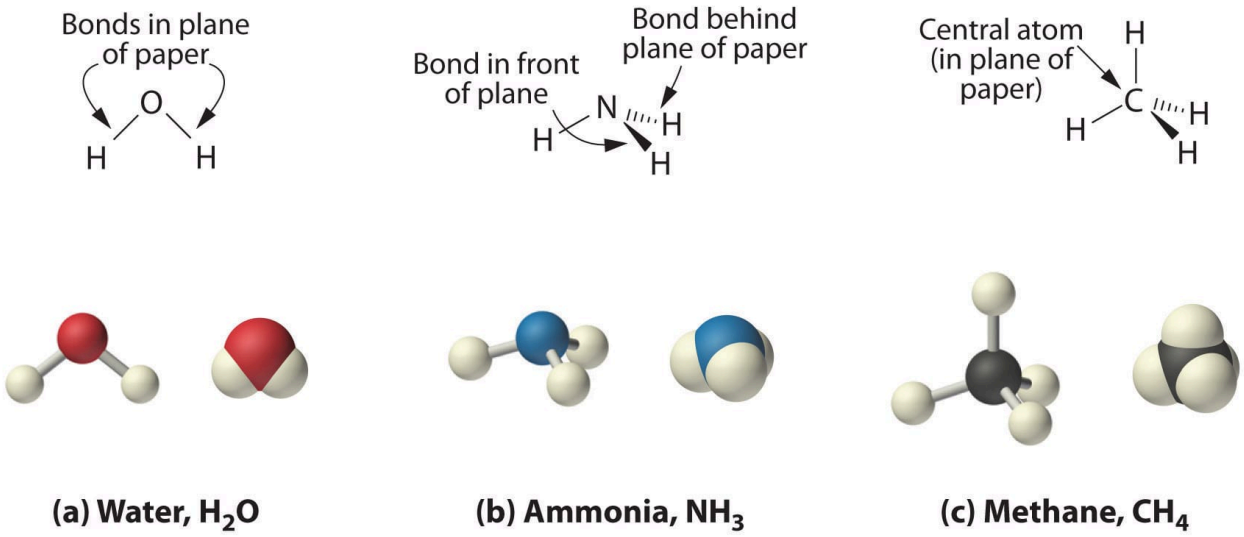


Figure 4.3.7 The three-dimensional structures, ball and stick models,and space filling models of water, ammonia, and methane. (a) Water is a bent shaped molecule, in which all three atoms lie in a plane. (b) In contrast, ammonia has a trigonal pyramidal structure, in which the three hydrogen atoms form the base of the pyramid and the nitrogen atom is at the vertex. (c) The four hydrogen atoms of methane form a tetrahedral structure; the carbon atom lies in the center.

Table 4.3.1 summarizes the shapes of molecules based on their number of electron groups and surrounding atoms. Notice that the bond angles are based on the total number of electron groups, but the molecular shape also takes into account whether the electron groups are atoms or lone pairs.

Table 4.3.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Approximate Bond Angles	Molecular Shape
2	2	0	linear	180°	linear
3	3	0	trigonal planar	120°	trigonal planar
3	2	1	trigonal planar	120°	bent
4	4	0	tetrahedral	109.5°	tetrahedral
4	3	1	tetrahedral	109.5°	trigonal pyramidal
4	2	2	tetrahedral	109.5°	bent

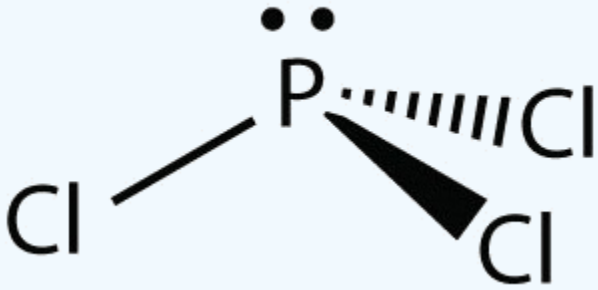
Example 4.3.2:

What is the molecular shape of each molecule?

a. PCl_3
b. NOF

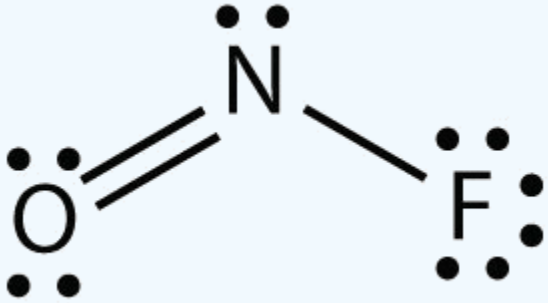
Solution

The first step is to draw the Lewis structure of the molecule. For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

Exercise 4.3.2

What is the molecular shape of CH_2Cl_2?

Answer available at the end of this section.

Exercise 4.3.3

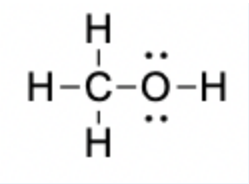
Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule.

Answer available at the end of this section.

When there are multiple central atoms we determine the molecular geometry for each of them individually.

Example 4.3.1

1. What are the bond angles and molecular geometry around the carbon atom in ethanol?
2. What are the bond angles and molecular geometry around the oxygen atom in ethanol?

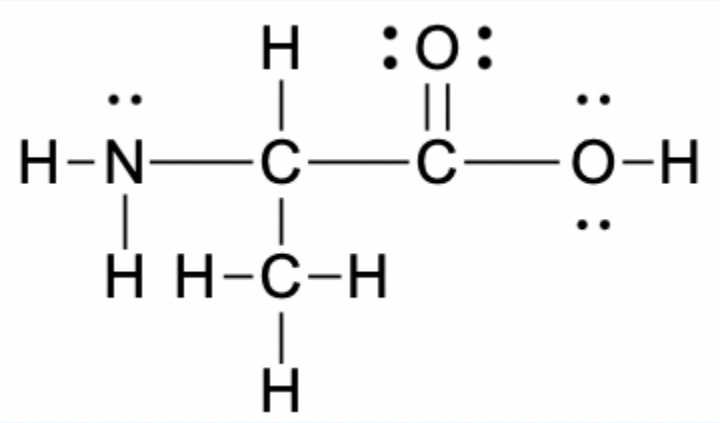


Solution

1. The C atom has four electron groups so it has bond angles of approximately **109.5°**. All four groups are bonding groups so the molecular geometry is **tetrahedral**.
2. The O atom also has four electron groups so it has bond angles of approximately **109.5°**. However, two group are bonding groups and two groups are nonbonding so its molecular geometry is **bent**.

Exercise 4.3.4

The Lewis structure for alanine, an amino acid, is shown below.



List the molecular geometries and bond angles for each of the four atoms in the middle (N, C, C, O).

Answer available at the end of this section.

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

Contributors and Attributions

-
- TextMap: Beginning Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Answers

Exercise 4.3.1

1. The carbon atom has **four electron groups**. **All four are bonding groups.**

2. The oxygen atom also has **four electron groups**, but **two are bonding groups and two are nonbonding groups**. Note: The C-O bond is shared between the carbon and oxygen atoms and it counts as an electron group for both of them.

Exercise 4.3.2

Tetrahedral

Exercise 4.3.3

Trigonal planar about both central C atoms

Exercise 4.3.4

1. N has three single bonds and one lone pair. It's molecular geometry is **trigonal pyramidal** and its angles are approximately **109.5°**.
2. The C on the left has four single bonds. It's molecular geometry is **tetrahedral** and its angles are approximately **109.5°**.
3. The C on the right has two singles bonds and a double bond. That is three bonding electron groups resulting in a molecular geometry of **trigonal planar.** Its bond angles are approximately **120°**.
4. The O atom on the right has two single bonds and two lone pairs so its molecular geometry is **bent** and its bond angles are approximately **109.5°**.

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4.4: Unequal Sharing- Polar Covalent Bonds

Learning Objectives

- Define electronegativity
- Determine the polarity of a covalent bond.

Electronegativity

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur) and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure 4.4.1.

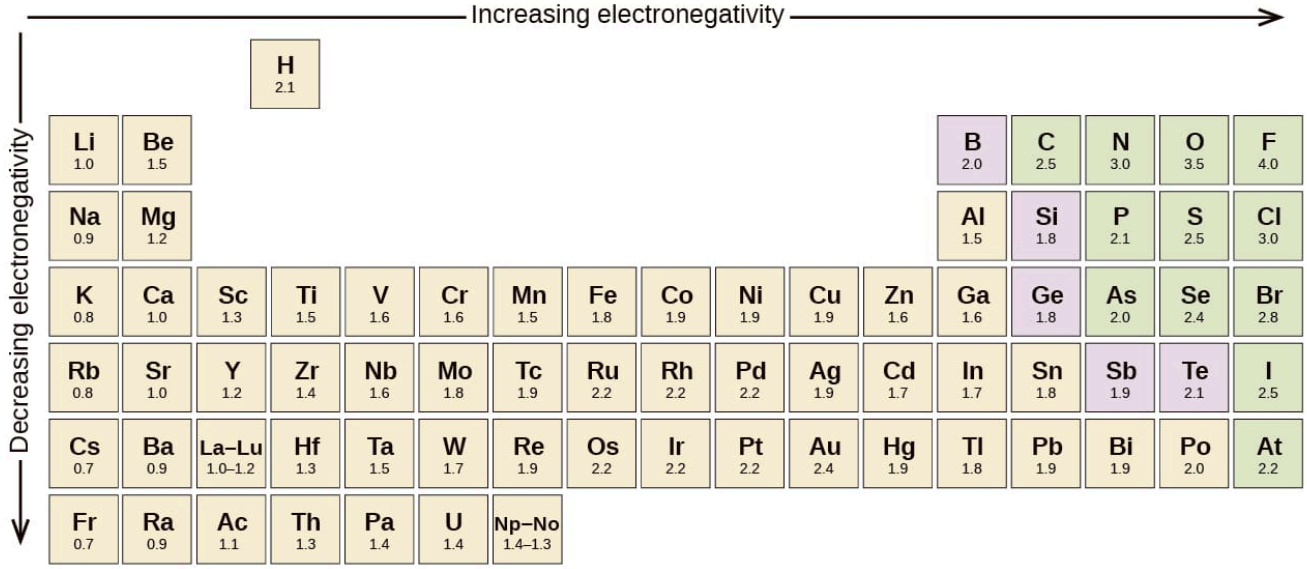


Figure 4.4.1 Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

Bond Polarity

Although polarity is really a continuum, it can be helpful to mark divisions in the continuum in order to talk about bonds with similar characteristics. The polarity of a covalent bond can be classified by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

Electronegativity Difference	Bond Type
0–0.4	nonpolar covalent
0.5–1.7	polar covalent

Electronegativity Difference	Bond Type
1.8 - 1.9	polar covalent if two nonmetals, ionic if a metal with a nonmetal
≥ 2.0	ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

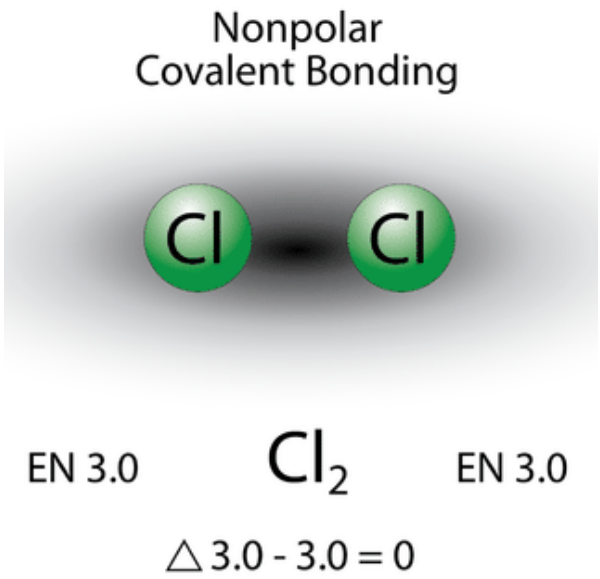


Figure 4.4.2 *A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.*

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the Cl_2 molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta EN = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and about 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

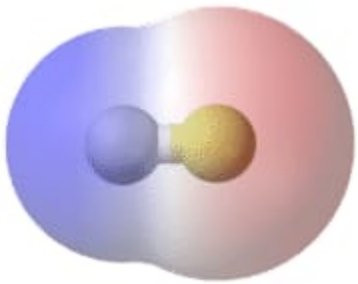


Figure 4.4.3 In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

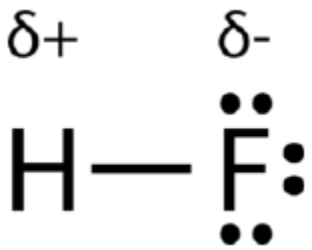


Figure $4.4.4$ Use of δ to indicate partial charge.

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

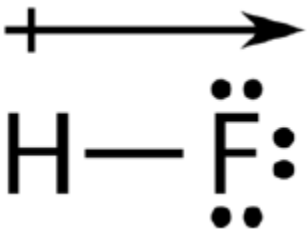


Figure $4.4.5$ Use of crossed arrow to indicate polarity

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Electronegativity differences in bonding using Pauling scale. Using differences in electronegativity to classify bonds as covalent, polar covalent, or ionic.

Example $4.4.1$: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure $4.4.1$, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 – 2.1 = 0.4. Thus we predict that this bond will be non polar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 – 2.1 = 1.4, so we predict that this bond will be definitely polar covalent.

Exercise 4.4.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a available at the end of this section.

Summary

- The type of bond (polar covalent, nonpolar covalent or ionic) between two atoms is determined by the differences in electronegativity.
- For atoms sharing a polar covalent bond, the atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge.

Contributors and Attributions

-
- StackExchange (thomij).
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Answers

Exercise 4.4.1
likely ionic
polar covalent

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4.5: Shapes and Properties- Polar and Nonpolar Molecules

Learning Objective

- Determine if a molecule is polar or nonpolar.

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is frequently useful to look at Lewis structures. **Nonpolar compounds** will be symmetric, meaning all of the sides around the central atom are identical - bonded to the same element with no unshared pairs of electrons. Notice that a tetrahedral molecule such as $\mathrm{CCl_4}$ is nonpolar Figure $(4.5.1$. Another nonpolar molecule shown below is boron trifluoride, BF_3. BF_3 is a trigonal planar molecule and all three peripheral atoms are the same.

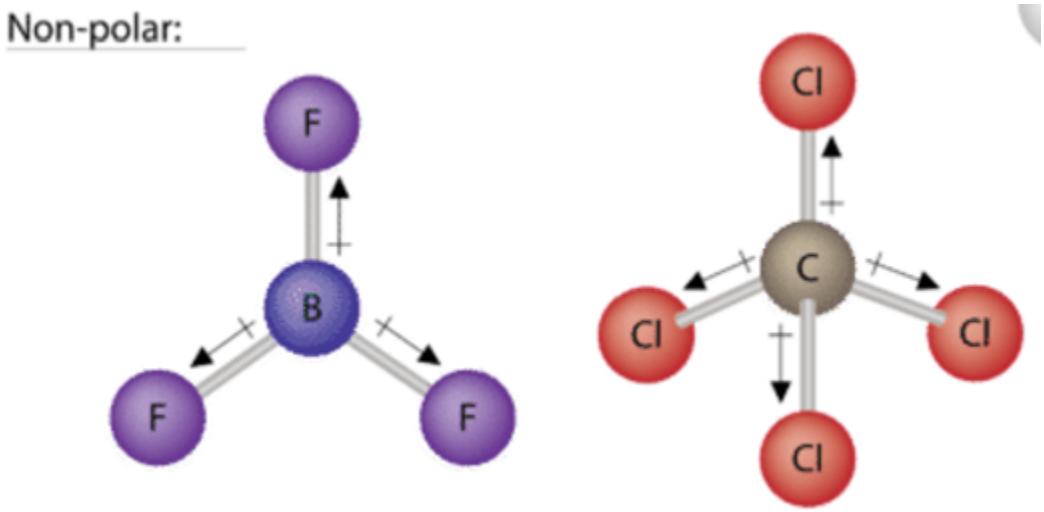
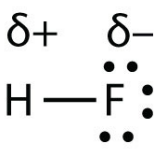


Figure $4.5.1$ Some examples of nonpolar molecules based on molecular geometry (BF_3 and CCl_4).

Polar molecules are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well - as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming you do, you can look at the structure of each one and decide if it is polar or not - *whether or not you know the individual atom electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule.



As mentioned in section 4.7, because the electrons in the bond are nearer to the F atom, this side of the molecule takes on a

partial negative charge, which is represented by δ– (δ is the lowercase Greek letter delta). The other side of the molecule, the H atom, adopts a partial positive charge, which is represented by δ+. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.



Figure 4.5.2 A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

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. The figure below shows 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 this way, they cancel out and the overall molecular polarity of CO_2 is zero.

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

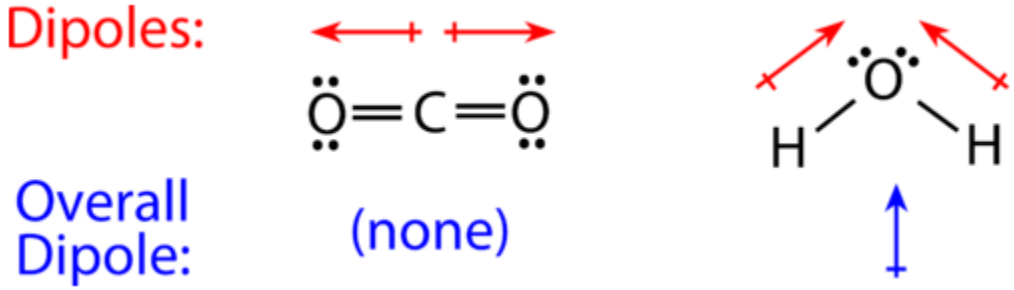


Figure 4.5.3 *The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO_2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.*

Three other polar molecules are shown below with the arrows pointing to the more electron dense atoms. Just like the water molecule, none of the bond moments cancel out.

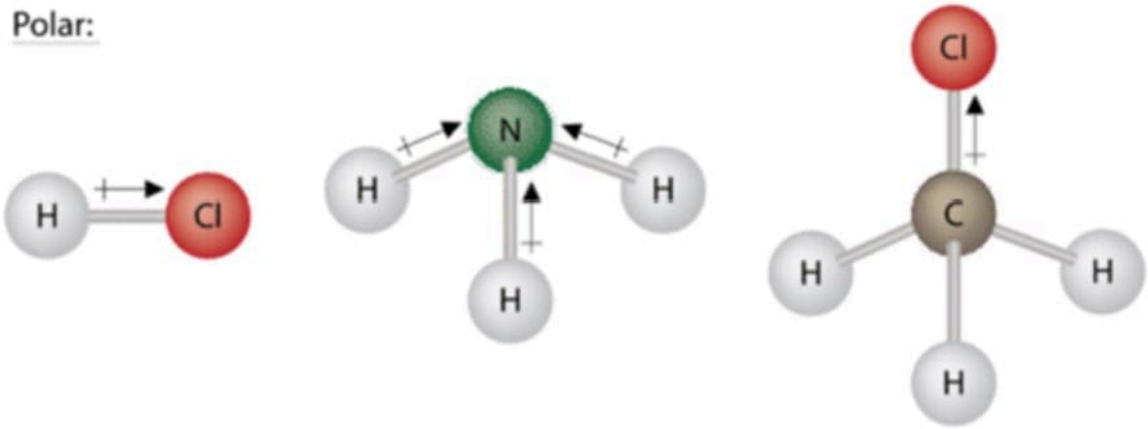


Figure 4.5.4 Some examples of polar molecules based on molecular geometry (HCl, NH_3 and CH_3Cl).

To summarize, to be polar, a molecule must:

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

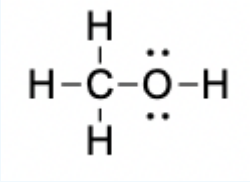
Steps to Identify Polar Molecules

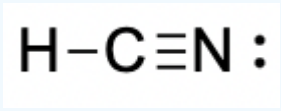
1. Draw the Lewis structure
2. Figure out the geometry (using VSEPR theory)
3. Visualize or draw the geometry
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it)
5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

Example 4.5.1:

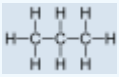
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: 

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
b. Methanol is polar. This is not a symmetric molecule. The $-\mathrm{OH}$ side is different from the other 3 $-\mathrm{H}$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

Exercise 4.5.1

Label the following as polar or nonpolar.

NH_3

Answer available at the end of this section.

Exercise 4.5.2

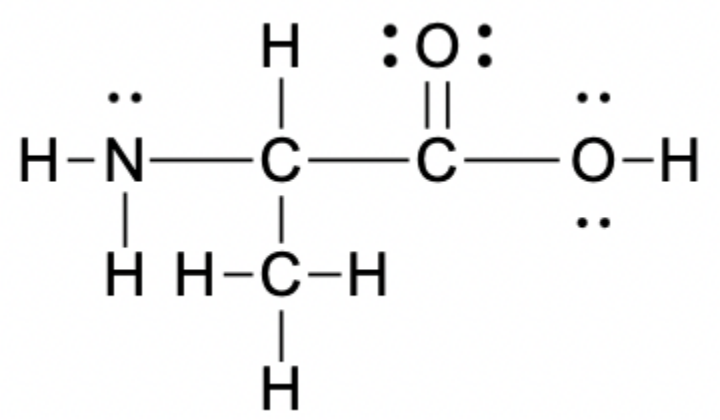
Label the following as polar or nonpolar.

CF_4

Answer available at the end of this section.

Exercise 4.5.3

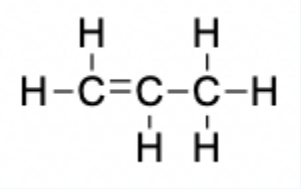
Label the following as polar or nonpolar.



Answer available at the end of this section.

Exercise 4.5.4

Label the following as polar or nonpolar.



Answer available at the end of this section.

Summary

- Nonpolar molecules are symmetric with no unshared electrons.
- Polar molecules are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded.

Contributors and Attributions

-

- StackExchange (thomij).

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Answers

Exercise 4.5.1

Polar

Exercise 4.5.2

Nonpolar. Note: Although this molecule contains several polar bonds, they are arranged symmetrically. The carbon atom is not bonded to different elements and it does not have any lone pairs.

Exercise 4.5.3

Polar. This molecule has polar bonds oriented in trigonal pyramidal, trigonal planar, and bent shapes. The partial charges are not arranged symmetrically around the outside of the molecule.

Exercise 4.5.4

Nonpolar. All compounds containing only carbon and hydrogen are nonpolar. Although the H atoms are not arranged symmetrically, the electronegativities of H and C are essentially the same. All bonds are nonpolar and there are no lone pairs so there are no partial charges on this molecule.

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5: INTERMOLECULAR FORCES AND SOLUTIONS

5.1: Forces between Molecules
5.2: Application of Intermolecular Forces: Melting and Boiling Points
5.3: Solutions - Homogeneous Mixtures
5.4: Passive Transport - Osmosis
5.5: Passive Transport - Tonicity
5.6: Osmotic Pressure and Osmolarity

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5.1: Forces between Molecules

Learning Objectives

- Describe the types of intermolecular forces possible between atoms or molecules in condensed phases (dispersion forces, dipole-dipole attractions, and hydrogen bonding)
- Identify the types of intermolecular forces experienced by specific molecules based on their structures

Solids and liquids have volumes that (mostly) do not change when compressed. In other words, you can try to squeeze them but they don't get smaller. On the other hand, gases are compressible. Solids also have a rigid shape, but the shapes of liquids and gases change to match their containers. These properties are illustrated in the Figure below. In the solid and liquid the particles (atoms or molecules) are touching one another, but in the liquid the particles are able to spread out along the bottom of the container. In the gas the particles are not touching and the gas fills up its container in all directions.

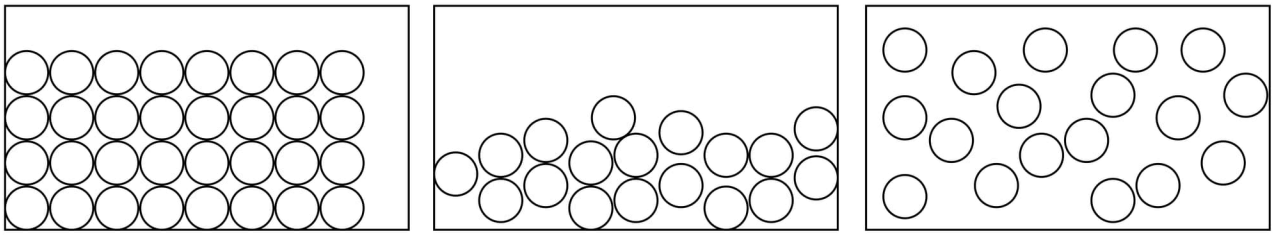


Figure 5.1.1: Comparison of the arrangement of particles in solids, liquids, and gases. Brightyellowjeans, CC BY-SA 4.0, via Wikimedia Commons

The particles that make up solids and liquids are touching one another due to intermolecular forces (IMF). The covalent bonds between atoms in a molecule, such as the bonds that hold H and O together in a water molecule, are intramolecular forces. The prefix intra means within whereas the prefix inter means between different copies. *Intra*molecular forces are those *within* the molecule that keep the molecule together. *Inter*molecular forces are the attractions *between* molecules, which determine many of the physical properties of a substance. Figure 5.1.2 illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules of energy. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

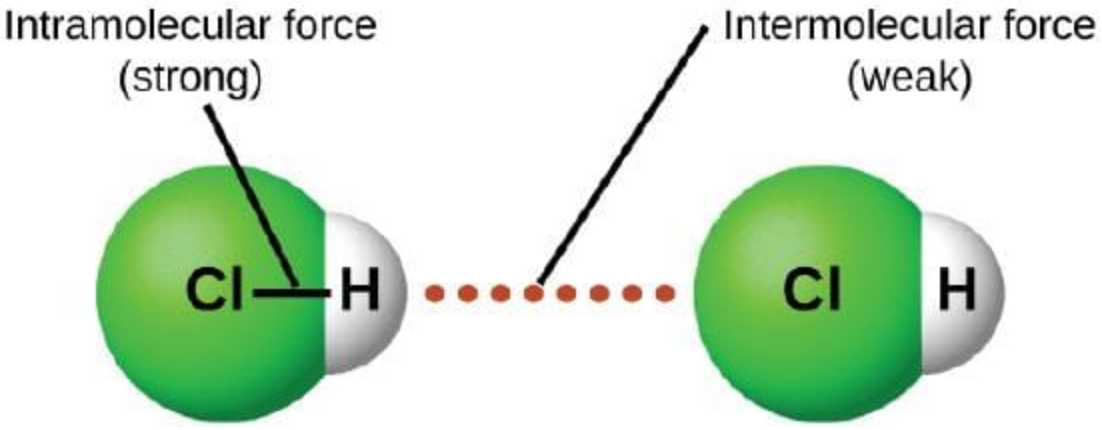


Figure 5.1.2 Intramolecular forces keep a molecule intact. Intermolecular forces hold multiple molecules together and determine many of a substance's properties.

Dipole–Dipole Forces

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

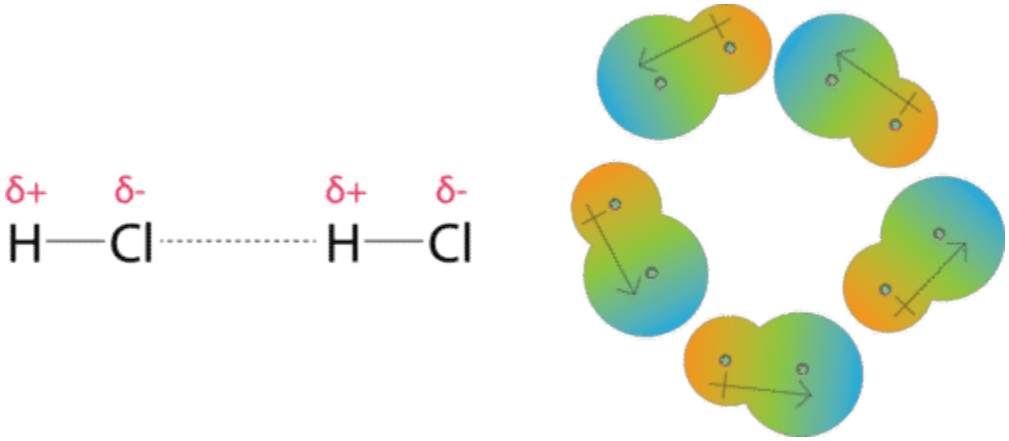


Figure 5.1.3: Dipole-dipole forces in hydrogen chloride. Dipole-dipole forces result from the attraction between the positive end of one dipole and the negative end of a neighboring dipole. Dipole-dipole forces are similar to ionic bonds, but because they involve only partial charges, they are much weaker.

Dispersion Forces

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

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



Figure 5.1.4 Random fluctuations in the electron density within the electron cloud of a helium atom results in a short-lived ("instantaneous") dipole. The attractive force between instantaneous dipoles and the resulting induced dipoles in neighboring molecules is called the London dispersion force.

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. For similar substances, London dispersion forces get stronger with increasing atomic or molecular size.

Hydrogen Bonds

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

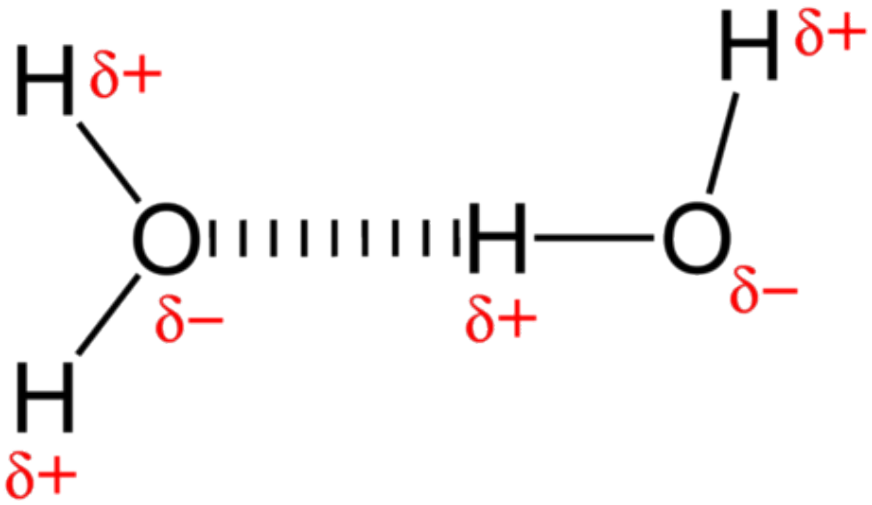


Figure 5.1.6 A hydrogen bond in water occurs between the hydrogen atom of one water molecule and the lone pair of electrons on the oxygen atom of a neighboring water molecule.

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

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

A molecule must contain an H-F, H-O, or H-N bond in order to initiate a hydrogen bond, but the H-F, H-O, or H-N bond itself is NOT the hydrogen bond.

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

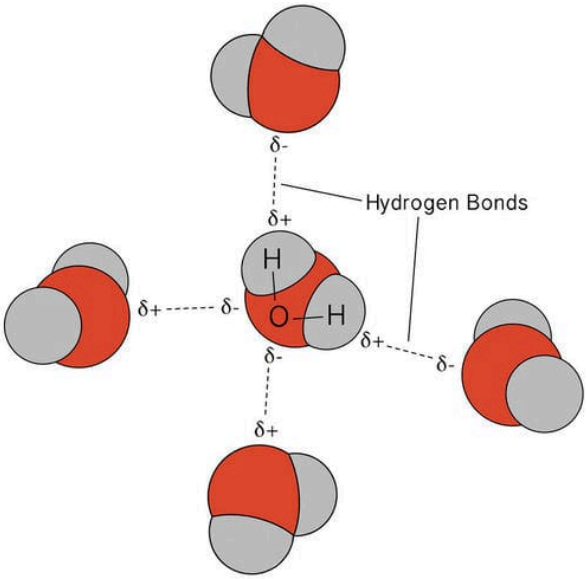


Figure 5.1.7Multiple hydrogen bonds occur simultaneously in water because of its bent shape and the presence of two very polar bonds per molecule.

In the liquid state, the hydrogen bonds of water can break and reform as the molecules flow from one place to another. When water is cooled, the molecules begin to slow down. Eventually, when water is frozen to ice, the hydrogen bonds become more rigid and form a well-defined network (Figure 5.1.8). The bent shape of the molecules leads to gaps in the hydrogen bonding network of ice. Ice has the very unusual property that its solid state is less dense than its liquid state. As a result, ice floats in liquid water. Virtually all other substances are denser in the solid state than in the liquid state. Hydrogen bonds also play a very important biological role in the physical structures of proteins and nucleic acids.

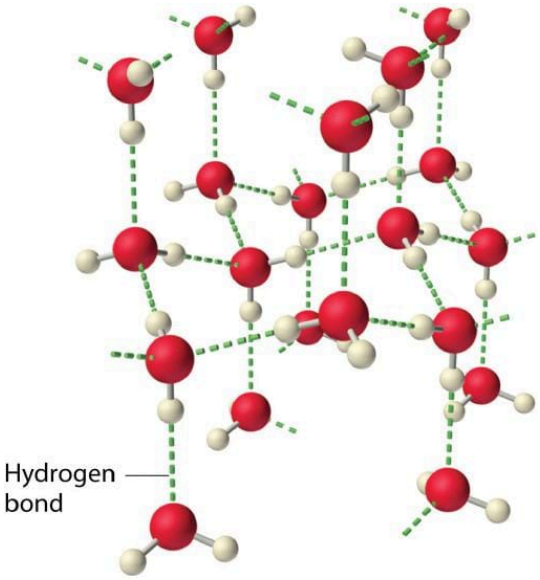
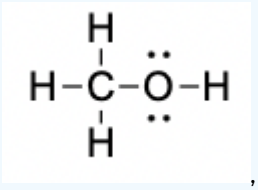
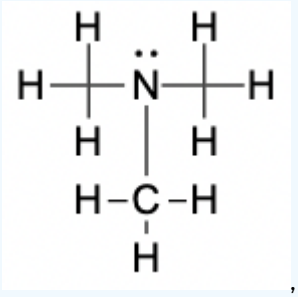


Figure 5.1.8 The Hydrogen-Bonded Structure of Ice.

Hydrogen bond formation requires both a hydrogen bond donor and a hydrogen bond acceptor.

Because ice is less dense than liquid water, rivers, lakes, and oceans freeze from the top down. In fact, the ice forms a protective surface layer that insulates the rest of the water, allowing fish and other organisms to survive in the lower levels of a frozen lake or sea. If ice were denser than the liquid, the ice formed at the surface in cold weather would sink as fast as it formed. Bodies of water would freeze from the bottom up, which would be lethal for most aquatic creatures. The expansion of water when freezing also explains why automobile or boat engines must be protected by “antifreeze” and why unprotected pipes in houses break if they are allowed to freeze.

Example 5.1.1

Considering  , C_2H_6, Xe, and  , which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Given: compounds

Asked for: formation of hydrogen bonds and structure

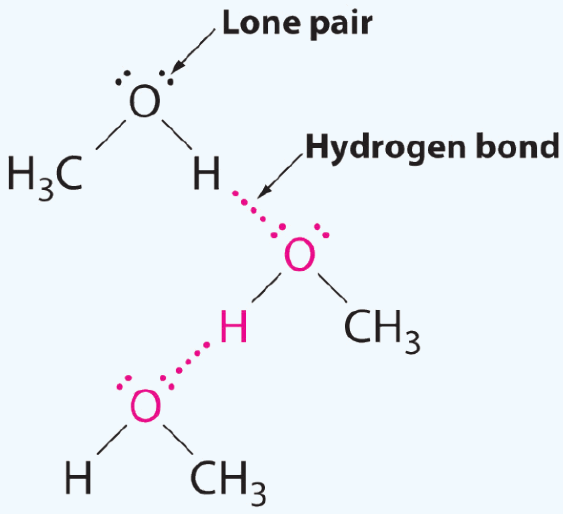
Strategy:

A. Identify the compounds with a hydrogen atom attached to O, N, or F. These are likely to be able to act as hydrogen bond donors.
B. Of the compounds that can act as hydrogen bond donors, identify those that also contain lone pairs of electrons, which allow them to be hydrogen bond acceptors. If a substance is both a hydrogen donor and a hydrogen bond acceptor, draw a structure showing the hydrogen bonding.

Solution:

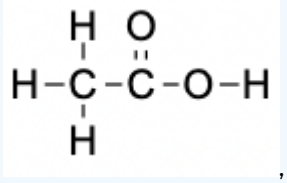
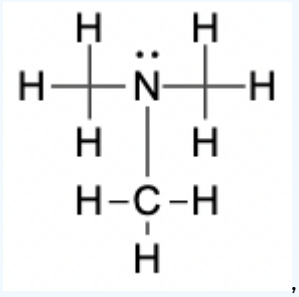
A Of the species listed, xenon (Xe), ethane (C_2H_6), and trimethylamine [$(CH_3)_3N$] do not contain a hydrogen atom attached to O, N, or F; hence they cannot act as hydrogen bond donors.

B The one compound that can act as a hydrogen bond donor, methanol (CH_3OH), contains both a hydrogen atom attached to O (making it a hydrogen bond donor) and two lone pairs of electrons on O (making it a hydrogen bond acceptor); methanol can thus form hydrogen bonds by acting as either a hydrogen bond donor or a hydrogen bond acceptor. The hydrogen-bonded structure of methanol is as follows:



Hydrogen bonding in methanol

Exercise 5.1.1

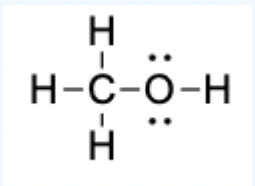
Considering  ,  , NH_3, and CH_3F, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Answer available at the end of this section.

Example 5.1.2:

Identify the most significant intermolecular force in each substance.

a. C_3H_8

b. 

c. H_2S

Solution

a. Because C–H bonds are nonpolar the most significant intermolecular force for this substance would be dispersion forces.

b. This molecule has an H atom bonded to an O atom, so it will experience hydrogen bonding.

c. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram and VSEPR indicate that it is bent, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.

Exercise 5.1.2

Identify the most significant intermolecular force in each substance.

a. HF
b. HCl

Answer a available at the end of this section.

Summary

- Molecules in liquids are held to other molecules by intermolecular interactions, which are weaker than the intramolecular interactions that hold the atoms together within molecules and polyatomic ions.
- **Dipole–dipole interactions** arise from the electrostatic interactions of the positive and negative ends of molecules with permanent dipole moments.
- **London dispersion forces** are due to the formation of **instantaneous dipole moments** in polar and nonpolar molecules as a result of short-lived fluctuations of electron charge distribution, which in turn cause the temporary formation of an **induced dipole** in adjacent molecules.
- **Hydrogen bonds** are especially strong dipole–dipole interactions between molecules that have hydrogen bonded to a highly electronegative atom, such as O, N, or F. The resulting partially positively charged H atom on one molecule (the *hydrogen bond donor*) can interact strongly with a lone pair of electrons of a partially negatively charged O, N, or F atom on adjacent molecules (the *hydrogen bond acceptor*).

Contributors and Attributions

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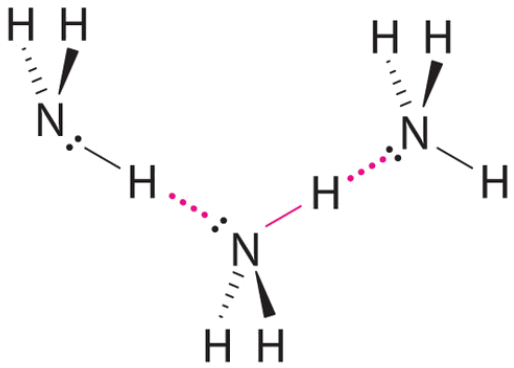
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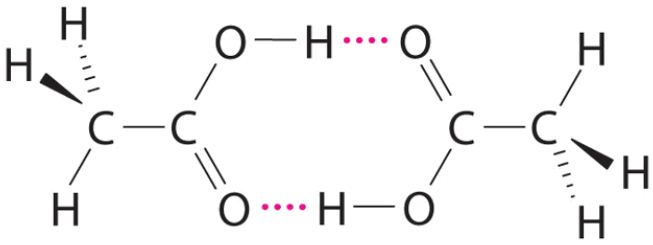
Answers

Exercise 5.1.1

CH_3CO_2H and NH_3;



Hydrogen bonding in ammonia



Hydrogen bonding in acetic acid

Exercise 5.1.2

hydrogen bonding

dipole-dipole interactions

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5.2: Application of Intermolecular Forces: Melting and Boiling Points

Learning Objectives

- Rank melting points and boiling points for molecules of approximately the same size
- Rank melting points and boiling points for atoms and molecules of different sizes

Introduction

The strength of intermolecular forces determines melting points (mp) and boiling points (bp) for substances. The stronger the IMF, the higher the melting point or boiling point.

This is because IMFs are an attractive force holding the particles together and energy (in the form of heat) must be added to weaken or break the forces/attractions. Melting occurs when IMFs are weakened and boiling occurs when the IMFs are broken and the particles are allowed to completely separate from one another.

Strength of Intermolecular Forces

The strength of IMFs increases as molecules become more polar (as measured by the dipole moment of the molecule), as shown in Table 5.2.1. Since the size of a molecule can also play a role, the table compares molecules of similar size as measured by the total number of electrons in the molecule.

Table 5.2.1: Relationships between the Dipole Moment and the Boiling Point for Organic Compounds of Similar Size

Compound	Total Number of Electrons	Dipole Moment (D)	Boiling Point (K)
C_3H_6 (cyclopropane)	24	0	240
CH_3OCH_3 (dimethyl ether)	26	1.30	248
CH_3CN (acetonitrile)	22	3.9	355

Example 5.2.1 Dipole-Dipole Forces and Their Effects

Predict which will have the higher boiling point: N_2 or CO. Explain your reasoning.

Solution

CO and N_2 are both small molecules with 14 electrons, so the polarity of the molecules will be the primary factor determining their boiling points. 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.

Exercise 5.2.1

Predict which will have the higher boiling point: ICl or Br_2. Explain your reasoning.

Answer available at the end of this section.

Size Determines the Strength of Dispersion Forces

Larger atoms and molecules exhibit stronger dispersion forces than do smaller atoms and molecules because they have more electrons that may temporarily shift one way or another. 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 5.2.1.

Table 5.2.2: Melting and Boiling Points of the Halogens

Halogen	Number of Electrons	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	18	72 pm	53 K	85 K
chlorine, Cl_2	34	99 pm	172 K	238 K
bromine, Br_2	70	114 pm	266 K	332 K
iodine, I_2	106	133 pm	387 K	457 K
astatine, At_2	170	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.

Shape Also Affects Dispersion Forces

The strengths of London dispersion forces also depend significantly on molecular shape because shape determines how much of one molecule can interact with its neighboring molecules at any given time. For example, part (b) in Figure 5.2.5 shows 2,2-dimethylpropane (neopentane) and *n*-pentane, both of which have the formula C_5H_{12}. Neopentane is almost spherical, with a small surface area for intermolecular interactions, whereas *n*-pentane has an extended conformation that enables it to come into close contact with other *n*-pentane molecules. As a result, the boiling point of neopentane (9.5°C) is more than 25°C lower than the boiling point of *n*-pentane (36.1°C). The red lines show that there is more contact between two n-pentane molecules than there is between two neopentane molecules, even though they are both C_5H_{12}.

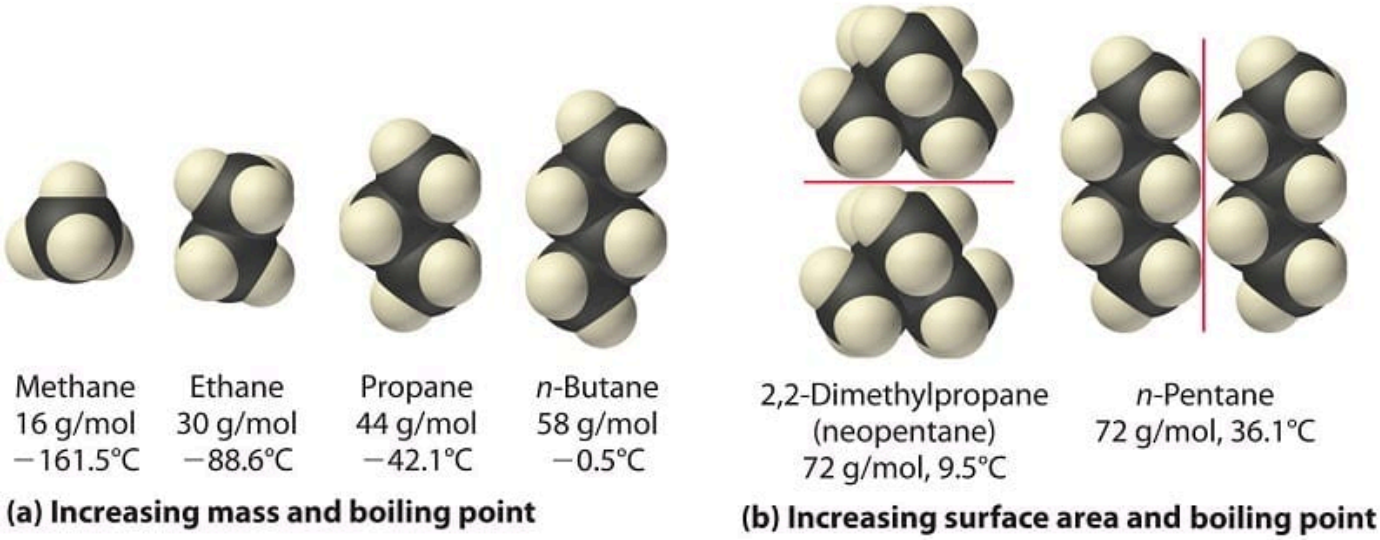


Figure 5.2.5 Mass and Surface Area Affect the Strength of London Dispersion Forces. (a) In this series of four simple alkanes, larger molecules have stronger London forces between them than smaller molecules and consequently higher boiling points. (b) Linear *n*-pentane molecules have a larger surface area and stronger intermolecular forces than spherical neopentane molecules. As a result, neopentane is a gas at room temperature, whereas *n*-pentane is a volatile liquid.

London dispersion forces

All molecules, whether polar or nonpolar, are attracted to one another by London dispersion forces in addition to any other attractive forces that may be present.

Example 5.2.2 London Forces and Their Effects

Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

Given: Molecular formulas

Asked for: Order of increasing boiling points

Strategy:

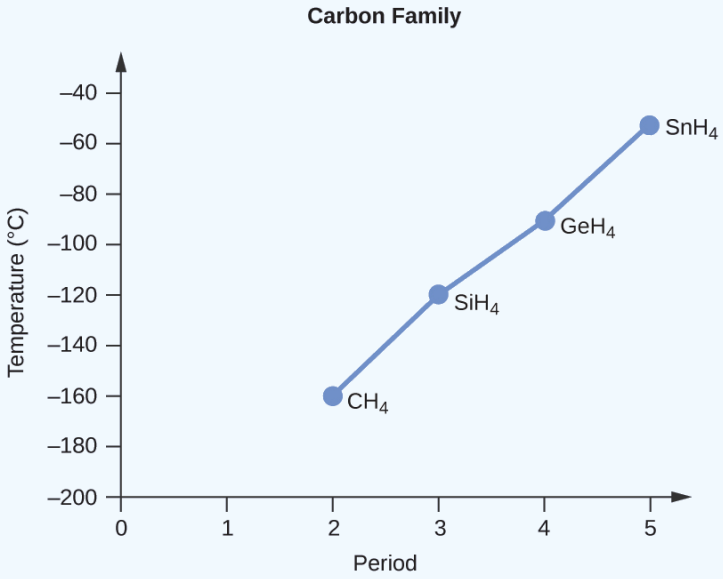
Determine the intermolecular forces in the compound. If the compounds can participate in the same type of IMFs, compare the size and surface area of compounds with similar types of intermolecular force. The substance with the weakest forces and the lowest lowest number of electrons and least surface area have the lowest boiling point.

Solution

Applying the skills acquired in the chapter on polarity 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 molecules of CH_4, SiH_4, GeH_4, and SnH_4 contain 10, 18, 36, and 54 electrons, 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 elements shows this prediction to be correct:



Comparing H-Bonding to Dispersion

The graph below is similar to the previous one. The Group 14 line (purple, bottom) shows the same boiling point data as the graph above. In Groups 14, 15, and 16 the same trend can be seen for hydrides from periods 3, 4, and 5. However, the boiling points for NH_3, HF, and H_2O are unusually high. This is because those molecules can form hydrogen bonds with their neighbors. Because hydrogen bonds are stronger interactions than dipole-dipole or dispersion forces, the boiling points are higher.

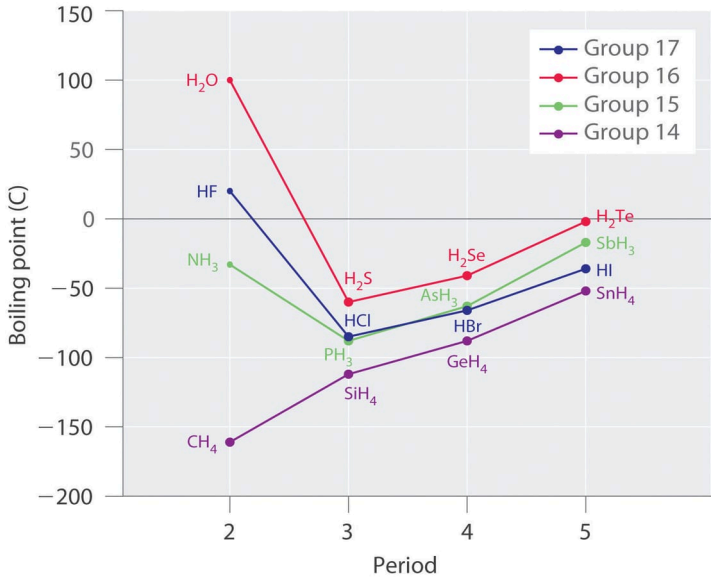


Figure 5.2.9 The Effects of Hydrogen Bonding on Boiling Points. These plots of the boiling points of the covalent hydrides of the elements of groups 14–17 show that the boiling points of the lightest members of each series for which hydrogen bonding is possible (HF, NH_3, and H_2O) are anomalously high for compounds with such low molecular masses.

Although hydrogen bonds are significantly weaker than covalent bonds, with typical dissociation energies of only 15–25 kJ/mol, they have a significant influence on the physical properties of a compound. Compounds such as HF can form only two hydrogen bonds at a time as can, on average, pure liquid NH_3. Consequently, even though their molecular masses are similar to that of water, their boiling points are significantly lower than the boiling point of water, which forms *four* hydrogen bonds at a time.

Exercise 5.2.2

Order the following hydrocarbons from lowest to highest boiling point: C_2H_6, C_3H_8, and C_4H_{10}.

Answer available at the end of this section.

Example 5.2.3: Buckyballs

Arrange C_{60} (buckminsterfullerene, which has a soccer ball structure), NaCl, He, Ar, and N_2O in order of increasing boiling points.

Given: compounds

Asked for: order of increasing boiling points

Strategy:

Identify the intermolecular forces in each compound and then arrange the compounds according to the strength of those forces. The substance with the weakest forces will have the lowest boiling point.

Solution:

Interactions are strongest for an ionic compound which has full positive and negative charges, so we expect NaCl to have the highest boiling point. To predict the relative boiling points of the other compounds, we must consider their polarity (for dipole–dipole interactions), their ability to form hydrogen bonds, and their sizes (for London dispersion forces). Helium is nonpolar and by far the smallest, so it should have the lowest boiling point. Argon and N_2O have very similar numbers of electrons (18 and 22, respectively), but N_2O is polar while Ar is not. Consequently, N_2O should have a higher boiling point. A C_{60} molecule is nonpolar, but it has 360 electrons, much greater than that of Ar or N_2O. Because the boiling points of nonpolar substances increase rapidly with size, C_{60} should boil at a higher temperature than the other nonionic substances. The predicted order is thus as follows, with actual boiling points in parentheses:

$$\text{He } (-269^\circ\text{C}) < \text{Ar } (-185.7^\circ\text{C}) < N_2O\ (-88.5^\circ\text{C}) < C_{60}\ (>280^\circ\text{C}) < \text{NaCl } (1465^\circ\text{C}).$$

Exercise 5.2.3

Arrange C_9H_{20}, Ne, CS_2, Cl_2, and KBr in order of decreasing boiling points.

Answer available at the end of this section.

Summary

- If molecules are approximately the same size (same total number of electrons), then the type of intermolecular force will determine the relative melting and boiling points.
  - Hydrogen bonding forces are the strongest and result in higher mp and bp
  - Dispersion forces are the weakest and result in low mp and bp
- If atoms and molecules have significantly different numbers of electrons, then:
  - Dispersion forces are highest in the atoms/molecules with the most electrons. This leads to higher mp and bp.
  - Dispersion forces are lowest in atoms/molecules with the fewest electrons. This leads to low mp and bp.

Answers

Exercise 5.2.1

ICl. ICl and Br_2 have the same number of electrons, 70, and therefore their polarity will be the primary factor determining their boiling points. ICl is polar and thus

participates in 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

Exercise 5.2.2

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

$$C_2H_6 < C_3H_8 < C_4H_{10}.$$

Exercise 5.2.3

KBr (1435°C) > C_9H_{20} (132.9°C for 2,4-dimethylheptane) > CS_2 (46.6°C) > Cl_2 (–34.6°C) > Ne (–246°C)

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5.3: Solutions - Homogeneous Mixtures

Learning Objectives

- Learn terminology involving solutions.
- Explain the significance of the statement "like dissolves like."
- Explain why certain substances dissolve in other substances.

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. By major and minor we mean whichever component has the greater or lesser presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, here we will confine the discussion to solutions for which the major component and the minor component are obvious.



Figure 5.3.1: Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water, while air is a solution of a gaseous solute ($\mathrm{O_2}$) in a gaseous solvent ($\mathrm{N_2}$). In all cases, however, the overall phase of the solution is the same phase as the solvent. Table 5.3.1 lists some common types of solutions, with examples of each.

Table 5.3.1: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol ($\mathrm{C_2H_5OH}$) in $\mathrm{H_2O}$ (alcoholic beverages)
liquid	solid	salt water
solid	gas	$\mathrm{H_2}$ gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

Example 5.3.1: Sugar and Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is **sucrose**, so it is **the solute**. **Water**—the majority component—is **the solvent**. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 5.3.1

A solution is made by dissolving 3.33 g of $HCl(g)$ in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.

Answer available at the end of this section.

Like Dissolves Like

A simple way to predict which compounds will dissolve in other compounds is the phrase "like dissolves like". What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve in water but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, a nonpolar molecule, does dissolve in water—it is this oxygen that the fish take in through their gills. The reason we can enjoy carbonated sodas is also due to a nonpolar compound that dissolves in water. Pepsi-cola and all the other sodas have carbon dioxide gas, CO_2, a nonpolar compound, dissolved in a sugar-water solution. In this case, to keep as much gas in solution as possible, the sodas are kept under pressure.

This general trend of "like dissolves like" is summarized in the following table:

Table 5.3.2: Summary of Solubilities

Solute (Polarity of Compound)	Solvent (Polarity of Compound)	Dominant Intermolecular Force	Is Solution Formed?
Polar	Polar	Dipole-Dipole Force and/or Hydrogen Bond	yes
Non-polar	Non-polar	Dispersion Force	yes
Polar	Non-polar		no
Non-polar	Polar		no
Ionic	Polar	Ion-Dipole	yes
Ionic	Non-polar		no

Note that every time charged particles (ionic compounds or polar substances) are mixed, a solution is formed. When particles with no charges (nonpolar compounds) are mixed, they will form a solution. However, if substances with charges are mixed with other substances without charges, a solution does not form. When an ionic compound is considered "insoluble", it doesn't necessarily mean the compound is completely untouched by water. All ionic compounds dissolve to some extent. An insoluble compound just doesn't dissolve in any noticeable or appreciable amount.

What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole

interactions, and hydrogen bonding (as described in Chapter 10). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water.



Figure 5.3.2: Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

Example 5.3.2: Polar and Nonpolar Solvents

Would $\mathrm{I_2}$ be more soluble in $\mathrm{CCl_4}$ or $\mathrm{H_2O}$? Explain your answer.

Solution

$\mathrm{I_2}$ is nonpolar. Of the two solvents, $\mathrm{CCl_4}$ is nonpolar and $\mathrm{H_2O}$ is polar, so $\mathrm{I_2}$ would be expected to be more soluble in $\mathrm{CCl_4}$.

Exercise 5.3.2

Would $\mathrm{C_3H_7OH}$ be more soluble in $\mathrm{CCl_4}$ or $\mathrm{H_2O}$? Explain your answer.

Answer available at the end of this section.

Example 5.3.3

Water is considered a polar solvent. Which substances should dissolve in water?

a. methanol ($\mathrm{CH_3OH}$)
b. sodium sulfate (\ce{Na2SO4}\))
c. octane ($\mathrm{C_8H_{18}}$)

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

a. 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*.
b. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.

c. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

Exercise 5.3.3: Toluene

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 available at the end of this section.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.

Answers

Exercise 5.3.1

solute: $HCl\,(g)$

solvent: CH_3OH

Exercise 5.3.2

H_2O, because both experience hydrogen bonding.

Exercise 5.3.3**: Toluene**

Octane (C_8H_{18}) will dissolve. It is also non-polar.

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5.4: Passive Transport - Osmosis

Learning Objectives

- Describe the process of osmosis and explain how concentration gradient affects osmosis

Osmosis and Semipermeable Membranes

Osmosis is the movement of water through a semipermeable membrane. The prefix "semi" means partially and "permeable" means something can pass through. Therefore, a semipermeable membrane lets some things pass through (small molecules such as water) and prevents other things from passing through (like larger solutes or ions). In osmosis water moves from the side of the membrane with a lower total concentration of solutes towards the side with a higher total concentration of solutes. We don't usually talk about the concentration of a solvent, but you could think of osmosis as the movement of water from the side of the membrane that has a higher concentration of water to the side with a lower concentration of water in order to even things out.

While **diffusion** transports various materials across membranes and within cells, **osmosis** transports only water across a membrane. The semipermeable membrane limits the diffusion of solutes in the water. There are proteins called aquaporins that form pores in cell membranes to make them semipermeable and facilitate water movement.

Example 5.4.1

A solution with a concentration of 1.15 M is separated from a solution with a concentration of 0.33 M by a semipermeable membrane. Which way will water flow? (Note: The identity of the solutes does not affect the direction of flow.)

Solution

- The scenario involves two solutions of different concentrations separated by a semipermeable membrane, therefore osmosis will occur.
- In osmosis, water moves from the solution with lower concentration towards the solution with higher concentration.
- **Water will move from 0.33 M solution into the 1.15 M solution**.
- As a result, the concentrations will become more similar. Adding water to the 1.15 M solution dilutes it and lowers its concentration. Removing water from the 0.33 M solution raises its concentration.

Exercise 5.4.1

A semipermeable membrane separates two solutions. One has a concentration of 0.5 % and the other solution has a concentration of 2 %. Which solution will lose water and which solution will gain water?

Answer available at the end of this section.

Mechanism of Osmosis

Osmosis is a special case of diffusion. Water, like other substances, moves from an area of high concentration to one of low concentration. An obvious question is what makes water move at all? Imagine a beaker with a semipermeable membrane separating the two sides or halves. On both sides of the membrane the water level is the same, but there are different concentrations of a dissolved substance, or solute, that cannot cross the membrane (otherwise the concentrations on each side would be balanced by the solute crossing the membrane). If the volume of the solution on both sides of the membrane is the same but the concentrations of solute are different, then there are different amounts of water, the solvent, on either side of the membrane. If there is more solute in one area, then there is less water; if there is less solute in one area, then there must be more water.

To illustrate this, imagine two full glasses of water. One has a single teaspoon of sugar in it, whereas the second one contains one-quarter cup of sugar. If the total volume of the solutions in both cups is the same, which cup contains more water? Because the large amount of sugar in the second cup takes up much more space than the teaspoon of sugar in the first cup, the first cup has more water in it.

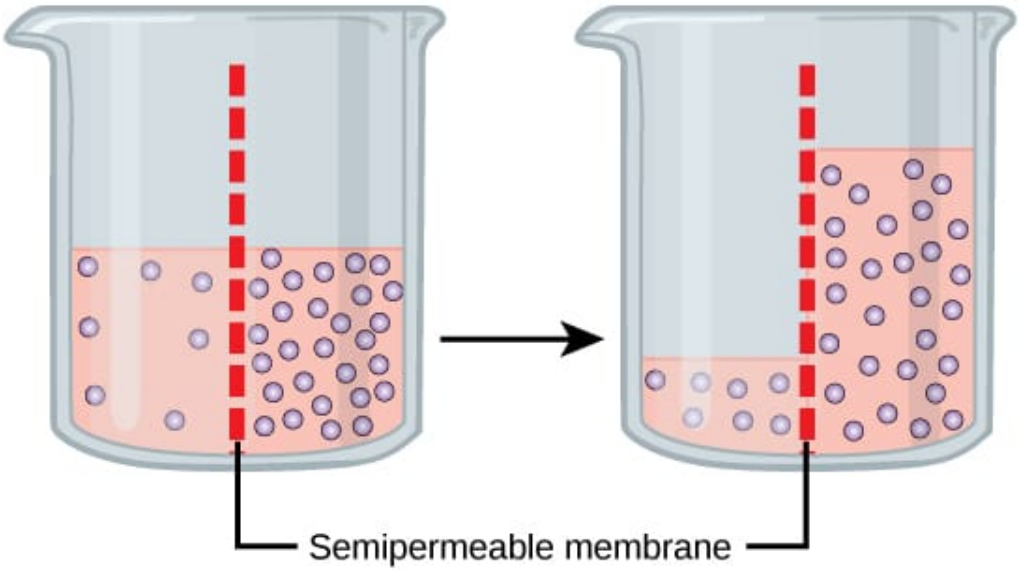


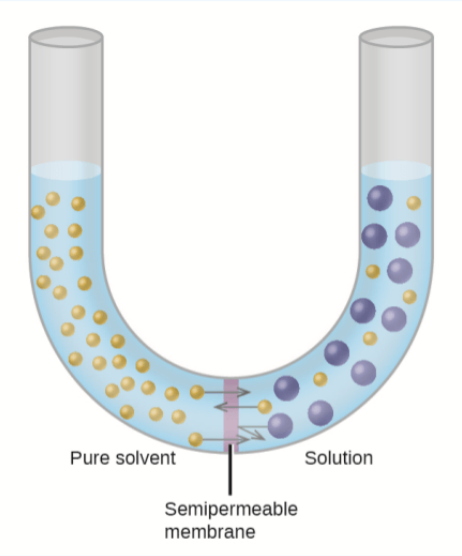
Figure 5.4.1: Osmosis: In osmosis, water always moves from an area of higher water concentration to one of lower concentration. In the diagram shown, the solute cannot pass through the selectively permeable membrane, but the water can.

Returning to the beaker example, recall that it has a mixture of solutes on either side of the membrane. A principle of diffusion is that the molecules move around and will spread evenly throughout the medium if they can. However, only the material capable of passing through the membrane will diffuse through it. In this example, the solute cannot diffuse through the membrane, but the water can. Water has a concentration gradient in this system. Thus, water will diffuse down its concentration gradient, crossing the membrane to the side where it is less concentrated. This diffusion of water through the membrane—osmosis—will continue until the concentration gradient of water goes to zero or until the hydrostatic pressure of the water balances the osmotic pressure. In the beaker example, this means that the level of fluid in the side with a higher solute concentration will go up.

Example 5.4.2

A tube with a semipermeable membrane in the middle contains water (pure solvent) on the left and a solution on the right. Initially the volume on each side of the membrane is the same. How will the volumes compare after osmosis occurs?

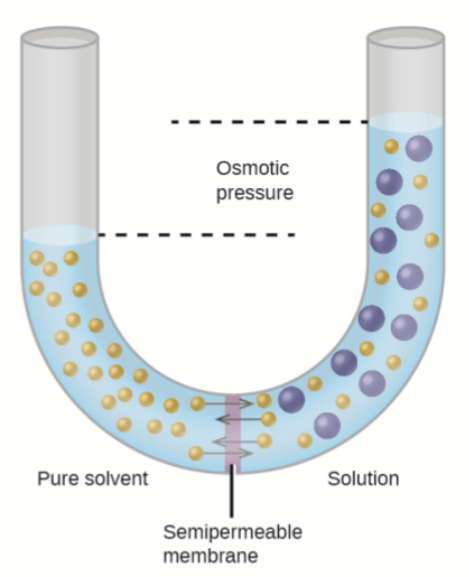
a. The volumes will remain equal
b. The volume on the left will be higher than on the right
c. The volume on the right will be higher than on the left



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Solution

Osmosis will occur. Water will move from the side with a higher concentration (more solute) towards the side with a lower concentration (less solute). Therefore, the correct answer is **(c) The volume on the right will be higher than on the left**.

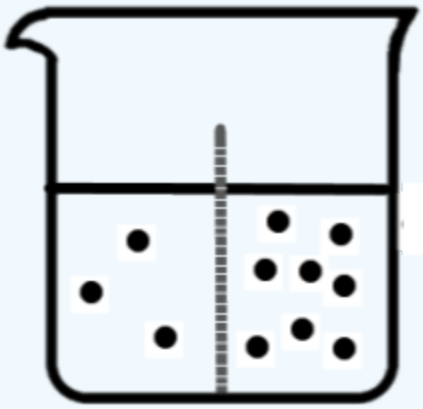


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Water (represented by yellow spheres in the figure) moved through the membrane. Solute particles (represented by larger, darker spheres in the figure) were too large to move through the membrane. Therefore there are now more particles on the right.

Exercise 5.4.2

In which direction will water flow in the beaker below? How will the volumes change?



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Answer available at the end of this section.

Key Points

- Osmosis occurs according to the concentration gradient of water across the membrane, which is inversely proportional to the concentration of solutes.
- Osmosis occurs until the concentrations on both sides of the membrane are the same.
- Osmosis occurs when there are different concentrations on two sides of a semipermeable membrane but the membrane does not allow diffusion of the solute.

Key Terms

- **solute**: Any substance that is dissolved in a liquid solvent to create a solution
- **osmosis**: The net movement of solvent molecules from a region of high solvent potential to a region of lower solvent potential through a partially permeable membrane
- **semipermeable membrane**: A type of biological membrane that will allow certain molecules or ions to pass through it by diffusion and occasionally by specialized facilitated diffusion

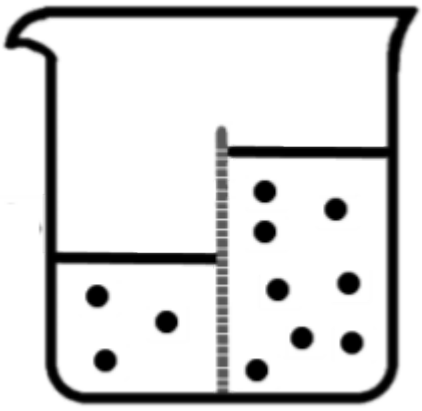
Answers

Exercise 5.4.1

The 0.5 % solution will lose water and the 2 % solution will gain water.

Exercise 5.4.2

Water will move from left to right (towards the more concentrated solution). The final volume on the right will be higher than that on the left.



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5.5: Passive Transport - Tonicity

Learning Objectives

- Define tonicity and describe its relevance to osmosis

Tonicity describes how an extracellular solution can change the volume of a cell by affecting osmosis. A solution's tonicity often directly correlates with the osmolarity of the solution. Osmolarity describes the total solute concentration of the solution. A solution with low osmolarity has a greater number of water molecules relative to the number of solute particles; a solution with high osmolarity has fewer water molecules with respect to solute particles. In a situation in which solutions of two different osmolarities are separated by a membrane permeable to water, though not to the solute, water will move from the side of the membrane with lower osmolarity (and more water) to the side with higher osmolarity (and less water). This effect makes sense if you remember that the solute cannot move across the membrane, and thus the only component in the system that can move—the water—moves along its own concentration gradient. An important distinction that concerns living systems is that osmolarity measures the number of particles (which may be molecules) in a solution. Therefore, a solution that is cloudy with cells may have a lower osmolarity than a solution that is clear if the second solution contains more dissolved molecules than there are cells.

Hypotonic Solutions

Three terms—hypotonic, isotonic, and hypertonic—are used to relate the osmolarity of a cell to the osmolarity of the extracellular fluid that contains the cells. In a hypotonic situation, the extracellular fluid has lower osmolarity than the fluid inside the cell, and water enters the cell. (In living systems, the point of reference is always the cytoplasm, so the prefix hypo- means that the extracellular fluid has a lower concentration of solutes, or a lower osmolarity, than the cell cytoplasm.) It also means that the extracellular fluid has a higher concentration of water in the solution than does the cell. In this situation, water will follow its concentration gradient and enter the cell, causing the cell to expand.

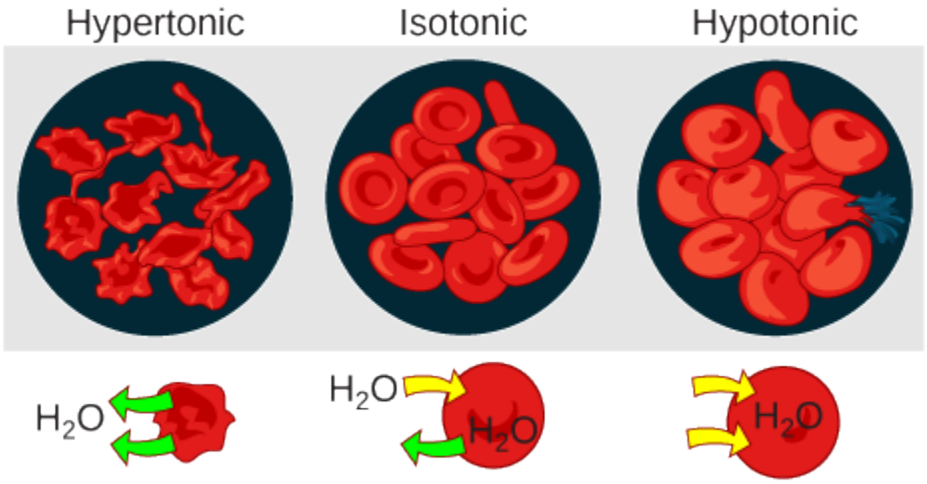


Figure 5.5.1: Changes in Cell Shape Due to Dissolved Solutes: Osmotic pressure changes the shape of red blood cells in hypertonic, isotonic, and hypotonic solutions.

Hypertonic Solutions

As for a hypertonic solution, the prefix hyper- refers to the extracellular fluid having a higher osmolarity than the cell's cytoplasm; therefore, the fluid contains less water than the cell does. Because the cell has a relatively higher concentration of water, water will leave the cell, and the cell will shrink.

Isotonic Solutions

In an isotonic solution, the extracellular fluid has the same osmolarity as the cell. If the osmolarity of the cell matches that of the extracellular fluid, there will be no net movement of water into or out of the cell, although water will still move in and out.

Blood cells and plant cells in hypertonic, isotonic, and hypotonic solutions take on characteristic appearances. Cells in an isotonic solution retain their shape. Cells in a hypotonic solution swell as water enters the cell, and may burst if the concentration gradient is large enough between the inside and outside of the cell. Cells in a hypertonic solution shrink as water exits the cell, becoming shriveled.

Example 5.5.1

What will happen if red blood cells are placed in a hypertonic solution?

Solution

The membranes of red blood cells are semipermeable, therefore osmosis can occur if the concentrations inside and outside the cell are different. Hypertonic means "higher concentration". In osmosis, water moves towards the more concentrated solution. Therefore, **water would move out of the cells and the cells would shrink and shrivel**.

Exercise 5.5.1

What will happen if red blood cells are placed in an isotonic solution?

Answer available at the end of this section.

Key Points

- Osmolarity describes the total solute concentration of a solution; solutions with a low solute concentration have a low osmolarity, while those with a high osmolarity have a high solute concentration.
- Water moves from the side of the membrane with lower osmolarity (and more water) to the side with higher osmolarity (and less water).
- In a hypotonic solution, the extracellular fluid has a lower osmolarity than the fluid inside the cell; water enters the cell.
- In a hypertonic solution, the extracellular fluid has a higher osmolarity than the fluid inside the cell; water leaves the cell.
- In an isotonic solution, the extracellular fluid has the same osmolarity as the cell; there will be no net movement of water into or out of the cell.

Key Terms

- **osmolarity**: The osmotic concentration of a solution, normally expressed as osmoles of solute per litre of solution.
- **hypotonic**: Having a lower osmotic pressure than another; a cell in this environment causes water to enter the cell, causing it to swell.
- **hypertonic**: having a greater osmotic pressure than another
- **isotonic**: having the same osmotic pressure

Answers

Exercise 5.5.1

Water would move in and out of the cells at equal rates and the cells would keep their normal size and shape.

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5.6: Osmotic Pressure and Osmolarity

Learning Objectives

- Calculate osmolarity when the concentration and identity of solute(s) are known

Osmotic Pressure

When two solutions of different concentration are present on either side of these membranes (called *semipermeable membranes*), there is a 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. This tendency is called **osmotic pressure**. External pressure can be exerted on a solution to counter the flow of solvent; the pressure required to halt the osmosis of a solvent is equal to the osmotic pressure of the solution.

Osmolarity (Osmol or Osm) is a way of reporting the total number of particles in a solution to determine osmotic pressure. It is defined as the molarity of a solute times the number of particles a formula unit of the solute makes when it dissolves (represented by i):

$$Osmol = M \times i \tag{5.6.1}$$

If more than one solute is present in a solution, the individual osmolarities are additive to get the total osmolarity of the solution. Solutions that have the same osmolarity have the same osmotic pressure. If solutions of differing osmolarities are present on opposite sides of a semipermeable membrane, solvent will transfer from the lower-osmolarity solution to the higher-osmolarity solution. Counterpressure exerted on the high-osmolarity solution will reduce or halt the solvent transfer. An even higher pressure can be exerted to force solvent from the high-osmolarity solution to the low-osmolarity solution, a process called *reverse osmosis*. Reverse osmosis is used to make potable water from saltwater where sources of fresh water are scarce.

Example 5.6.3

A 0.50 M NaCl aqueous solution and a 0.30 M $Ca(NO_3)_2$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

Solution

The solvent will flow into the solution of higher osmolarity. The NaCl solute separates into two ions—Na^+ and Cl^-—when it dissolves, so its osmolarity is as follows:

$$\text{Osmol (NaCl)} = 0.50\ \text{M} \times 2 = 1.0\ \text{Osmol}$$

The $Ca(NO_3)_2$ solute separates into three ions—one Ca^{2+} and two NO_3^-—when it dissolves, so its osmolarity is as follows:

$$\text{Osmol [Ca(NO}_3)_2\text{]} = 0.30\ \text{M} \times 3 = 0.90\ \text{osmol}$$

The osmolarity of the $Ca(NO_3)_2$ solution is lower than that of the NaCl solution, so water will transfer through the membrane from the $Ca(NO_3)_2$ solution to the NaCl solution.

Exercise 5.6.3

A 1.5 M $C_6H_{12}O_6$ aqueous solution and a 0.40 M $Al(NO_3)_3$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

Answer available at the end of this section.

Multiple Solutes

Osmotic pressure is based on the total concentration of a solution. Many solutions, especially in the body, contain multiple solutes. The osmolarity of a solution is the sum of the contributions from each solute.

$$Osmol = (M_1 \times i_1) + (M_2 \times i_2) + ...$$

Example 5.6.2

A solution is 0.2 M glucose ($C_6H_{12}O_6$) and 0.04 M NaCl. What is the osmolarity of the solution?

Solution

Glucose is a molecular compound so its value for i = 1. The compound NaCl is ionic and it consists of two ions, Na^+ and Cl^-, so its value of i = 2

$$Osmol = (M_{C_6H_{12}O_6} \times i_{C_6H_{12}O_6}) + (M_{NaCl} \times i_{NaCl})$$

$$Osmol = (0.2\ M \times 1) + (0.04\ M \times 2)$$

Osmolarity = 0.28 Osm

Exercise 5.6.2

What is the osmolarity of a solution that contains 0.15 M $CaCl_2$, 0.10 M $NaNO_3$, and 0.20 M CH_4O?

Answer available at the end of this section.

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 one form of dialysis, called *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. The osmotic pressure of the blood forces waste molecules and excess water through the membrane into the sterile solution. 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 5.6.3: 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 osmosis 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.

Cell walls are semipermeable membranes, so the osmotic pressures of the body's fluids have important biological consequences. If solutions of different osmolarity exist on either side of the cells, solvent (water) may pass into or out of the cells, sometimes with disastrous results. Consider what happens if red blood cells are placed in a *hypotonic* solution, meaning a solution of lower osmolarity than the liquid inside the cells. The cells swell up as water enters them, disrupting cellular activity and eventually causing the cells to burst. This process is called *hemolysis*. If red blood cells are placed in a *hypertonic* solution, meaning one having a higher osmolarity than exists inside the cells, water leaves the cells to dilute the external solution, and the red blood cells shrivel and die. This process is called *crenation*. Only if red blood cells are placed in *isotonic* solutions that have the same osmolarity as exists inside the cells are they unaffected by negative effects of osmotic pressure. **Glucose solutions of about 0.31 M, or sodium chloride solutions of about 0.16 M, are isotonic with blood plasma.**

The concentration of an isotonic sodium chloride (NaCl) solution is only half that of an isotonic glucose ($C_6H_{12}O_6$) solution because NaCl produces two ions when a formula unit dissolves, while molecular $C_6H_{12}O_6$ produces only one particle when a formula unit dissolves. The osmolarities are therefore the same even though the concentrations of the two solutions are different.

Osmotic pressure explains why you should not drink seawater if you are abandoned in a life raft in the middle of the ocean. Its osmolarity is about three times higher than most bodily fluids. You would actually become thirstier as water from your cells was drawn out to dilute the salty ocean water you ingested. Our bodies do a better job coping with hypotonic solutions than with hypertonic ones. The excess water is collected by our kidneys and excreted.

Osmotic pressure effects are used in the food industry to make pickles from cucumbers and other vegetables and in brining meat to make corned beef. It is also a factor in the mechanism of getting water from the roots to the tops of trees!

Career Focus: Perfusionist

A perfusionist is a medical technician trained to assist during any medical procedure in which a patient's circulatory or breathing functions require support. The use of perfusionists has grown rapidly since the advent of open-heart surgery in 1953.

Most perfusionists work in operating rooms, where their main responsibility is to operate heart-lung machines. During many heart surgeries, the heart itself must be stopped. In these situations, a heart-lung machine keeps the patient alive by aerating the blood with oxygen and removing carbon dioxide. The perfusionist monitors both the machine and the status of the blood, notifying the surgeon and the anesthetist of any concerns and taking corrective action if the status of the blood becomes abnormal.

Despite the narrow parameters of their specialty, perfusionists must be highly trained. Certified perfusion education programs require a student to learn anatomy, physiology, pathology, chemistry, pharmacology, math, and physics. A college degree is usually required. Some perfusionists work with other external artificial organs, such as hemodialysis machines and artificial livers.

Summary

- Osmolarity is the total concentration of solutes in a solution
- i represents the number of particles that result per unit of solute.
  - i = 1 for molecular compounds
  - i = the number of ions per formula unit for ionic compounds
- Osmolarity = i x Molarity
  - If there are multiple solutes, add the results

Answers

Exercise $5.6.3$

Osmol $C_6H_{12}O_6$ = 1.5; Osmol $Al(NO_3)_3$ = 1.6

The solvent flows from $C_6H_{12}O_6$ solution (lower osmolarity) to $\mathrm{Al(NO_3)_3}$ solution (higher osmolarity).

Exercise $5.6.2$

i = 3 for $CaCl_2$ (Ca^{2+}, Cl^-, and Cl^-)

i = 2 for $NaNO_3$ (Na^+ and NO_3^-)

i = 1 for CH_4O (only nonmetals = molecular compound)

$Osmol = (0.15\ M \times 3) + (0.10\ M \times 2) + (0.20\ M \times 1)$

Osmolarity = 0.85 Osm

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6: ORGANIC CHEMISTRY

6.1: Introduction to Organic Chemistry
6.2: Representing Organic Molecules
6.3: Classifying Organic Molecules
6.4: Distinguishing Isomers

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6.1: Introduction to Organic Chemistry

Learning Objectives

- To define organic chemistry
- To classify compounds as organic or inorganic when given the chemical formula

Organic vs Inorganic

In chemistry the word “organic” doesn’t mean the same thing that it means at the supermarket. In the past, it was thought that some compounds could only be produced by living things and these were classified as organic. It turns out that is incorrect! Compounds found in living things CAN be produced synthetically. The first one to be created in the lab was urea which is found in urine. Since then, many more compounds, like ascorbic acid (vitamin C), have been produced in the lab.

So, a different definition of organic compounds was needed. Now we classify chemicals as organic if they contain the elements carbon and hydrogen bonded together. Methane (CH_4), sucrose (a sugar), proteins, and DNA are all examples of organic compounds because they contain C-H bonds. Pesticides are often organic compounds, so the definition of “organic” in chemistry is quite different from the way the word is used in daily life! Carbon dioxide (CO_2), water (H_2O), and calcium carbonate ($CaCO_3$) are not organic compounds because they contain either carbon or hydrogen, not both. Baking soda, which is sodium bicarbonate, $NaHCO_3$, is tricky because it contains both carbon and hydrogen, but they are not bonded together. The carbon is only bonded to oxygen and the H atom is also bonded to one of the oxygen atoms. So baking soda is inorganic.

Organic molecules do not necessarily come from living things under the modern definition of the term. Compounds containing carbon and hydrogen can be produced in the lab or may come indirectly from living things like petroleum (crude oil) that is produced when dead organisms are subjected to heat and pressure over a long period of time.

Exercise 6.1.1

Classify each compound as organic or inorganic.

1. NaCl
2. C_2H_6
3. $C_6H_{12}O_6$
4. K_2CO_3
5. $C_2H_5NO_2$

Answer 1 available at the end of this section.

The Importance of Carbon in Organic Chemistry

Why carbon? Carbon is very versatile! It has four valence electrons so it forms four covalent bonds. These could be four single bonds, a double bond and two single bonds, two double bonds, or a triple bond and a single bond. That means that carbon can have many different molecular geometries. What is the molecular geometry around carbon for each of the bonding patterns? (see VSEPR)

Carbon atoms can also form long chains or rings. Because of this versatility, carbon is found in a wide variety of molecules that perform a wide variety of functions. Its compounds are the basis of life on Earth.

Summary

- Organic compounds contain carbon and hydrogen bonded together (and often other elements as well)
- Organic molecules may contain long chains, rings, branched carbon chains, etc.

Answers

Exercise 6.1.1

Inorganic because it does not contain carbon and hydrogen
Organic because it contains both carbon and hydrogen
Organic because it contains both carbon and hydrogen
Inorganic
Organic

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6.2: Representing Organic Molecules

Learning Objectives

- To draw organic structures using structural formulas, condensed structural formulas, and skeletal formulas

Depending on the size of molecules, how much detail we need to show, and the resources available (paper and pencil vs. Microsoft Word vs. a program designed to draw chemical structures), we use different methods of representing organic molecules on paper. This section shows some of the formulas that will be used in this chapter.

Molecular Formulas

You are already familiar with molecular formulas which indicate the number of each atom, e.g. CH_4, CH_3O, etc. This is the simplest way to represent a molecule. Unfortunately, molecular formulas cannot uniquely identify big organic molecules because the only information provided is the number of atoms for each element. Organic molecules can have many atoms of just a few different elements with different bonding patterns or in chains of various sizes. In order to distinguish molecules that have the same number of atoms bonded in a different way, more detailed formulas are needed to uniquely identify a molecule.

Structural Formulas

Structural formulas *can* uniquely identify an organic molecule because they show all the atoms and how they are bonded to one another. They are the same as the Lewis structures you have already learned, except that lone pairs are omitted. You are expected to be able to tell where the lone pairs are because we keep working with the same elements over and over again in organic chemistry. Hydrogen and carbon never get any lone pairs. Nitrogen usually has one lone pair and oxygen usually has two lone pairs.

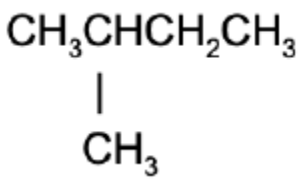
Condensed Structural Formulas

Some organic structures are quite large, and it can be time consuming to draw all of the atoms and bonds. In addition, chains branching off the main chain can get in the way of one another in a two-dimensional drawing that represents a three-dimensional structure. Condensed structural formulas are like a hybrid between molecular formulas and structural formulas. They group hydrogen atoms with carbon atoms because you always know how they will be bonded (hydrogens can only form single bonds). Therefore, you can concentrate on how the carbon atoms and any other atoms (oxygen, nitrogen, etc.) are bonded.

Example: Propane, C_3H_8, would be written as $CH_3CH_2CH_3$

Single bonds between carbon atoms may or may not be shown, so an equivalent condensed structural formula for propane would be CH_3-CH_2-CH_3. Any double or triple bonds must be shown explicitly as in the molecules propene, CH_2=$CHCH_3$, and acetylene, HC≡CH.

Branches are shown coming off the main chain, either above or below. The molecule 2-methylbutane can be represented by the following condensed structural formula:

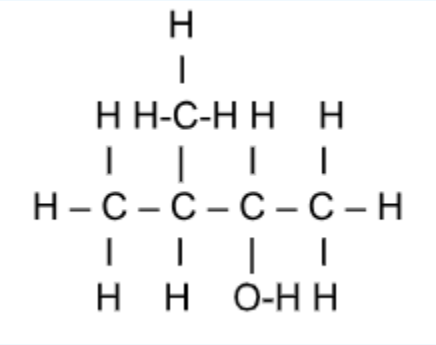


In large organic molecules with repeating structures, such as polymers which are discussed at the end of this chapter, identical groups may be written together. For example, n-hexane, $CH_3CH_2CH_2CH_2CH_2CH_3$, can be written as $CH_3(CH_2)_4CH_3$ where the four

CH_2 groups in a row are grouped together in parentheses with a subscript telling how many there are.

Example 6.2.1

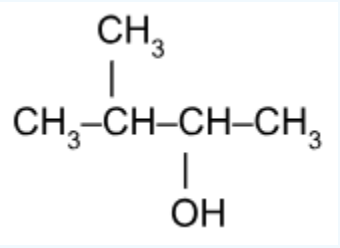
Draw the condensed structural formula for this compound:



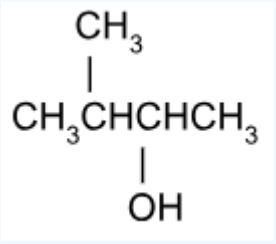
Solution

Look at each carbon (or oxygen, nitrogen, etc.) atom one at a time. Write the number of hydrogen atoms bonded to the atom. Adding bonds between the groups can make the structure easier to read, but they are not required. (Do show bonds going to groups written above or below the main chain).

1. The carbon on the left has three H atoms, so we write CH_3.
2. The second carbon atom has only one H atom, CH. Also bonded to the second carbon atom is a carbon atom with three H atoms, CH_3.
3. The third carbon atom has one H and an OH group is also attached.
4. The fourth carbon atom has three H atoms, CH_3.



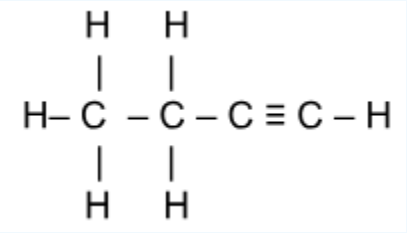
or



Notice that the carbon atoms are not grouped, only the hydrogen atoms. Bonds may or may not be shown, but all atoms are shown.

Exercise 6.2.1

Draw the condensed structural formula for this compound.



Answer available at the end of this section.

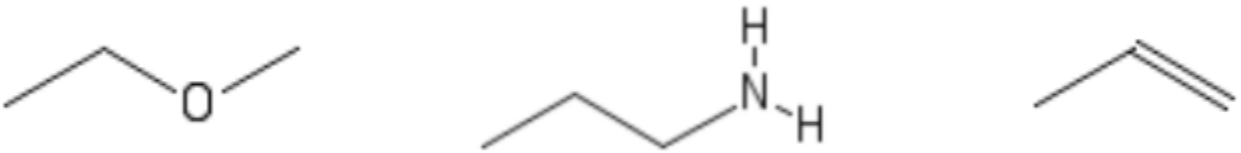
Skeletal Formulas

Skeletal formulas omit the hydrogens that are bonded to carbon. This means less writing, but also means that the reader is expected to know that carbon always forms four bonds. In addition, the carbon atoms are frequently represented as points at the ends of lines. Any other atoms present, like oxygen and nitrogen, must be shown explicitly (the hydrogens attached to these atoms should be shown). The following is a skeletal structure for 1-butane or CH_3-CH_2-CH_2-CH_3:



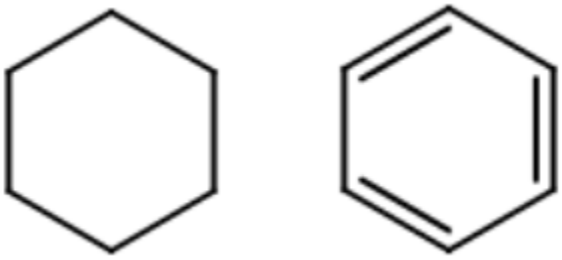
The four carbon atoms are represented by the ends of the three lines: the terminus on the left, the two vertices in the middle, and the terminus on the right. Because the carbon on the left is shown bonded to one thing, the neighboring carbon, the reader knows that there must be three hydrogen atoms singly bonded to the carbon. Likewise, because the second carbon atom is at the corner between two lines, it is known that it has two single bonds to carbon atoms. Therefore it must also have two hydrogen atoms.

Any other atoms present, or any double and triple bonds, must be shown explicitly as in the examples below.



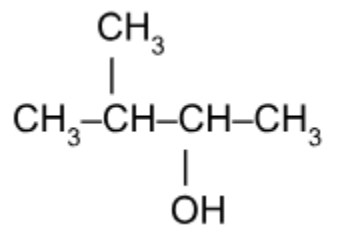
From left to right these skeletal structures represent: CH_3-CH_2-O-CH_3, CH_3-CH_2-CH_2-NH_2, and CH_3-CH=CH_2. Double check that you understand these drawings by counting the number of carbon atoms. Notice that when another atom is present, such as the oxygen atom in the first example and the nitrogen atom in the second example, the end of the line does *not* indicate a carbon atom. That is, there are only three carbon atoms in each of the examples above. Also double check that you can count the hydrogen atoms when they are not included in the drawing.

Remember that double bonds must be shown in skeletal structures. Six-membered rings of carbon are common in organic chemistry. The structures below represent cyclohexane, which contains all single bonds, and benzene which has alternating double and single bonds.



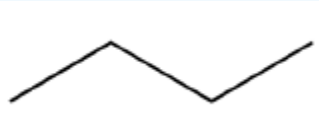
Example 6.2.2

Draw the skeletal formula for this compound:

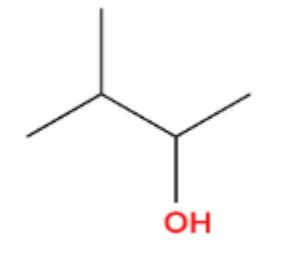


Solution

First count the number of carbon atoms that are bonded in a row without branches or backtracking. Draw a zigzag line with one end or corner for each carbon atom.



At the second carbon atom there is another carbon branching off. At the third carbon there is an OH group. Add these to the skeletal structure. Remember that H atoms are not shown if they are bonded to carbon, but are shown if they are bonded to anything else.



Exercise 6.2.2

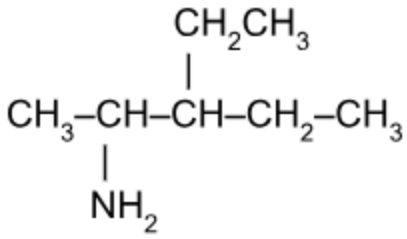
Draw the skeletal formula for this compound:

CH_3-CH_2-C≡CH

Answer available at the end of this section.

Exercise 6.2.3

Draw the skeletal formula for this compound:



Answer available at the end of this section.

Summary

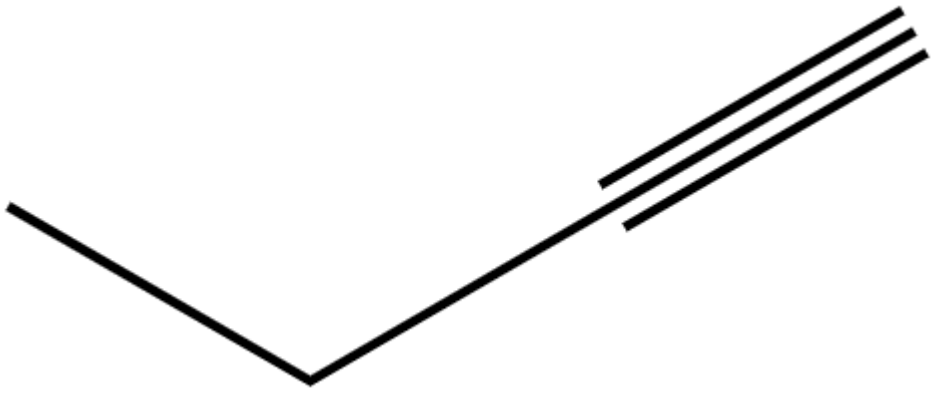
- Different types of drawings can be used to represent organic compounds.
- Structural formulas show all atoms and all bonds.
- Condensed structural formulas show all of the atoms, but not all of the bonds. Hydrogen atoms are grouped with the atom they are bonded to.
- In skeletal formulas carbon atoms all bonds are shown, but not all atoms are shown. The symbol C is never written. The symbol H is only written when bonded to an atom other than carbon.

Answers

Exercise 6.2.1

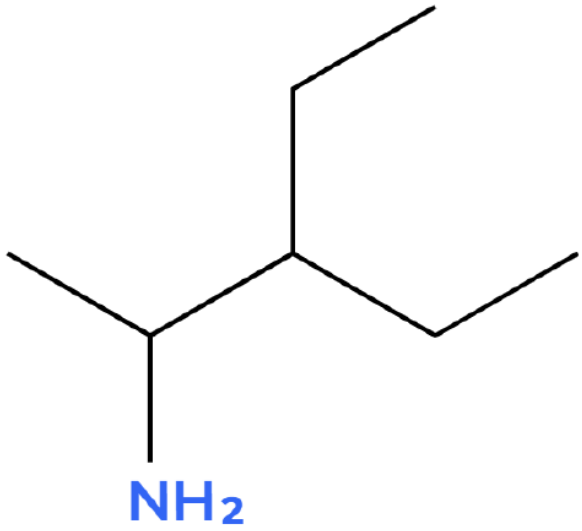
CH_3-CH_2-C≡CH or $CH_3CH_2C{\equiv}CH$

Exercise 6.2.2



There are carbon atoms at the end, at the corner, at the start of the triple bond, and at the end of the triple bond. The skeletal structure is showing that the molecule has a linear molecular geometry at the third carbon.

Exercise 6.2.3



Remember that H atoms must be shown if they are bonded to anything other than carbon. This structure has five carbon atoms in a row. The other two carbons are branching off of the third carbon.

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6.3: Classifying Organic Molecules

Learning Objectives

- To view a chemical structure for an organic compound and classify it based on functional groups

There are a lot of possible organic compounds! Even just carbon and hydrogen can be combined in many ways depending on how many carbons are in the chain or ring, whether there are any branches coming off the chain, and whether there are any double or triple bonds (compared to only single bonds). Chemists classify organic compounds based on bonding patterns in their structures. These patterns are important because structure determines function. That is, these patterns in structure determine what properties a compound will have and the types of chemical reactions in which it can participate. Therefore, by grouping organic compounds based on their structures, we are also classifying them based on their physical and chemical properties.

Hydrocarbons

If an organic molecule contains only carbon and hydrogen, then it is a hydrocarbon. Hydrocarbons are further classified based on what types of bonds they contain as shown in the table below.

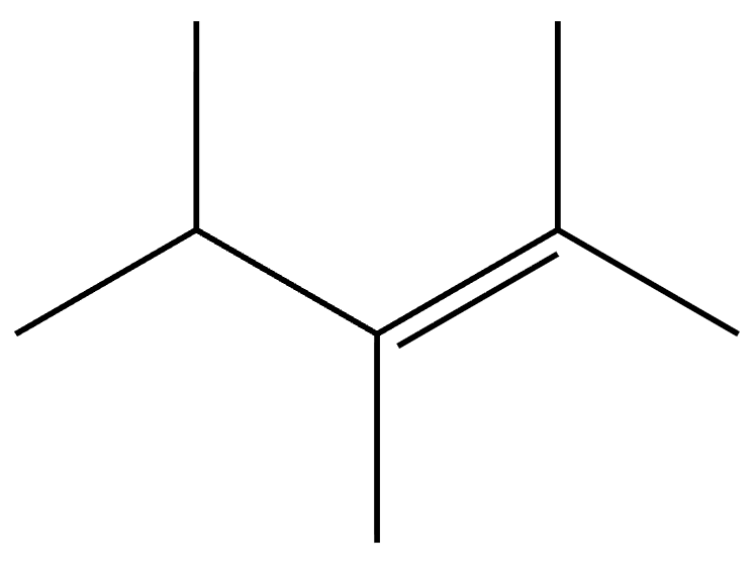
Hydrocarbon Category	Type of bonds
Alkane	All single bonds
Alkene	At least one double bond
Alkyne	At least one triple bond
Aromatic	Contains alternating single and double bonds

Notice that there is only one letter that is different in the first three names! You will find that suffixes are very important in organic chemistry. You should also remember that hydrogen can only form single bonds, so the double and triple bonds in alkenes and alkynes refer to C=C or C≡C bonds. Because this is only an introduction to organic chemistry, we will not deal with compounds that contain both double and triple bonds.

Alkanes are saturated hydrocarbons, meaning they have the maximum number of hydrogen atoms. Alkenes and alkynes are unsaturated hydrocarbons. Because they have at least one double or triple bond, they have fewer hydrogen atoms per carbon atom than alkanes do. The concept of saturated and unsaturated molecules will come up again in the Biochemistry chapter when we discuss saturated and unsaturated fatty acids.

Example 6.3.1

Classify the hydrocarbon below as an alkane, alkene, alkyne, or aromatic hydrocarbon.

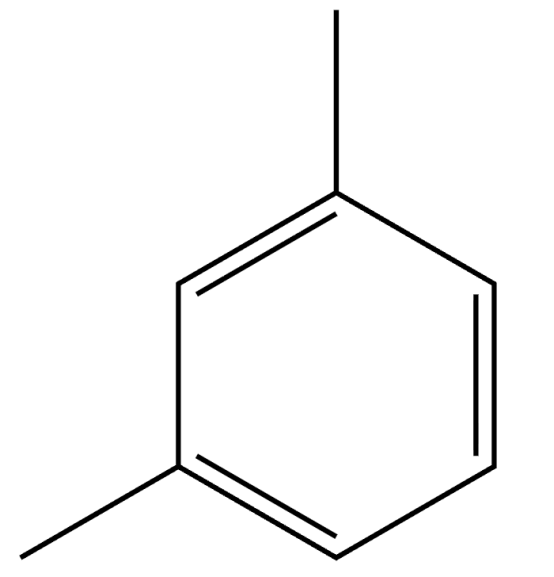


Solution

This compound contains only carbon and hydrogen and it has a double bond. It is an alkene. An alkene does not only have double bonds; frequently they have many single bonds as well. We distinguish between hydrocarbons that have at least one double bond and those that do not because double bonds are very reactive. Alkenes do reactions that alkanes do not.

Exercise 6.3.1

Classify this compound.



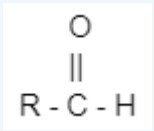
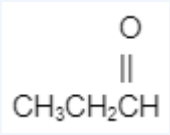
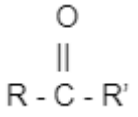
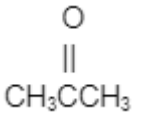
Answer available at the end of this section.

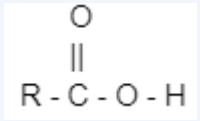
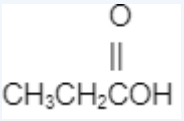
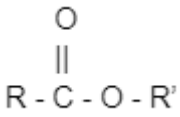
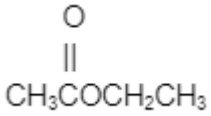
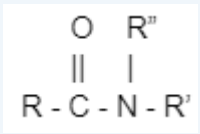
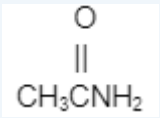
Functional Groups

Functional groups are what we call the structural patterns that we use to classify organic compounds. Remember that the

hydrocarbons discussed in the previous section are just one subgroup of organic chemicals. Organic compounds commonly contain other elements in addition to carbon and hydrogen such as oxygen and nitrogen. As you will see in the table below, functional groups are a pattern of certain elements bonded together with certain bonds (single, double, or triple). Because functional groups are only a portion of a molecule, we use a generic symbol to represent the rest of the molecule. Just like you commonly use x as a variable in algebra, we use R as a variable group in organic chemistry. So, H-O-R means that there is a hydrogen atom singly bonded to an oxygen atom and that oxygen atom is singly bonded to something else which has not been specified. R- will usually represent a chain of carbons and hydrogens. The example H-O-R could represent H-O-CH_3 or H-O-CH_2CH_3 or H-O-$CH_2CH_2CH_3$, etc. If the functional group connects to the rest of the structure in multiple places, then the symbol R can be modified with primes. Use R if there is one attachment, R' for a second attachment, and R'' for a third chain attached to the functional group.

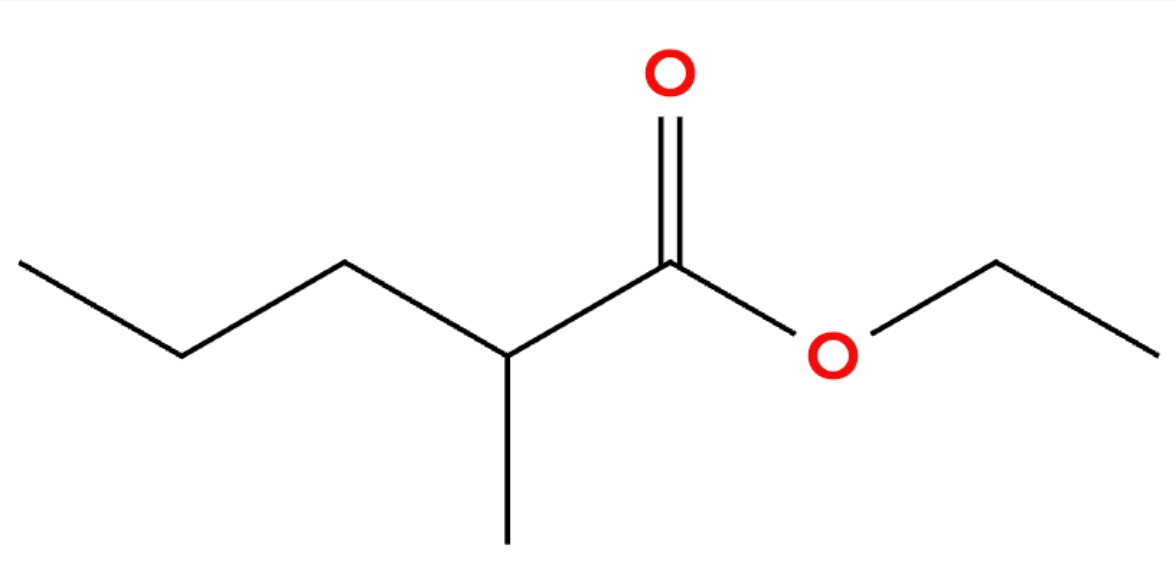
The table below lists the functional groups you will be expected to recognize for this course. Each functional group has a name that represents a structure that is commonly found as a portion of some organic molecules. Remember that these functional groups give the molecule certain properties, e.g. the ability to dissolve in water rather than a non-polar solvent, and the ability to participate in certain chemical reactions, e.g. the addition of a chlorine atom. Finally, you will see a suffix listed for each functional group. You will see these suffixes again later in this chapter when you learn how to name organic molecules. The suffix of the name indicates the functional group found in the molecule. Thus, when you see the name of a new molecule you can quickly get an idea of some of its properties based on its suffix.

Functional Group Name	Structure	Suffix(es)	Example
Alcohol	R-OH R cannot be H	-ol	Ethanol CH_3CH_2OH
Ether	R-O-R' R/R' cannot be H	-oxy -ane (two parts)	Methoxyethane CH_3-O-CH_2CH_3
Amine		-amine	Ethanamine $CH_3CH_2NH_2$
Aldehyde	 or abbreviated CHO	-al	Ethanal 
Ketone	 R/R' cannot be H	-one	Propanone 

Functional Group Name	Structure	Suffix(es)	Example
Carboxylic Acid	 or abbreviated COOH	-ic acid (two words)	Propanoic acid 
Ester	 R' cannot be H	-yl -oate (two words)	Ethyl ethanoate 
Amide		-amide	Ethanamide 

Example 6.3.1

Classify this compound based on its functional group.



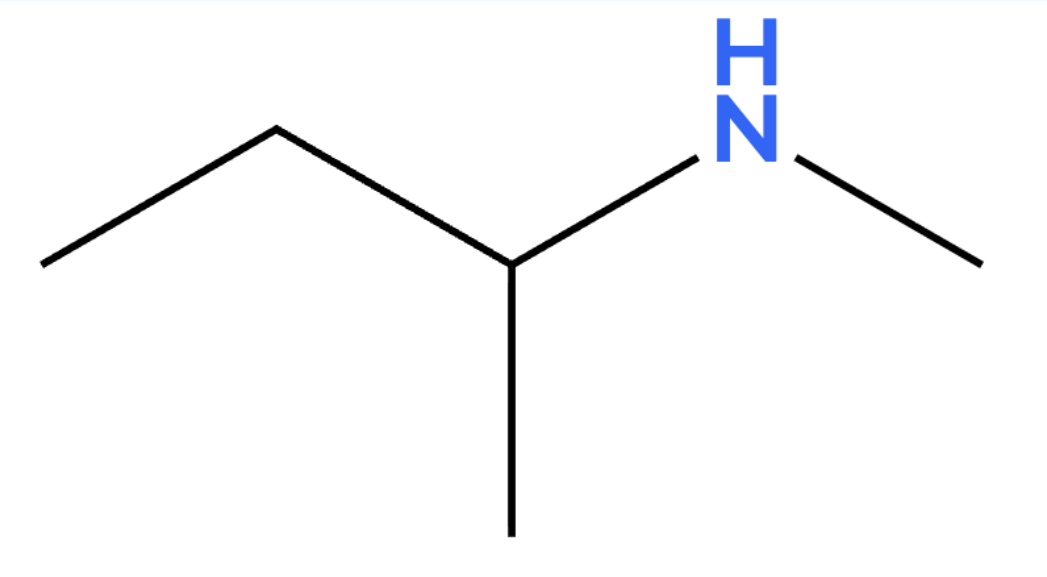
Solution

There are five functional groups listed above that have a C=O. To distinguish between them, look to see what elements are bonded immediately next to the C=O. In this case there is a carbon on one side and an oxygen atom on the other; that narrows the options down to either carboxylic acid or ester.

Both carboxylic acids and esters have a carbon with a double bond to one oxygen atom and a single bond to another. The difference is what is on the other side of the single-bonded oxygen. Carboxylic acids have an OH next to the C=O whereas esters have at least one carbon atom on the other side of the single-bonded oxygen. This compound is an ester.

Exercise 6.3.1

Classify this compound based on its functional group.



Answer available at the end of this section.

Summary

- Hydrocarbons contain only carbon and hydrogen. They can be further classified by whether they contain double bonds, triple bonds, or only single bonds.
- A functional group is a portion of a chemical structure that gives the compound certain properties.

Answers

Exercise 6.3.1

This is an aromatic hydrocarbon. Although it has double bonds, it is not considered an alkene because this specific pattern of alternating double bonds is very stable. The three double bonds in this ring are not reactive like the double bonds in an alkene, so this pattern of six carbons in a ring with alternating double bonds gets its own category: aromatic ring.

Exercise 6.3.1

This compound contains a nitrogen atom, so that narrows it down to either an amine or an amide. An amide has a C=O immediately next to the N but an amine does not. Therefore, this is an amine.

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6.4: Distinguishing Isomers

Learning Objectives

- To view two or more chemical structures and determine if they represent the same compound or isomers
- To distinguish between structural isomers, geometric isomers, and optical isomers

Isomers

For organic molecules with only a few carbon atoms such as CH_4, C_2H_6, and C_3H_8, there is only one possible arrangement that contains all of the atoms listed in the molecular formula and follows the bonding rules to make a stable structure.

Now consider a molecular formula with more atoms: C_5H_{12}. This is an alkane with all single bonds. However, the molecular formula does not tell us if all of the carbon atoms are in one chain or whether there are any branches.

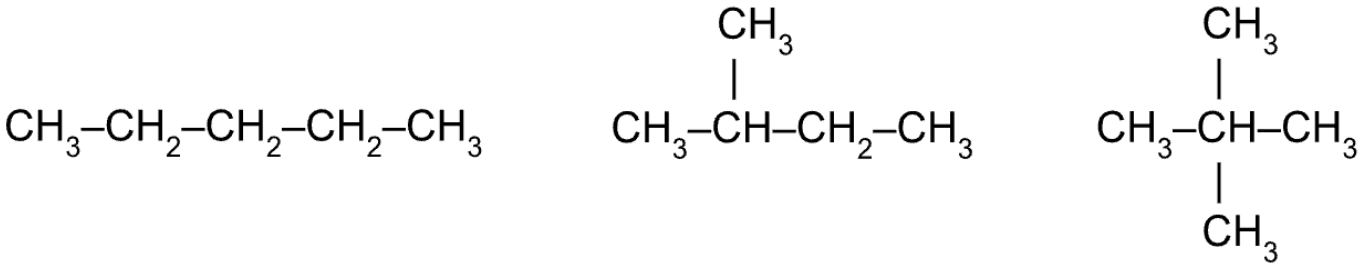
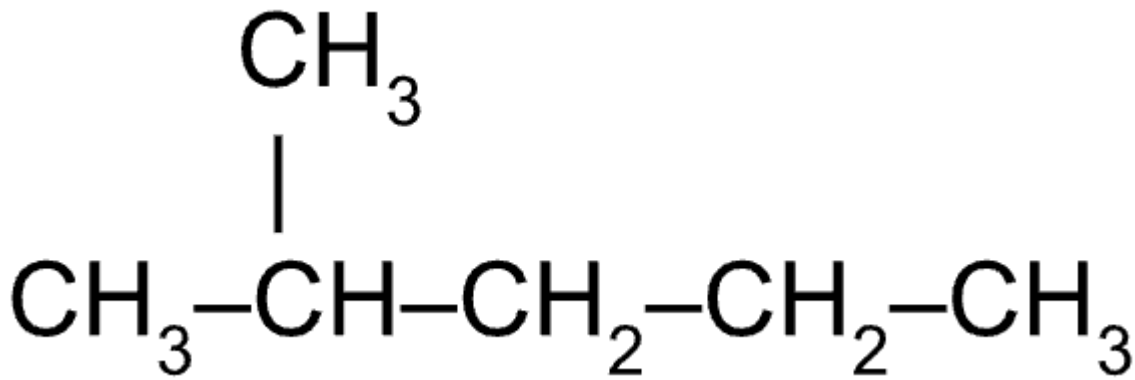


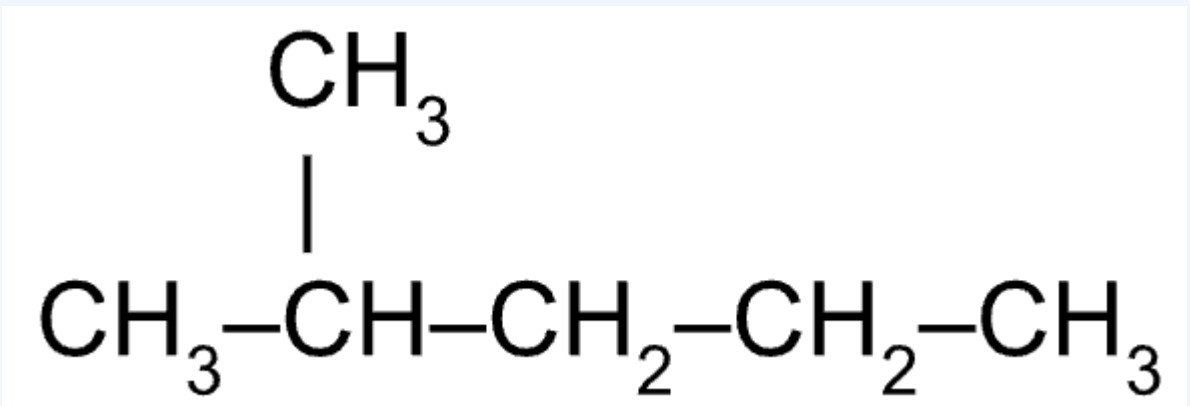
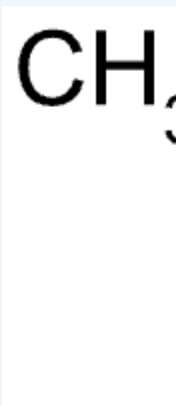
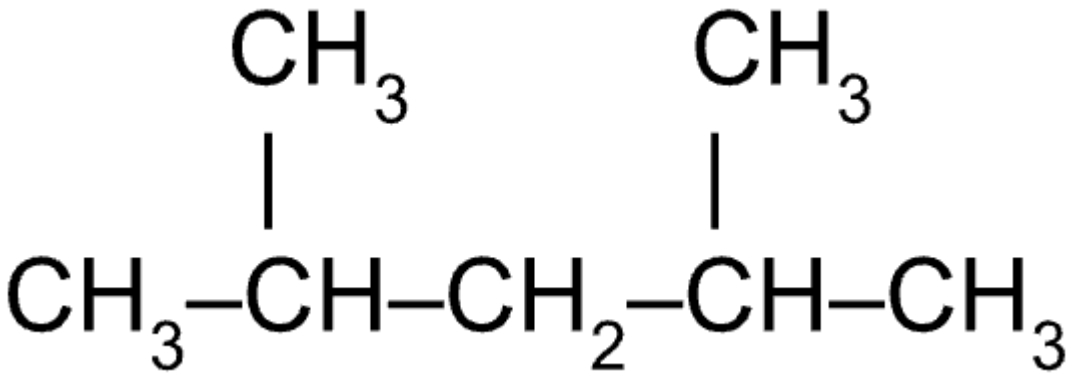
Figure 6.4.1: Three isomers of C_5H_{12}.

The three compounds above are isomers. The prefix "iso" means "same". Isomers are two or more compounds that have the same atoms, but arranged in a different way. The different arrangements can affect properties such as melting point, boiling point, and which bonds are most likely to change in chemical reactions.

Before we go farther, remember from your previous experience drawing Lewis structures that they do not accurately represent the three-dimensional structure of a molecule. This is important in distinguishing isomers because the same molecule can be drawn differently.

Isomers must have different bonding, not just different drawings.

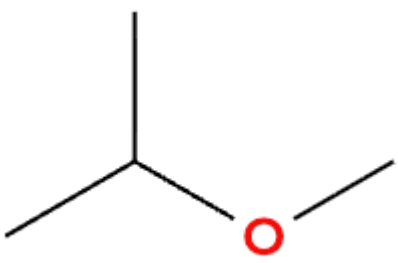
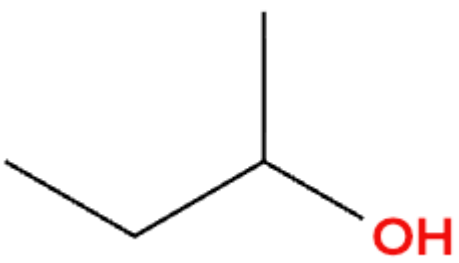
Structure 1	Structure 2
	

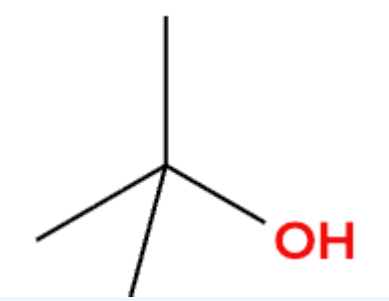
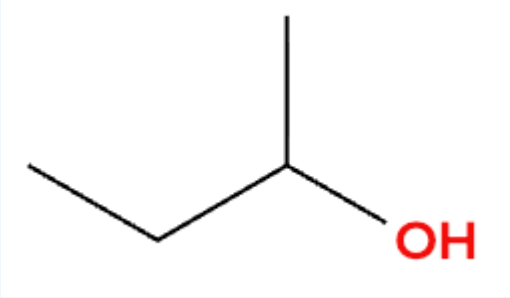
Structure 1	Structure 2
	
	

Structural Isomers

Structural isomers have atoms bonded in a different order. For instance, there may be branches in different places such as in Figure 6.4.1.

Structural isomers may or not have the same functional groups.

Structure 1	Structure 2	Relationship
		Structural Isom They are both $C_4H_{10}O$, but the atoms are bonc different places result Structure an ether and Structure 2 is a alcohol.

Structure 1	Structure 2	Relationship
		Structural Isomers. They both $C_4H_{10}O$, t the atoms are bonded in diffe places. Both Structure 1 and Structure 2 are alcohols, but th reactions they undergo are different becau the different arrangement o carbon atoms. Note: Structure this row is also isomer of Struc in row 1. There be many isome a compound.
		Structural Isom Both are C_4H_8 Structure 1 has double bond at end of the carb chain whereas Structure 2 it is the middle. Note: The diffe may be easier using Condens Structural Form Structure 1, CH_3-CH_2-CH has one CH_3 g and one CH gr whereas Struct CH_3-CH=CH-C has two CH_3 gr and two CH gr

Stereoisomers

Not all isomers are structural isomers. Unlike structural isomers, stereoisomers have the same number of each type of atom and the atoms are bonded in the same order. The difference between stereoisomers their 3D arrangement. Geometric isomers and optical isomers are subcategories of stereoisomers.

Geometric Isomers

Geometric isomers are sets of molecules with the same bonding patterns, but some restriction in rotation that makes the atoms or groups point in different directions in the different versions. In other words, geometric isomers have a different 3D arrangement around a bond that cannot rotate.

Just like in a model kit, double bonds in molecules cannot rotate. Therefore the two molecules below are isomers, not the same compound.

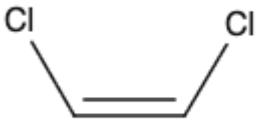
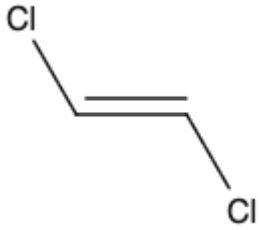
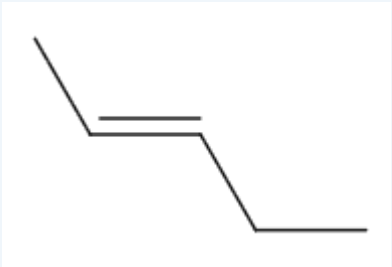
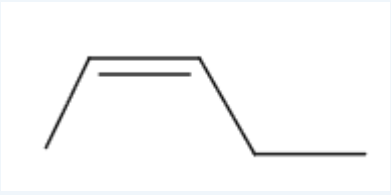


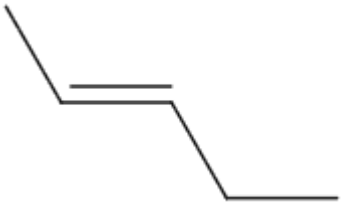
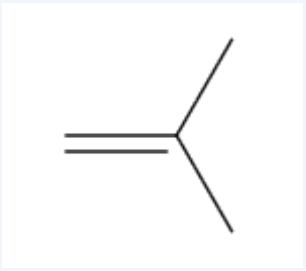
Figure 6.4.2: The two geometric isomers of C_4H_8. The structure on the left is the trans isomer and the structure on the right is the cis isomer.

The structure on the left in Figure 6.4.2 has one end of the carbon chain pointing up from the double bond and the other pointing down from the double bond in a zigzag shape. This is called the trans isomer. "Trans" means across and the carbon chain is going across the double bond: up on one side and down on the other. Because the double bond cannot rotate, this compound cannot flip or rotate to get the same 3D arrangement as the structure on the right.

The structure on the right in Figure 6.4.2 has both ends of the carbon chain pointing up. This is the cis isomer. "Cis" means same and both ends of the carbon chain are pointing up (or down, depending on your perspective) relative to the double bond.

Additional examples of geometric isomers

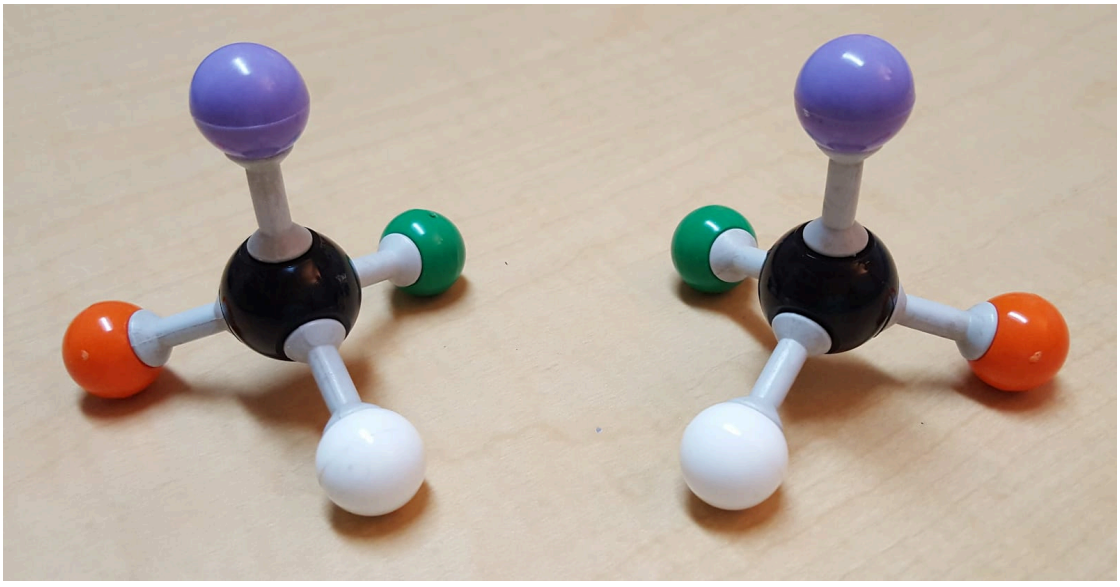
Structure 1	Structure 2	Relationship
		Geometric isomers. In this case we compare the 3D arrangement of the chlorine atoms rather than carbon chains. Structure 1 is the cis isomer and Structure 2 is the trans isomers. Geometric always come in pairs: one cis isomer and one trans isomer with everything else the same.
		Geometric isomers. It does not matter that the length of the chain is different on the left and right of the double bond. Structure 1 is the trans isomer and Structure 2 is the trans isomer.

Structure 1	Structure 2	Relationship
		Not isomers. These two structures have a different number of atoms.
		Structural isomers. Both of these structures are C_4H_8, but the atoms are bonded in a different order. In Structure 1 the double bond is on the end of the chain. It is neither a cis nor trans isomer because the carbon chain goes neither up nor down on the left (and it goes both up and down on the right). Note: Geometric isomers must have atoms bonded in the same order. Everything must be the same except that one isomer is the cis configuration and the other is trans.

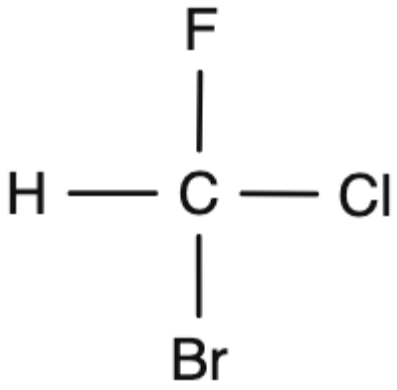
Optical Isomers (Enantiomers)

The last category of isomers that we will examine are optical isomers which are pairs of molecules called enantiomers.

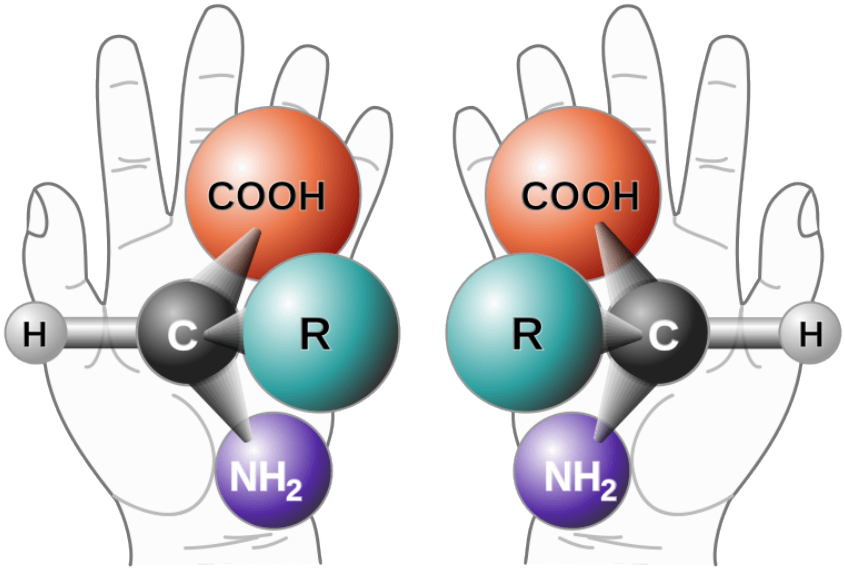
At first glance, the two models below may seem to represent the same molecule.



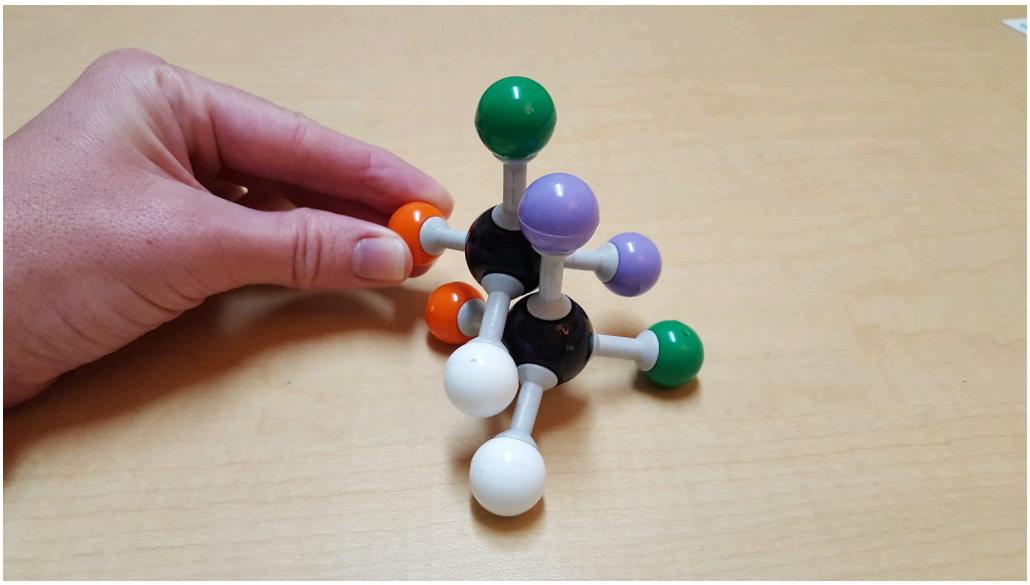
In fact, they do have the same molecular formula, CHBrClF. However, there is a difference between them. Enantiomers are pairs of chiral molecules (pronounced Ki-ral like in kayak). A chiral molecule has a carbon atom with four different groups or atoms attached to it such as:



Molecules with these **chiral centers**, as the carbon atom in this molecule is called, have non-superimposable mirror images. The image below shows a chiral amino acid (the R group is not the same as any of the other three groups) emphasizing that the molecule has a mirror image, just like your hands.



Just as your hands are not superimposable, neither are the two molecules in the picture above. You cannot overlap your two hands with the palms facing the same direction and have the thumbs and fingers line up. If you line up the fingers and thumbs, then your palms are facing in opposite directions. The same is true for chiral molecules. In the photo below, I am able to align the carbon, hydrogen, and bromine atoms (black, white, and orange), but when I do so, the fluorine and chlorine atoms (purple and green) do not match up.



There is no way that I can turn the models to align all five atoms. The only way to make them the same would be two break two of the bonds and reform them in different positions. The two molecules in the photo are enantiomers. They have the same chemical formula and the same bonding pattern (in each case H, Br, Cl, and F are attached to the carbon atom with single bonds), but they are nonsuperimposable mirror images of one another.

Enantiomers generally have the same physical and chemical properties. However, they can interact differently with enzymes in the body. There are twenty naturally occurring amino acids of which 19 are chiral. All 19 of these have the groups arranged around the chiral center in the same way which we designate L- as in L-alanine or L-leucine. Sugars are also chiral and all of the ones that humans can digest are the same form and are designated D-, such as D-glucose. Amino acids and sugars will be introduced in the next chapter, Introduction to Biochemistry, but the difference between the L- and D- designations is a topic for a more advanced course.

Summary

- There is not one correct way to represent a chemical structure on paper. Images of the same compound may appear flipped or rotated. Also remember that single bonds can rotate and that 2D representations of structures usually do not show the correct bond angles.
- Isomer means "same parts". Structures are isomers of one another if they have the same number of atoms of each element, but arranged in a different way.
- In structural isomers, atoms are bonded in a different order. For example, one isomer may have six carbon atoms in a single chain whereas another may have five carbon atoms in a chain with a sixth branching off of that chain. The location of double bonds may also be different in structural isomers. For instance, one isomer may have a double bond between the first and second carbons while a structural isomer may have the double bond between the second and third carbons.
- Stereoisomers have all of the same atoms bonded in the same order, but with a different 3D arrangement. In geometric isomers, there is a different 3D arrangement around a double bond. One isomer has a cis configuration and the other one is trans. Optical isomers have a different 3D arrangement around a chiral center which is a carbon atom bonded to four different atoms or groups.

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7: BIOCHEMISTRY

Biochemistry is the study of the chemicals and chemical processes found in living things. In this chapter you will learn about some classes of chemicals that you have heard of before: Carbohydrates, Lipids, Nucleic Acids, and Proteins. A lot of topics that you have studied in other chapters are relevant here. Many biochemicals can be identified by combinations of functional groups that you studied in the chapter on Organic Chemistry (for example, amino acids contain both amine and carboxylic acid groups). Lipids on the other hand, are defined by their solubility in nonpolar solvents. They are either mostly or entirely nonpolar molecules. The double helix of DNA is held together by hydrogen bonds. Finally, many of the biochemicals we will study are polymers (or the monomers that make them up). Polymer means "many parts", so expect to see some big molecules in this chapter!

7.1: Carbohydrates

7.2: Lipids

7.3: Proteins

7.4: Nucleic Acids

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7.1: Carbohydrates

Learning Objectives

- To identify a chemical structure as a carbohydrate
- To classify carbohydrates as monosaccharides, disaccharides, or polysaccharides

Recognizing & Classifying Carbohydrates

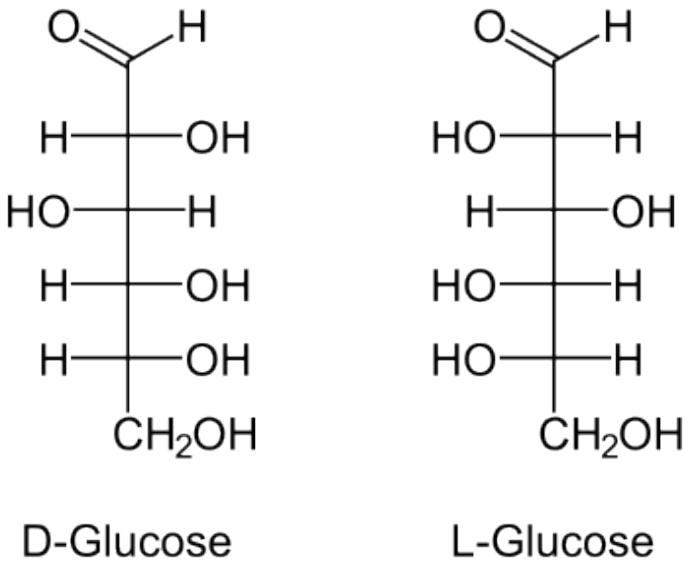
Carbohydrates contain only three elements: carbon, hydrogen, and oxygen. They also contain a limited number of functional groups. Each carbohydrate contains several alcohol groups and at least one aldehyde or ketone group. Despite the similarities between carbohydrate molecules, there are still many different carbohydrates due to the wide variety of sizes among carbohydrates. Because some carbohydrates are polymers these biomolecules can be classified as monosaccharides, disaccharides, and polysaccharides.

In addition to being the building blocks of larger carbohydrates, monosaccharides are often found in fruits, cane sugar, and honey. Examples of these molecules, also called simple sugars, include fructose and glucose. Monosaccharides can be classified by the number of carbon atoms (e.g. 3 = tri, 4 = tetra, etc.) and whether they contain an aldehyde or ketone functional group (an aldo- or keto- prefix is used). As you will see in the table below and throughout the carbohydrate section, names of these molecules often end in -ose.

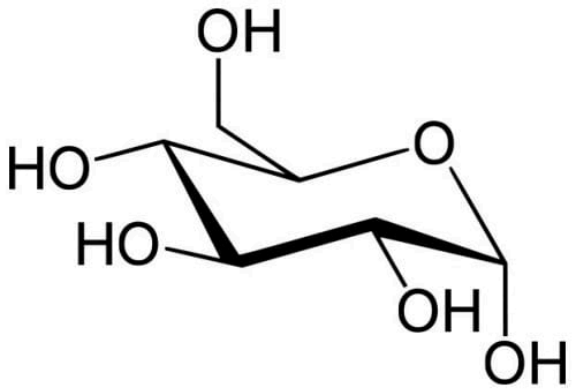
Classification of monosaccharides

Number of Carbon Atoms	Aldose	Ketose
3	aldotriose	ketotriose
4	aldotetrose	ketotetrose
5	aldopentose	ketopentose
6	aldohexose	ketohexose
7	aldoheptose	ketoheptose

The images shown are called Fischer projections. They are similar to the skeletal structures introduced in the Organic Chemistry chapter, but also provide information about chirality. At each intersection of the lines there is a carbon atom; all other atoms are shown explicitly. The L- and D- forms of glucose are two different isomers (enantiomers).



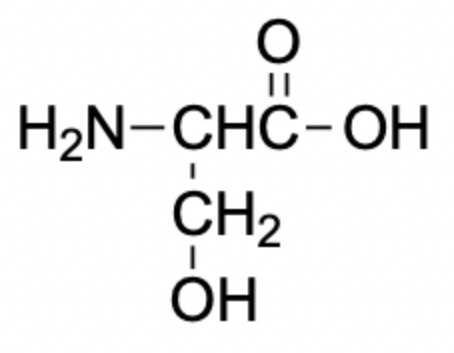
Each monosaccharide has several chiral carbon atoms. (Can you identify them in the examples above? Remember that a chiral center is a carbon atom with four different groups attached.) To distinguish between different forms prefixes L- and D- are used, for example L-glucose and D-glucose. Interestingly, most monosaccharides found in nature are the D- form. Monosaccharides often exist in a cyclic form rather than the straight chain forms. The cyclic form of D-glucose is shown below.

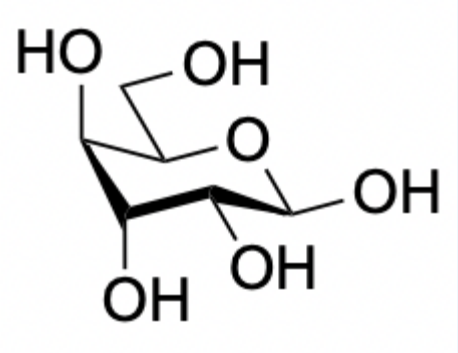


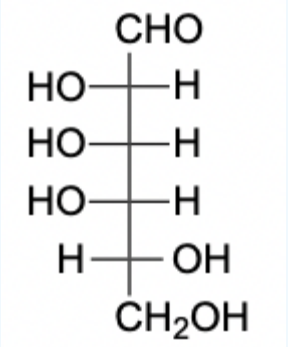
The ring is formed by a reaction that forms a single bond between the oxygen on carbon 5 and carbon 1 (where the aldehyde group was). All of the same atoms are present, but the hydrogen atom that had been part of the hydroxyl (-OH) group on carbon 5 is now bonded to the oxygen atom on carbon 1 (formally the carbonyl oxygen).

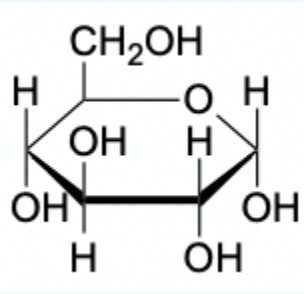
Exercise 7.1.1

Which of the following are monosaccharides? Choose all that apply.





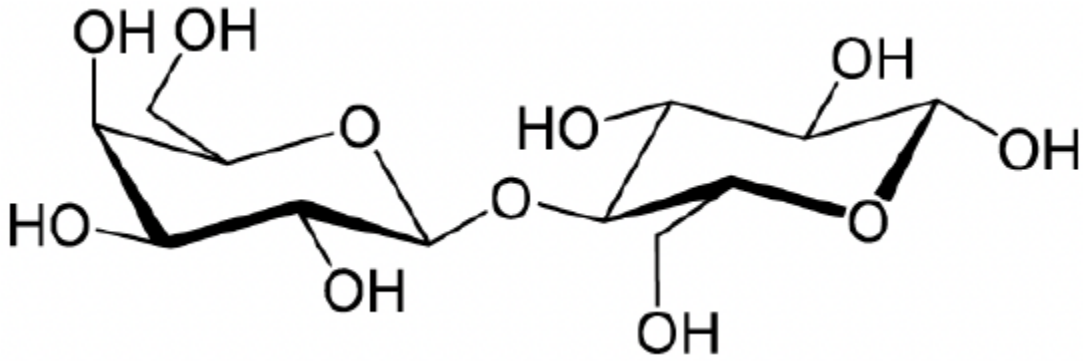




Answer available at the end of this section.

Disaccharides

As the name implies, disaccharides are carbohydrates made from two monosaccharides. A common example is lactose which contains the monomers glucose and galactose.



Oligosaccharides

Oligosaccharides are moderately large carbohydrates consisting of 3 - 10 monosaccharides bonded together. Each monosaccharide is in its cyclic form and the rings are bonded together by an oxygen atom with two single bonds: -O-. Examples of oligosaccharides include the A and B antigens on red blood cells that determine blood type.

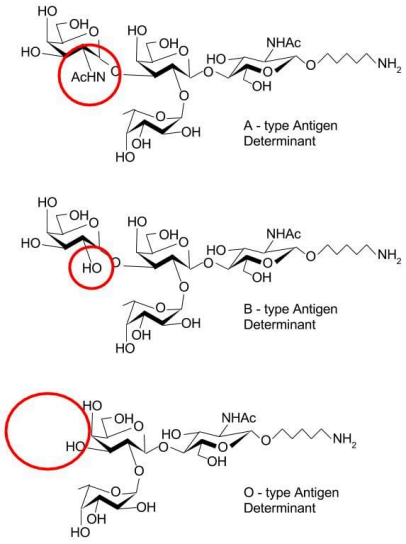
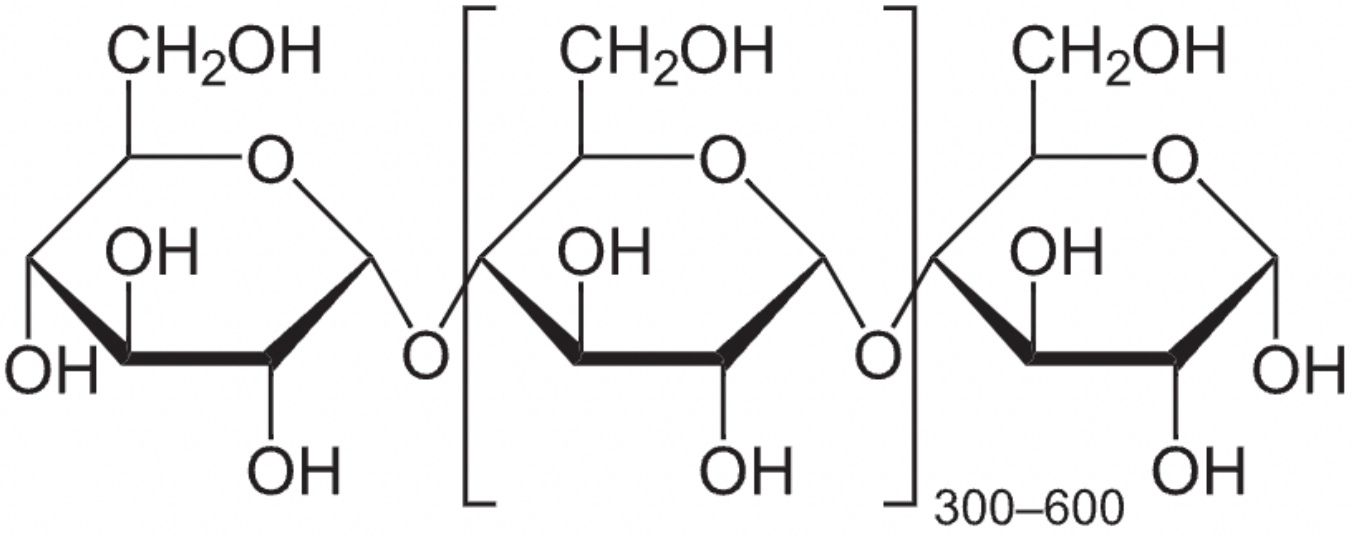


Figure 7.1.1: Oligosaccharides that form antigens and result in type A, B, AB, and O blood.

Polysaccharides

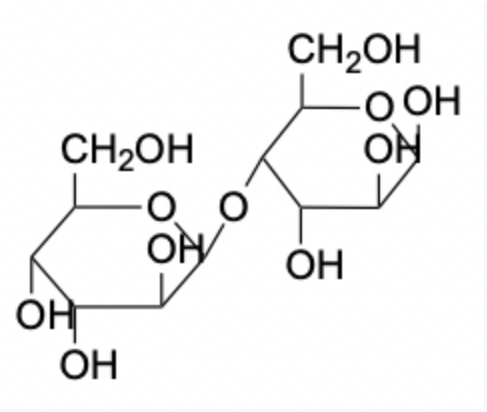
Polysaccharides are natural polymers consisting of many monosaccharide units. Examples include starch, cellulose, and glycogen. Starch is a complex carbohydrate found in potatoes and other foods. Cellulose gives structure to cell walls in plants and cannot be digested by humans because we cannot break the bond between the monosaccharide units. Glycogen is a polysaccharide produced in the body to store excess glucose for later use.

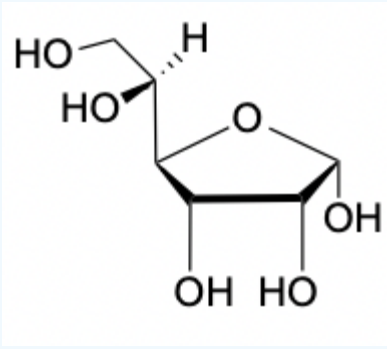
Amylose, a starch, is a polymer made from glucose monomers. Notice that the figure below has a glucose monomer in brackets with a subscript of 300-600 indicating that there are hundreds of glucose monomers in a chain of amylose.

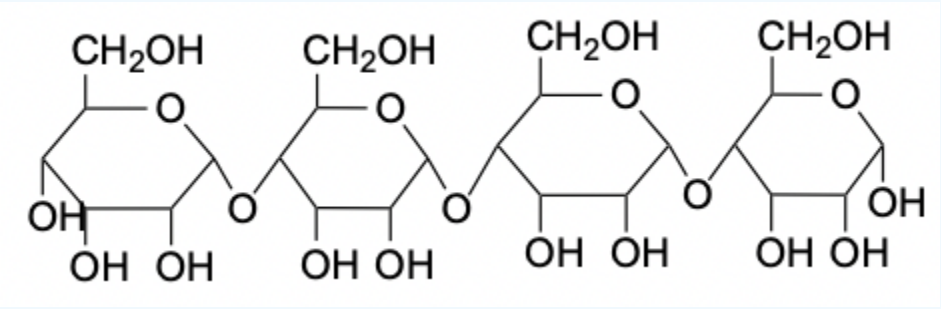


Exercise 7.1.2

Classify the following carbohydrate as a monosaccharide, disaccharide, or oligosaccharide.







Answer available at the end of this section.

Summary

- Carbohydrates generally contain carbon, hydrogen, and oxygen with an approximately equal number of carbon and oxygen atoms. They have many hydroxy groups (OH).
- Monosaccharides may be cyclic or acyclic. If cyclic, they contain only one ring. Monosaccharides have either an aldehyde or a ketone functional group in addition to many hydroxy groups.
- Disaccharides consist of two monosaccharides bonded together. They have two rings connected by an oxygen atom with two single bonds.
- Oligosaccharides and polysaccharides contain many monosaccharides bonded together. Each monosaccharide is a ring, typically of 5 - 6 atoms, and the rings are bonded together by the same -O- pattern found in disaccharides.

Answers

Exercise 7.1.1

Structures 2, 3, and 4 are monosaccharides. Carbohydrates have the elements C, H, and O with approximately equal numbers of C and O. Almost every carbon atom has an OH group. Remember that the carboxylic acid group in Structure 1 is not the same as the alcohol functional group. Monosaccharides may be either cyclic or acyclic.

Exercise 7.1.2

- Structure 1 is a disaccharide because it has two monosaccharides bonded together (each of the rings is a monosaccharide).
- Structure 2 is a monosaccharide it is a carbohydrate consisting of only one ring.
- Structure 3 is an oligosaccharide because it has more than two monosaccharides bonded together.

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7.2: Lipids

Learning Objectives

- To look at the chemical structure of a biomolecule and identify it as a lipid
- To distinguish between fatty acids, phospholipids, triacylglycerols, and steroids

Identifying Lipids & Types of Lipids

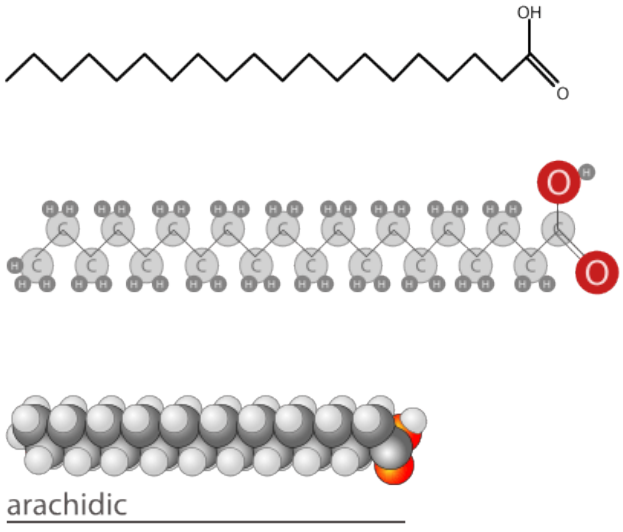
Unlike other classes of biomolecules, lipids have a wide variety of structures. The thing that they have in common is a property rather than a structure: **lipids are soluble in nonpolar solvents.** Therefore, you can generally recognize lipids by looking for a large nonpolar region, i.e. lots of C-C and C-H bonds.

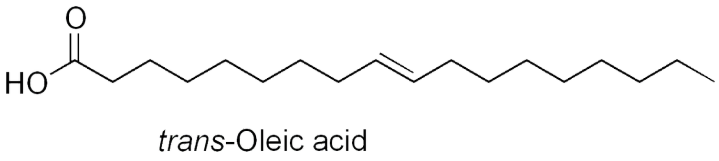
Lipids can generally be sorted into two categories: those that are made from fatty acids and those that are not. In this course we will examine triacylglycerols and phospholipids (which are made from fatty acids) and steroids (which are not).

Fatty Acids and Triacylglycerols

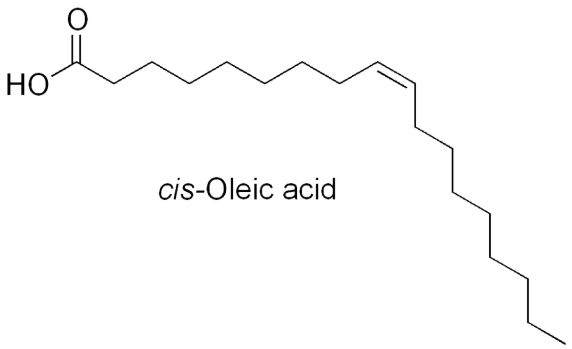
Fatty acids are long molecules composed of a carbon chain (with an appropriate number of hydrogen atoms) and a carboxylic acid group on one end. The long chain, consisting of 12 - 20 carbon atoms, gives the fatty acid its ability to dissolve in nonpolar solvents. The carboxylic acid end allows the fatty acid to react with glycerol to form triacylglycerides.

Relative melting points for fatty acids can be estimated based on the length of their carbon chains and the number of cis double bonds. As you saw with hydrocarbons, longer carbon chains can experience stronger (London) dispersion forces and therefore require more heat to melt and therefore have higher melting points. Saturated fatty acids have higher melting points than unsaturated fatty acids because the molecules can align better to form stronger dispersion forces. Double bonds, especially cis- double bonds, cause the chains to bend. Compare the structures of arachidic acid on the left and the two oleic acid structures on the right.





trans-Oleic acid

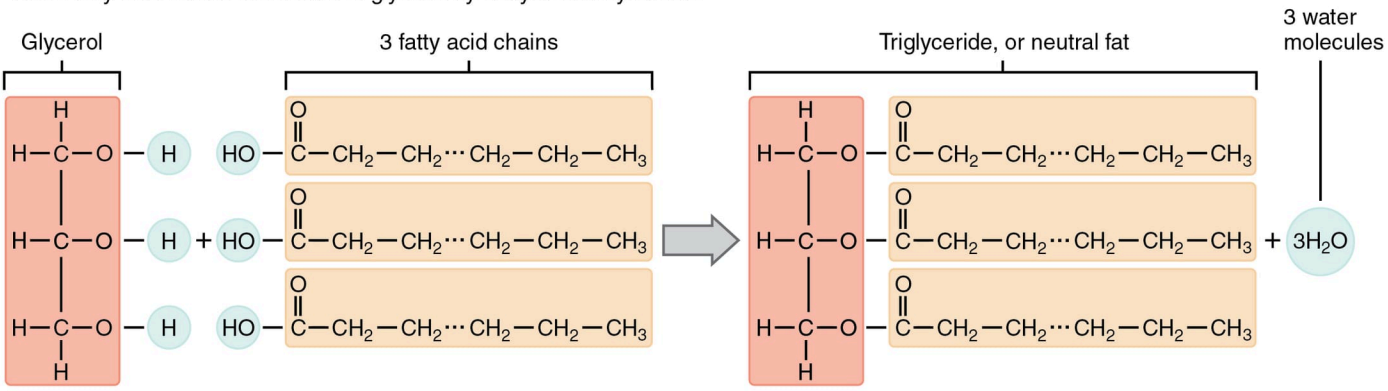


cis-Oleic acid

Arachidic acid is a saturated fatty acid where all the C-C bonds are single bonds. The oleic acid fatty acids are unsaturated due to their double bonds. Of these molecules the cis-oleic acid has the lowest melting point (13.4 ℃) because the molecules cannot line up well. The dispersion forces are therefore weaker in cis-oleic acid compared to trans-oleic acid (melting point = 43 ℃) and arachidic acid (melting point = 75.5 ℃).

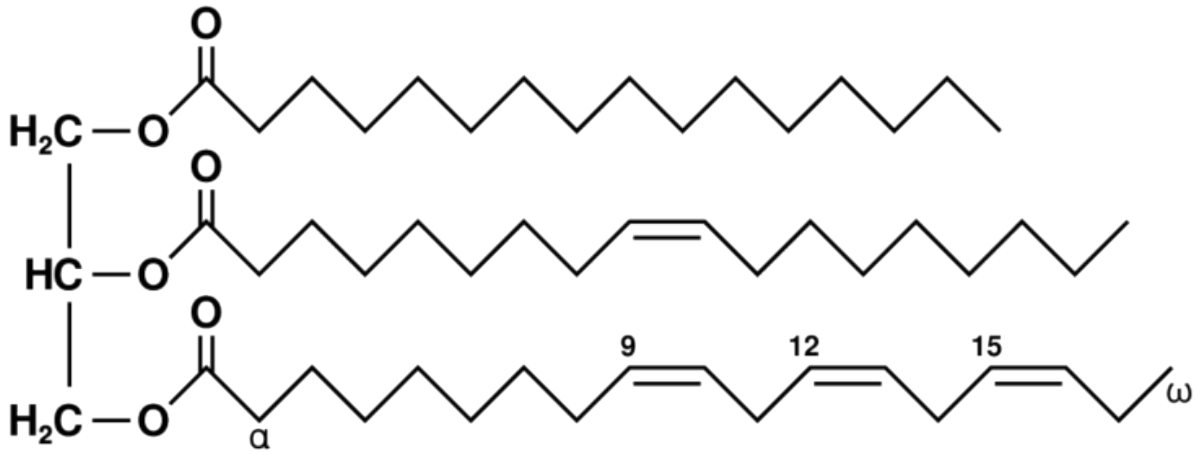
The triacylglycerols formed when fatty acids react with glycerol are also called triglycerides and are frequently measured in blood tests to determine heart disease risk. The chemical equation below shows how glycerol (a molecule with three carbons and an alcohol functional group on each carbon) reacts with three fatty acids to form a triacylglycerol (or triglyceride) and three water molecules.

Three fatty acid chains are bound to glycerol by dehydration synthesis.



Source: Anatomy & Physiology, Connexions Web site. http://cnx.org/content/col11496/1.6/, Jun 19, 2013.

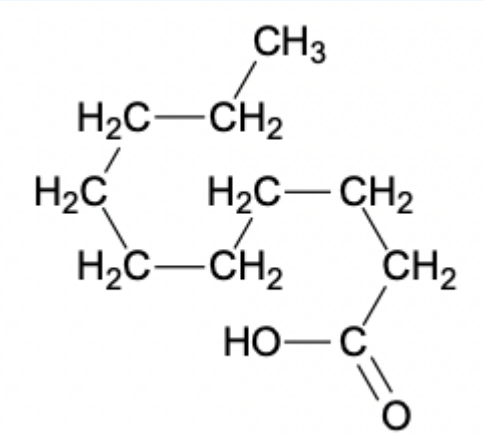
The structure below shows a triacylglycerol. The three carbons on the left came from the glycerol. Each one reacted with a fatty acid, which is where the “tri” comes from in triglyceride. The remnants of the fatty acids’ carboxylic acid groups can be seen connected to the three carbon atoms. Towards the right are three long nonpolar carbon chains. Remember that these line structures have a carbon atom at each corner and at the end of the line. The triacylglycerol has three ester functional groups which formed when the alcohol groups of glycerol reacted with the carboxylic acid groups of the fatty acids (the extra oxygen and hydrogen atoms formed water as shown in the chemical equation above).

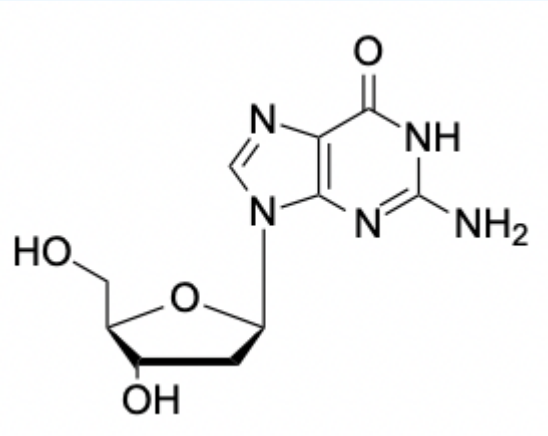


If the carbon chains contain one or more double bonds, such as the second and third chains in the figure above, then they are unsaturated fatty acids. If all the bonds are single bonds, as in the first carbon chain, then it is a saturated fatty acid. That is, it is saturated with hydrogen atoms.

Exercise 7.2.1

Classify each of the following as a lipid or non-lipid.



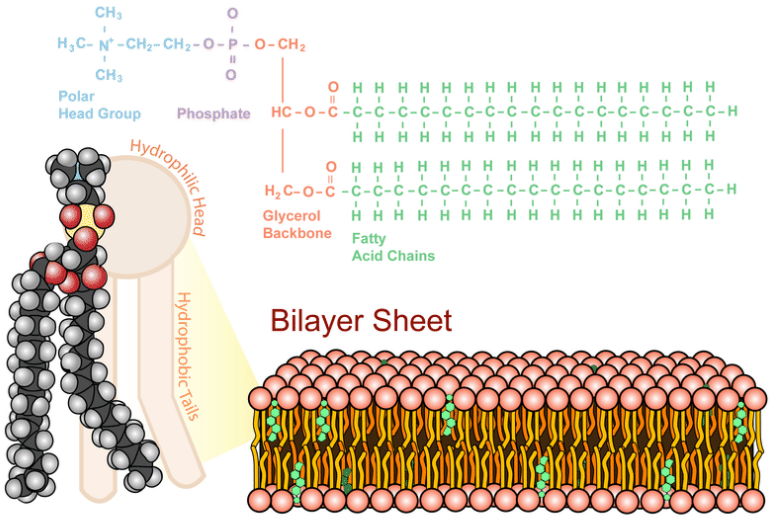


Answer available at the end of this section.

Phospholipids

Phospholipids have a structure similar to triacylglycerides. Both are formed from a glycerol molecule and fatty acids, but whereas the triacylglycerols contain three fatty acid chains, the phospholipids contain two fatty acids and a phosphate/alcohol group.

In the structure shown below, the two fatty acid chains are shown at the bottom of the picture (black and white, hydrophobic tails). The phosphate group is shown near the top in yellow (phosphorus) and red (oxygen) along with a polar head group (hydrophilic head).

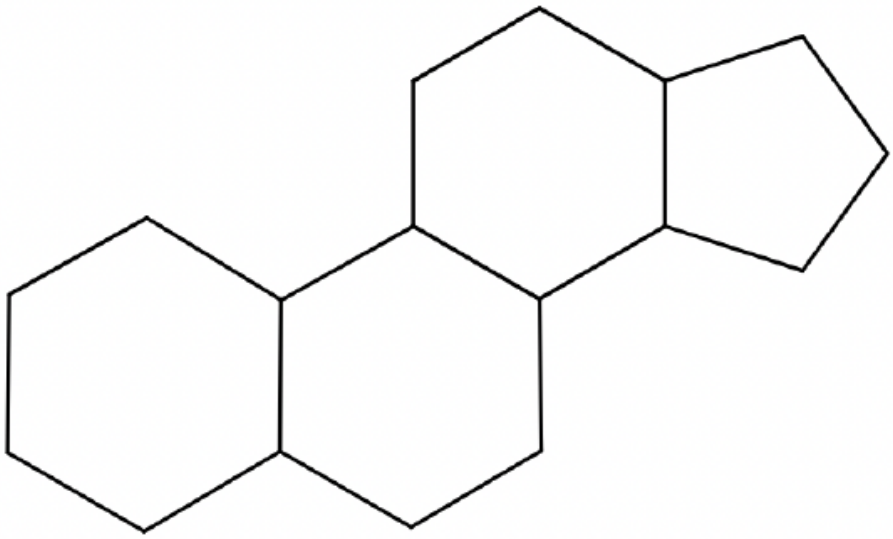


Credit: Mariana Ruiz Villarreal (LadyofHats) for the CK-12 Foundation

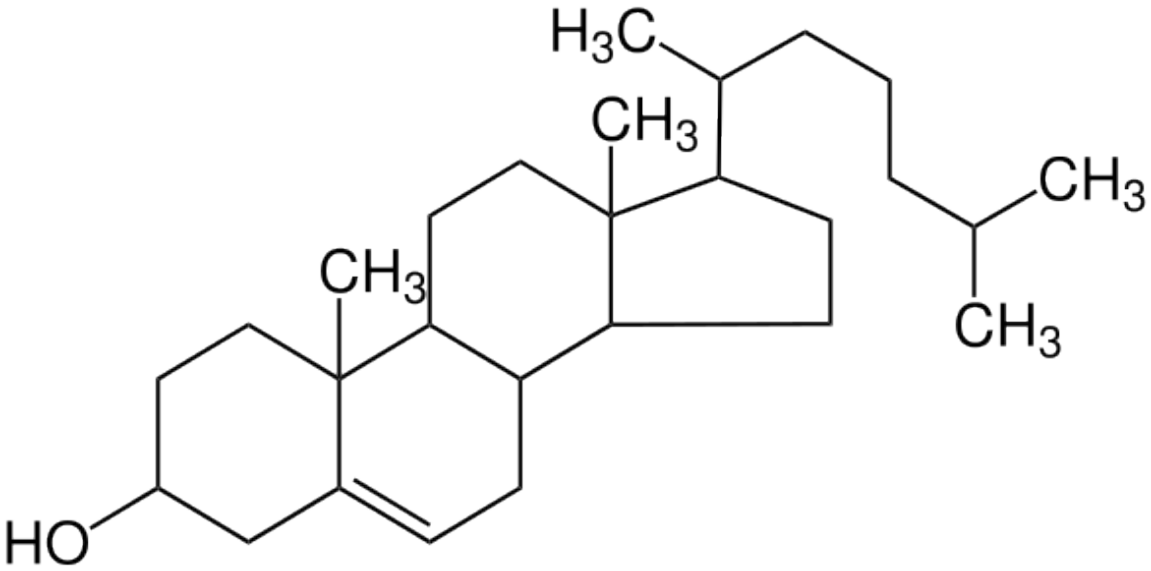
Also shown in the figure above is how phospholipids constitute a majority of the molecules in the membranes of cells. Two layers of phospholipids align with the nonpolar tails pointing towards one another. The polar heads of the molecules face the extracellular fluid outside the cell and the cytosol inside the cell, both of which are aqueous and therefore polar.

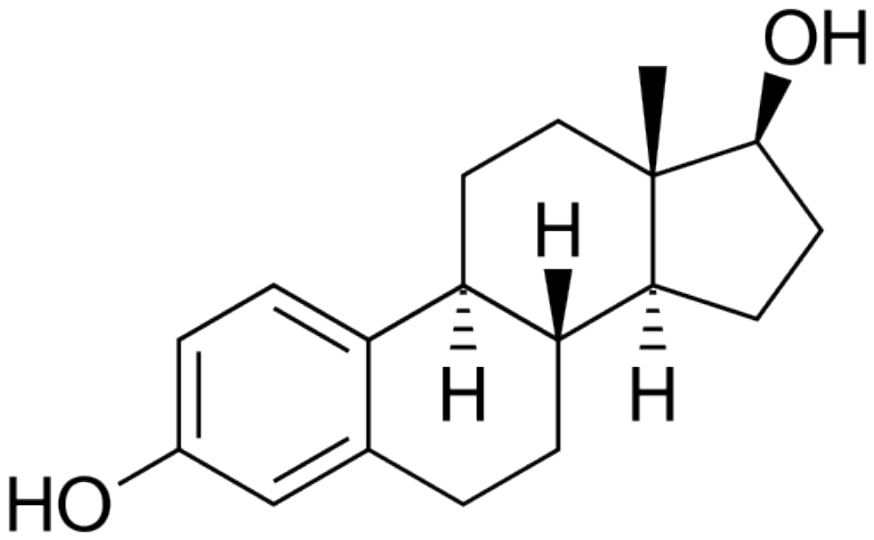
Steroids

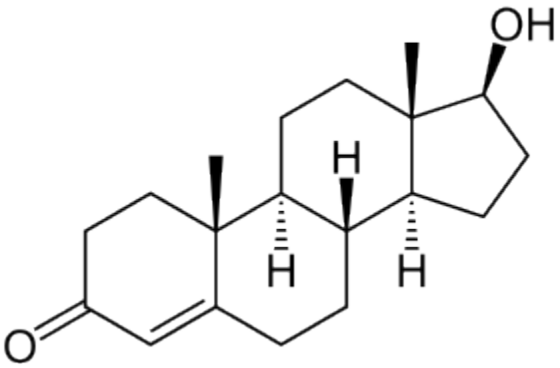
Though they are lipids, the structure of steroids is quite different than triacylglycerols or phospholipids. The basic structure, shown below, is a series of fused rings: three six-membered rings and a 5-membered ring.



A variety of different functional groups can be attached to this backbone to form different steroids. The three diagrams below show three common steroids: cholesterol (top), estrogen (bottom left), and testosterone (bottom right).



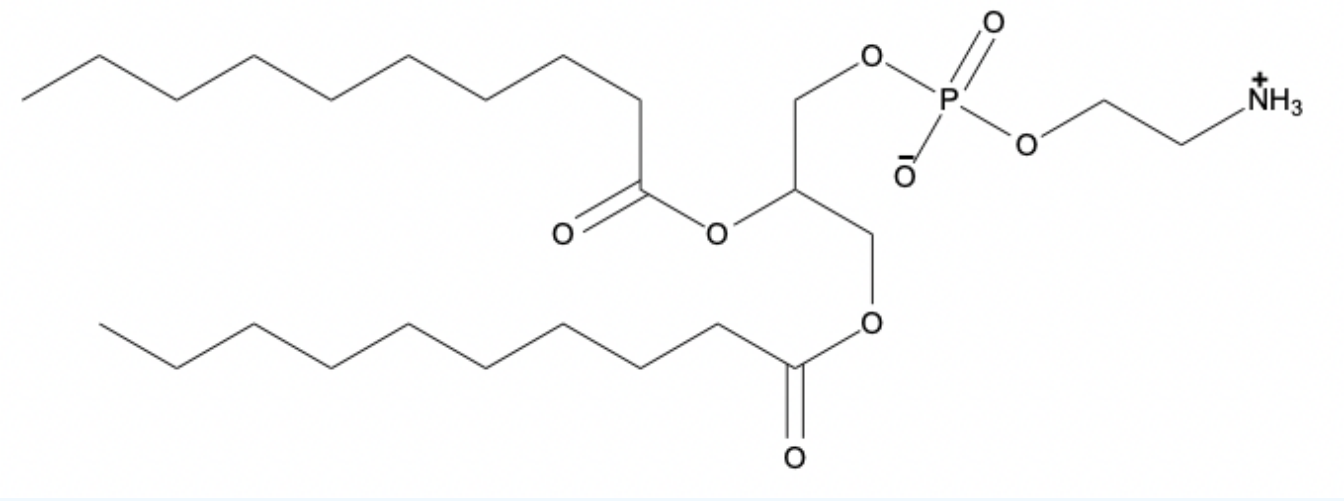


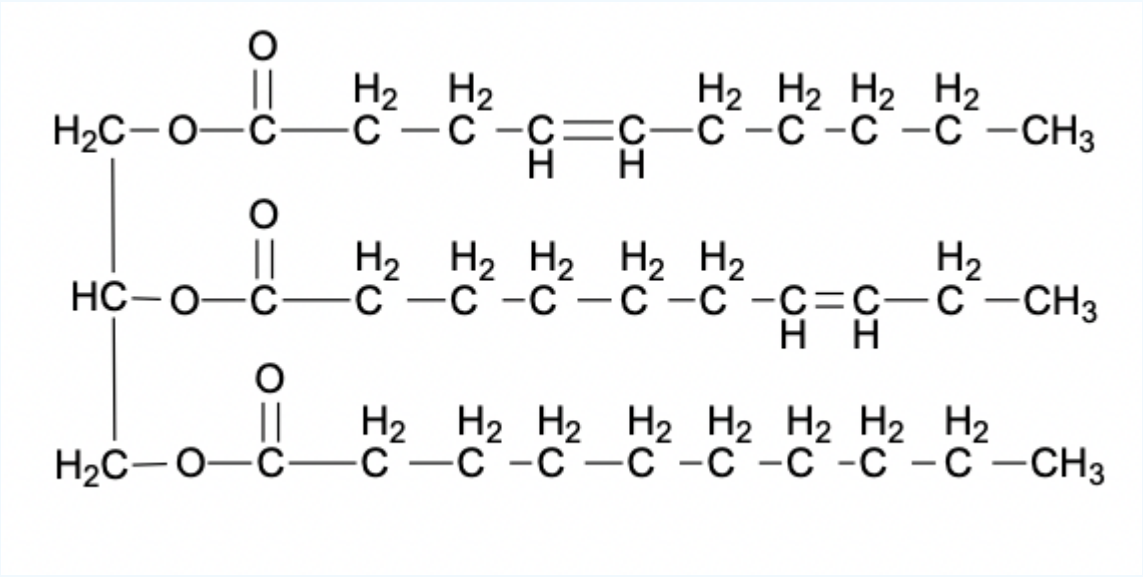


Notice that all three have the four fused rings characteristic of steroids, however, each has a different combination of functional groups.

Exercise 7.2.2

Classify each structure as a fatty acid, phospholipid, triacylglycerol, or steroid.





Answer available at the end of this section.

Summary

- Lipids are defined as biomolecules that are soluble in nonpolar solvents. Therefore, their structures have many more carbon and hydrogen atoms than any other element (because C-C and C-H bonds are nonpolar)
- Fatty acids contain a carboxylic acid functional group and a carbon chain of 10 - 20 atoms. The carbon chain may have all single bonds or may have C=C bonds with a cis configuration
- Phospholipids (specifically glycerophospholipids) contain two ester groups and two long carbon chains (from fatty acids). They also have a polar head that contains P and several atoms with higher electronegativities.
- Triacylglycerols contain three ester groups and three long carbon chains.
- Steroids consist of four rings fused together, i.e. they share a side. Three are six-membered rings and one is a five-membered ring. The functional groups vary which enables different steroids to perform different functions in the body.

Answers

Exercise 7.2.1

- Structure 1 is a lipid. It has many more atoms of carbon and hydrogen than anything else. Specifically it is a fatty acid because it contains a carboxylic acid functional group with a long carbon chain.
- Structure 2 is not a lipid. It has too many highly electronegative atoms (O and N). Therefore, it is not soluble in nonpolar solvents.

Exercise 7.2.2

- Structure 1 is a phospholipid. It contains a PO_4 group as well as two ester functional groups each with a long carbon chain.
- Structure 2 is a triacylglycerol. It has three ester groups and three long carbon chains.

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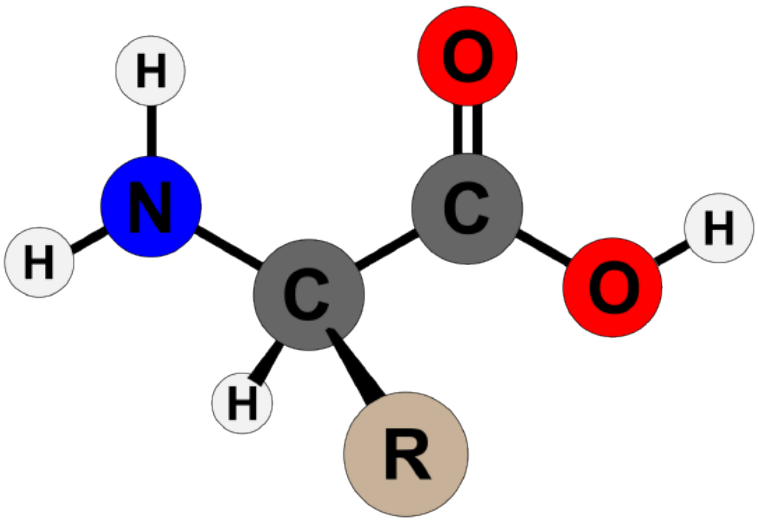
7.3: Proteins

Learning Objectives

- To identify structures of amino acids, peptides, and proteins

Amino acids

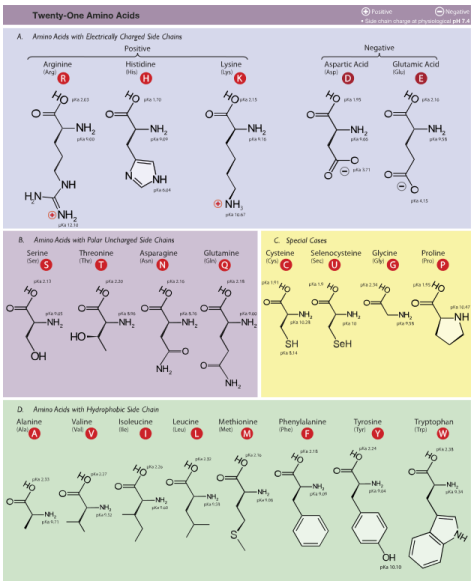
There are 20 naturally occurring amino acids. They have the general format shown below.



Credit: Amino Acid Structure by TechGuy78

These molecules are called amino acids because they all share two functional groups in common: an amine group (-NH2) and a carboxylic acid group (-COOH). With the exception of proline, they all have the same backbone: NH2-CH-COOH. They each have a different R- group or side chain attached to the first carbon atom.

The twenty naturally occurring amino acids (plus selenocysteine) are shown below. Can you identify the backbone in each image? The rest of the molecule is the side chain.



Credit: Amino Acids by Dan Cojocari

There is a wide variety of side chains and they give the amino acids different properties. Some are hydrophobic (from the Greek for water and fear) and cannot dissolve in water because they are nonpolar. Others are hydrophilic (from the Greek for water and love) and can dissolve in water. This affects protein folding (discussed below) because hydrophobic amino acids tend to be found on the inside of proteins whereas hydrophilic proteins tend to be found on the outer edges of proteins (near aqueous solutions).

The Peptide Bond

Amino acids are monomers that form proteins, naturally occurring polymers. The bond between amino acids is called a peptide bond. It forms between the carboxylic acid group of one amino acid and the amine group of another amino acid. Thus, the backbones of the amino acids get connected to one another and the side chains do not connect to one another. The resulting chain, or protein, looks like a charm bracelet, a long backbone with various different attachments at each amino acid. The general form of the reaction is:

$$\text{amino acid 1} + \text{amino acid 2} \rightarrow \text{dipeptide} + \text{water}$$

This reaction is called condensation because water is produced. To see where the water comes from, examine the structures involved:

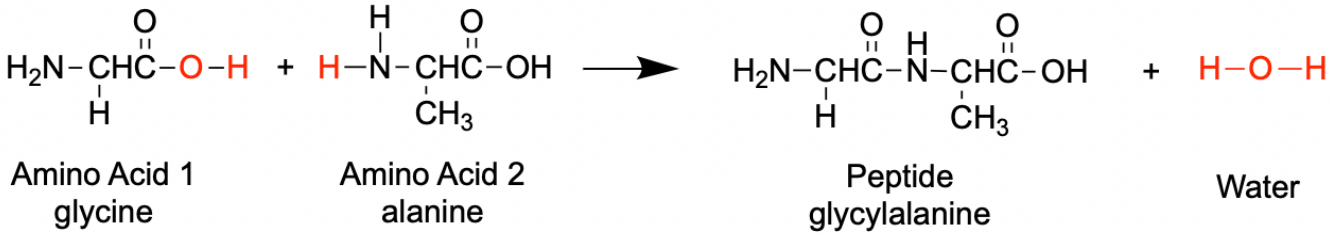


Figure 7.3.1: Reaction of two amino acids to form a dipeptide.

When the reaction occurs, the carboxylic acid group loses the OH and the amine group loses an H. Together, these form water. The peptide bond forms between the carbon of amino acid 1 and the nitrogen of amino acid 2. The peptide bond can also be called an amide bond because after the carboxylic acid group and amine group react an amide functional group is formed.

When this reaction repeats, additional amino acids are added to the chain. Because water is removed, the backbone of a protein has the pattern N-C-C-N-C-C-N-C-C (there is an amine group on the first amino acid and a carboxylic acid group on the last amino acid). Therefore, you can determine how many amino acids are in a peptide by counting how many times the -N-C-C- pattern repeats.

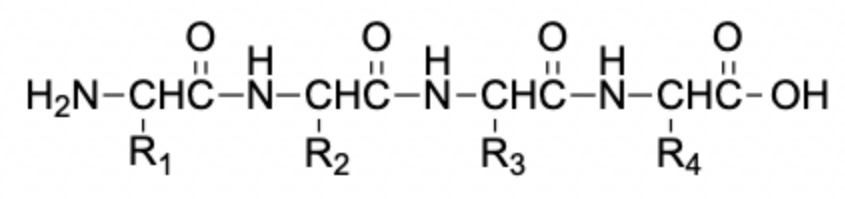
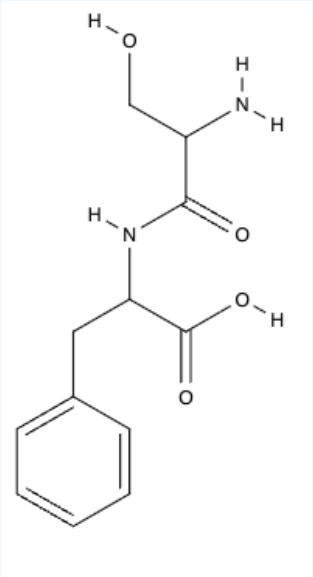


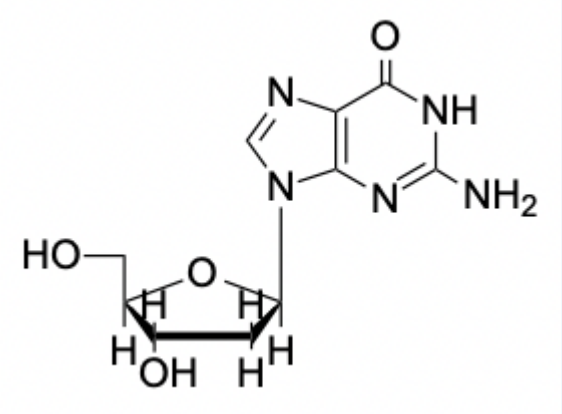
Figure 7.3.1: A tetrapeptide which consists of four amino acids bonded together.

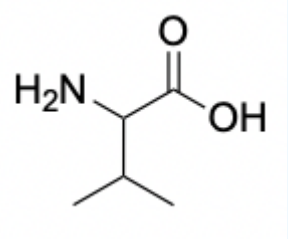
Peptide sequences are generally written with the N-terminus, the amine end, on the left and the C-terminus, the carboxylic acid end, on the right.

Exercise 7.3.1

Determine whether each of the following are amino acids, peptides, or neither.







Answer available at the end of this section.

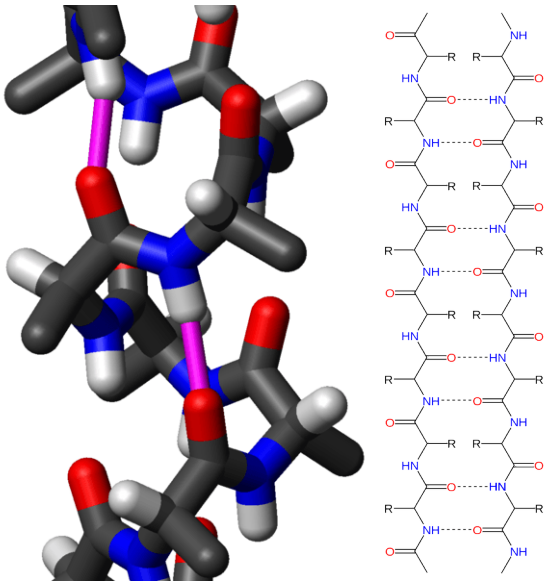
Protein Structure

Because proteins are very large molecules, we can think about their structure at four different levels. Sometimes we are interested in the order in which the amino acids are connected in the chain. This is the primary structure, 1°, of the protein.

-Trp-Tyr-Ser-Ala-Leu-

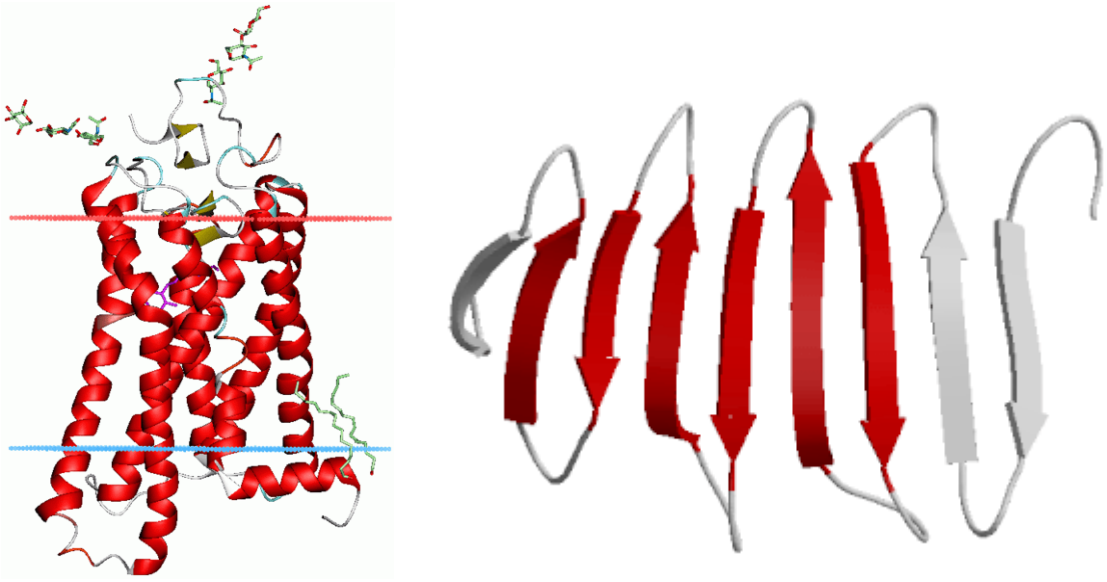
It lists the amino acids present in the correct order (from N-terminus to C-terminus), but tells nothing about the three-dimensional shape of the protein.

The next level of protein structure is the secondary structure, 2°. This indicates a local 3D structure, or in other words, interactions that are occurring in a portion of the protein. It describes the structure in only a portion of the protein chain. Therefore, a protein may contain many secondary structures. Some common secondary structures are the alpha helix (⍺-helix) and the beta-sheet (⍺-sheet). In these secondary structures hydrogen bonds hold different parts of the protein backbone together in a spiral (the alpha helix) or in strands that go back and forth (the beta sheet).



Credit: Pi Helix by WillowW

On the left above is an alpha helix where the pink lines indicate hydrogen bonds between N-H and C=O portions of the backbone. To the right is a portion of a beta-sheet where hydrogen bonds are indicated by dashed lines that connect two portions of the same protein (it folds over along itself). This pattern continues with the protein backbone folding again and again as shown using a ribbon diagram in the image on the right below.



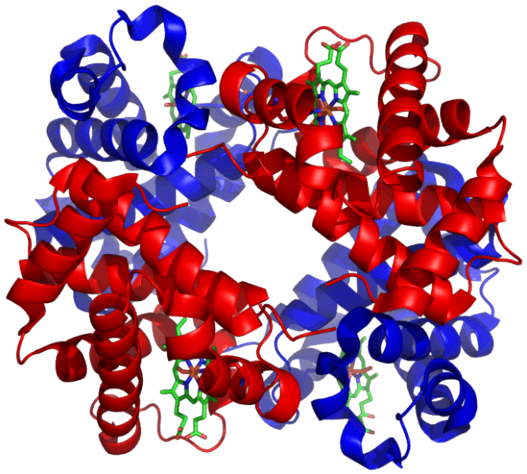
Credit: Protein Image from OPM database by Andrei Lomize

The figure to the left above shows an entire protein. The backbone in portions of the structure is shown as a ribbon to emphasize the local, or secondary, structure: an alpha-helix. Other portions of the backbone are shown as tubes. Similarly, the figure on the right shows a beta-sheet with the portions of the backbone that are hydrogen bonding represented with ribbons.

Thus, you can see that the secondary structure does not indicate the overall shape of a protein, just the structure in a particular portion of the backbone. The tertiary structure (3°) *does* refer to the entire 3D shape of a protein strand. Because the tertiary structure of each protein is unique, a collaborative effort has been made to create a record of all known protein structures. The result is the Protein Data Bank.

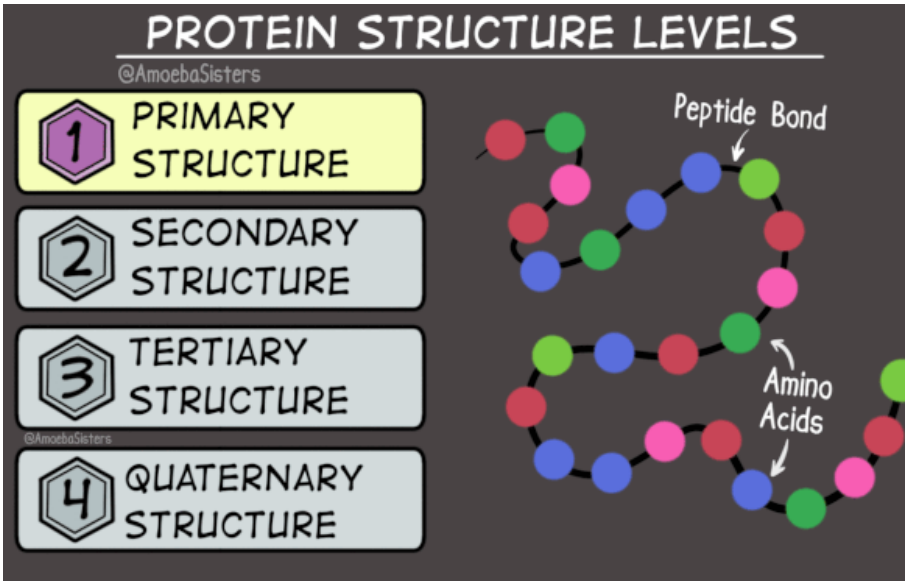
It may seem as if no additional levels of structure would be needed. However, some proteins are composed of multiple strands

interacting with one another. The overall arrangement of these strands is the quaternary structure (4°). Many proteins do not have a quaternary structure because they consist of only one chain. Hemoglobin is an example of a protein that does have a quaternary structure because it has four strands. Its quaternary structure can be seen below.



Credit: Structure of haemoglobin by Richard Wheeler.

The four strands that make up this protein are called subunits and are shown in red and blue in the picture.



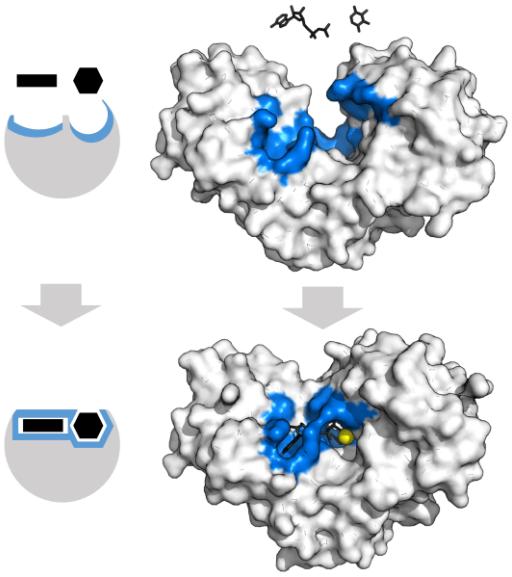
Enzymes

Enzymes are proteins that act as catalysts. They speed up reactions that would otherwise take too long within the body. This is preferable to another method of speeding up reactions: increasing the temperature! Instead, mammals keep their body temperatures regulated and enzymes perform well in this range.

There are many different enzymes in the body because they each catalyze specific chemical reactions. Name of enzymes typically end with the suffix -ase and the rest of the name often provides a clue as to the function of the enzyme, e.g. the enzyme lactase breaks lactose down into two simple sugars and DNA polymerase is used in the synthesis of DNA (a polymer).

Enzymes provide a location for one or more reactants to bind while the reaction is occurring and lower the activation energy of the reaction. In these reactions the reactants are called substrates and the location on the enzyme where the reaction occurs is called the active site.

The picture below shows the interaction between substrates and an enzyme for a reaction that follows the induced-fit model. In the induced-fit model of enzyme catalysis, when the substrate(s) bind in the active site the protein changes structure slightly to fit around the substrate(s).



Credit: Hexokinase induced fit by Thomas Shafee

The top of the diagram is the “before” picture where the enzyme and two substrate molecules are initially separate from one another. In both the block diagram on the left and the molecular model on the right the active site is highlighted in blue and the substrate molecules are shown in black. In the bottom half of the diagram, the substrates are bound to the enzyme and the active site has closed around them. After the reaction is completed (not shown) the product will leave the active site and the enzyme will return to its original conformation to be used again by new substrate molecules. Thus enzymes, like all catalysts, are not consumed in chemical reactions and can be reused by other substrate molecules.

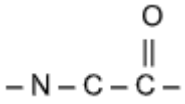
Another model of enzyme function is the lock-and-key model. This model is similar to the induced-fit model in that the substrate(s) bind to an active site in the enzyme, but there is no change in the structure of the enzyme. The substrate(s) fit into the active site like a key fits into a lock.

In order to control reactions occurring in the body, inhibitors can be produced. Inhibitors prevent enzymes from catalyzing chemical reactions, for instance when enough product has already been produced. There is a delicate balance between the amounts of substrates, products, and inhibitors so that the body functions properly.

Some inhibitors disrupt the chemical reactions by stealing the active site. These are called competitive inhibitors and they have a shape similar to the substrate so that they can fit into the active site. There are a limited number of enzyme molecules present and whenever an inhibitor molecule binds with one there is one less opportunity for the substrate to react. Increasing the concentration of inhibitor molecules slows down the reaction whereas increasing the concentration of substrate molecules overcomes the inhibition. Thus, as concentrations of substrates and inhibitors change, the amount of product being produced changes.

Other inhibitors do not bind to the active site, they have a totally different shape than the substrates that they inhibit.

Summary

- Amino acids have a backbone consisting of  (H atoms not shown)
- Amino acid backbones bond together to create peptides (approximately 2 - 100 amino acids) and proteins (hundreds to thousands of amino acids)
- Each amino acid has a different atom or group of atoms bonded to the central carbon of the backbone

Answers

Exercise 7.3.1

- Structure 1 is a peptide. It contains the pattern N-C-C-N-C-C twice. The side chains or R groups for the amino acids are CH_2OH and CH_2 with an aromatic ring.
- Structure 2 is neither an amino acid nor a peptide. It contains N atoms and a C=O, but they are not separated by one carbon atom. The N atoms are almost always found in a chain rather than in rings.
- Structure 3 is an amino acid. The N-C-C- pattern happens once. The side chain or R group is $CH_2(CH_3)_2$.

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7.4: Nucleic Acids

Learning Objectives

- Identify a nucleic acid when shown its structure
- Distinguish between nucleobases, nucleosides, nucleotides, RNA, and DNA

Nucleic acids store and process genetic information. They include DNA and RNA.

Building Blocks of Nucleic Acids

Just as amino acids with the same backbone but different side chains can be joined together in a different order to make each protein in an organism, different nucleotides also bond together to make nucleic acids. Like amino acids, all nucleotides have certain components in their backbones plus another structure that gives each nucleotide a unique identity.

Nucleobases

Nucleobases, frequently just referred to as bases, are the portion of a nucleotide with the most variation. They are analogous to the side chains in amino acids. Nucleobases consist of 1 - 2 rings each containing several nitrogen atoms.

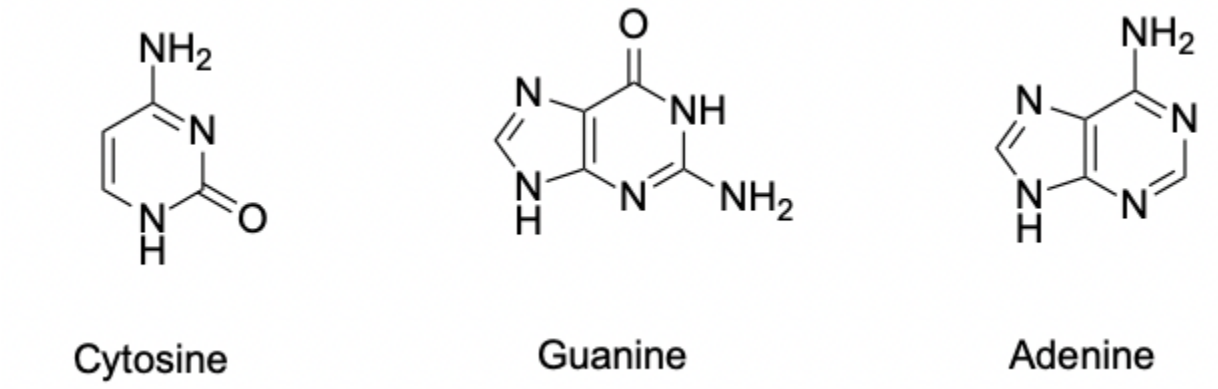


Figure 7.4.1: Examples of nucleobases. This is not a complete list, but these are the three that are found in both RNA and DNA.

The bases adenine (A), cytosine (C), and guanine (G) are found in both RNA and DNA. An additional bases, uracil (U), is found in RNA while thymine (T) is found in DNA. Therefore, each nucleic acid contains four types of bases: A, C, G, and U in RNA and A, C, G, and T in DNA.

Notice that the nucleobases have several N-H bonds. This helps them to hydrogen bond to one another in base pairs. The structure of each base determines which other base it can partner with. Guanine (G) can only hydrogen bond with cytosine (C). Adenine (A) can hydrogen bond with either uracil (U) in RNA or thymine (T) in DNA.

Nucleosides

Each nucleobase can bond to a monosaccharide to make a nucleoside. The two possible monosaccharides are ribose and deoxyribose.

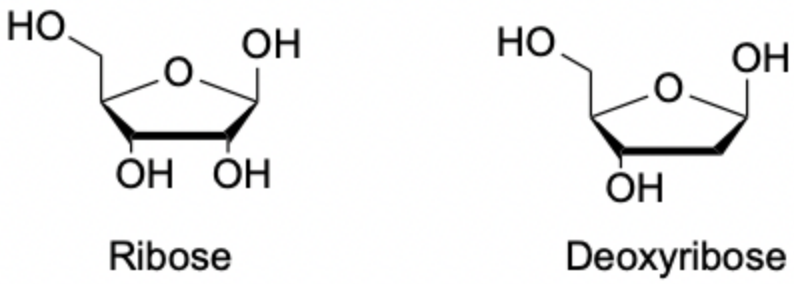


Figure 7.4.2: Ribose is the monosaccharide found in RNA (and in nucleosides and nucleotides associated with RNA) whereas deoxyribose is found in DNA and its nucleosides and nucleotides. The only difference is at the second carbon atom clockwise from the O in the ring. If an oxygen atom is present at the second carbon, the monosaccharide is ribose. If it is absent, it is deoxyribose.

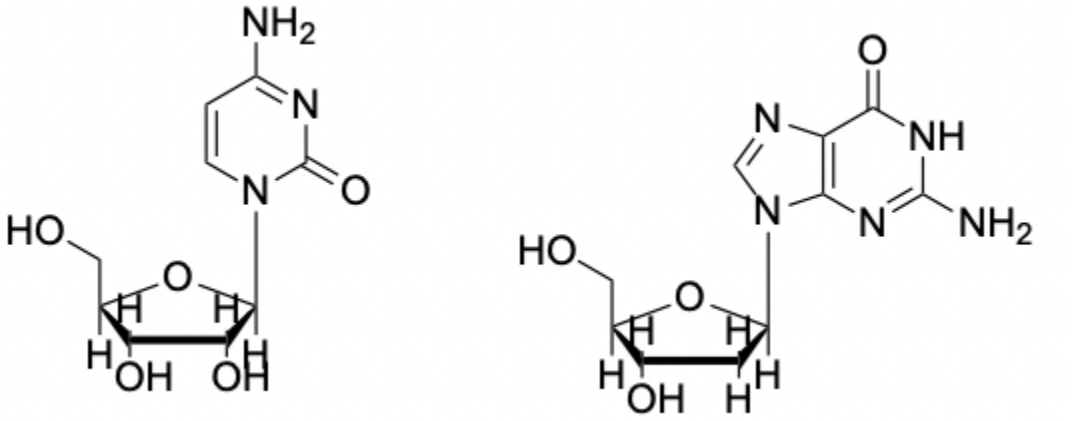


Figure 7.4.3: Examples of nucleosides. Nucleosides consist of a nucleobase (either 1 or 2 rings, each with several N atoms) and a monosaccharide. The structure on the left has ribose as its monosaccharide; the structure on the right contains deoxyribose.

Nucleotides

RNA and DNA are polymers where the monomers are nucleic acids. In other words, nucleotides are the building blocks of nucleic acids. A nucleotide consists of a nucleobase, a monosaccharide, and a phosphate group. That means that a nucleotide is equivalent to a nucleoside bonded to a phosphate group.

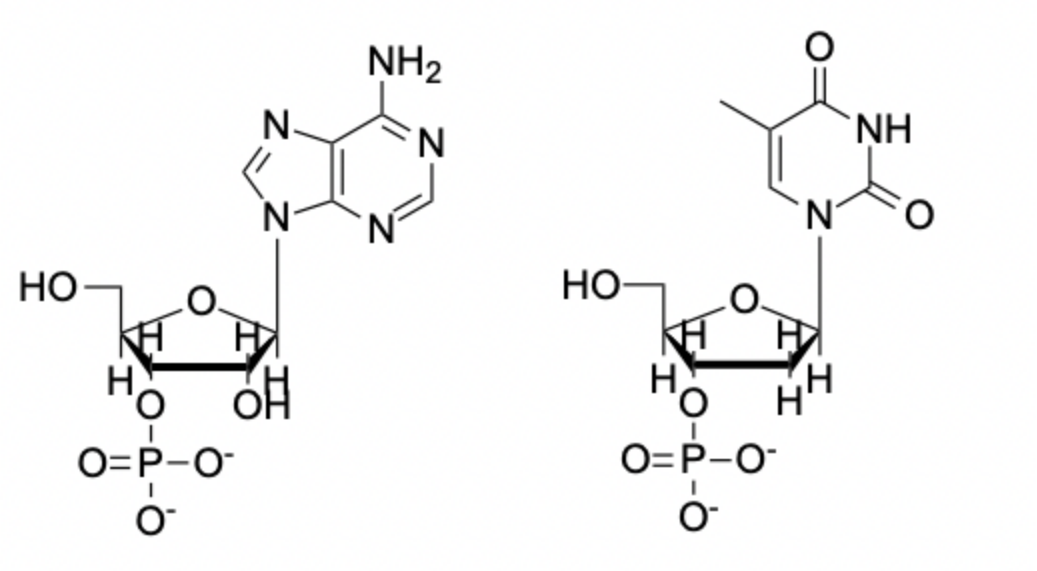
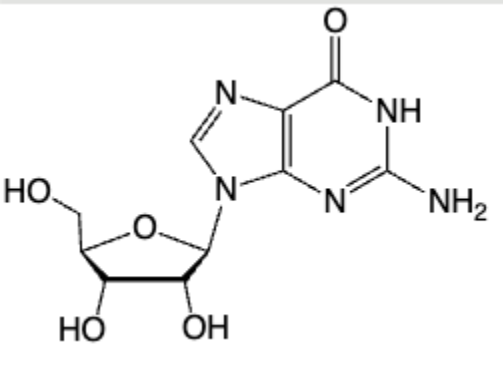
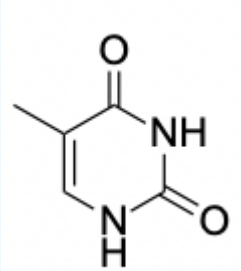


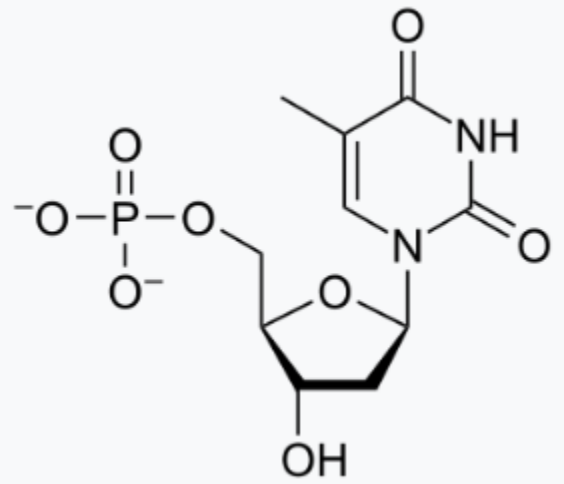
Figure 7.4.4: Two nucleotides. The base is bonded to the first carbon of the monosaccharide and the phosphate is bonded to either the third or fifth carbon of the monosaccharide.

Exercise 7.4.1

Identify each of the following as a nucleobase, nucleoside, or nucleotide.







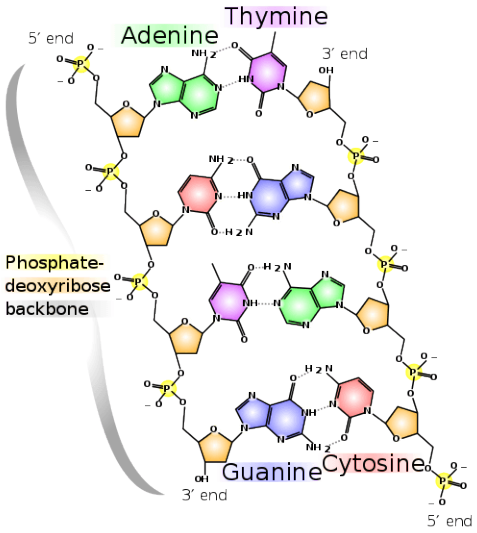
Structure 1 image credit: NEUROtiker - Own work, Public Domain. Structure 3 image credit: By Hbf878 - Own work, CC0.

Answer available at the end of this section.

Structure of RNA and DNA

DNA is an acronym (rather than a chemical formula) for **d**eoxyribo**n**ucleic **a**cid. It stores the genetic information for an organism and also passes that information to the next generation. RNA is **r**ibo**n**ucleic **a**cid and it is used as a template for the production of proteins based on the genetic information. The similarity between the names **ribonucleic acid** and deoxy**ribonucleic acid** reveals a similarity of their structures. We will first examine the similarities, followed by the differences.

Both DNA and RNA have backbones made of alternating sugar and phosphate groups. In the diagram of DNA below, the phosphate groups are highlighted in yellow and they are bonded to the sugar groups highlighted in orange. Notice that these form a chain or polymer down the left and right sides of the image.



Credit: DNA chemical structure by Madeleine Price Ball.

Connected to the backbone at each sugar group are bases. The four bases in the picture above can be grouped into two categories: adenine (A) and guanine (G) have a double ring structure whereas thymine (T) and cytosine (C) have a single ring structure. All of the bases have several N - H bonds and can participate in hydrogen bonding (there are also polar C=O groups on some of the bases). This hydrogen bonding holds the two strands of DNA together. Specifically, A and T always H-bond with one another (sharing two H-bonds) and G and C always H-bond with one another (3 H-bonds). Because of differences in their structures, other pairings of these four bases are not possible.

There are three differences between the structures of DNA and RNA:

1. The sugars in the backbone of DNA and RNA differ by one oxygen atom. The deoxy- prefix in DNA indicates that the sugar is like ribose, but with one less oxygen atom.

2. DNA is double stranded but RNA contains only a single strand

3. In addition to the bases adenine (A), guanine (G), and cytosine (C) which are found in both DNA and RNA, DNA contains the base thymine (T) whereas RNA contains uracil (U).

Processes Involving DNA and RNA

DNA replicates itself within cells (makes another copy of itself). In addition, DNA is used as a template for the synthesis of mRNA which is used as a template for the creation of proteins. So, the main purpose of DNA in your body is to store the information that is ultimately used to make the proteins that serve a wide variety of functions throughout the body.

Exercise 7.4.2

First classify the following relevant to carbohydrates, lipids, nucleic acids, or proteins. Then try to describe the structure more specifically.

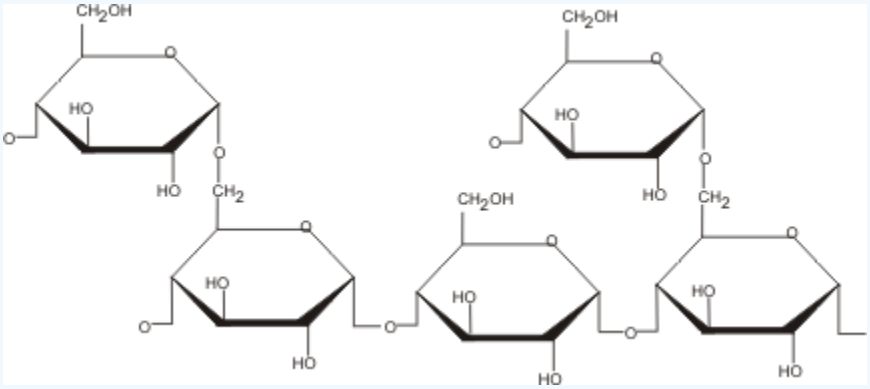
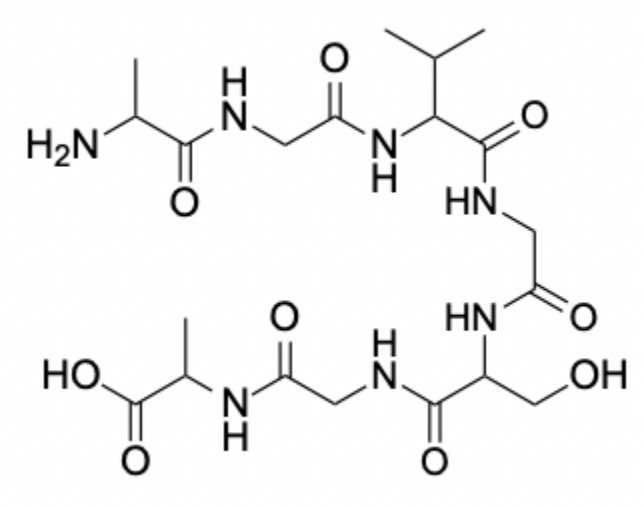


Image credit: Boumphreyfr, CC BY-SA 3.0, via Wikimedia Commons

Answer available at the end of this section.

Exercise 7.4.3

First classify the following relevant to carbohydrates, lipids, nucleic acids, or proteins. Then try to describe the structure more specifically.



Answer available at the end of this section.

Summary

- Nucleobases, a component of nucleic acids, consist of 1 - 2 rings containing several N atoms. This can be helpful in identifying the structures of nucleic acids and their component parts.
- Nucleosides are made of a nucleobase and a monosaccharide (either ribose or deoxyribose)
- Nucelotides are made of a nucleobase, a monosaccharide, and a phosphate group (PO_4)
- RNA and DNA are both nucleic acids made of many nucleotides. The monosaccharides and phosphate groups of the

nucleotides form the backbone of these structures. RNA has one backbone whereas DNA has two. The monosaccharide in RNA is ribose (RNA stands for ribonucleic acid); the monosaccharide in DNA (deoxyribonucleic acid) is deoxyribose. Deoxyribose has one fewer O atom than ribose. Both DNA and RNA can contain the G, C, and A nucleobases. Only RNA contains U and only DNA contains T.

Answers

Exercise 7.4.1

- Structure 1 is a nucleoside. It contains a nucleobase (the two rings with N atoms) and a monosaccharide (ribose in this case).
- Structure 2 is a nucleobase. It is a ring with two N atoms. It is not bonded to a monosaccharide or phosphate.
- Structure 3 is a nucleotide. It contains a nucleobase with two N atoms, a monosaccharide (deoxyribose in this case), and the monosaccharide is bonded to a phosphate group.

Exercise 7.4.2

This molecule belongs to the carbohydrate class of biomolecules. Five monosaccharides are shown, with the structure unfinished in four places indicating that the pattern continues. Therefore it is a polysaccharide.

Exercise 7.4.3

This molecule belongs to the protein class of biomolecules. It contains the amino acid backbone -N-C-C- where there is a double bond to O on the second carbon and the R group on the first carbon varies. In this molecule the -N-C-C- pattern repeats indicating that this is a peptide rather than a single amino acids. This peptide contains seven amino acids because the -N-C-C- repeats seven times.

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8: WRITING CHEMICAL EQUATIONS

8.1: Chemical Equations
8.2: Precipitation Reactions
8.3: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

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8.1: Chemical Equations

Learning Objectives

- Define chemical reaction.
- Use a balanced chemical equation to represent a chemical reaction.

Water (H_2O) is composed of hydrogen and oxygen. Suppose we imagine a process in which we take some elemental hydrogen (H_2) and elemental oxygen (O_2) and let them react to make water. The statement

"*hydrogen and oxygen react to make water*"

is one way to represent that process, which is called a chemical reaction. Figure 8.1.1 shows a rather dramatic example of this very reaction.

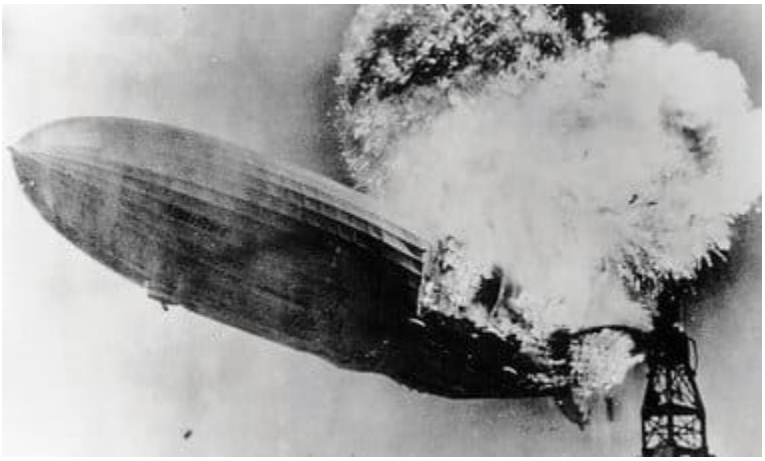


Figure 8.1.1: The Formation of Water. Hydrogen and oxygen combine to form water. Here, the hydrogen gas in the zeppelin SS *Hindenburg* reacts with oxygen in the air to make water. Source: Photo courtesy of the US Navy. For a video of this see www.youtube.com/watch?v=CgWHbpMVQ1U.

To simplify the writing of reactions, we use formulas instead of names when we describe a reaction. We can also use symbols to represent other words in the reaction. A **plus sign** connects the initial substances (and final substances, if there is more than one), and an **arrow** ($\rightarrow$) represents the chemical change:

$$\mathrm{H_2 + O_2 \ \rightarrow H_2O} \tag{8.1.1}$$

This statement is one example of a **chemical equation**, an abbreviated way of using symbols to represent a **chemical change**. The substances on the left side of the arrow are called **reactants**, and the substances on the right side of the arrow are called **products**. It is not uncommon to include a phase label with each formula—(s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for a substance dissolved in water, also known as an *aqueous solution*. If we included phase labels for the reactants and products, under normal environmental conditions, the reaction would be as follows:

$$\mathrm{H_2\,(g) + O_2\,(g) \ \rightarrow H_2O\ (\ell)} \tag{8.1.2}$$

> Chemical equations can also be used to describe physical changes. We will see examples of this soon.

This equation is still not complete because *it does not satisfy the* ***law of conservation of matter***. Count the **number of atoms of each element** on each side of the arrow. On the reactant side, there are two H atoms and two O atoms; on the product side, there are two H atoms and only one oxygen atom. The equation is **not balanced** because the number of oxygen atoms on each

side is not the same (Figure 8.1.2).

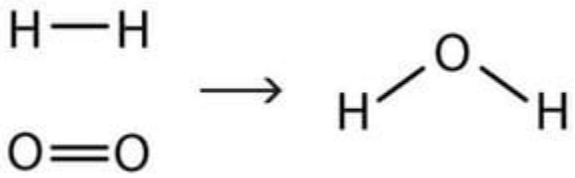


Figure 8.1.2: Balanced—Yes or No?. By counting the atoms of each element, we can see that the reaction is not balanced as written.

To make this chemical equation conform to the law of conservation of matter, we must revise the amounts of the reactants and the products as necessary to get the same number of atoms of a given element on each side. Because every substance has a characteristic chemical formula, we cannot change the chemical formulas of the individual substances. For example, we cannot change the formula for elemental oxygen to O. However, we can assume that different numbers of reactant molecules or product molecules may be involved. For instance, perhaps two water molecules are produced, not just one:

$$\mathrm{H_2\,(g) + O_2\,(g)\ \rightarrow 2\,H_2O\ (\ell)} \tag{8.1.3}$$

The **2** preceding the formula for water is called a **coefficient**. It implies that two water molecules are formed. There are now two oxygen atoms on each side of the equation.

> This point is so important that we should repeat it. You **cannot** change the formula of a chemical substance to balance a chemical reaction! You **must** use the proper chemical formula of the substance.

Unfortunately, by inserting the coefficient 2 in front of the formula for water, we have also changed the number of hydrogen atoms on the product side as well. As a result, we no longer have the same number of hydrogen atoms on each side. This can be easily fixed, however, by putting a coefficient of 2 in front of the diatomic hydrogen reactant:

$$\mathrm{2\,H_2\,(g) + O_2\,(g)\ \rightarrow 2\,H_2O\ (\ell)} \tag{8.1.4}$$

Now we have four hydrogen atoms and two oxygen atoms on each side of the equation. The **law of conservation of matter is satisfied** because we now have the same number of atoms of each element in the reactants and in the products. We say that the reaction is **balanced** (Figure 8.1.3). The diatomic oxygen has a coefficient of 1, which typically is not written but assumed in balanced chemical equations.

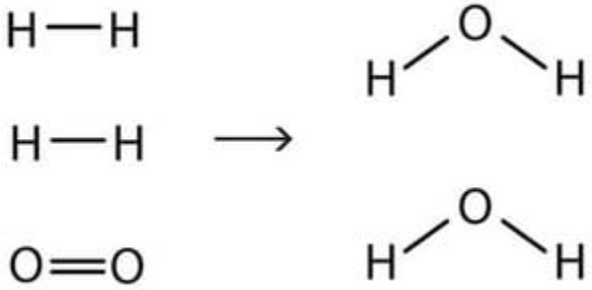


Figure 8.1.3: Balanced—Yes or No?. By counting the atoms of each element, we can see that the reaction is now balanced.

Proper chemical equations should be balanced. Writing balanced reactions is a chemist's way of acknowledging the law of conservation of matter.

Example 8.1.1

Is each chemical equation balanced?

a. $\mathrm{2\,Na\,(s) + O_2\,(g) \longrightarrow 2\,Na_2O\,(s)}$

b. $CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(\ell)$
c. $AgNO_3(aq) + 2\,KCl(aq) \longrightarrow AgCl(s) + KNO_3(aq)$

Solution

a. By counting, we find two sodium atoms and two oxygen atoms in the reactants and four sodium atoms and two oxygen atoms in the products. This equation is not balanced.
b. The reactants have one carbon atom, four hydrogen atoms, and four oxygen atoms. The products have one carbon atom, four hydrogen atoms, and four oxygen atoms. This equation is balanced.
c. The reactants have one silver atom, one nitrogen atom, three oxygen atoms, two potassium atoms, and two chlorine atoms. The products have one silver atom, one chlorine atom, one potassium atom, one nitrogen atom, and three oxygen atoms. Because there are different numbers of chlorine and potassium atoms, this equation is not balanced.

Exercise 8.1.1

Is each chemical equation balanced?

a. $2\,Hg(l) + O_2(g) \longrightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2\,O_2(g) \longrightarrow 2\,CO_2(g) + 2\,H_2O(l)$
c. $Mg(NO_3)_2(s) + 2\,Li(s) \longrightarrow Mg(s) + 2\,LiNO_3(s)$.

Answer a: available at the end of this section.

How does one balance a chemical equation, starting with the correct formulas of the reactants and products? Basically, a **back-and-forth approach** is adopted, counting the number of atoms of one element on one side, checking the number of atoms of that element on the other side, and changing a coefficient if necessary. Then check another element, going back and forth from one side of the equation to another, until each element has the same number of atoms on both sides of the arrow. In many cases, it does not matter which element is balanced first and which is balanced last, as long as all elements have the same number of atoms on each side of the equation.

Below are guidelines for writing and balancing chemical equations.

Writing and Balancing Chemical Equations

1. Determine the correct chemical formulas for each reactant and product. Write the skeleton equation.
2. Count the number of atoms of each element that appears as a reactant and as a product. If a polyatomic ion is unchanged on both sides of the equation, count it as a unit.
3. Balance each element one at a time by placing coefficients in front of the formulas. No coefficient is written for a 1. It is best to begin by balancing elements that only appear in one chemical formula on each side of the equation. NEVER change the subscripts in a chemical formula - you can only balance equations by using coefficients.
4. Check each atom or polyatomic ion to be sure that they are equal on both sides of the equation.
5. Make sure that all coefficients are in the lowest possible ratio. If necessary, reduce to the lowest ratio.

For example, to balance the equation

$$CH_4 + Cl_2 \rightarrow CCl_4 + HCl$$

Step 1: Write the skeleton equation with the correct formulas.

$$CH_4 + Cl_2 \rightarrow CCl_4 + HCl \tag{8.1.5}$$

Step 2: Count the number of each atom or polyatomic ion on both sides of the equation.

Reactants	Products
1 C atom	1 C atom
4 H ions	1 H ions
2 Cl atom	5 Cl atoms

Step 3: We find that both sides are already balanced with one carbon atom. So we proceed to balance the hydrogen atoms. We find that the reactant side has four hydrogen atoms, so the product side must also have four hydrogen atoms. This is balanced by putting a 4 in front of the HCl:

$$\mathrm{CH_4 + Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{8.1.6}$$

Reactants	Products
1 C atom	1 C atom
4 H ions	4 H ions
2 Cl atom	8 Cl atoms

Now each side has four hydrogen atoms. The product side has a total of eight chlorine atoms (four from the $\mathrm{CCl_4}$ and four from the four molecules of HCl), so we need eight chlorine atoms as reactants. Because elemental chlorine is a diatomic molecule, we need four chlorine molecules to get a total of eight chlorine atoms. We add another 4 in front of the $\mathrm{Cl_2}$ reactant:

$$\mathrm{CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl} \tag{8.1.7}$$

Reactants	Products
1 C atom	1 C atom
4 H ions	4 H ions
8 Cl atom	8 Cl atoms

Step 3: Now we check: each side has one carbon atom, four hydrogen atoms, and eight chlorine atoms. The chemical equation is balanced. And, the coefficients are in the lowest possible ratio.

Example 8.1.2

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

Step 1: Write the skeleton equation with the correct formulas.

$$\mathrm{Pb(NO_3)_2\,(aq) + NaCl\,(aq) \rightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$$

Step 2: Count the number of each atom or polyatomic ion on both sides of the equation.

Reactants	Products
1 Pb atom	1 Pb atom
2 $\mathrm{NO_3^-}$ ions	1 $\mathrm{NO_3^-}$ ions
1 Na atom	1 Na atom
1 Cl atom	2 Cl atoms

Step 3: Solve. The nitrate ions and the chlorine atoms are unbalanced. Start by placing a 2 in front of the NaCl. This increases the reactant counts to 2 Na atoms and 2 Cl atoms. Then place a 2 in front of the $\mathrm{NaNO_3}$. The result is:

$$\mathrm{Pb(NO_3)_2\,(aq) + 2NaCl\,(aq) \rightarrow 2NaNO_3\,(aq) + PbCl_2\,(s)}$$

Step 4: The new count for each atom and polyatomic ion becomes:

Reactants	Products
1 Pb atom	1 Pb atom
2 NO_3^- ions	2 NO_3^- ions
2 Na atom	2 Na atom
2 Cl atom	2 Cl atoms

Step 5: Think about the result.

The equation is now balanced since there are equal numbers of atoms of each element on both sides of the equation. And, the coefficients are in the lowest possible ratio.

Exercise 8.1.2

Fermentation is a biochemical process that enables yeast cells to live in the absence of oxygen. Humans have exploited it for centuries to produce wine and beer and make bread rise. In fermentation, sugars such as glucose ($C_6H_{12}O_6$) are converted to ethanol (C_2H_5OH) and carbon dioxide CO_2. Write a balanced chemical reaction for the fermentation of glucose.



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

Answer available at the end of this section.

Key Takeaway

Chemical reactions are represented by chemical equations that list reactants and products. Proper chemical equations are balanced; the same number of each element's atoms appears on each side of the equation.

Answers

Exercise 8.1.1

balanced

O is not balanced; the 4 atoms of oxygen on the left does not balance with the 6 oxygen atoms on the right

balanced

Exercise 8.1.2

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

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8.2: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product—a precipitate—when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate, a reddish precipitate of silver dichromate is produced.

$$\mathrm{AgNO_3\,(aq) + K_2Cr_2O_7\,(aq) \;\rightarrow Ag_2Cr_2O_7\,(s) + KNO_3\,(aq)} \tag{8.2.1}$$

This unbalanced equation has the general form of an exchange reaction:

$$\mathrm{AC + BD \;\rightarrow \underset{insoluble}{AD} + BC} \tag{8.2.2}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called **double-displacement reactions**. Precipitation reactions are used to isolate metals that have been extracted from their ores, and to recover precious metals for recycling.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video: Mixing potassium dichromate and silver nitrate together to initiate a precipitation reaction (Equation 8.2.1).

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of aqueous $NaCl$ solution is mixed with 500 mL of aqueous KBr solution, the final solution has a volume of 1.00 L and contains $\mathrm{Na^+\,(aq)}$, $\mathrm{Cl^-\,(aq)}$, $\mathrm{K^+\,(aq)}$, and $\mathrm{Br^-\,(aq)}$. As you will see in (Figure 8.2.1), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

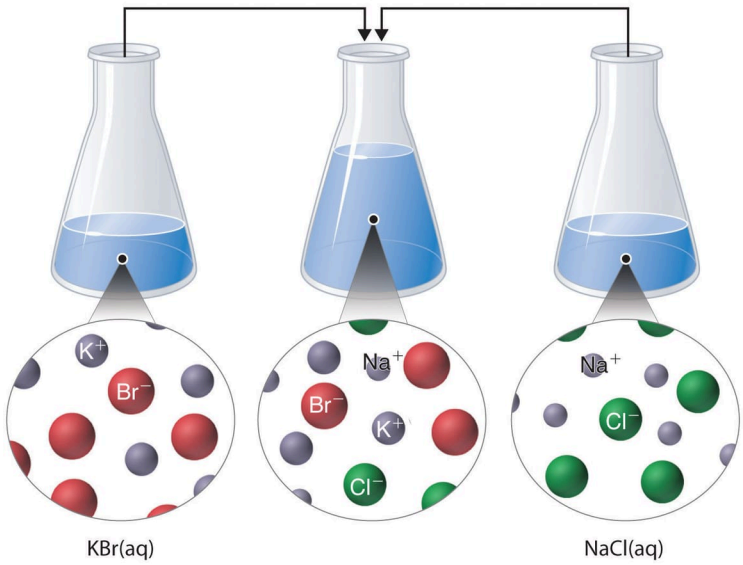


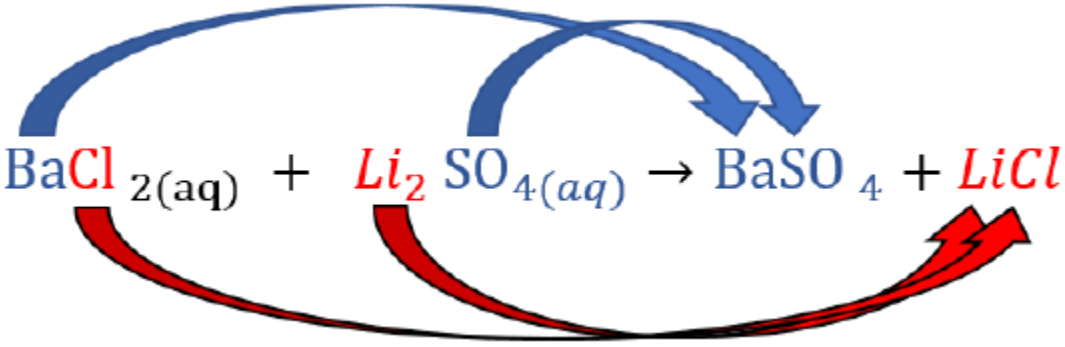
Figure 8.2.1: The Effect of Mixing Aqueous KBr and NaCl Solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurred.

Predict what will happen when aqueous solutions of barium chloride and lithium sulfate are mixed.

Change the partners of the anions and cations on the reactant side to form new compounds (products):



Chemical equation of the reactants barium chloride and lithium sulfate forming the products barium sulfate and lithium chloride.

Because barium chloride and lithium sulfate are strong electrolytes, each dissociates completely in water to give a solution that contains the constituent anions and cations. Mixing the two solutions *initially* gives an aqueous solution that contains $\mathrm{Ba^{2+}}$, $\mathrm{Cl^-}$, $\mathrm{Li^+}$, and $\mathrm{SO_4{}^{2-}}$ ions. The only possible exchange reaction is to form LiCl and $\mathrm{BaSO_4}$.

Correct the formulas of the products based on the charges of the ions.

No need to correct the formula as both compounds already have their charges balanced.

$$\mathrm{BaCl_2\,(aq) + Li_2SO_4\,(aq) \;\rightarrow BaSO_4 + LiCl}$$

Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.

$$\mathrm{BaCl_2\,(aq) + Li_2SO_4\,(aq)\ \rightarrow BaSO_4\,(s) + LiCl\,(aq)}$$

Table 7.5.1 from the previous section shows that LiCl is soluble in water, but $\mathrm{BaSO_4}$ is not soluble in water.

Balance the equation:

$$\mathrm{BaCl_2\,(aq) + Li_2SO_4\,(aq)\ \rightarrow BaSO_4\,(s) + 2\,LiCl\,(aq)}$$

Although soluble barium salts are toxic, $\mathrm{BaSO_4}$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a "barium milkshake" or a "barium enema"—a suspension of very fine $\mathrm{BaSO_4}$ particles in water.



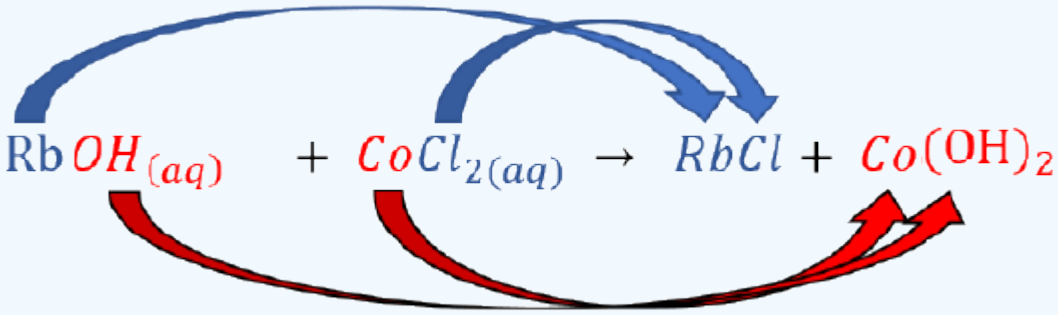
Figure 8.2.2: An x-ray of the digestive organs of a patient who has swallowed a "barium milkshake." A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. (Public Domain; Glitzy queen00 via Wikipedia).

Example 8.2.1

Predict what will happen if aqueous solutions of rubidium hydroxide and cobalt(II) chloride are mixed.

Solution

Solutions to Example 7.6.1

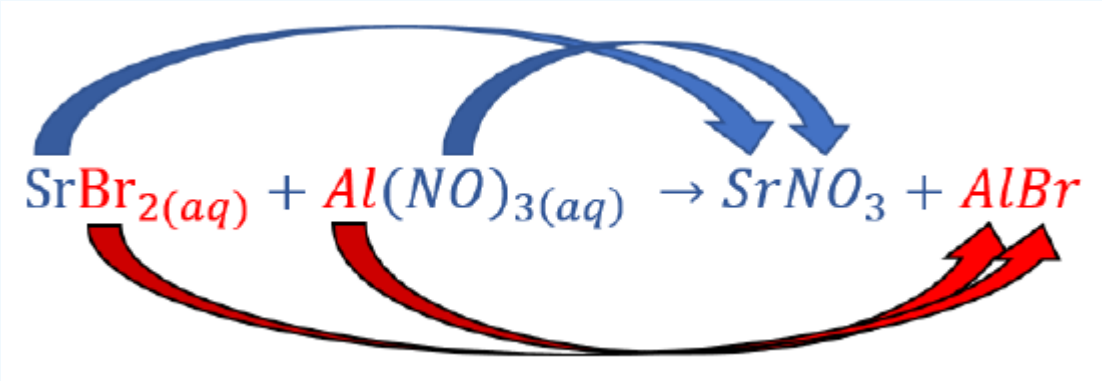
Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	 Chemical equation of the reactants rubidium hydroxide and cobalt(II) chloride forming the products rubidium chloride and cobalt hydroxide.
Correct the formulas of the products based on the charges of the ions.	$\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl + Co(OH)_2}$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl\,(aq) + Co(OH)_2\,(s)}$
Balance the equation.	Coefficients already balanced. $\mathrm{RbOH\,(aq) + CoCl_2\,(aq) \rightarrow RbCl\,(aq) + Co(OH)_2\,(s)}$

Example 8.2.2

Predict what will happen if aqueous solutions of strontium bromide and aluminum nitrate are mixed.

Solution

Solutions for Example 7.6.2

Steps	Example
Change the partners of the anions and cations on the reactant side to form new compounds (products).	 Chemical equation of the reactants strontium bromide and aluminum nitrate forming the products strontium nitrate and aluminum bromide.
Correct the formulas of the products based on the charges of the ions.	$\mathrm{SrBr_2\,(aq) + Al(NO_3)_3\,(aq)\ \rightarrow Sr(NO_3)_2 + AlBr_3}$
Refer to the solubility rules table to determine insoluble products which will therefore form a precipitate.	$\mathrm{SrBr_2\,(aq) + Al(NO_3)_3\,(aq)\ \rightarrow Sr(NO_3)_2\,(aq) + AlBr_3\,(aq)}$ According to Table 7.5.1 from the previous section, both $\mathrm{AlBr_3}$ (rule 4) and $\mathrm{Sr(NO_3)_2}$ (rule 2) are soluble.
If all possible products are soluble, then no net reaction will occur.	$\mathrm{SrBr_2\,(aq) + Al(NO_3)_3\,(aq)\ \rightarrow}$ **NO REACTION**

Exercise 8.2.2

Using the information in Table 7.5.1 from the previous section, predict what will happen in each case involving strong electrolytes.

a. An aqueous solution of strontium hydroxide is added to an aqueous solution of iron(II) chloride.
b. Solid potassium phosphate is added to an aqueous solution of mercury(II) perchlorate.
c. Solid sodium fluoride is added to an aqueous solution of ammonium formate.
d. Aqueous solutions of calcium bromide and cesium carbonate are mixed.

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Summary

In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed. To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

Contributions & Attributions

- Modified by Joshua Halpern (Howard University)

Answers

Exercise 8.2.2

$Fe(OH)_2$ precipitate is formed.

Exercise 8.2.2

$Hg_3(PO_4)_2$ precipitate is formed.

Exercise 8.2.2

No Reaction.

Exercise 8.2.2

$CaCO_3$ is precipitate formed.

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8.3: Writing Chemical Equations for Reactions in Solution- Molecular, Complete Ionic, and Net Ionic Equations

A typical precipitation reaction occurs when an aqueous solution of barium chloride is mixed with one containing sodium sulfate. The **complete chemical equation** can be written to describe what happens, and such an equation is useful in making chemical calculations.

$$\underbrace{\mathrm{BaCl_2\,(aq) + Na_2SO_4\,(aq) \longrightarrow BaSO_4\,(s) + 2\,NaCl\,(aq)}}_{\text{Complete Chemical Equation}} \tag{8.3.1}$$

However, Equation 8.3.1 does not really represent the microscopic particles (that is, the ions) present in the solution. Below is the complete ionic equation:

$$\underbrace{\mathrm{Ba^{2+}\,(aq)} + \overbrace{\mathrm{2Cl^-(aq)}}^{\text{spectator}} + \overbrace{\mathrm{2Na^+(aq)}}^{\text{spectator}} + \mathrm{SO_4{}^{2-}\,(aq) \longrightarrow BaSO_4\,(s)} + \overbrace{\mathrm{2Na^+(aq)}}^{\text{spectator}} + \overbrace{\mathrm{Cl^-(aq)}}^{\text{spectator}}}_{\text{Complete Ionic Equation}} \tag{8.3.2}$$

Equation 8.3.2 is rather cumbersome and includes so many different ions that it may be confusing. In any case, we are often interested in the independent behavior of ions, not the specific compound from which they came. A precipitate of $\mathrm{BaSO_4\,(s)}$ will form when *any* solution containing $\mathrm{Ba^{2+}\,(aq)}$ is mixed with *any* solution containing $\mathrm{SO_4{}^{2-}\,(aq)}$ (provided concentrations are not extremely small). This happens independently of the $\mathrm{Cl^-\,(aq)}$ and $\mathrm{Na^+\,(aq)}$ ions in Equation 8.3.2. These ions are called **spectator ions** because they do not participate in the reaction. When we want to emphasize the independent behavior of ions, a **net ionic equation** is written, omitting the spectator ions. For precipitation of $\mathrm{BaSO_4}$ the net ionic equation is

$$\underbrace{\mathrm{Ba^{2+}\,(aq) + SO_4{}^{2-}\,(aq) \longrightarrow BaSO_4\,(s)}}_{\text{Net Ionic Equation}} \tag{8.3.3}$$

Example 8.3.1

a. When a solution of $\mathrm{AgNO_3}$ is added to a solution of $\mathrm{CaCl_2}$, insoluble AgCl precipitates. Write three equations (complete chemical equation, complete ionic equation, and net ionic equation) that describe this process.
b. Write the balanced net ionic equation to describe any reaction that occurs when the solutions of $\mathrm{Na_2SO_4}$ and $\mathrm{NH_4I}$ are mixed.

Solution

Solutions to Example 7.7.1

Equation Type	Example 8.3.1a	Example 8.3.1b
Complete Chemical Equation	$\mathrm{2\,AgNO_3\,(aq) + CaCl_2\,(aq) \longrightarrow\ 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)}$ The proper states and formulas of all products are written and the chemical equation is balanced.	$\mathrm{Na_2SO_4\,(aq) + 2\,NH_4I\,(aq) \longrightarrow 2\,NaI\,(aq) + (NH_4)_2SO_4\,(aq)}$ Both products are aqueous so there is no net ionic equation that can be written.
Complete Ionic Equation	$\mathrm{2\,Ag^+\,(aq) + 2\,NO_3{}^-\,(aq) + Ca^{2+}\,(aq) + 2\,Cl^-\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca^{2+}\,(aq) + 2\,NO_3{}^-\,(aq)}$ AgCl is a solid so it does not break up into ions in solution.	

Equation Type	Example 8.3.1a	Example 8.3.1b
Net Ionic Equation	$\mathrm{Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)}$ All spectator ions are removed. The chemical equation is written using the lowest common coefficients.	NaI and $\mathrm{(NH_4)_2SO_4}$ are both soluble. No net ionic equation. There is no reaction.

The occurrence or nonoccurrence of precipitates can be used to detect the presence or absence of various species in solution. A $\mathrm{BaCl_2}$ solution, for instance, is often used as a test for the presence of $\mathrm{SO_4^{2-}(aq)}$ ions. There are several insoluble salts of Ba, but they all dissolve in dilute acid except for $\mathrm{BaSO_4}$. Thus, if $\mathrm{BaCl_2}$ solution is added to an unknown solution which has previously been acidified, the occurrence of a white precipitate is proof of the presence of the $\mathrm{SO_4^{2-}}$ ion.



Figure 8.3.1: The three common silver halide precipitates: AgI, AgBr and AgCl (left to right). The silver halides precipitate out of solution, but often form suspensions before settling. (CC BY-SA 3.0; Cychr).

$\mathrm{AgNO_3}$ solutions are often used in a similar way to test for halide ions. If $\mathrm{AgNO_3}$ solution is added to an acidified unknown solution, a white precipitate indicates the presence of $\mathrm{Cl^-}$ ions, a cream-colored precipitate indicates the presence of $\mathrm{Br^-}$ ions, and a yellow precipitate indicates the presence of $\mathrm{I^-}$ ions (Figure 8.3.1). Further tests can then be made to see whether perhaps a mixture of these ions is present. When $\mathrm{AgNO_3}$ is added to tap water, a white precipitate is almost always formed. The $\mathrm{Cl^-}$ ions in tap water usually come from the $\mathrm{Cl_2}$ which is added to municipal water supplies to kill microorganisms.

Exercise 8.3.1

Write balanced net ionic equations to describe any reaction that occurs when the following solutions are mixed.

a. $\mathrm{K_2CO_3 + SrCl_2}$
b. $\mathrm{FeSO_4 + Ba(NO_3)_2}$

Answer a available at the end of this section.

Answer b available at the end of this section.

Precipitates are also used for quantitative analysis of solutions, that is, to determine the amount of solute or the mass of solute in a given solution. For this purpose it is often convenient to use the first of the three types of equations described above. Then

the rules of stoichiometry may be applied.

Contributions & Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

Answers

Exercise 8.3.1

$$\mathrm{Sr^{2+}(aq) + {CO_3}^{2-}(aq) \longrightarrow SrCO_3(s)}$$

Exercise 8.3.1

$$\mathrm{Ba^{2+}(aq) + {SO_4}^{2-}(aq) \longrightarrow Ba(SO_4)(s)}$$

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9: MOLES AND STOICHIOMETRY

9.1: The Mole

9.2: Atomic and Molar Masses

9.3: Mole-Mass Conversions

9.4: Mole-Mole Relationships in Chemical Reactions

9.5: Mole-Mass and Mass-Mass Problems

9.6: Solution Concentration- Molarity

9.7: Molality

9.8: Reaction Stoichiometry

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9.1: The Mole

Learning Objectives

- To define the mole unit.

Figure 9.1.1 shows that 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.

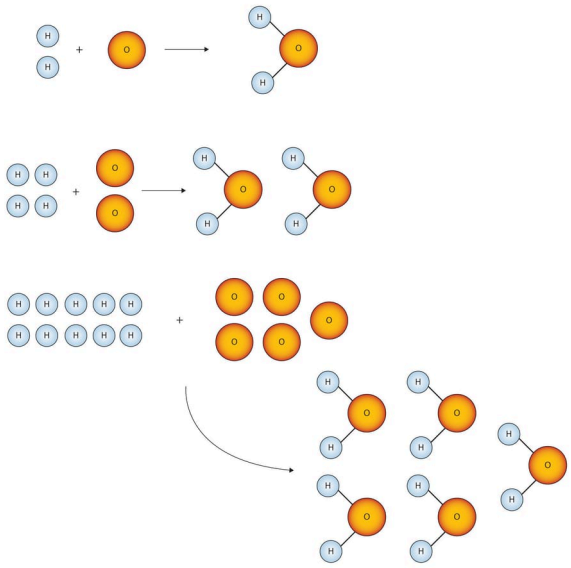


Figure 9.1.1 Water Molecules. The ratio of hydrogen atoms to oxygen atoms used to make water molecules is always 2:1, no matter how many water molecules are being made.

One problem we have, however, is that 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)? We do it by using mass rather than by counting individual atoms.

A hydrogen atom has a mass of approximately 1 atomic mass unit (amu). An oxygen atom has a mass of approximately 16 amu. The ratio of the mass of an oxygen atom to the mass of a hydrogen atom is therefore approximately 16:1.

If we have 2 atoms of each element, the ratio of their masses is approximately 32:2, which reduces to 16:1—the same ratio. If we have 12 atoms of each element, the ratio of their total masses is approximately (12 × 16):(12 × 1), or 192:12, which also reduces to 16:1. If we have 100 atoms of each element, the ratio of the masses is approximately 1,600:100, which again reduces to 16:1. As long as we have equal numbers of hydrogen and oxygen atoms, the ratio of the masses will always be 16:1.

The same consistency is seen when ratios of the masses of other elements are compared. For example, the ratio of the masses of silicon atoms to equal numbers of hydrogen atoms is always approximately 28:1, while the ratio of the masses of calcium atoms to equal numbers of lithium atoms is approximately 40:7.

So we have established that the masses of atoms are constant with respect to each other, as long as we have the same number of each type of atom. Consider a more macroscopic example. If a sample contains 40 g of Ca, this sample has the same number of atoms as there are in a sample of 7 g of Li. What we need, then, is a number that represents a convenient quantity of atoms so we can relate macroscopic quantities of substances. Clearly even a dozen small to easily quantify the number of atoms in a sample because atoms themselves are so small. We need a number that represents billions and billions of atoms.

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.

Notice that we are applying the mole unit to different types of **chemical entities**. The word *mole* represents a number of things, 6.022×10^{23} of them, but does not by itself specify what "they" are. The chemical entities can be ***atoms***, ***molecules***, ***formula units*** and ***ions***. This specific information needs to be specified accurately. Most students find this confusing hence, we need to review the composition of elements, covalent and ionic compounds.

Most elements are made up of individual ***atoms***, such as helium (He). However, some elements consist of ***molecules***, such as the diatomic elements, nitrogen (N_2), hydrogen (H_2), oxygen (O_2), etc. One mole of He consists of 6.022×10^{23} He ***atoms*** but one mole of nitrogen contains 6.022×10^{23} N_2 ***molecules***. The basic units of molecular compounds are ***molecules*** as well. The molecules of "compounds" consist of different kinds of atoms while the molecules of "elements" consist of only one type of atom. For example, the molecules of ammonia (NH_3) consist of nitrogen and hydrogen atoms while N_2 molecules have N atoms only. Compounds that are ionic, like NaCl, are represented by ionic formulas. One mole of NaCl, for example, refers to 6.022×10^{23} ***formula units*** of NaCl. And, one formula unit of NaCl consists of one sodium ***ion*** and one chloride ***ion***. Figure 6.1.2 summarizes the basic units of elements, covalent and ionic compounds

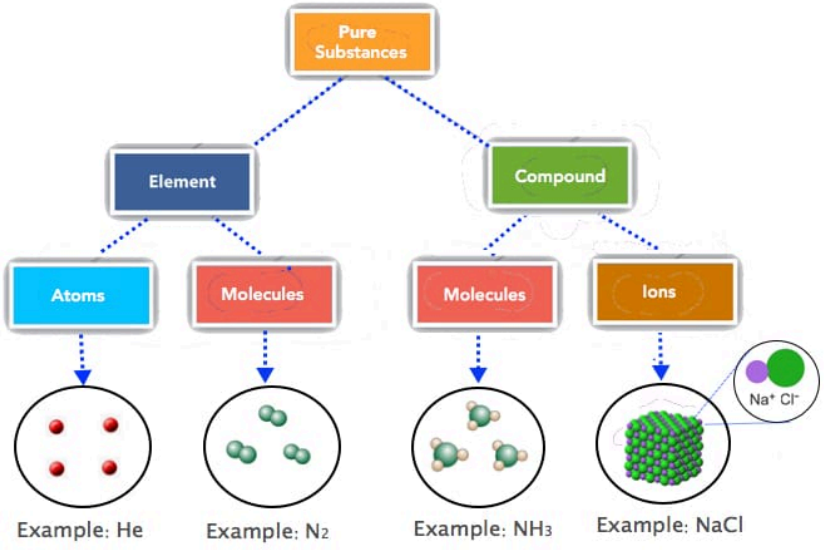


Figure 9.1.2: The basic units of elements (atoms or molecules), covalent compounds (molecules) and ionic compounds (formula units of ions).

Conversion Between Moles and Atoms, Molecules and Ions

We can use the mole relationship and the chemical formula to convert back and forth between the moles and the number of chemical entities (atoms, molecules or ions).

Because 1 N_2 molecule contains 2 N atoms, 1 mol of N_2 molecules (6.022×10^{23} molecules) has 2 mol of N atoms. How many nitrogen atoms are in 0.45 mol N_2?

$$1\ mole\ N_2 = 2\ mole\ N$$

$$1\ mole\ N = 6.022 \times 10^{23}\ atoms\ N$$

$$0.45\ mol\ N_2 \times \frac{2\ mol\ N}{1\ mol\ N_2} \times \frac{6.022 \times 10^{23}\ atoms\ N}{1\ mol\ N} = 5.4 \times 10^{23}\ atoms\ N$$

Using formulas to indicate how many atoms of each element we have in a substance, we can relate the number of moles of molecules to the number of moles of atoms. For example, in 1 mol of ethanol (C_2H_6O), we can construct the following relationships (Table 9.1.1):

Table 9.1.1: Molecular Relationships

1 Molecule of C_2H_6O Has	1 Mol of C_2H_6O Has	Molecular Relationships
2 C atoms	2 mol of C atoms	$\frac{2\ \text{mol C atoms}}{1\ \text{mol } C_2H_6O\ \text{molecules}}$ or $\frac{1\ \text{mol } C_2H_6O\ \text{molecules}}{2\ \text{mol C atoms}}$
6 H atoms	6 mol of H atoms	$\frac{6\ \text{mol H atoms}}{1\ \text{mol } C_2H_6O\ \text{molecules}}$ or $\frac{1\ \text{mol } C_2H_6O\ \text{molecules}}{6\ \text{mol H atoms}}$
1 O atom	1 mol of O atoms	$\frac{1\ \text{mol O atoms}}{1\ \text{mol } C_2H_6O\ \text{molecules}}$ or $\frac{1\ \text{mol } C_2H_6O\ \text{molecules}}{1\ \text{mol O atoms}}$

The following example illustrates how we can use these relationships as conversion factors.

Example 9.1.1

If a sample consists of 2.5 mol of ethanol (C_2H_6O), how many moles of carbon atoms, hydrogen atoms, and oxygen atoms does it have?

Solution

Using the relationships in Table 9.1.1, we apply the appropriate conversion factor for each element:

$$2.5\ \text{mol } C_2H_6O\ \text{molecules} \times \underbrace{\frac{2\ \text{mol C atoms}}{1\ \text{mol } C_2H_6O\ \text{molecules}}}_{\text{conversion factor}} = 5.0\ \text{mol C atoms}$$

Note how the unit *mol C_2H_6O molecules* cancels algebraically. Similar equations can be constructed for determining the number of H and O atoms:

$$2.5\ \text{mol } C_2H_6O\ \text{molecules} \times \frac{6\ \text{mol H atoms}}{1\ \text{mol } C_2H_6O\ \text{molecules}} = 15\ \text{mol H atoms}$$

$$2.5\ \text{mol } C_2H_6O\ \text{molecules} \times \frac{1\ \text{mol O atoms}}{1\ \text{mol } C_2H_6O\ \text{molecules}} = 2.5\ \text{mol O atoms}$$

Exercise 9.1.1

If a sample contains 6.75 mol of Na_2SO_4, how many moles of sodium atoms, sulfur atoms, and oxygen atoms does it have?

Answer available at the end of this section.

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. For example, if we are dealing with element X, the mole relationship is expressed as follows:

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms}$$

We can convert this relationship into two possible conversion factors shown below:

$$\frac{1 \text{ mol X}}{6.022 \times 10^{23} \text{ X atoms}} \text{ or } \frac{6.022 \times 10^{23} \text{ X atoms}}{1 \text{ mol X}}$$

If the number of "atoms of element X" is given, we can convert it into "moles of X" by multiplying the given value with the conversion factor at the left. However, if the number of "mol of X" is given, the appropriate conversion factor to use is the one at the right.

If we are dealing with a molecular compound (such as C_4H_{10}), the mole relationship is expressed as follows:

'_' allowed only in math mode

If working with ionic compounds (such as NaCl), the mole relationship is expressed as follows:

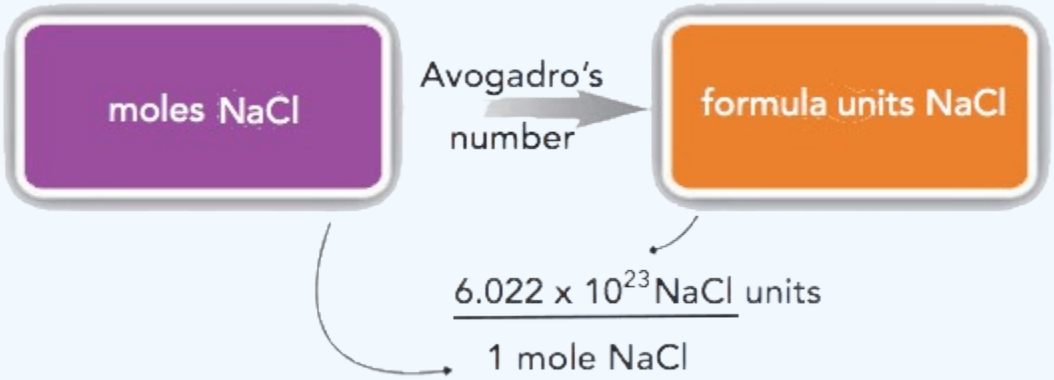
$$1 \text{ mol NaCl} = 6.022 \times 10^{23} \text{ NaCl formula units}$$

Example 9.1.2

How many formula units are present in 2.34 mol of NaCl? How many ions are in 2.34 mol?

Solution

Typically in a problem like this, we start with what we are given and apply the appropriate conversion factor. Here, we are given a quantity of 2.34 mol of NaCl, to which we can apply the definition of a **mole as a conversion factor**:



$$2.34 \text{ mol NaCl} \times \frac{6.022 \times 10^{23} \text{ NaCl units}}{1 \text{ mol NaCl}} = 1.41 \times 10^{24} \text{ NaCl units}$$

Because there are two ions per formula unit, there are

$$1.41 \times 10^{24} \text{ NaCl units} \times \frac{2 \text{ ions}}{\text{NaCl units}} = 2.82 \times 10^{24} \text{ ions}$$

in the sample.

Exercise 9.1.2

How many molecules are present in 16.02 mol of C_4H_{10}? How many atoms are in 16.02 mol?

Answer available at the end of this section.

Key Takeaway

- A mole is 6.022×10^{23} things.

Answers

Exercise 9.1.1

13.5 mol Na, 6.75 mol S and 27 mol O.

Exercise 9.1.2

9.647×10^{24}molecules, 1.351×10^{26} atoms.

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9.2: Atomic and Molar Masses

Learning Objectives

- To learn how the masses of moles of atoms and molecules are expressed.

Now that we have introduced the mole and practiced using it as a conversion factor, we ask the obvious question: why is the mole *that particular* number of things? Why is it 6.022×10^{23} and not 1×10^{23} or even 1×10^{20}?

The number in a mole, Avogadro's number, is related to the relative sizes of the **atomic mass unit** and **gram mass units**. Whereas one hydrogen atom has a mass of approximately 1 u, **1 mol of H atoms** has a mass of approximately **1 *gram***. And whereas one sodium atom has an approximate mass of 23 u, 1 mol of Na atoms has an approximate mass of 23 *grams*.

One mole of a substance has the same mass in grams that one atom or molecule has in atomic mass units. The numbers in the periodic table that we identified as the atomic masses of the atoms not only tell us the mass of one atom in u but also tell us the mass of 1 mol of atoms in grams.

One mole of a substance has the same mass in grams that one atom or molecule has in atomic mass units.

Example 9.2.1: Moles to Mass Conversion with Elements

What is the mass of each quantity?

a. 1 mol of Al atoms
b. 2 mol of U atoms

Solution

a. One mole of Al atoms has a mass in grams that is numerically equivalent to the atomic mass of aluminum. The periodic table shows that the atomic mass (rounded to two decimal points) of Al is 26.98, so 1 mol of Al atoms has a mass of 26.98 g.
b. According to the periodic table, 1 mol of U has a mass of 238.0 g, so the mass of 2 mol is twice that, or 476.0 g.

Exercise 9.2.1: Moles to Mass Conversion with Elements

What is the mass of each quantity?

a. 1 mol of Au atoms
b. 5 mol of Br atoms

Answer a available at the end of this section.

Answer b available at the end of this section.

The mole concept can be extended to masses of formula units and molecules as well. The mass of 1 mol of molecules (or formula units) in grams is numerically equivalent to the mass of one molecule (or formula unit) in atomic mass units. For example, a single molecule of O_2 has a mass of 32.00 u, and 1 mol of O_2 molecules has a mass of 32.00 g. As with atomic mass unit–based masses, to obtain the mass of 1 mol of a substance, we simply sum the masses of the individual atoms in the formula of that substance. The mass of 1 mol of a substance is referred to as its **molar mass**, whether the substance is an element, an ionic compound, or a covalent compound.

Example 9.2.2: Moles to Mass Conversion with Compounds

What is the mass of 1 mol of each substance?

a. NaCl
b. bilirubin ($C_{33}H_{36}N_4O_6$), the principal pigment present in bile (a liver secretion)

Solution a

1. Summing the molar masses of the atoms in the NaCl formula unit gives

Solutions to Example 6.2.2, Summing the molar masses of the atoms in the NaCl formula unit gives

1 Na molar mass:	22.99 g
1 Cl molar mass:	35.45 g
Total:	58.44 g

The mass of 1 mol of NaCl is 58.44 g.

Solution b

Multiplying the molar mass of each atom by the number of atoms of that type in bilirubin's formula and adding the results, we get

Solutions to Example 6.2.2, Multiplying the molar mass of each atom by the number of atoms of that type in bilirubin's formula and adding the results, we get

33 C molar mass:	33 × 12.01 g	396.33 g
36 H molar mass:	36 × 1.01 =	36.36 g
4 N molar mass:	4 × 14.01 =	56.04 g
6 O molar mass:	6 × 16.00 =	96.00 g
Total:		584.73 g

The mass of 1 mol of bilirubin is 584.73 g.

Exercise 9.2.2: Moles to Mass Conversion with Compounds

What is the mass of 1 mol of each substance?

a. barium sulfate ($BaSO_4$), used to take X rays of the gastrointestional tract
b. adenosine ($C_{10}H_{13}N_5O_4$), a component of cell nuclei crucial for cell division

Answer a available at the end of this section.

Be careful when counting atoms. In formulas with polyatomic ions in parentheses, the subscript outside the parentheses is applied to every atom inside the parentheses. For example, the molar mass of $Ba(OH)_2$ requires the sum of 1 mass of Ba, 2 masses of O, and 2 masses of H:

The molar mass of Ba(OH)2 requires the sum of 1 mass of Ba, 2 masses of O, and 2 masses of H:

1 Ba molar mass:	1 × 137.3 g =	137.3 g
2 O molar mass:	2 × 16.00 g =	32.00 g

1 Ba molar mass:	1 × 137.3 g =	137.3 g
2 H molar mass:	2 × 1.01 g =	2.02 g
Total:		171.32 g

Because molar mass is defined as the mass for 1 mol of a substance, we can refer to molar mass as grams per mole (g/mol). The division sign (/) implies “per,” and “1” is implied in the denominator. Thus, the molar mass of bilirubin can be expressed as 584.73 g/mol, which is read as “five hundred eighty four point seventy three grams per mole.”

Key Takeaway

- The mass of moles of atoms and molecules is expressed in units of grams.

Answers

Exercise 9.2.1: Moles to Mass Conversion with Elements

197.0 g

Exercise 9.2.1: Moles to Mass Conversion with Elements

5 mol Br atoms x 79.90 g/mol = 399.5 g

Exercise 9.2.2: Moles to Mass Conversion with Compounds

233.36 g

267.28 g

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9.3: Mole-Mass Conversions

Learning Objectives

- To convert quantities between mass units and mole units.

A previous Example stated that the mass of 2 mol of U is twice the molar mass of uranium. 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.

The simplest type of manipulation using molar mass as a conversion factor is a mole-mass conversion (or its reverse, a mass-mole conversion). In such a conversion, we use the molar mass of a substance as a conversion factor to convert mole units into mass units (or, conversely, mass units into mole units).

We also established that 1 mol of Al has a mass of 26.98 g (Example $9.3.1$). 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:

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

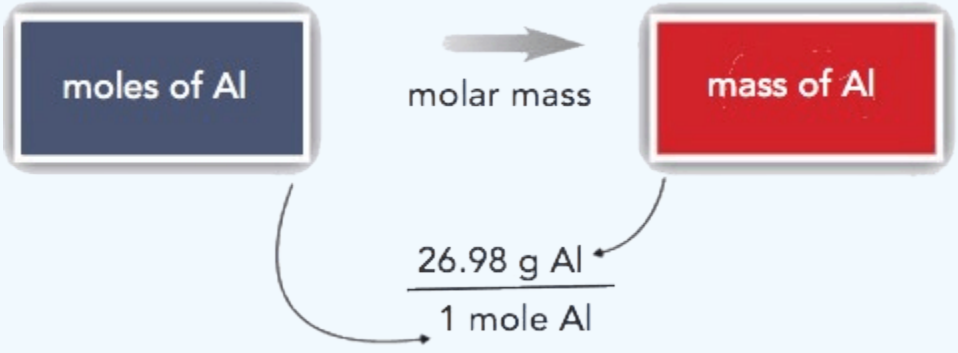
The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do “by eye.”

Example $9.3.1$

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 9.3.1

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 available at the end of this section.

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 9.3.1 is a chart for determining what conversion factor is needed, and Figure 9.3.2 is a flow diagram for the steps needed to perform a conversion.

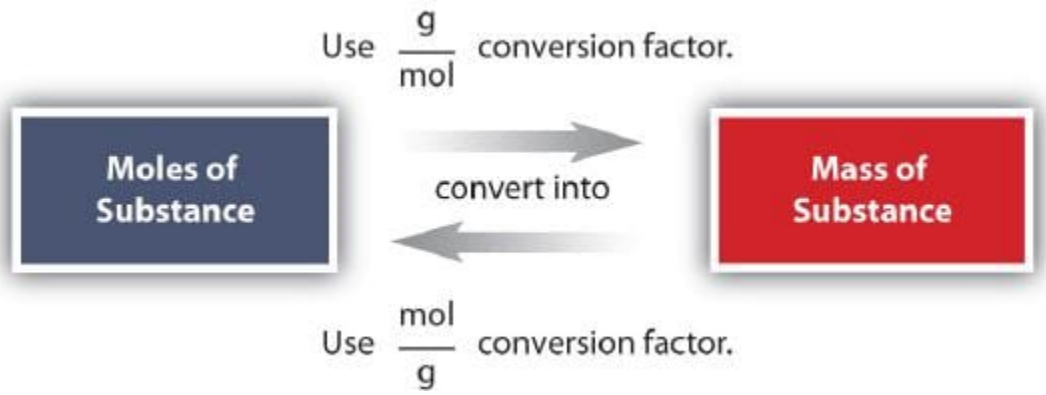


Figure 9.3.1 A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles.

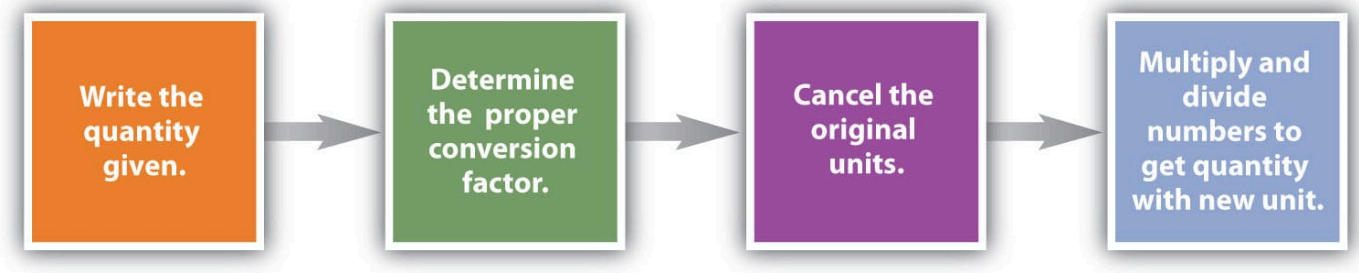


Figure 9.3.2 A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

Example 9.3.2

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 molar mass of bilirubin, which we can determine from its chemical formula:

Solutions to Example 6.3.2

33 C molar mass:	33 × 12.01 g =	396.33 g
36 H molar mass:	36 × 1.01 g =	36.36 g
4 N molar mass:	4 × 14.01 g =	56.04 g
6 O molar mass:	6 × 16.00 g =	96.00 g
Total:		584.73 g

The molar mass 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 9.3.2:

$$0.00655 \text{ mol bilirubin} \times \frac{584.73 \text{ g bilirubin}}{\text{mol bilirubin}} = 3.83 \text{ g bilirubin}$$

The mol bilirubin unit cancels. The biochemist needs 3.83 g of bilirubin.

Exercise 9.3.2

A chemist needs 457.8 g of $KMnO_4$ to make a solution. How many moles of $KMnO_4$ is that?

Answer available at the end of this section.

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 9.3.1: 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

Mineral	Male (age 19–30 y)		Female (age 19–30 y)	
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 9.3.1 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

- It is possible to convert between moles of material and mass of material.

Answers

Exercise 9.3.1

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

Exercise 9.3.2

$$457.8 \text{ g } KMnO_4 \times \frac{1 \text{ mol } KMnO_4}{158.04 \text{ g } KMnO_4} = 2.897 \text{ mol } KMnO_4$$

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9.4: Mole-Mole Relationships in Chemical Reactions

Learning Objectives

- To use a balanced chemical reaction to determine molar relationships between the substances.

Previously, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). Here, we will extend the meaning of the coefficients in a chemical equation.

Consider the simple chemical equation

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

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4\,H_2 + 2\,O_2 \longrightarrow 4\,H_2O$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22\,H_2 + 11\,O_2 \longrightarrow 22\,H_2O$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ H_2 + 6.022 \times 10^{23}\ O_2 \rightarrow 12.044 \times 10^{23}\ H_2O$$

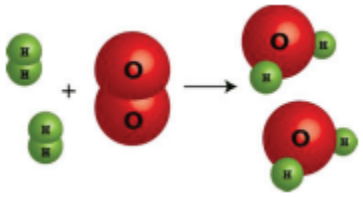
These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ \text{mol}\ \ H_2 + 1\ \ \text{mol}\ \ O_2 \longrightarrow 2\ \text{mol}\ \ H_2O$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

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

Now we interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? *Balanced chemical equations are balanced not only at the molecular level but also in terms of molar amounts of reactants and products*. Thus, we can read this reaction as “two moles of hydrogen react with one mole of oxygen to produce two moles of water.”



2 molecules H_2 1 molecule O_2 2 molecules H_2O

2 moles H_2 1 mole O_2 2 moles H_2O

2 x 2.02 g=4.04 g H_2 32.0 g O_2 2 x 18.02 g=36.04 g H_2O

Figure 9.4.1: This representation of the production of water from oxygen and hydrogen show several ways to interpret the quantitative information of a chemical reaction.

By the same token, the ratios we constructed to describe molecules reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

$$\frac{2\ \text{mol}\ H_2}{1\ \text{mol}\ O_2}\ \text{or}\ \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2}$$

$$\frac{2\ \text{mol}\ H_2O}{1\ \text{mol}\ O_2}\ \text{or}\ \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O}$$

$$\frac{2\ \text{mol}\ H_2}{2\ \text{mol}\ H_2O}\ \text{or}\ \frac{2\ \text{mol}\ H_2O}{2\ \text{mol}\ H_2}$$

We can use these ratios to determine what amount of a substance, in moles, will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical reactions is called ***stoichiometry*.** The ratio of coefficients in a balanced chemical equation, used in computations relating amounts of reactants and products is called the **stoichiometric factor.**

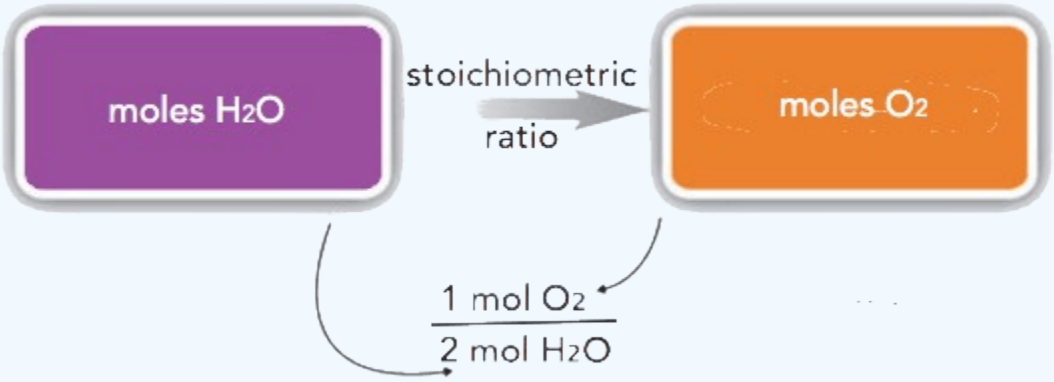
Example 9.4.1

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O? The balanced equation is as follows:

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

Solution

Because we are dealing with quantities of H_2O and O_2, we will use the **stoichiometric ratio** that relates those two substances. Because we are given an amount of H_2O and want to determine an amount of O_2, we will use the ratio that has H_2O in the denominator (so it cancels) and O_2 in the numerator (so it is introduced in the answer). Thus,



$$27.6\ \text{mol}\ H_2O \times \frac{1\ \text{mol}\ O_2}{2\ \text{mol}\ H_2O} = 13.8\ \text{mol}\ O_2$$

To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.

Exercise 9.4.1

Using $2H_2 + O_2 \rightarrow 2H_2O$, how many moles of hydrogen react with 3.07 mol of oxygen to produce H_2O?

Answer available at the end of this section.

Key Takeaway

- The balanced chemical reaction can be used to determine molar relationships between substances.

Answers

Exercise 9.4.1

$$3.07 \text{ mol } O_2 \times \frac{2 \text{ mol } H_2}{1 \text{ mol } O_2} = 6.14 \text{ mol } H_2$$

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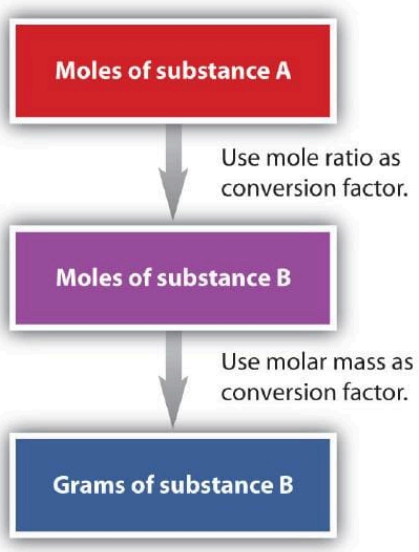
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9.5: Mole-Mass and Mass-Mass Problems

Learning Objectives

- To convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.

We have established that a balanced chemical equation is balanced in terms of moles as well as atoms or molecules. We have used balanced equations to set up ratios, now in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions, such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that ability to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:



Collectively, these conversions are called mole-mass calculations.

As an example, consider the balanced chemical equation

$$Fe_2O_3 + 3\,SO_3 \rightarrow Fe_2(SO_4)_3$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. Then using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

As usual, we start with the quantity we were given:

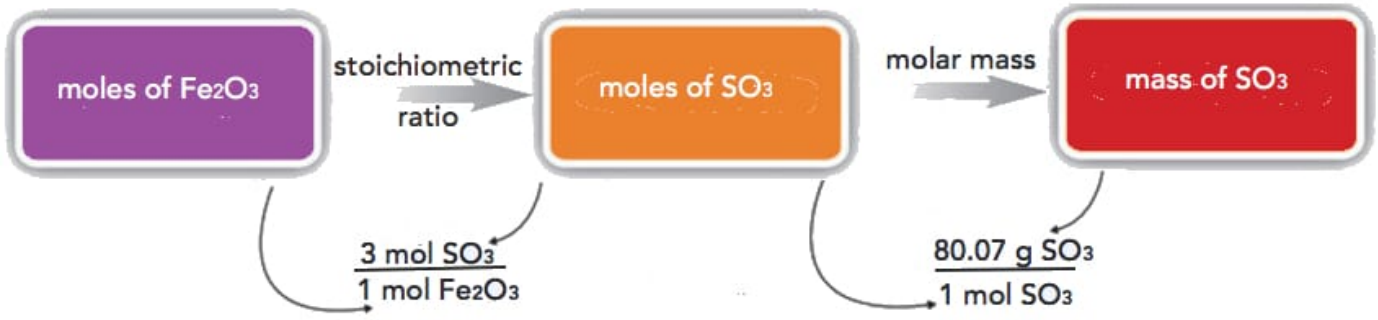
$$3.59\ \text{mol}\ Fe_2O_3 \times \frac{3\ \text{mol}\ SO_3}{1\ \text{mol}\ Fe_2O_3} = 10.77\ \text{mol}\ SO_3$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$10.77\ \text{mol}\ SO_3 \times \frac{80.07\ \text{g}\ SO_3}{1\ \text{mol}\ SO_3} = 862.4\ \text{g}\ SO_3$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner.

The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:



$$3.59 \text{ mol } Fe_2O_3 \times \underbrace{\frac{3 \text{ mol } SO_3}{1 \text{ mol } Fe_2O_3}}_{\text{converts to moles of } SO_3} \times \underbrace{\frac{80.07 \text{ g } SO_3}{1 \text{ mol } SO_3}}_{\text{converts to grams of } SO_3} = 862.4 \text{ g } SO_3$$

We get exactly the same answer when combining all the math steps together as we do when we calculate one step at a time.

Example 9.5.1

How many grams of CO_2 are produced if 2.09 mol of HCl are reacted according to this balanced chemical equation?

$$CaCO_3 + 2\,HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

Solution

Our strategy will be to convert from moles of HCl to moles of CO_2 and then from moles of CO_2 to grams of CO_2. We will need the molar mass of CO_2, which is 44.01 g/mol. Performing these two conversions in a single-line gives 46.0 g of CO_2:

$$2.09 \text{ mol HCl} \times \underbrace{\frac{1 \text{ mol } CO_2}{2 \text{ mol HCl}}}_{\text{from the coefficients of the balanced equation}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2} = 46.0 \text{ g } CO_2$$

The molar ratio between CO_2 and HCl comes from the balanced chemical equation.

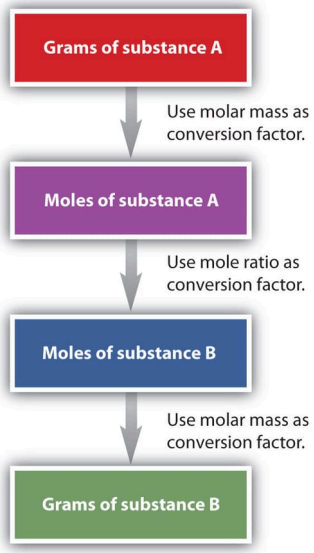
Exercise

How many grams of glucose ($C_6H_{12}O_6$) are produced if 17.3 mol of H_2O are reacted according to this balanced chemical equation?

$$6\,CO_2 + 6\,H_2O \rightarrow C_6H_{12}O_6 + 6\,O_2$$

Answer available at the end of this section.

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques.

Example 9.5.2: Chlorination of Carbon

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows:

$$CH_4 + 4\,Cl_2 \rightarrow CCl_4 + 4\,HCl$$

How many grams of HCl are produced by the reaction of 100.0 g of CH_4?

Solution

First, let us work the problem in stepwise fashion. We begin by converting the mass of CH_4 to moles of CH_4, using the molar mass of CH_4 (16.05 g/mol) as the conversion factor:

$$100.0\text{ g } CH_4 \times \frac{1\text{ mol } CH_4}{16.05\text{ g } CH_4} = 6.231\text{ mol } CH_4$$

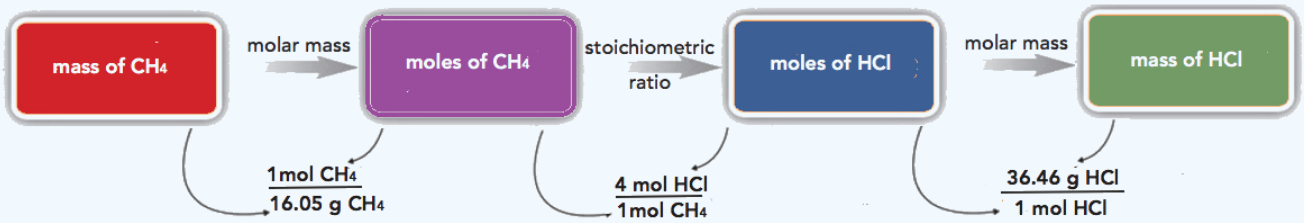
Note that we inverted the molar mass so that the gram units cancel, giving us an answer in moles. Next, we use the balanced chemical equation to determine the ratio of moles CH_4 and moles HCl and convert our first result into moles of HCl:

$$6.231 \text{ mol } CH_4 \times \frac{4 \text{ mol HCl}}{1 \text{ mol } CH_4} = 24.92 \text{ mol HCl}$$

Finally, we use the molar mass of HCl (36.46 g/mol) as a conversion factor to calculate the mass of 24.92 mol of HCl:

$$24.92 \text{ mol HCl} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 908.5 \text{ g HCl}$$

In each step, we have limited the answer to the proper number of significant figures. If desired, we can do all three conversions on a single line:



$$100.0 \text{ g } CH_4 \times \frac{1 \text{ mol } CH_4}{16.05 \text{ g } CH_4} \times \frac{4 \text{ mol HCl}}{1 \text{ mol } CH_4} \times \frac{36.46 \text{ g HCl}}{1 \text{ mol HCl}} = 908.7 \text{ g HCl}$$

This final answer is slightly different from our first answer because only the final answer is restricted to the proper number of significant figures. In the first answer, we limited each intermediate quantity to the proper number of significant figures. As you can see, both answers are essentially the same.

Exercise 9.5.2: Oxidation of Propanal

The oxidation of propanal (CH_3CH_2CHO) to propionic acid (CH_3CH_2COOH) has the following chemical equation:

$$CH_3CH_2CHO + 2K_2Cr_2O_7 \rightarrow CH_3CH_2COOH + \text{other products}$$

How many grams of propionic acid are produced by the reaction of 135.8 g of $K_2Cr_2O_7$?

Answer available at the end of this section.

To Your Health: The Synthesis of Taxol

Taxol is a powerful anticancer drug that was originally extracted from the Pacific yew tree (*Taxus brevifolia*). As you can see from the accompanying figure, taxol is a very complicated molecule, with a molecular formula of $C_{47}H_{51}NO_{14}$. Isolating taxol from its natural source presents certain challenges, mainly that the Pacific yew is a slow-growing tree, and the equivalent of six trees must be harvested to provide enough taxol to treat a single patient. Although related species of yew trees also produce taxol in small amounts, there is significant interest in synthesizing this complex molecule in the laboratory.

After a 20-year effort, two research groups announced the complete laboratory synthesis of taxol in 1994. However, each synthesis required over 30 separate chemical reactions, with an overall efficiency of less than 0.05%. To put this in perspective, to obtain a single 300 mg dose of taxol, you would have to begin with 600 g of starting material. To treat the 26,000 women who are diagnosed with ovarian cancer each year with one dose, almost 16,000 kg (over 17 tons) of starting material must be converted to taxol. Taxol is also used to treat breast cancer, with which 200,000 women in the United States are diagnosed every year. This only increases the amount of starting material needed.

Clearly, there is intense interest in increasing the overall efficiency of the taxol synthesis. An improved synthesis not only will be easier but also will produce less waste materials, which will allow more people to take advantage of this potentially life-saving drug.

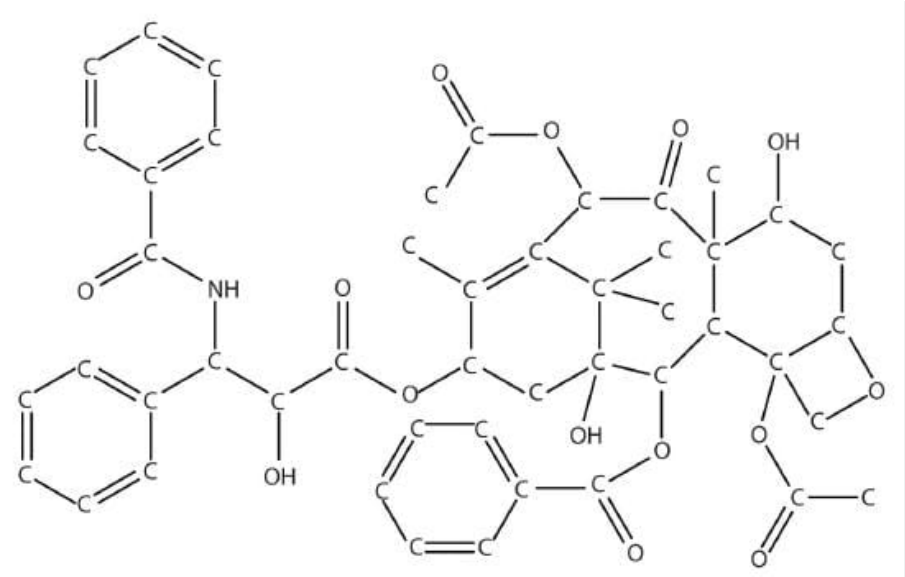


Figure 9.5.1 The Structure of the Cancer Drug Taxol. Because of the complexity of the molecule, hydrogen atoms are not shown, but they are present on every atom to give the atom the correct number of covalent bonds (four bonds for each carbon atom).

Key Takeaway

- A balanced chemical equation can be used to relate masses or moles of different substances in a reaction.

Answers

Exercise

$$17.3 \text{ mol } H_2O \times \frac{1 \text{ mol } C_6H_{12}O_6}{6 \text{ mol } H_2O} \times \frac{180.18 \text{ g } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6} = 520 \text{ g } C_6H_{12}O_6$$

Exercise 9.5.2: Oxidation of Propanal

$$135.8 \text{ g } K_2Cr_2O_7 \times \frac{1 \text{ mol } K_2Cr_2O_7}{294.20 \text{ g } K_2Cr_2O_7} \times \frac{1 \text{ mol } CH_3CH_2COOH}{2 \text{ mol } K_2Cr_2O_7} \times \frac{74.09 \text{ g } CH_3CH_2COOH}{1 \text{ mol } CH_3CH_2COOH} = 17.10 \text{ g } CH_3CH_2COOH$$

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9.6: Solution Concentration- Molarity

Learning Objectives

- Use molarity to determine quantities in chemical reactions.
- Use molarity as a conversion factor in calculations.

Another way of expressing concentration is to give the number of moles of solute per unit volume of solution. Of all the quantitative measures of concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{9.6.1}$$

The symbol for molarity is M or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[\mathrm{Ag}^+]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to perform calculations with, but the most difficult to make in the lab. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

It is important to remember that “mol” in this expression refers to moles of solute and that “L” refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is

$$\frac{1.5\ \text{mol NaCl}}{0.500\ \text{L solution}} = 3.0\ \text{M NaCl}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” This is read as "a 3.00 ***molar*** sodium chloride solution," meaning that there are 3.00 moles of NaOH dissolved per one liter of solution.

> Be sure to note that molarity is calculated as the total volume of the **entire** solution, not just volume of solvent! The solute contributes to total volume.

If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L?

Step 1: First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ \cancel{g HCl} \times \frac{1\ mol\ HCl}{36.5\ \cancel{g HCl}} = 0.614\ mol\ HCl$$

Step 2: Now we can use the definition of molarity to determine a concentration:

$$M\ =\ \frac{0.614\ mol\ HCl}{1.56 L\ solution} = 0.394\ M HCl$$

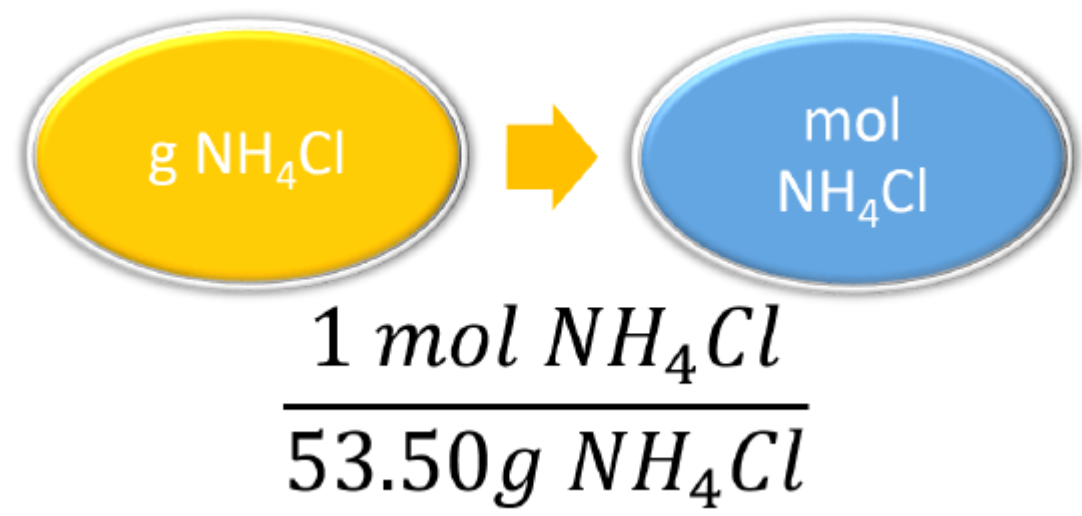
Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

Example 9.6.1

A solution is prepared by dissolving $42.23\ \mathrm{g}$ of $\mathrm{NH_4Cl}$ into enough water to make $500.0\ \mathrm{mL}$ of solution. Calculate its molarity.

Solution

Solutions to Example 13.6.1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Mass $= 42.23\ \text{g}\ \text{NH}_4\text{Cl}$ Volume solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$ Find: Molarity = ? M
List other known quantities.	Molar mass $\text{NH}_4\text{Cl} = 53.50\ \text{g/mol}$
Plan the problem.	1. The mass of the ammonium chloride is first converted to moles.  2. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters. $\text{M} = \frac{\text{mol NH}_4\text{Cl}}{\text{L solution}}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23\ \cancel{\text{g NH}_4\text{Cl}} \times \frac{1\ \text{mol NH}_4\text{Cl}}{53.50\ \cancel{\text{g NH}_4\text{Cl}}} = 0.7893\ \text{mol NH}_4\text{Cl}$ $\frac{0.7893\ \text{mol NH}_4\text{Cl}}{0.5000\ \text{L solution}} = 1.579\ \text{M}$
Think about your result.	The molarity is $1.579\ \text{M}$, meaning that a liter of the solution would contain $1.579\ \text{mol}\ \text{NH}_4\text{Cl}$. Four significant figures are appropriate.

Exercise $9.6.1A$

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

Exercise $9.6.1B$

What is the concentration, in $\mathrm{mol/L}$, where $137\ \mathrm{g}$ of NaCl has been dissolved in enough water to make $500\ \mathrm{mL}$ of solution?

Answer available at the end of this section.

Using Molarity in Calculations

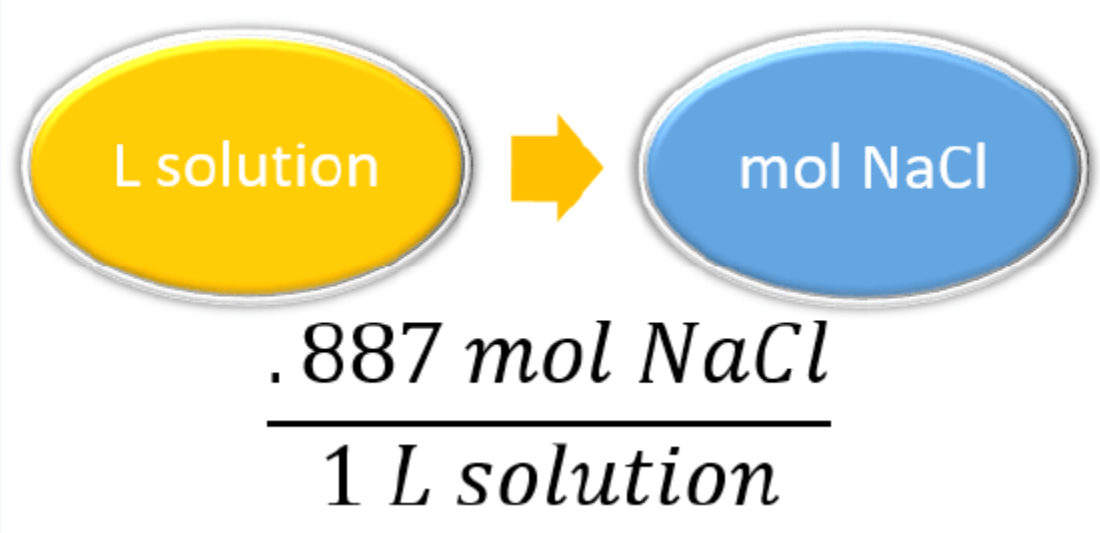
Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition.

Example $9.6.2$: Determining Moles of Solute, Given the Concentration and Volume of a Solution

For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

Solution



$$0.108\ \cancel{L\ solution} \times \frac{0.887\ mol\ NaCl}{1\ \cancel{L\ solution}} = 0.0958\ mol\ NaCl$$

If we used the definition approach, we get the same answer, but now we are using conversion factor skills.

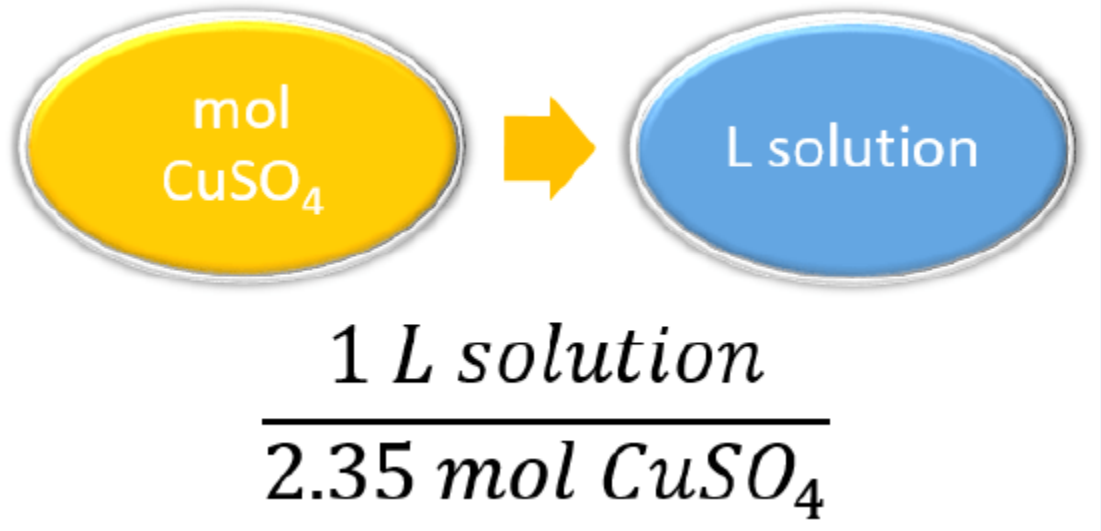
Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 9.6.3: Determining Volume of a Solution, Given the Concentration and Moles of Solute

Using concentration as a conversion factor, how many liters of 2.35 M CuSO4 are needed to obtain 4.88 mol of CuSO4?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:



$$4.88\ \cancel{mol\ CuSO_4} \times \frac{1\ L\ solution}{2.35\ \cancel{mol\ CuSO_4}} = 2.08\ L\ of\ solution$$

In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams. The following example illustrates this.

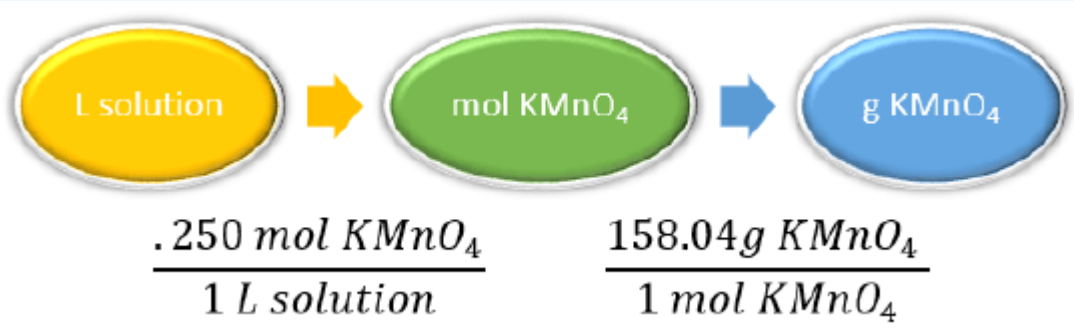
Example 9.6.4

A chemist needs to prepare $3.00\ \text{L}$ of a $0.250\ \text{M}$ solution of potassium permanganate (KMnO_4). What mass of KMnO_4 does she need to make the solution?

Solution

Solutions to Example 13.6.4

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: Molarity $= 0.250\ \text{M}$ Volume $= 3.00\ \text{L}$ Find: Mass $\text{KMnO}_4 = ?\ \text{g}$

Steps for Problem Solving	
List other known quantities.	Molar mass $\mathrm{KMnO_4} = 158.04\ \mathrm{g/mol}$ 0.250 mol $KMnO_4$ to 1 L of $KMnO_4$ solution
Plan the problem.	
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $\mathrm{mol\ KMnO_4} = 0.250\ \mathrm{M\ KMnO_4} \times 3.00\ \mathrm{L} = 0.750\ \mathrm{mol\ KMnO_4}$ $\cancel{3.00\ L\ solution} \times \frac{0.250\ \cancel{\mathrm{mol\ KMnO_4}}}{1\ \cancel{L\ solution}} \times \frac{158.04\ \mathrm{g\ KMnO_4}}{1\ \cancel{\mathrm{mol\ KMnO_4}}} = 119\ \mathrm{g\ KMnO_4}$
Think about your result.	When $119\ \mathrm{g}$ of potassium permanganate is dissolved into water to make $3.00\ \mathrm{L}$ of solution, the molarity is $0.250\ \mathrm{M}$.

Exercise $9.6.4A$

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer available at the end of this section.

Exercise $9.6.4B$

Answer the problems below using concentration as a conversion factor.

a. What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?
b. What volume of 1.50 M HCl solution contains 10.0 g of hydrogen chloride?

Answer a available at the end of this section.

How to Indicate Concentration

- Square brackets are often used to represent concentration, e.g., [NaOH] = 0.50 M.
- Use the capital letter M for molarity, not a lower case m (this is a different concentration unit called *molality*).

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Watch as the Flinn Scientific Tech Staff demonstrates "How To Prepare Solutions."

It is important to note that there are many different ways you can set up and solve your chemistry equations. Some students prefer to answer multi-step calculations in one long step, while others prefer to work out each step individually. Neither method is necessarily better or worse than the other method—whichever makes the most sense to ***you*** is the one that you should use. In this text, we will typically use unit analysis (also called dimension analysis or factor analysis).

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.

Answers

Exercise $9.6.1A$

1.57 M $C_6H_{12}O_6$

Exercise $9.6.1B$

4.69 M NaCl

Exercise $9.6.4A$

1.74 L

Exercise $9.6.4B$

6.12 g

183 mL or 0.183L

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9.7: Molality

Learning Objectives

- To calculate molarity and use it in calculations
- To compare and contrast molarity and molality

For many purposes, molarity is very convenient concentration unit. However, when we want to know the concentration of solute present in situations where there are temperature changes, molarity will not work. The volume of the solution will change somewhat with temperature, enough to affect the accuracy of data observations and calculations. Another unit is needed, one not affected by the temperature of the material we are studying.

Molality

Another way to express the concentration of a solution is by its molality. The **molality** (m) of a solution is the moles of solute divided by the kilograms of solvent. A solution that contains $1.0\ \text{mol}$ of NaCl dissolved into $1.0\ \text{kg}$ of water is a "one-molal" solution of sodium chloride. The symbol for molality is a lower-case m written in italics.

$$\text{Molality}\,(m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{\text{mol}}{\text{kg}}$$

Molality differs from molarity only in the denominator. While molarity is based on the liters of solution, molality is based on the kilograms of solvent. Concentrations expressed in molality are used when studying properties of solutions related to vapor pressure and temperature changes. Molality is used because its value does not change with changes in temperature. The volume of a solution, on the other hand, is slightly dependent upon temperature. This can be helpful when studying the human body since molarity would be different at room temperature vs body temperature but the molality of a solution would be the same at both temperatures.

*Careful! There is only one letter that is different in the words molarity and molality, but they are different units.

Molarity vs. Molality

Aspect/ Property	Molarity	Molality
Unit for numerator	moles	moles
Substance for numerator	solute	solute
Property and unit for denominator	volume in L	mass in kg
Substance for denominator	solution (both solute and solvent)	solvent
Benefit	It is usually easy to measure the volume of solution in a lab, even for someone who did not prepare the sample themselves.	Volume changes with temperature but mass and moles do not, so molality stays constant at high and low temperatures.
Summary	**moles of solute per L of solution**	**moles of solute per kg of solvent**

Example 9.7.1

Determine the molality of a solution prepared by dissolving $28.60\ \text{g}$ of glucose $(\text{C}_6\text{H}_{12}\text{O}_6)$ into $250\ \text{g}$ of water.

Solution

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

Known

- Mass solute $= 28.60\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6$
- Mass solvent $= 250\ \text{g} = 0.250\ \text{kg}$
- Molar mass $\text{C}_6\text{H}_{12}\text{O}_6 = 180.18\ \text{g/mol}$

Unknown

Convert grams of glucose to moles and divide by the mass of the water in kilograms.

Step 2: Solve.

Convert mass of solute to moles of solute:

$$28.60\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{1\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6}{180.18\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6} = 0.1587\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6$$

Solve for the molality:

$$Molality = \frac{moles\ solute}{mass\ solvent} = \frac{0.1587\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6}{0.250\ \text{kg}\ \text{H}_2\text{O}} = 0.63\ m\ \text{C}_6\text{H}_{12}\text{O}_6$$

Step 3: Think about your result.

The answer represents the moles of glucose per kilogram of water and has two significant figures.

Exercise 9.7.1

What is the mass of salt, NaCl, in a 0.98 *m* solution prepared with 1.5 kg of water?

Answer available at the end of this section.

Molality and molarity are closely related in value for dilute aqueous solutions because the density of those solutions is relatively close to $1.0\ \text{g/mL}$. This means that $1.0\ \text{L}$ of solution has nearly a mass of $1.0\ \text{kg}$. As the solution becomes more concentrated, its density will not be as close to $1.0\ \text{g/mL}$ and the molality value will be different than the molarity. For solutions with solvents other than water, the molality will be very different than the molarity. Make sure that you are paying attention to which quantity is being used in a given problem.

Summary

- The molality (m) of a solution is the moles of solute divided by the kilograms of solvent.
- The value of molality does not change with temperature.

Answers

Exercise 9.7.1

$$Molality = \frac{moles\ solute}{mass\ solvent}$$

This can be rearranged to:

$$moles = Molality \times mass\ solvent = (0.98\ m)(1.5\ kg) = 1.47\ mol\ NaCl$$

Finally, convert moles to mass using the molar mass of NaCl (calculated from values on the periodic table):

$$1.47\ mol\ NaCl \times \frac{58.44\ g\ NaCl}{1\ mol\ NaCl} = 86\ g\ NaCl$$

Alternatively, the problem could be solved using unit analysis. Start with the simple unit (kg). Remember that *m* stands for mol/kg so it is a compound unit and can be split into two lines to make a conversion factor.

$$1.5\ kg \times \frac{0.98\ mol\ NaCl}{1\ kg} \times \frac{58.44\ g\ NaCl}{1\ mol\ NaCl} = 86\ g\ NaCl$$

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9.8: Reaction Stoichiometry

Learning Objectives

- Explain the concept of stoichiometry as it pertains to chemical reactions
- Use balanced chemical equations to derive stoichiometric factors relating amounts of reactants and products
- Perform stoichiometric calculations involving mass, moles, and solution molarity

A balanced chemical equation provides a great deal of information in a very succinct format. Chemical formulas provide the identities of the reactants and products involved in the chemical change, allowing classification of the reaction. Coefficients provide the relative numbers of these chemical species, allowing a quantitative assessment of the relationships between the amounts of substances consumed and produced by the reaction. These quantitative relationships are known as the reaction's stoichiometry, a term derived from the Greek words *stoicheion* (meaning "element") and *metron* (meaning "measure"). In this module, the use of balanced chemical equations for various stoichiometric applications is explored.

The general approach to using stoichiometric relationships is similar in concept to the way people go about many common activities. Cooking, for example, offers an appropriate comparison. Suppose a recipe for making eight pancakes calls for 1 cup pancake mix, $\frac{3}{4}$ cup milk, and one egg. The "equation" representing the preparation of pancakes per this recipe is

$$1 \text{ cup mix} + \frac{3}{4} \text{ cup milk} + 1 \text{ egg} \rightarrow 8 \text{ pancakes}$$

If two dozen pancakes are needed for a big family breakfast, the ingredient amounts must be increased proportionally according to the amounts given in the recipe. For example, the number of eggs required to make 24 pancakes is

$$24 \ \cancel{\text{pancakes}} \times \frac{1 \text{ egg}}{8 \ \cancel{\text{pancakes}}} = 3 \text{ eggs}$$

Balanced chemical equations are used in much the same fashion to determine the amount of one reactant required to react with a given amount of another reactant, or to yield a given amount of product, and so forth. The coefficients in the balanced equation are used to derive stoichiometric factors that permit computation of the desired quantity. To illustrate this idea, consider the production of ammonia by reaction of hydrogen and nitrogen:

$$\mathrm{N_2}(g) + 3\,\mathrm{H_2}(g) \rightarrow 2\,\mathrm{NH_3}(g)$$

This equation shows that ammonia molecules are produced from hydrogen molecules in a 2:3 ratio, and stoichiometric factors may be derived using any amount (number) unit:

$$\frac{2\ NH_3\ molecules}{3\ H_2\ molecules} \text{ or } \frac{2\ doz\ NH_3\ molecules}{3\ doz\ H_2\ molecules} \text{ or } \frac{2\ mol\ NH_3\ molecules}{3\ mol\ H_2\ molecules}$$

These stoichiometric factors can be used to compute the number of ammonia molecules produced from a given number of hydrogen molecules, or the number of hydrogen molecules required to produce a given number of ammonia molecules. Similar factors may be derived for any pair of substances in any chemical equation.

Example 9.8.1: Moles of Reactant Required in a Reaction

How many moles of $\mathrm{I_2}$ are required to react with 0.429 mol of Al according to the following equation (see Figure 9.8.2)?

$$2\,\mathrm{Al} + 3\,\mathrm{I_2} \rightarrow 2\,\mathrm{AlI_3}$$

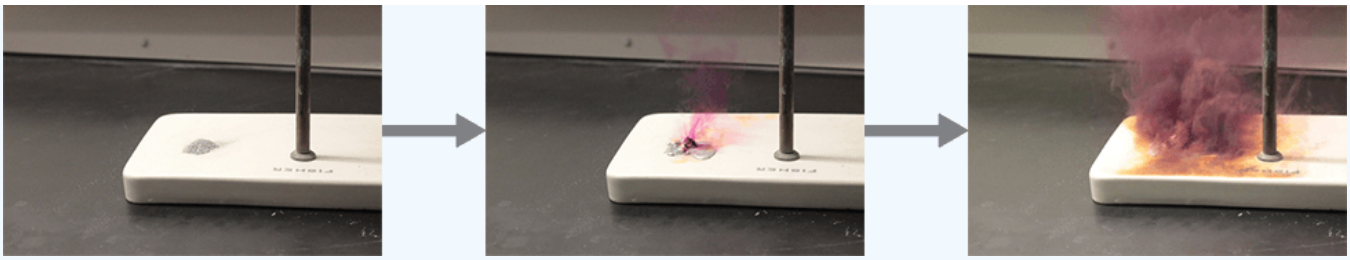
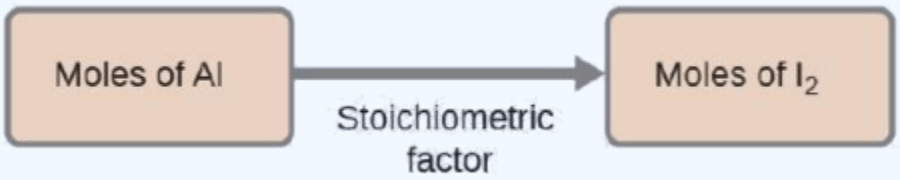


Figure 9.8.1: Aluminum and iodine react to produce aluminum iodide. The heat of the reaction vaporizes some of the solid iodine as a purple vapor. (credit: modification of work by Mark Ott)

Three pictures showing the progression of the reaction between iodine and aluminum. The first diagram shows only a pile of silver powder while the second diagram shows small amount of vapor forming. The final picture shows a large release of purple vapor.

Solution

Referring to the balanced chemical equation, the stoichiometric factor relating the two substances of interest is $\frac{3\ mol\ I_2}{2\ mol\ Al}$. The molar amount of iodine is derived by multiplying the provided molar amount of aluminum by this factor:



$$\text{mol I}_2 = 0.429\ \cancel{\text{mol Al}} \times \frac{3\ \text{mol I}_2}{2\ \cancel{\text{mol Al}}}$$

$$= 0.644\ \text{mol I}_2$$

Exercise 9.8.1

How many moles of $Ca(OH)_2$ are required to react with 1.36 mol of H_3PO_4 to produce $Ca_3(PO_4)_2$ according to the equation $3\ \text{Ca(OH)}_2 + 2\ \text{H}_3\text{PO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + 6\ \text{H}_2\text{O}$?

Answer available at the end of this section.

Example 9.8.2: Number of Product Molecules Generated by a Reaction

How many carbon dioxide molecules are produced when 0.75 mol of propane is combusted according to this equation?

$$\text{C}_3\text{H}_8 + 5\ \text{O}_2 \rightarrow 3\ \text{CO}_2 + 4\ \text{H}_2\text{O}$$

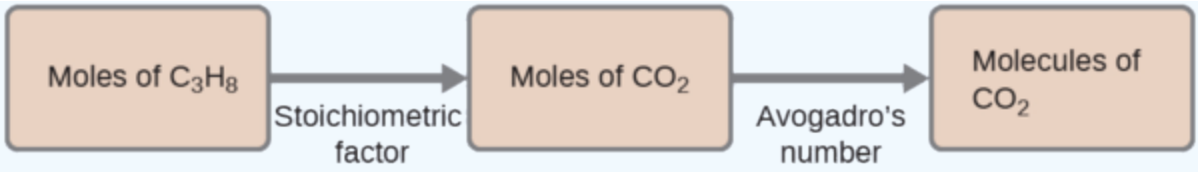
Solution

The approach here is the same as for Example 9.8.1, though the absolute number of molecules is requested, not the number of moles of molecules. This will simply require use of the moles-to-numbers conversion factor, Avogadro's number.

The balanced equation shows that carbon dioxide is produced from propane in a 3:1 ratio:

\[\ce{\dfrac{3\: mol\: CO_2}{1\: mol\: C_3H_8}} \nonumber \]

Using this stoichiometric factor, the provided molar amount of propane, and Avogadro's number,



$$0.75 \ \cancel{\text{mol } C_3H_8} \times \frac{3 \ \cancel{\text{mol } CO_2}}{1 \ \cancel{\text{mol } C_3H_8}} \times \frac{6.022 \times 10^{23} \ CO_2 \text{ molecules}}{\cancel{\text{mol } CO_2}} = 1.4 \times 10^{24} \ CO_2 \text{ molecules}$$

Exercise 9.8.1

How many NH_3 molecules are produced by the reaction of 4.0 mol of $Ca(OH)_2$ according to the following equation:

$$(NH_4)_2SO_4 + Ca(OH)_2 \rightarrow 2\,NH_3 + CaSO_4 + 2\,H_2O$$

Answer available at the end of this section.

These examples illustrate the ease with which the amounts of substances involved in a chemical reaction of known stoichiometry may be related. Directly measuring numbers of atoms and molecules is, however, not an easy task, and the practical application of stoichiometry requires that we use the more readily measured property of mass.

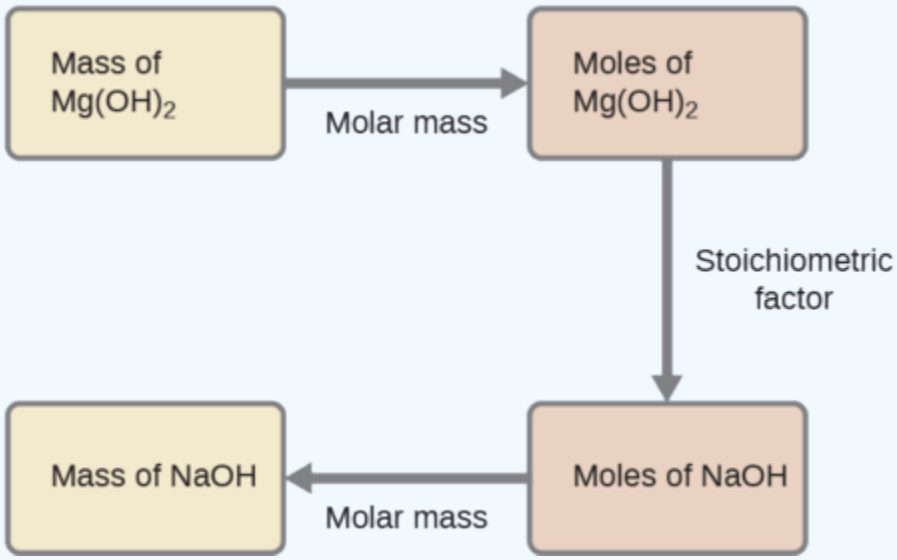
Example 9.8.3: Relating Masses of Reactants and Products

What mass of sodium hydroxide, NaOH, would be required to produce 16 g of the antacid milk of magnesia [magnesium hydroxide, $Mg(OH)_2$] by the following reaction?

$$MgCl_2(aq) + 2\,NaOH(aq) \rightarrow Mg(OH)_2(s) + 2\,NaCl(aq)$$

Solution

The approach used previously in Examples 9.8.1 and 9.8.2 is likewise used here; that is, we must derive an appropriate stoichiometric factor from the balanced chemical equation and use it to relate the amounts of the two substances of interest. In this case, however, masses (not molar amounts) are provided and requested, so additional steps of the sort learned in the previous chapter are required. The calculations required are outlined in this flowchart:



$$16\ \cancel{\text{g Mg(OH)}_2} \times \frac{1\ \cancel{\text{mol Mg(OH)}_2}}{58.3\ \cancel{\text{g Mg(OH)}_2}} \times \frac{2\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol Mg(OH)}_2}} \times \frac{40.0\ \text{g NaOH}}{\cancel{\text{mol NaOH}}} = 22\ \text{g NaOH}$$

Exercise 9.8.3

What mass of gallium oxide, Ga_2O_3, can be prepared from 29.0 g of gallium metal? The equation for the reaction is $4\ \text{Ga} + 3\ \text{O}_2 \rightarrow 2\ \text{Ga}_2\text{O}_3$.

Answer available at the end of this section.

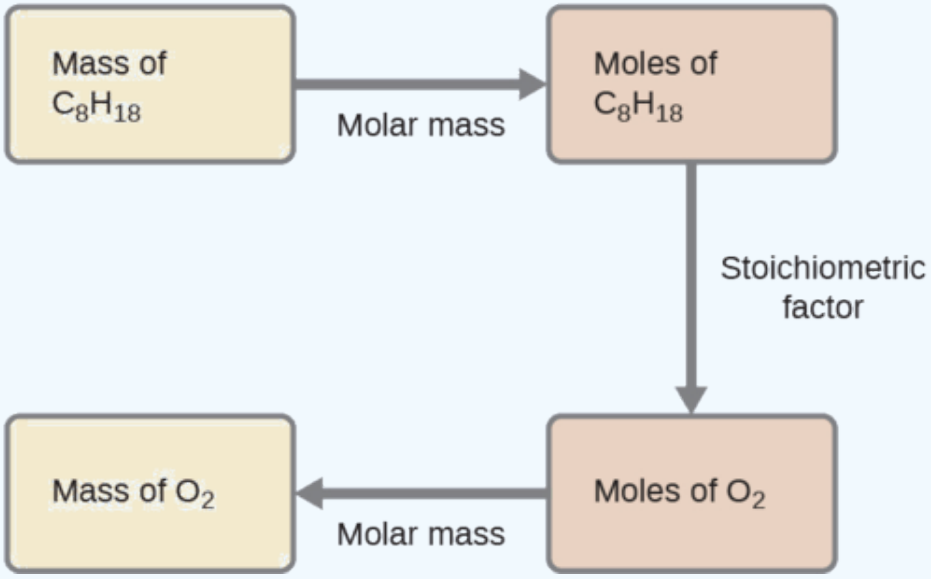
Example 9.8.4: Relating Masses of Reactants

What mass of oxygen gas, O_2, from the air is consumed in the combustion of 702 g of octane, C_8H_{18}, one of the principal components of gasoline?

$$2\ \text{C}_8\text{H}_{18} + 25\ \text{O}_2 \rightarrow 16\ \text{CO}_2 + 18\ \text{H}_2\text{O}$$

Solution

The approach required here is the same as for the Example 9.8.3, differing only in that the provided and requested masses are both for reactant species.



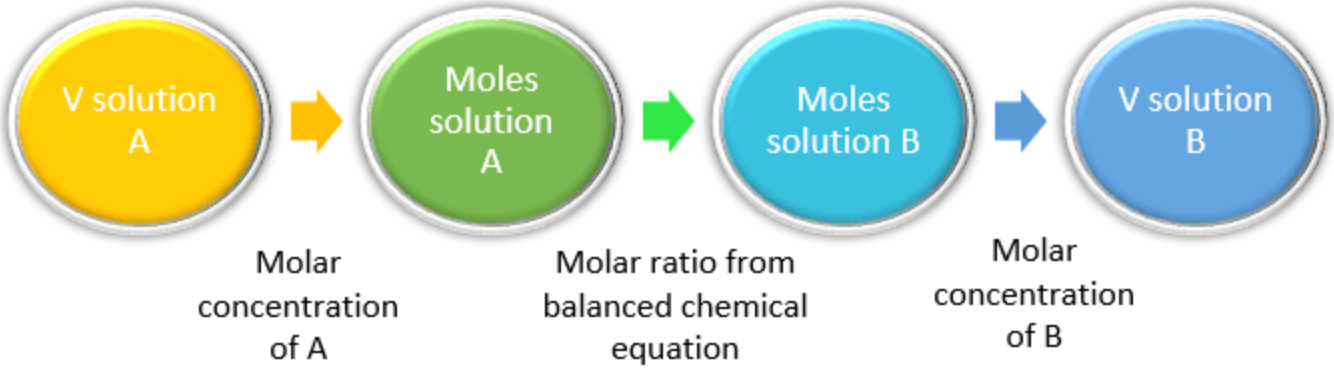
$$702\ \cancel{\text{g C}_8\text{H}_{18}} \times \frac{1\ \cancel{\text{mol C}_8\text{H}_{18}}}{114.23\ \cancel{\text{g C}_8\text{H}_{18}}} \times \frac{25\ \cancel{\text{mol O}_2}}{2\ \cancel{\text{mol C}_8\text{H}_{18}}} \times \frac{32.00\ \text{g O}_2}{\cancel{\text{mol O}_2}} = 2.46 \times 10^3\ \text{g O}_2$$

Exercise 9.8.4

What mass of CO is required to react with 25.13 g of Fe_2O_3 according to the equation $\text{Fe}_2\text{O}_3 + 3\ \text{CO} \rightarrow 2\ \text{Fe} + 3\ \text{CO}_2$?

Answer available at the end of this section.

We can use the concept of molarity to directly calculate the number of moles of reactants or products that will be formed, and therefore their amounts (i.e. volume of solutions or mass of precipitates).



As an example, lead (II) nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, lead (II) chloride.

$$\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow PbCl_2\,(s) + 2\,NaNO_3\,(aq)} \tag{9.8.1}$$

In the reaction shown above, if we mixed 0.123 L of a 1.00 M solution of NaCl with 1.50 M solution of $\mathrm{Pb(NO_3)_2}$, we could calculate the volume of $\mathrm{Pb(NO_3)_2}$ solution needed to completely precipitate the $\mathrm{Pb^{2+}}$ ions.

The molar concentration can also be expressed as the following:

$$1.00\,M\,\mathrm{NaCl} = \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}}$$

and

$$1.50\,M\,\mathrm{Pb(NO_3)_2} = \frac{1.50\ mol\ \mathrm{Pb(NO_3)_2}}{1\ L\ \mathrm{Pb(NO_3)_2 solution}}$$

First, we must examine the reaction stoichiometry in the balanced reaction (Equation 9.8.1). In this reaction, one mole of $\mathrm{Pb(NO_3)_2}$ reacts with two moles of NaCl to give one mole of $\mathrm{PbCl_2}$ precipitate. Thus, the concept map utilizing the stoichiometric ratios is:



so the volume of lead (II) nitrate that reacted is calculated as:

$$0.123\ L\ \mathrm{NaCl\,solution} \times \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}} \times \frac{1\ mol\ \mathrm{Pb(NO_3)_2}}{2\ mol\ \mathrm{NaCl}} \times \frac{1\ L\ \mathrm{Pb(NO_3)_2\ solution}}{1.5\ mol\ \mathrm{Pb(NO_3)_2}} = 0.041\ \mathrm{Pb(NO_3)_2}\ L\ \mathrm{solution}$$

This volume makes intuitive sense for two reasons: (1) the number of moles of $\mathrm{Pb(NO_3)_2}$ required is half of the number of moles of NaCl, based off of the stoichiometry in the balanced reaction (Equation 9.8.1); (2) the concentration of $\mathrm{Pb(NO_3)_2}$ solution is 50% greater than the NaCl solution, so less volume is needed.

Example 9.8.1

What volume (in L) of 0.500 M sodium sulfate will react with 275 mL of 0.250 M barium chloride to completely precipitate all $\mathrm{Ba^{2+}}$ in the solution?

$Na_2SO_4\,(aq) + BaCl_2\,(aq) \longrightarrow BaSO_4\,(s) + 2NaCl\,(aq)$

Solution

Steps for Problem Solving	Example 9.8.1
Identify the "given" information and what the problem is asking you to "find."	Given: 275 mL $BaCl_2$ 0.250 M $BaCl_2$ or $\frac{0.250\ mol BaCl_2}{1\ L\ BaCl_2\ solution}$ 0.500 M Na_2SO_4 or $\frac{0.500\ mol Na_2SO_4}{1\ L\ Na_2SO_4\ solution}$ Find: Volume Na_2SO_4 solution.
List other known quantities.	1 mol of Na_2SO_4 to 1 mol $BaCl_2$ 1000 mL = 1 L
Prepare a concept map and use the proper conversion factor.	mL $BaCl_2$ solution → L $BaCl_2$ solution → mol $BaCl_2$ solution → mol Na_2SO_4 solution $\frac{1\ L}{1000\ mL}$ $\frac{0.250\ mol\ BaCl_2}{1\ L\ BaCl_2 solution}$ $\frac{1\ mol\ Na_2SO_4}{1\ mol\ BaCl_2}$ $\frac{1\ L\ Na_2SO_4\ solution}{0.500\ mol}$
Cancel units and calculate.	$275\ \cancel{mL\ BaCl_2\ solution} \times \frac{1\ \cancel{L}}{1000\ \cancel{mL}} \times \frac{0.250\ \cancel{mol\ BaCl_2}}{1\ \cancel{L\ BaCl_2\ solution}} \times \frac{1\ \cancel{mol\ Na_2SO_4}}{1\ \cancel{mol\ BaCl_2}} \times \frac{1\ L\ Na_2SO_4\ solution}{0.500\ \cancel{mol Na_2SO_4}}$ = 0.1 sulfate
Think about your result.	The lesser amount (almost half) of sodium sulfate is to be expected as it is more concentrated than barium chloride. Also, the units are correct.

Exercise 9.8.1

What volume of 0.250 M lithium hydroxide will completely react with 0.500 L of 0.250 M of sulfuric acid solution?

$H_2SO_4\,(aq) + 2\,LiOH\,(aq) \longrightarrow H_2O\,(l) + 2_{24}\,LiSO(aq)$

Answer available at the end of this section.

These examples illustrate just a few instances of reaction stoichiometry calculations. Numerous variations on the beginning

and ending computational steps are possible depending upon what particular quantities are provided and sought (volumes, solution concentrations, and so forth). Regardless of the details, all these calculations share a common essential component: the use of stoichiometric factors derived from balanced chemical equations. Figure 9.8.2 provides a general outline of the various computational steps associated with many reaction stoichiometry calculations.

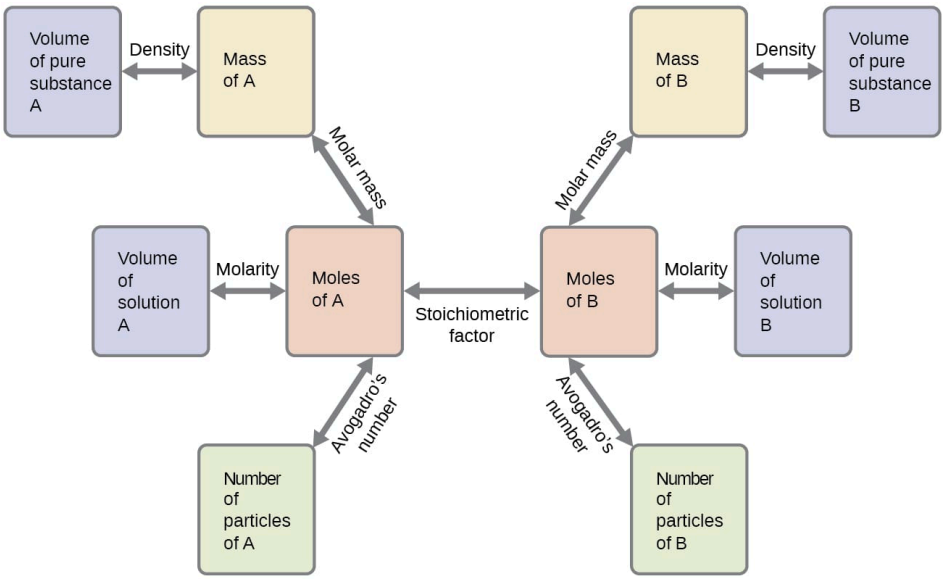


Figure 9.8.2: The flowchart depicts the various computational steps involved in most reaction stoichiometry calculations.

Schematic flowchart showing how density is used for converting volume of pure substance to mass, molar mass for mass and moles, molarity for moles and volume of solution, Avogadro's number for moles and number of particles, and stoichiometric factor for relating moles of one substance to another.

Summary

A balanced chemical equation may be used to describe a reaction's stoichiometry (the relationships between amounts of reactants and products). Coefficients from the equation are used to derive stoichiometric factors that subsequently may be used for computations relating reactant and product masses, molar amounts, and other quantitative properties.

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- Henry Agnew (UC Davis)

Glossary

stoichiometric factor
: ratio of coefficients in a balanced chemical equation, used in computations relating amounts of reactants and products

stoichiometry
: relationships between the amounts of reactants and products of a chemical reaction

Answers

Exercise 9.8.1

2.04 mol

Exercise 9.8.1

4.8×10^{24} NH_3 molecules

Exercise 9.8.3

39.0 g

Exercise 9.8.4

13.22 g

Exercise 9.8.1

1.00 L $LiOH$ solution

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10: TYPES OF CHEMICAL REACTIONS

10.1: Energy and Chemical and Physical Change
10.2: Oxidation-Reduction (Redox) Reactions
10.3: Chemical Equilibrium
10.4: Shifting Equilibria - Le Chatelier's Principle

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10.1: Energy and Chemical and Physical Change

Learning Objectives

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

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

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

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



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

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The **law of conservation**

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

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

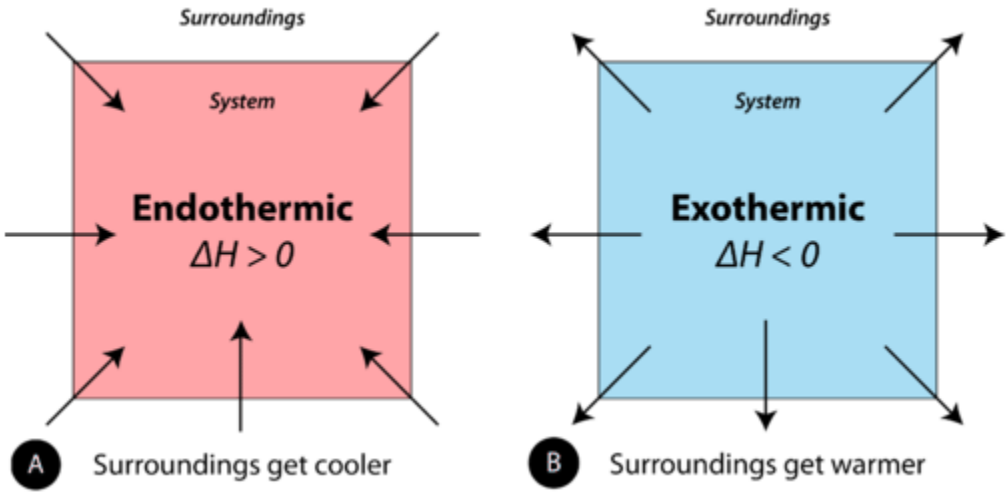


Figure 10.1.1: (A) Endothermic reaction. (B) Exothermic reaction.

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

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

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

Example 10.1.1

Label each of the following processes as endothermic or exothermic.

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

Solution

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

Exercise 10.1.1

Label each of the following processes as endothermic or exothermic.

a. water vapor condensing
b. gold melting

Answer (a) available at the end of this section.

Answer (b) available at the end of this section.

Summary

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

Answers

Exercise 10.1.1
exothermic

Exercise 10.1.1
endothermic

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10.2: Oxidation-Reduction (Redox) Reactions

Learning Objectives

- To identify oxidation and reduction in a redox reaction.
- To identify oxidizing agents and reducing agents.

When zinc metal is submerged into a quantity of aqueous HCl, the following reaction occurs (Figure $10.2.1$):

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow H_2\,(g) + ZnCl_2\,(aq)} \tag{10.2.1}$$

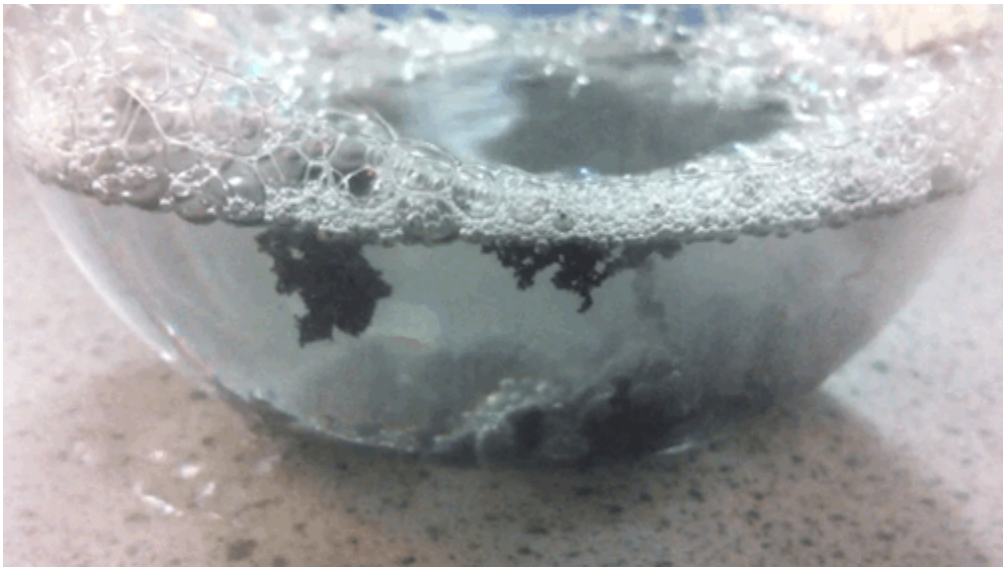


Figure $10.2.1$: Zinc Metal plus Hydrochloric Acid. It is fairly obvious that zinc metal reacts with aqueous hydrochloric acid! The bubbles are hydrogen gas (right side of Equation $10.2.1$).

Because some of the substances in this reaction are aqueous, we can separate them into ions:

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) + 2\,Cl^-\,(aq) \longrightarrow H_2\,(g) + Zn^{2+}\,(aq) + 2\,Cl^-\,(aq)}$$

Viewed this way, the net reaction seems to be a charge transfer between zinc and hydrogen atoms. (There is no net change experienced by the chloride ion.) In fact, electrons are being transferred from the zinc atoms to the hydrogen atoms (which ultimately make a molecule of diatomic hydrogen), changing the charges on both elements.

To understand electron-transfer reactions like the one between zinc metal and hydrogen ions, chemists separate them into two parts: one part focuses on the loss of electrons, and one part focuses on the gain of electrons. The loss of electrons is called oxidation. The gain of electrons is called reduction. Because any loss of electrons by one substance must be accompanied by a gain in electrons by something else, oxidation and reduction always occur together. As such, electron-transfer reactions are also called oxidation-reduction reactions, or simply **redox reactions**. The atom that loses electrons is **oxidized**, and the atom that gains electrons is **reduced**. Also, because we can think of the species being oxidized as causing the reduction, the species being oxidized is called the **reducing agent**, and the species being reduced is called the **oxidizing agent**.

Because batteries are used as sources of electricity (that is, of electrons), all batteries are based on redox reactions.

Example $10.2.1$: Silver Ions

$$\mathrm{3\,Ag^+ + Al \longrightarrow 3\,Ag + Al^{3+}}$$

Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

Solution

- The **substance oxidized** is the reactant that had undergone oxidation, the one for which the charge increased: Al
- The **substance reduced** is the reactant that had undergone reduction, the one for which the charge decreased: Ag^+
- The **reducing agent** is the same as the substance oxidized: Al
- The **oxidizing agent** is the same as the substance reduced: Ag^+

Exercise 10.2.1

$Ca^{2+} + 2K \rightarrow Ca + 2K^+$

Identify the substance oxidized, substance reduced, reducing agent and reducing agent.

Answer available at the end of this section.

To Your Health: Redox Reactions and Pacemaker Batteries

All batteries use redox reactions to supply electricity because electricity is basically a stream of electrons being transferred from one substance to another. Pacemakers—surgically implanted devices for regulating a person's heartbeat—are powered by tiny batteries, so the proper operation of a pacemaker depends on a redox reaction.

Pacemakers used to be powered by NiCad batteries, in which nickel and cadmium (hence the name of the battery) react with water according to this redox reaction:

$$\mathrm{Cd\,(s) + 2\,NiOOH\,(s) + 2\,H_2O(\ell) \longrightarrow Cd(OH)_2\,(s) + 2\,Ni(OH)\ 2\,(s)}$$

The cadmium is oxidized, while the nickel atoms in NiOOH are reduced. Except for the water, all the substances in this reaction are solids, allowing NiCad batteries to be recharged hundreds of times before they stop operating. Unfortunately, NiCad batteries are fairly heavy batteries to be carrying around in a pacemaker. Today, the lighter lithium/iodine battery is used instead. The iodine is dissolved in a solid polymer support, and the overall redox reaction is as follows:

$$\mathrm{2\,Li\,(s) + I_2\,(s) \longrightarrow 2\,LiI\,(s)}$$

Lithium is oxidized, and iodine is reduced. Although the lithium/iodine battery cannot be recharged, one of its advantages is that it lasts up to 10 years. Thus, a person with a pacemaker does not have to worry about periodic recharging; about once per decade a person requires minor surgery to replace the pacemaker/battery unit. Lithium/iodine batteries are also used to power calculators and watches.



Figure 10.2.1: A small button battery like this is used to power a watch, pacemaker, or calculator. (CC BY-SA; Gerhard H Wrodnigg via Wikipedia)

Oxidation and reduction can also be defined in terms of changes in composition. The original meaning of **oxidation** was "**adding oxygen**," so when oxygen is added to a molecule, the molecule is being oxidized. The reverse is true for **reduction**: if a molecule **loses oxygen** atoms, the molecule is being reduced. For example, the acetaldehyde (CH_3CHO) molecule takes on an oxygen atom to become acetic acid (CH_3COOH).

$$2\,CH_3CHO + O_2 \longrightarrow 2\,CH_3COOH \tag{10.2.2}$$

Thus, acetaldehyde is being oxidized.

Similarly, **reduction and oxidation can be defined in terms of the gain or loss of hydrogen atoms**. If a molecule adds hydrogen atoms, it is being reduced. If a molecule loses hydrogen atoms, the molecule is being oxidized. For example, in the conversion of acetaldehyde into ethanol (CH_3CH_2OH), hydrogen atoms are added to acetaldehyde, so the acetaldehyde is being reduced:

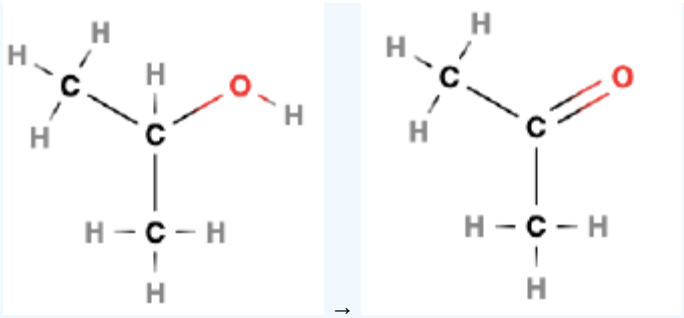
$$CH_3CHO + H_2 \longrightarrow CH_3CH_2OH \tag{10.2.3}$$

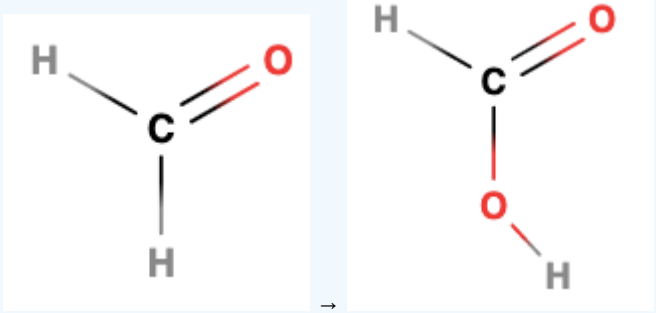
Process	Change in oxygen (some reactions)	Change in hydrogen (some reactions)
Oxidation	gain	lose
Reduction	lose	gain

Example 10.2.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $N_2 \rightarrow NH_3$

b.  →

c.  →

Solution

a. Hydrogen is being added to the original reactant molecule, so reduction is occurring.
b. Hydrogen is being removed from the original reactant molecule, so oxidation is occurring.
c. Oxygen is being added to the original reactant molecule, so oxidation is occurring.

Exercise 10.2.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $CH_4 \rightarrow CO_2 + H_2O$
b. $NO_2 \rightarrow N_2$
c. $CH_2{=}CH_2 \rightarrow CH_3CH_3$

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer a: available at the end of this section.

Concept Review Exercises

1. Give two different definitions for oxidation and reduction.
2. Give an example of each definition of oxidation and reduction.

Answers

1. Oxidation is the loss of electrons or the addition of oxygen; reduction is the gain of electrons or the addition of hydrogen.
2. $Zn \rightarrow Zn^{2+} + 2e^-$ (oxidation); $C_2H_4 + H_2 \rightarrow C_2H_6$ (reduction) (answers will vary)

Key Takeaway

Chemical reactions in which electrons are transferred are called oxidation-reduction, or redox, reactions. Oxidation is the loss of electrons. Reduction is the gain of electrons. Oxidation and reduction always occur together, even though they can be written as separate chemical equations.

Exercises

1. Identify the oxidizing and reducing agents.

    a. $3Mg + 2AlCl_3 \rightarrow 2Al + 3MgCl_2$
    b. $H_2O_2 + H_2 \rightarrow 2H_2O$

2. Identify the oxidizing and reducing agents.

    a. $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$
    b. $2K + 2H_2O \rightarrow 2KOH + H_2$

Answers

1.

    a. oxidizing agent: $AlCl_3$; reducing agent: Mg

    b. oxidizing agent: H_2O_2; reducing agent: H_2

2.

    a. oxidizing agent: O_2; reducing agent: C_2H_6
    b. oxidizing agent: H_2O; reducing agent: K

Answers

Exercise 10.2.1

Reduction: $Ca^{2+} \rightarrow Ca$

Oxidation: $K \rightarrow K^{+}$

- The **substance oxidized** is the reactant that had undergone oxidation (increased charge): K
- The **substance reduced** is the reactant that had undergone reduction (decreased charge): Ca^{2+}
- The **reducing agent** is the same as the substance oxidized: K
- The **oxidizing agent** is the same as the substance reduced: Ca^{2+}

Potassium has been used as a reducing agent to obtain various metals in their elemental form.

Exercise 10.2.2

Oxygen is being added. Oxidation is occurring.

Exercise 10.2.2

Oxygen is being removed. Reduction is occurring.

Exercise 10.2.2

Hydrogen is being added. Reduction is occurring.

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10.3: Chemical Equilibrium

Learning Objectives

- Define *chemical equilibrium*.
- 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):

$$H2 + I2 \rightarrow 2HI$$

This is simply 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 H2 and the I2 react to make HI.

However, this is not the case. The reverse chemical reaction is also taking place:

$$2HI \rightarrow H2 + I2$$

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, any continued overall reaction stops: the reaction has reached **chemical equilibrium** (sometimes just spoken as *equilibrium*; plural *equilibria*), 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 10.3.1

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 10.3.1

Write the equilibrium equation between elemental hydrogen and elemental oxygen as reactants and water as the product.

Answer available at the end of this section.

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. Also, because 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, however.

Key Takeaways

- Chemical reactions eventually reach equilibrium, a point at which forward and reverse reactions balance each other's progress.
- Chemical equilibria are dynamic: the chemical reactions are always occurring; they just cancel each other's progress.

Answers

Exercise 10.3.1

$$2H_2 + O_2 + \rightleftharpoons 2H_2O$$

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10.4: Shifting Equilibria - Le Chatelier's Principle

Learning Objectives

- Define *Le Chatelier's principle*.
- Predict the direction of shift for an equilibrium under stress.

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 Chatelier'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. If reactant or product is removed, the equilibrium shifts to make more reactant or product, respectively, to make up for the loss.

Example 10.4.1

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.
2. If NH_3 is added, there is now more product, so the reaction will shift toward reactants to reduce the added NH_3.
3. If NH_3 is removed, there is now less product, so the reaction will shift toward products to replace the product removed.

Exercise 10.4.1

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
2. toward reactants

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.

How does an equilibrium react to a change in pressure? Pressure changes do not markedly affect the solid or liquid phases. However, pressure strongly impacts the gas phase. Le Chatelier's principle implies that a pressure increase shifts an equilibrium to the side of the reaction with the fewer number of moles of gas, while a pressure decrease shifts an equilibrium to the side of the reaction with the greater number of moles of gas. If the number of moles of gas is the same on both sides of the reaction, pressure has no effect.

Example 10.4.2

What is the effect on this equilibrium if pressure is increased?

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Solution

According to Le Chatelier's principle, if pressure is increased, then the equilibrium shifts to the side with the fewer number of moles of gas. This particular reaction shows a total of 4 mol of gas as reactants and 2 mol of gas as products, so the reaction shifts toward the products side.

Exercise 10.4.2

What is the effect on this equilibrium if pressure is decreased?

$$3O_2(g) \rightleftharpoons 2O_3(g)$$

Answer available at the end of this section.

What is the effect of temperature changes on an equilibrium? It depends on whether the reaction is endothermic or exothermic. Recall that *endothermic* means that energy is absorbed by a chemical reaction, while *exothermic* means that energy is given off by the reaction. As such, energy can be thought of as a reactant or a product, respectively, of a reaction:

endothermic: energy + reactants → products

exothermic: reactants → products + energy

Because temperature is a measure of the energy of the system, increasing temperature can be thought of as adding energy. The reaction will react as if a reactant or a product is being added and will act accordingly by shifting to the other side. For example, if the temperature is increased for an endothermic reaction, essentially a reactant is being added, so the equilibrium shifts toward products. Decreasing the temperature is equivalent to decreasing a reactant (for endothermic reactions) or a product (for exothermic reactions), and the equilibrium shifts accordingly.

Example 10.4.3

Predict the effect of increasing the temperature on this equilibrium.

$$PCl_3 + Cl_2 \rightleftharpoons PCl_5 + 60kJ$$

Solution

Because energy is listed as a product, it is being produced, so the reaction is exothermic. If the temperature is increasing, a product is being added to the equilibrium, so the equilibrium shifts to minimize the addition of extra product: it shifts back toward reactants.

Exercise 10.4.3

Predict the effect of decreasing the temperature on this equilibrium.

$$N_2O_4 + 57kJ \rightleftharpoons 2NO_2$$

Answer available at the end of this section.

In the case of temperature, the value of the equilibrium has changed because the K_{eq} is dependent on temperature. That is why equilibria shift with changes in temperature.

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. However, a catalyst does not affect the extent or position of a reaction at equilibrium. It helps a reaction achieve equilibrium faster.

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 10.4.1 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 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.
- The direction of shift can be predicted for changes in concentrations, temperature, or pressure.
- Catalysts do not affect the position of an equilibrium; they help reactions achieve equilibrium faster.

Exercise 10.4.1

1. Define *Le Chatelier's principle*.
2. What is meant by a stress? What are some of the ways an equilibrium can be stressed?
3. Given this equilibrium, predict the direction of shift for each stress.

$$H_2(g) + I_2(s) + 53kJ \rightleftharpoons 2HI(g)$$

   a. decreased temperature
   b. increased pressure
   c. removal of HI
4. Given this equilibrium, predict the direction of shift for each stress.

$$H_2(g) + F_2(g) \rightleftharpoons 2HF(g) + 546kJ$$

   a. increased temperature
   b. addition of H2
   c. decreased pressure
5. Given this equilibrium, predict the direction of shift for each stress.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + 196kJ$$

   a. removal of SO3
   b. addition of O2
   c. decreased temperature
6. Given this equilibrium, predict the direction of shift for each stress.

$$CO_2(g) + C(s) + 171kJ \rightleftharpoons 2CO(g)$$

   a. addition of CO
   b. increased pressure
   c. addition of a catalyst
7. The synthesis of NH3 uses this chemical reaction.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + 92kJ$$

   Identify three stresses that can be imposed on the equilibrium to maximize the amount of NH3.
8. The synthesis of CaCO3 uses this chemical reaction.

$$CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s) + 180kJ$$

   Identify three stresses that can be imposed on the equilibrium to maximize the amount of CaCO3.

Answers

1. When an equilibrium is stressed, the equilibrium shifts to minimize that stress.
2.
3. a. toward reactants

- toward reactants
- toward products
-
- a. toward products
  b. toward products
  c. toward products
-
- increased pressure, decreased temperature, removal of NH_3

Answers

Exercise 10.4.2

Reaction shifts toward reactants.

Exercise 10.4.3

Equilibrium shifts toward reactants.

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11: ACIDS, BASES, AND BUFFERS

11.1: Acids and Bases in Water
11.2: Reactions of Acids and Bases
11.3: Water - Both an Acid and a Base
11.4: The Strengths of Acids and Bases
11.5: The pH Scale
11.6: Buffers

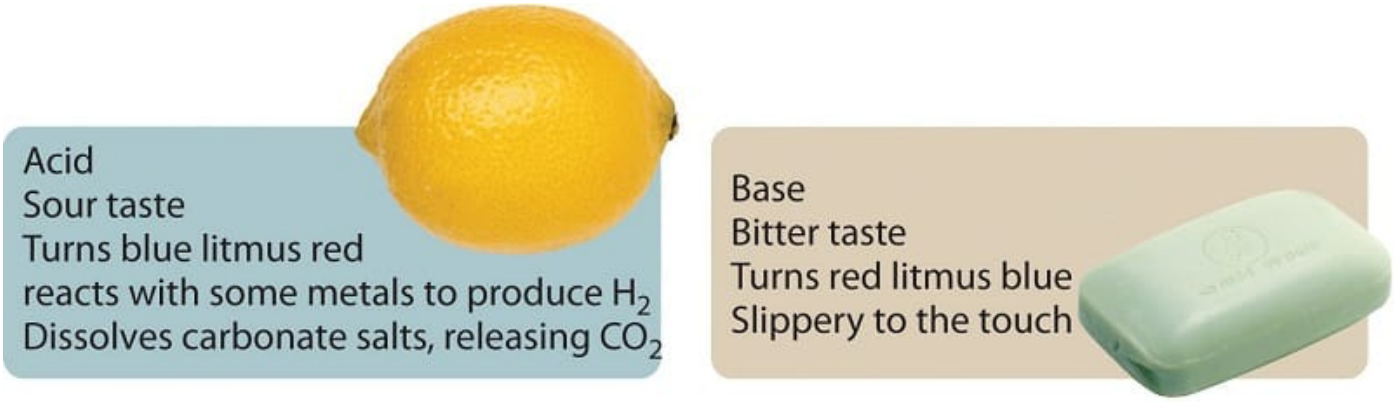
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11.1: Acids and Bases in Water

Learning Objectives

- To recognize a compound as an Arrhenius acid or an Arrhenius base.
- To describe characteristics of acids and bases.

One way to define a class of compounds is by describing the various characteristics its members have in common. In the case of the compounds known as acids, the common characteristics include a sour taste, the ability to change the color of the vegetable dye *litmus* to red, and the ability to dissolve certain metals and simultaneously produce hydrogen gas. For the compounds called bases, the common characteristics are a slippery texture, a bitter taste, and the ability to change the color of litmus to blue. Acids and bases also react with each other to form compounds generally known as salts.



Although we include their tastes among the common characteristics of acids and bases, we never advocate tasting an unknown chemical!

Arrhenius Acids and Bases

Chemists prefer, however, to have definitions for acids and bases in chemical terms. The Swedish chemist Svante Arrhenius developed the first chemical definitions of acids and bases in the late 1800s. Arrhenius defined an **acid** as a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution. Many acids are simple compounds that release a hydrogen cation into solution when they dissolve. Similarly, Arrhenius defined a **base** as a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution. Many bases are ionic compounds that have the hydroxide ion as their anion, which is released when the base dissolves in water.

Table 11.1.1*: Formulas and Names for Some Acids and Bases*

Acids		Bases	
Formula	Name	Formula	Name
HCl(aq)	hydrochloric acid	NaOH(aq)	sodium hydroxide
H_2S(aq)	hydrosulfuric acid	KOH(aq)	potassium hydroxide
$HC_2H_3O_2$(aq)	acetic acid	$Mg(OH)_2$(aq)	magnesium hydroxide
HNO_2(aq)	nitrous acid	$Ca(OH)_2$(aq)	calcium hydroxide
H_3PO_4(aq)	phosphoric acid	NH_3(aq)	ammonia

Acids		Bases	
$H_2CO_3(aq)$	carbonic acid		

Many bases and their aqueous solutions are named using the normal rules of ionic compounds that were presented previously; that is, they are named as hydroxide compounds. For example, sodium hydroxide (NaOH) is both an ionic compound and a base. However, aqueous solutions of acids have their own naming rules which are not covered here. Table 11.1.1 lists some acids and bases and their names. Note that molecular formulas for acids have hydrogen written first, as if it were the cation, while most bases have the negative hydroxide ion, if it appears in the formula, written last.

> The name oxygen comes from the Latin meaning “acid producer” because its discoverer, Antoine Lavoisier, thought it was the essential element in acids. Lavoisier was wrong, but it is too late to change the name now.

Brønsted-Lowry Acids and Bases

Notice that one base listed in Table 11.1.1—ammonia—does not have hydroxide as part of its formula. How does this compound increase the amount of hydroxide ion in aqueous solution? Instead of dissociating into hydroxide ions, ammonia molecules react with water molecules by taking a hydrogen ion from the water molecule to produce an ammonium ion and a hydroxide ion:

$$NH_{3(aq)} + H_2O_{(\ell)} \rightarrow NH^+_{4(aq)} + OH^-_{(aq)} \tag{11.1.1}$$

Because this reaction of ammonia with water causes an increase in the concentration of hydroxide ions in solution, ammonia satisfies the Arrhenius definition of a base. Many other nitrogen-containing compounds are bases because they too react with water to produce hydroxide ions in aqueous solution.

Ammonia (NH_3) increases the hydroxide ion concentration in aqueous solution by reacting with water rather than releasing hydroxide ions directly. In fact, the Arrhenius definitions of an acid and a base focus on hydrogen ions and hydroxide ions. Are there more fundamental definitions for acids and bases?

In 1923, the Danish scientist Johannes Brønsted and the English scientist Thomas Lowry independently proposed new definitions for acids and bases. Rather than considering both hydrogen and hydroxide ions, they focused on the hydrogen ion only. A Brønsted-Lowry acid is a compound that supplies a hydrogen ion in a reaction. A Brønsted-Lowry base, conversely, is a compound that accepts a hydrogen ion in a reaction. Thus, the Brønsted-Lowry definitions of an acid and a base focus on the movement of hydrogen ions in a reaction, rather than on the production of hydrogen ions and hydroxide ions in an aqueous solution.

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$NH_3(aq) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

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:

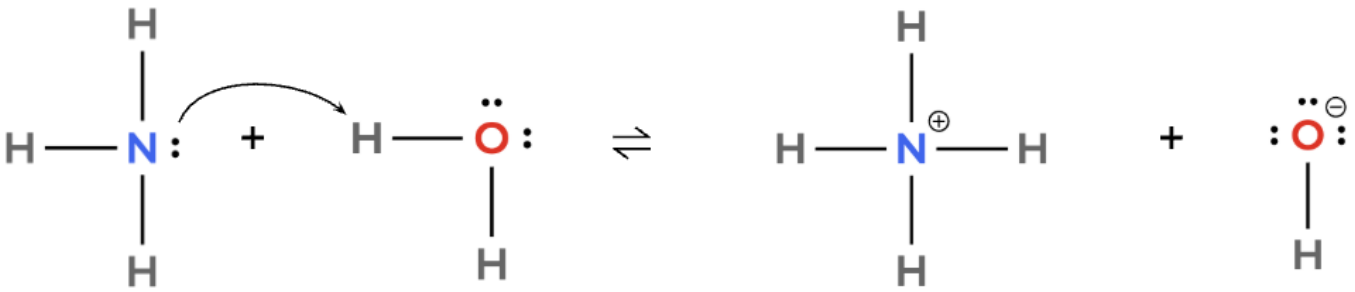


Figure 11.1.1: Reaction of ammonia and water. The curved arrow shows how the lone pair of valence electrons on the nitrogen becomes a bond between N and H. Because the ammonia gained H^+ (a proton) and not a complete H atom, the product now has a positive charge. Water lost H^+ but kept all the electrons, so it becomes OH^- with a negative charge.

Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

> A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than H_3O^+. It is simpler, however, to use H_3O^+.

With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$HCl(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

We can depict this process using Lewis electron dot diagrams:

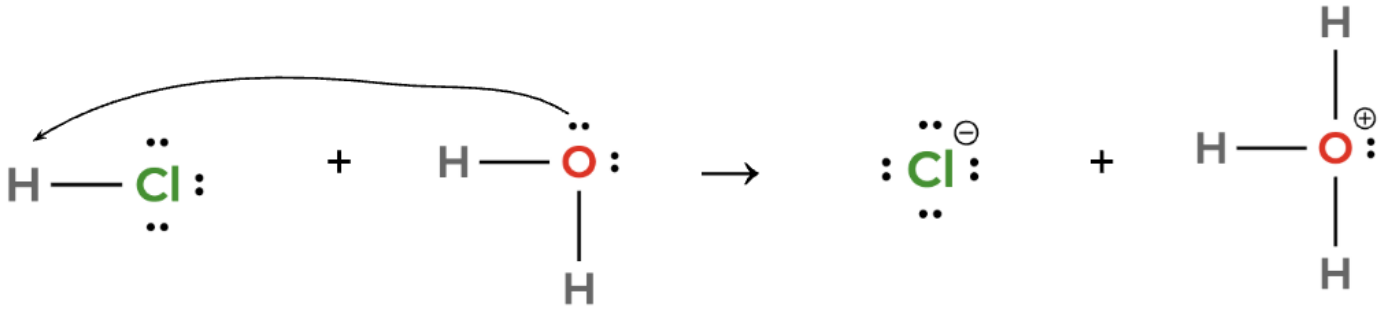


Figure 11.1.2: Reaction of HCl and water. The curved arrow shows how a lone pair of valence electrons on the oxygen atom becomes a bond between O and H. Because the water gained H^+ (a proton) and not a complete H atom, the product now has a positive charge. The acid lost H^+ but kept all the electrons, so it becomes H_3O^+ with a positive charge.

Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, **HCl acts as a Brønsted-Lowry acid**; as a hydrogen ion acceptor, **H_2O is a Brønsted-Lowry base**. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

> All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

Example 11.1.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$C_6H_5NH_2(aq) + H_2O(\ell) \rightarrow C_6H_5NH_3^+(aq) + OH^-(aq)$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

Exercise 11.1.1

Caffeine ($C_8H_{10}N_4O_2$) is a stimulant found in coffees and teas. When dissolved in water, it can accept a proton from a water molecule. Write the chemical equation for this process and identify the Brønsted-Lowry acid and base.

Answer available at the end of this section.

The Brønsted-Lowry definitions of an acid and a base can be applied to chemical reactions that occur in solvents other than water. The following example illustrates.

Example 11.1.2

Sodium amide ($NaNH_2$) dissolves in methanol (CH_3OH) and separates into sodium ions and amide ions (NH_2^-). The amide ions react with methanol to make ammonia and the methoxide ion (CH_3O^-). Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

Solution

The equation for the reaction is between NH_2^- and CH_3OH to make NH_3 and CH_3O^- is as follows:

$$NH_2^-(solv) + CH_3OH(\ell) \rightarrow NH_3(solv) + CH_3O^-(solv)$$

The label *(solv)* indicates that the species are dissolved in some solvent, in contrast to *(aq)*, which specifies an aqueous (H_2O) solution. In this reaction, we see that the NH_2^- ion accepts a proton from a CH_3OH molecule to make an NH_3 molecule. Thus, as the proton acceptor, NH_2^- is the Brønsted-Lowry base. As the proton donor, CH_3OH is the Brønsted-Lowry acid.

Exercise 11.1.2

Pyridinium chloride (C_5H_5NHCl) dissolves in ethanol (C_2H_5OH) and separates into pyridinium ions ($C_5H_5NH^+$) and chloride ions. The pyridinium ion can transfer a hydrogen ion to a solvent molecule. Write a balanced chemical equation for this process and identify the Brønsted-Lowry acid and base.

Answer available at the end of this section.

How to Recognize Acids and Bases

Acids and bases can be inorganic or organic. Inorganic acids and bases will often be represented by molecular formulas, as in Table 11.1.1. Organic acids and bases can be recognized based on their functional group when the structure is known.

Acids

As mentioned above, molecular formulas for acids start with H. This lets you know that the substance is an acid, and also indicates how many of the hydrogen atoms are acidic, that is how many can be lost as H^+ in reactions. For example, the molecular formula for acetic acid is $HC_2H_3O_2$. Although there are four hydrogen atoms in the molecule, the format indicates that only one is acidic. Acetic acid can lose dissociate (separate) into H^+ and $C_2H_3O_2^-$. The other three hydrogen atoms are not acidic; they are not lost during reactions. This is indicated by the fact that they are not written first in the molecular formula. Compare this to hydrosulfuric acid, H_2S, and phosphoric acid, H_3PO_4. In these molecules all the hydrogen atoms are acidic so they are all written first in the molecular formula.

For organic acids you may be provided with a picture of the structure (condensed structural formula or skeletal structure) instead of a molecular formula. In that case the presence of a carboxylic acid functional group will indicate that the substance is an acid.

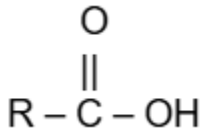


Figure 11.1.2: Carboxylic acid functional group

The carboxylic acid functional group can also be abbreviated COOH as in CH_3CH_2COOH.

Bases

Inorganic bases are often metal hydroxides such as NaOH and $Ca(OH)_2$. Pay attention when you see OH in a formula! Metals bonded to hydroxide ions (OH^-) are bases, but the abbreviation COOH represents a carboxylic acid and other organic compounds containing an OH are alcohols which are neither an acid nor a base.

Amines are organic bases. Remember that the amine functional group contains a nitrogen that is not bonded to C=O. The number of hydrogen atoms bonded to the nitrogen in an amine can vary, so they should not be used for identification. For example, CH_3NHCH_3 and $CH_3CH_2NH_2$ are both amines.

There are additional inorganic bases with a variety of formulas. An easy one to recognize is ammonia, NH_3, because it is similar to an amine.

Concept Review Exercises

1. Give the Arrhenius definitions of an acid and a base.
2. What is neutralization?

Answers

1. Arrhenius acid: a compound that increases the concentration of hydrogen ion (H^+) in aqueous solution; Arrhenius base: a compound that increases the concentration of hydroxide ion (OH^-) in aqueous solution.
2. The reaction of an acid and a base

Key Takeaway

- An Arrhenius acid increases the H^+ ion concentration in water, while an Arrhenius base increases the OH^- ion concentration in water.

Exercises

1. Give two examples of Arrhenius acids.
2. Give two examples of Arrhenius bases.
3. List the general properties of acids.
4. List the general properties of bases.
5. Write a balanced chemical equation for the neutralization of $Ba(OH)_2(aq)$ with $HNO_3(aq)$.
6. Write a balanced chemical equation for the neutralization of $H_2SO_4(aq)$ with $Cr(OH)_3(aq)$.
7. Gastric juice, the digestive fluid produced in the stomach, contains hydrochloric acid, HCl. Milk of Magnesia, a suspension of solid $Mg(OH)_2$ in an aqueous medium, is sometimes used to neutralize excess stomach acid. Write a complete balanced equation for the neutralization reaction.
8. Identify the salt produced in each acid-base reaction below. Then, balance the equation.
   a. $2HCl + Sr(OH)_2 \rightarrow 2H_2O + ??$
   b. KNO_3; $HNO_3 + KOH \rightarrow ?? + H_2O$
   c. $HF + Ca(OH)_2$ ---> $?? + H_2O$
9. Hydrazoic acid (HN_3) can be neutralized by a base. Write the balanced chemical equation for the reaction between hydrazoic acid and calcium hydroxide.
10. Citric acid ($H_3C_6H_5O_7$) has three hydrogen atoms that can form hydrogen ions in solution. Write the balanced chemical equation for the reaction between citric acid and sodium hydroxide.

Answers

1. HCl and HNO_3 (answers will vary)
2. NaOH and $Ca(OH)_2$ (answers will vary)
3. Sour taste, react with metals, react with bases, and turn litmus red
4. Bitter taste, feels slippery, react with acids and turn litmus blue

5. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O$

6. $3H_2SO_4(aq) + 2Cr(OH)_3(aq) \rightarrow Cr_2(SO_4)_3(aq) + 6H_2O$

7. $Mg(OH)_2 + 2HCl$ --> $MgCl_2 + 2H_2O$

8. a. $SrCl_2$; $2HCl + Sr(OH)_2 \rightarrow 2H_2O + SrCl_2$

   b. KNO_3; $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

   c. CaF_2; $2HF + Ca(OH)_2 \rightarrow CaF_2 + 2H_2O$

9. $2HN_3(aq) + Ca(OH)_2 \rightarrow Ca(N_3)_2 + 2H_2O$

10. $H_3C_6H_5O_7(aq) + 3NaOH(aq) \rightarrow Na_3C_6H_5O_7(aq) + 3H_2O$

Answers

Exercise 11.1.1

$C_8H_{10}N_4O_2(aq) + H_2O(\ell) \rightarrow C_8H_{11}N_4O_2^+(aq) + OH^-(aq)$

B-L base B-L acid

Exercise 11.1.2

$C_5H_5NH^+(solv) + C_2H_5OH(\ell) \rightarrow C_5H_5N(solv) + C_2H_5OH_2^+(solv)$

B-L acid B-L base

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11.2: Reactions of Acids and Bases

Learning Objectives

1. Recognize a compound as a Brønsted-Lowry acid or a Brønsted-Lowry base.
2. Illustrate the proton transfer process that defines a Brønsted-Lowry acid-base reaction.

Neutralization

As we noted previously, acids and bases react chemically with each other to form *salts*. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen ion containing compound and the base is a hydroxide ion containing compound, water is also a product. The general reaction is as follows:

$$\text{acid + base} \rightarrow \text{water + salt}$$

The reaction of acid and base to make water and a salt is called **neutralization**. Like any chemical equation, a neutralization chemical equation must be properly balanced. For example, the neutralization reaction between sodium hydroxide and hydrochloric acid is as follows:

$$NaOH(aq) + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(\ell)} \tag{11.2.1}$$

with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2H_2O_{(\ell)} \tag{11.2.2}$$

Once a neutralization reaction is properly balanced, we can use it to perform stoichiometry calculations, such as the ones we practiced earlier.

There are a number of examples of acid-base chemistry in everyday life. One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter "rises." Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2CO_3(aq) \tag{11.2.3}$$

$$H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$$

Example 11.2.2

Nitric acid [HNO_3(aq)] can be neutralized by calcium hydroxide [$Ca(OH)_2$(aq)]. Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.

Solution

HNO_3 is the acid, so it loses H^+. In acid-base neutralization reactions any metals will always be spectator ions. In this case the positive calcium ion will be attracted to the nitrate ion that remains when nitric acid loses H^+. This ionic compound is the salt and it has the formula $Ca(OH)_2$ because calcium ions have a +2 charge while hydroxide ions have a -1 charge. The base is OH^- and as a base its job is to accept H^+. The number of H atoms increases by 1 and the charge increases by 1 because OH^- is gaining an H^+, not an H atom. OH^- + H^+ form H_2O. Because there are two OH^- ions in the formula for $Ca(OH)_2$, we need two moles of HNO_3 to provide H^+ ions. The balanced chemical equation is as follows:

$Ca(OH)_2(aq) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2H_2O(\ell)$

Exercise 11.2.2

Hydrocyanic acid [HCN(aq)] can be neutralized by potassium hydroxide [KOH(aq)]. Write a balanced chemical equation for the reaction between these two compounds and identify the salt it produces.

Answer available at the end of this section.

Stomach Antacids

Our stomachs contain a solution of roughly 0.03 *M* HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

$$CaCO_3(s) + 2HCl(aq) \rightleftharpoons CaCl_2(aq) + H_2O(l) + CO_2(g)$$

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction:

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that :

$$H_3O^+ + OH^- \rightleftharpoons 2H_2O(l)$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect. Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in concert with magnesium hydroxide to balance the side effects of the two substances.

Conjugate Acid-Base Pairs

According to the Bronsted-Lowry theory of acids and bases, an acid is a proton donor and a base is a proton acceptor. Once an acid has given up a proton, the part that remains is called the acid's **conjugate base**. This species is a base because it can accept a proton (to re-form the acid). The conjugate base of HF (first example below) is fluoride ion, F^-.

$$\text{Acid} = \text{H}^+ + \text{Conjugate base of Acid}^-$$

$$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$$

$$\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$$

$$\text{NH}_4^+ \rightleftharpoons \text{H}^+ + \text{NH}_3$$

Similarly, the part of the base that remains after a base accepts a proton is called the base's **conjugate acid**. This species is an acid because it can give up a proton (and thus re-form the base). The conjugate acid of fluoride ion, F^- (first example below) is HF.

$$\text{H}^+ + \text{Base} = \text{Conjugate acid of Base}^+$$

$$\text{H}^+ + \text{F}^- \rightleftharpoons \text{HF}$$

$$H^+ + OH^- \rightleftharpoons H_2O$$

$$H^+ + H_2O \rightleftharpoons H_3O^+$$

$$H^+ + NH_3 \rightleftharpoons NH_4^+$$

To summarize, the conjugate base of HF is fluoride ion, F^-, and the conjugate acid of fluoride ion, F^-, is HF. The HF and F^- pair is referred to as a conjugate acid-base pair. The difference in the formulas of a conjugate acid-base pair is H^+. The table below lists conjugate acid-base pairs for your reference so that you can figure out the strategy of identifying them by looking at the pattern in the examples. For any given acid or base, you should be able to give its conjugate base or conjugate acid. The formula of an acid's conjugate base is generated by **removing a proton** (H^+) from the acid formula. The formula of the base's conjugate acid is formed by **adding a proton** (H^+) to the formula of the base.

Table 11.2.1. *Conjugate acid-base pairs.*

Conjugate Acid	Conjugate Base
H_3O^+	H_2O
H_2O	OH^-
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
NH_4^+	NH_3
NH_3	NH_2^-
CH_3COOH	CH_3COO^-
$CH_3NH_3^+$	CH_3NH_2

Example 11.2.3 : Conjugate Pairs

Write the formula of the conjugate base of (a) HCl and (b) HCO_3^-.

Write the formula of the conjugate acid of (c) CH_3NH_2 and (d) OH^-.

Solution:

A conjugate base is formed by removing a proton (H^+). A conjugate acid is formed by accepting a proton (H^+).

a. After HCl donates a proton, a Cl^- ion is produced, and so Cl^- is the conjugate base.
b. After hydrogen carbonate ion, HCO_3^-, donates a proton, its conjugate base, CO_3^{2-} is produced.
c. After accepting a proton (H^+), CH_3NH_2 is converted to $CH_3NH_3^+$, its conjugate acid.
d. After accepting a proton (H^+), OH^- is converted to H_2O, its conjugate acid.

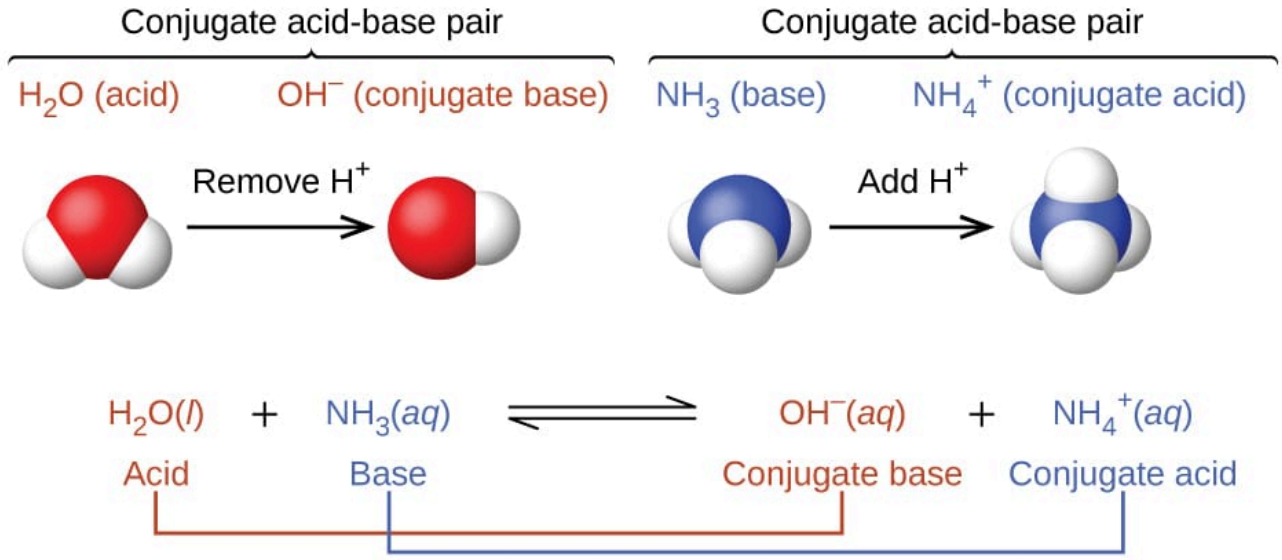
11.2.3Exercise : Conjugate Pairs

Write the formula of the conjugate base of (a) HNO_2 and (b) H_2CO_3.

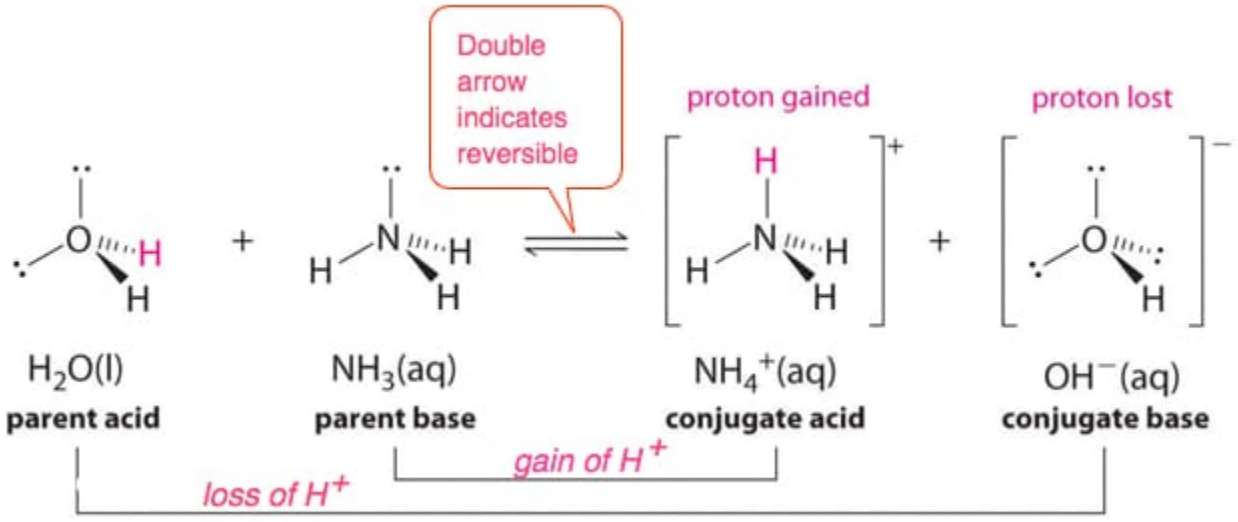
Write the formula of the conjugate acid of (c) C_6H5NH_2 and (d) HCO_3^-.

Answer available at the end of this section.

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The **conjugate acid–base pairs** for this reaction are NH_4^+/NH_3 and H_2O/OH^-. This means that the conjugate acid of the base NH_3 is NH_4^+ while the conjugate base of the acid NH_4^+ is NH_3. Similarly, the conjugate base of the acid H_2O is OH^-, and the conjugate acid of the base OH^- is H_2O.

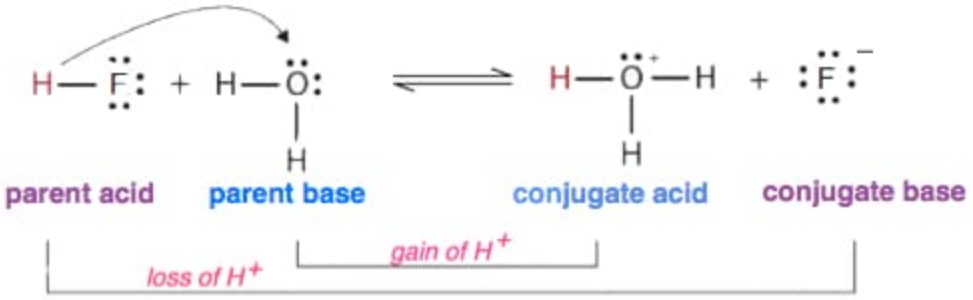


In the **forward** reaction, the parent acid is H_2O and and the parent base is NH_3 (shown in the illustration below). The acid H_2O loses a proton (H^+) to form its conjugate base OH^-. The base NH_3 gains a proton, to produce its conjugate acid NH_4^+. In the **reverse** reaction, the acid NH_4^+ loses a proton (H^+) to form its conjugate base NH_3. The base OH^- gains a proton, to produce its conjugate acid H_2O.



When hydrogen fluoride (HF) dissolves in water and ionizes, protons are transferred from hydrogen fluoride (parent acid) molecules to water (parent base) molecules, yielding hydronium ions (conjugate acid of water) and fluoride ions (conjugate base of HF):

$$\mathrm{HF} + H_2O \rightleftharpoons H_3O^+ + \mathrm{F}^-$$



Example 11.2.3

Identify the conjugate acid-base pairs in this equilibrium.

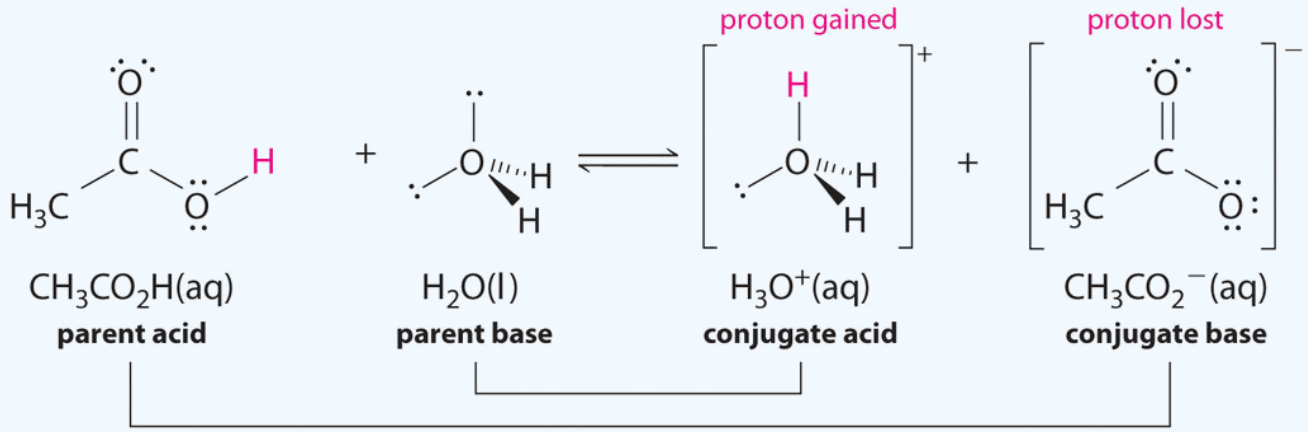
$$CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-$$

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).



Example 11.2.4

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H2O and OH^-, where H2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base.

The other pair consists of $(CH_3)_3N$ and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and $(CH_3)_3N$ is the conjugate base.

Exercise 11.2.4

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

Answer: available at the end of this section.

The use of conjugate acid-base pairs allows us to make a very simple statement about relative strengths of acids and bases. *The stronger an acid, the weaker its conjugate base*, and, conversely, *the stronger a base, the weaker its conjugate acid.*

To Your Health: Brønsted-Lowry Acid-Base Reactions in Pharmaceuticals

There are many interesting applications of Brønsted-Lowry acid-base reactions in the pharmaceutical industry. For example, drugs often need to be water soluble for maximum effectiveness. However, many complex organic compounds are not soluble or are only slightly soluble in water. Fortunately, those drugs that contain proton-accepting nitrogen atoms (and there are a lot of them) can be reacted with dilute hydrochloric acid [HCl(aq)]. The nitrogen atoms—acting as Brønsted-Lowry bases—accept the hydrogen ions from the acid to make an ion, which is usually much more soluble in water. The modified drug molecules can then be isolated as chloride salts:

$$RN(sl\ aq) + HCl(aq) \rightarrow RNH^+(aq) + Cl^-(aq) \rightarrow RNHCl(s)$$

where RN represents some organic compound containing nitrogen. The label *(sl aq)* means "slightly aqueous," indicating that the compound RN is only slightly soluble. Drugs that are modified in this way are called *hydrochloride salts*. Examples include the powerful painkiller codeine, which is commonly administered as codeine hydrochloride. Acids other than hydrochloric acid are also used. Hydrobromic acid, for example, gives *hydrobromide salts*. Dextromethorphan, an ingredient in many cough medicines, is dispensed as dextromethorphan hydrobromide. The accompanying figure shows another medication (lidocaine) as a hydrochloride salt.

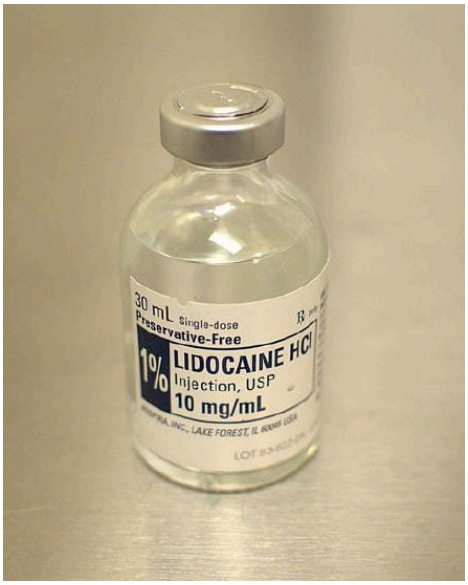


Figure used with permission from Wikipedia.

Concept Review Exercises

1. What is neutralization?

Answers

1. The reaction of an acid and a base

Exercises

1. Write a balanced chemical equation for the neutralization of $Ba(OH)_2(aq)$ with $HNO_3(aq)$.
2. Write a balanced chemical equation for the neutralization of $H_2SO_4(aq)$ with $Cr(OH)_3(aq)$.
3. Gastric juice, the digestive fluid produced in the stomach, contains hydrochloric acid, HCl. Milk of Magnesia, a suspension of solid $Mg(OH)_2$ in an aqueous medium, is sometimes used to neutralize excess stomach acid. Write a complete balanced equation for the neutralization reaction.
4. Identify the salt produced in each acid-base reaction below. Then, balance the equation.
   a. $2HCl + Sr(OH)_2 \rightarrow 2H_2O + ??$
   b. KNO_3; $HNO_3 + KOH \rightarrow ?? + H_2O$
   c. $HF + Ca(OH)_2$ ---> $?? + H_2O$
5. Hydrazoic acid (HN_3) can be neutralized by a base. Write the balanced chemical equation for the reaction between hydrazoic acid and calcium hydroxide.
6. Citric acid ($H_3C_6H_5O_7$) has three hydrogen atoms that can form hydrogen ions in solution. Write the balanced chemical equation for the reaction between citric acid and sodium hydroxide.

Answers

1. $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O$
2. $3H_2SO_4(aq) + 2Cr(OH)_3(aq) \rightarrow Cr_2(SO_4)_3(aq) + 6H_2O$
3. $Mg(OH)_2 + 2HCl$ --> $MgCl_2 + 2H_2O$
4. a. $SrCl_2$; $2HCl + Sr(OH)_2 \rightarrow 2H_2O + SrCl_2$

   b. KNO_3; $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

   c. CaF_2; $2HF + Ca(OH)_2 \rightarrow CaF_2 + 2H_2O$

1. $2HN_3(aq) + Ca(OH)_2 \rightarrow Ca(N_3)_2 + 2H_2O$
2. $H_3C_6H_5O_7(aq) + 3NaOH(aq) \rightarrow Na_3C_6H_5O_7(aq) + 3H_2O$

Answers

Exercise 11.2.2

$KOH(aq) + HCN(aq) \rightarrow KCN(aq) + H_2O(\ell)$

KCN is the salt.

a. NO_2^- is the conjugate base of HNO_2.

b. HCO_3^- is the conjugate base of H_2CO_3

c. $C_6H5NH_3^+$ is the conjugate acid of C_6H5NH_2.

d. H_2CO_3 is the conjugate acid of HCO_3^-

Exercise 11.2.4

H2O (acid) and OH^- (base); $NH2^-$ (base) and NH3 (acid)

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11.3: Water - Both an Acid and a Base

Learning Objectives

- To write chemical equations for water acting as an acid and as a base.

Water (H_2O) is an interesting compound in many respects. Here, we will consider its ability to behave as an acid or a base.

In some circumstances, a water molecule will accept a proton and thus act as a **Brønsted-Lowry base**. We saw an example in the dissolving of HCl in H_2O:

$$\mathrm{HCl} + \mathrm{H_2O}(\ell) \rightarrow \mathrm{H_3O^+(aq)} + \mathrm{Cl^-(aq)} \tag{11.3.1}$$

In other circumstances, a water molecule can donate a proton and thus act as a **Brønsted-Lowry acid**. For example, in the presence of the amide ion (see Example 4 in Section 10.2), a water molecule donates a proton, making ammonia as a product:

$$H_2O(\ell) + NH_2^-(aq) \rightarrow OH^-(aq) + NH_3(aq) \tag{11.3.2}$$

In this case, NH_2^- is a Brønsted-Lowry base (the proton acceptor).

So, depending on the circumstances, H_2O can act as either a Brønsted-Lowry acid or a Brønsted-Lowry base. 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. A substance that can either donate or accept a proton, depending on the circumstances, is called an **amphiprotic** compound.

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:

$$H_2O(\ell) + H_2O(\ell) \rightarrow H_3O^+(aq) + OH^-(aq) \tag{11.3.3}$$

This process is called the **autoionization of water** (Figure 11.3.1) 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.)

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.

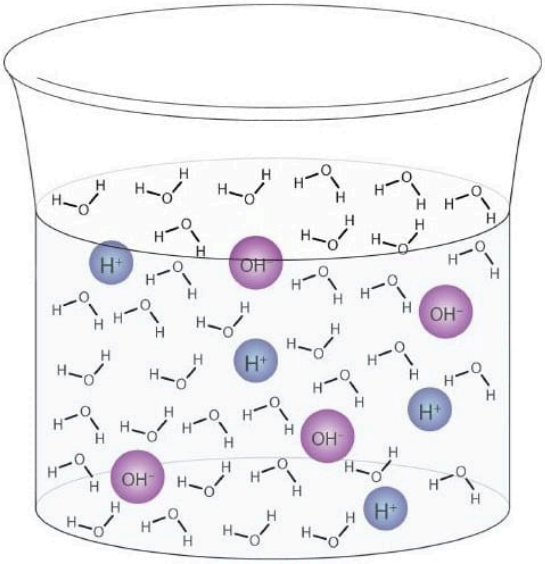


Figure 11.3.1 Autoionization. A small fraction of water molecules—approximately 6 in 100 million—ionize spontaneously into hydronium ions and hydroxide ions. This picture necessarily overrepresents the amount of autoionization that really occurs in pure water.

Example 11.3.1

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $H_2O(\ell) + NO_2^-(aq) \rightarrow HNO_2(aq) + OH^-(aq)$
2. $HC_2H_3O_2(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)$

Solution

1. In this reaction, the water molecule donates a proton to the NO_2^- ion, making $OH^-(aq)$. As the proton donor, H_2O acts as a Brønsted-Lowry acid.
2. In this reaction, the water molecule accepts a proton from $HC_2H_3O_2$, becoming $H_3O^+(aq)$. As the proton acceptor, H_2O is a Brønsted-Lowry base.

Exercise 11.3.2

Identify water as either a Brønsted-Lowry acid or a Brønsted-Lowry base.

1. $HCOOH(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + HCOO^-(aq)$
2. $H_2O(\ell) + PO_4^{3-}(aq) \rightarrow OH^-(aq) + HPO_4^{2-}(aq)$

Answer available at the end of this section.

Key Takeaway

- Water molecules can act as both an acid and a base, depending on the conditions.

Answers

Exercise 11.3.2

1. H_2O acts as the proton acceptor (Brønsted-Lowry base)

2. H_2O acts as the proton donor (Brønsted-Lowry acid)

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11.4: The Strengths of Acids and Bases

Learning Objectives

- Describe the difference between strong and weak acids and bases.
- Describe how a chemical reaction reaches chemical equilibrium.

Acids and bases do not all demonstrate the same degree of chemical activity in solution. Different acids and bases have different strengths. In other words, some are strong electrolytes and some are weak electrolytes.

Strong and Weak Acids

Let us consider the strengths of acids first. A small number of acids ionize completely in aqueous solution. For example, when HCl dissolves in water, every molecule of HCl separates into a hydronium ion and a chloride ion:

$$\mathrm{HCl\,(g) + H_2O\,(l) \xrightarrow{\sim 100\,\%} H_3O^+\,(aq) + Cl^-\,(aq)} \tag{11.4.1}$$

HCl(aq) is one example of a **strong acid**, which is a compound that is essentially **100% ionized** in aqueous solution. There are very few strong acids. The important ones are listed in Table 11.4.1.

Table 11.4.1: Strong Acids and Bases (All in Aqueous Solution)

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	$Mg(OH)_2$
H_2SO_4	$Ca(OH)_2$
$HClO_4$	

By analogy, a strong base is a compound that is essentially 100% ionized in aqueous solution. As with acids, there are only a few strong bases, which are also listed in Table 11.4.1.

If an acid is **not listed** in Table 11.4.1, it is likely a **weak acid**, which is a compound that is not 100% ionized in aqueous solution. Similarly, a weak base is a compound that is not 100% ionized in aqueous solution. For example, acetic acid ($HC_2H_3O_2$) is a weak acid. The ionization reaction for acetic acid is as follows:

$$\mathrm{HC_2H_3O_2\,(aq) + H_2O(\ell) \rightarrow H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)} \tag{11.4.2}$$

Depending on the concentration of $HC_2H_3O_2$, the ionization reaction may occur only for 1%–5% of the acetic acid molecules.

Looking Closer: Household Acids and Bases

Many household products are acids or bases. For example, the owner of a swimming pool may use muriatic acid to clean the pool. Muriatic acid is another name for hydrochloric acid [$\mathrm{HCl\,(aq)}$]. Vinegar has already been mentioned as a dilute solution of acetic acid [$\mathrm{HC_2H_3O_2\,(aq)}$]. In a medicine chest, one may find a bottle of vitamin C tablets; the chemical name of vitamin C is ascorbic acid ($\mathrm{HC_6H_7O_6}$).

One of the more familiar household bases is ammonia (NH_3), which is found in numerous cleaning products. As we mentioned previously, ammonia is a base because it increases the hydroxide ion concentration by reacting with water:

$$\mathrm{NH_3\,(aq) + H_2O(\ell)\ \rightarrow NH_4^+\,(aq) + OH^-\,(aq)} \tag{11.4.3}$$

Many soaps are also slightly basic because they contain compounds that act as Brønsted-Lowry bases, accepting protons from water and forming excess hydroxide ions. This is one reason that soap solutions are slippery.



Figure 11.4.1: (left) Bottles of alkaline drain cleaners containing sodium hydroxide can dissolve greases and hair. (right) Chemical burns caused by sodium hydroxide solution photographed 44 hours after exposure. Figures used with permission from Wikipedia.

Perhaps the most dangerous household chemical is the lye-based drain cleaner. Lye is a common name for sodium hydroxide, although it is also used as a synonym for potassium hydroxide. Lye is an extremely caustic chemical that can react with grease, hair, food particles, and other substances that may build up and form a clog in a pipe. Unfortunately, lye can also attack tissues and other substances in our bodies. Thus, when we use lye-based drain cleaners, we must be very careful not to touch any of the solid drain cleaner or spill the water it was poured into. Safer, nonlye drain cleaners use peroxide compounds to react on the materials in the clog and clear the drain.

Chemical Equilibrium in Weak Acids and Bases

The behavior of **weak acids and bases** illustrates a key concept in chemistry. Does the chemical reaction 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 $\mathrm{HC_2H_3O_2\,(aq)}$ is as follows:

$$\mathrm{HC_2H_3O_2\,(aq) + H_2O(\ell)\ \rightarrow H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)} \tag{11.4.4}$$

The reverse process also begins to occur:

$$\mathrm{H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)\ \rightarrow HC_2H_3O_2\,(aq) + H_2O(\ell)} \tag{11.4.5}$$

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:

$$\mathrm{HC_2H_3O_2\,(aq) + H_2O(\ell)\ \rightleftharpoons H_3O^+\,(aq) + C_2H_3O_2^-\,(aq)} \tag{11.4.6}$$

The $\rightleftharpoons$ implies that **both the forward and reverse reactions** are occurring, and their effects cancel each other out. A 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*.

Example 11.4.1: Partial Ionization

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

a. $HNO_2(aq)$
b. $C_5H_5N(aq)$

Solution

a. $HNO_2(aq) + H_2O(\ell) \rightleftharpoons NO_2^-(aq) + H_3O^+(aq)$
b. $C_5H_5N(aq) + H_2O(\ell) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq)$

Exercise 11.4.1

Write the equilibrium chemical equation for the partial ionization of each weak acid or base.

a. $HF_{(aq)}$
b. $AgOH_{(aq)}$
c. $CH_3NH_2(aq)$

Answer available at the end of this section.

Key Takeaways

- Acids and bases can be strong or weak depending on the extent of ionization in solution.
- Most chemical reactions reach equilibrium at which point there is no net change.

Concept Review Exercises

1. Explain the difference between a strong acid or base and a weak acid or base.
2. Explain what is occurring when a chemical reaction reaches equilibrium.

Answers

1. A strong acid or base is 100% ionized in aqueous solution; a weak acid or base is less than 100% ionized.
2. The overall reaction progress stops because the reverse process balances out the forward process.

Exercises

1. Name a strong acid and a weak acid. (Hint: use Table 10.4.1.)
2. Name a strong base and a weak base. (Hint: use Table 10.4.1.)
3. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution.
    1. HF
    2. $HC_2H_3O_2$
    3. HCl
    4. $HClO_4$
4. Is each compound a strong acid or a weak acid? Assume all are in aqueous solution.
    1. H_2SO_4
    2. HSO_4^-

3. HPO_4^{2-}
4. HNO_3

5. Is each compound a strong base or a weak base? Assume all are in aqueous solution.
   1. NH_3
   2. NaOH
   3. $Mg(OH)_2$
   4. $Cu(OH)_2$
6. Is each compound a strong base or a weak base? Assume all are in aqueous solution.
   1. KOH
   2. H_2O
   3. $Fe(OH)_2$
   4. $Fe(OH)_3$
7. Write the chemical equation for the equilibrium process for each weak acid in Exercise 3.
8. Write the chemical equation for the equilibrium process for each weak acid in Exercise 4.
9. Write the chemical equation for the equilibrium process for each weak base in Exercise 5.
10. Write the chemical equation for the equilibrium process for each weak base in Exercise 6.

Answers

1. strong acid: HCl; weak acid: $HC_2H_3O_2$ (answers will vary)
2. strong base: NaOH; weak base: NH_3 (answers will vary)

3. 1. weak
   2. weak
   3. strong
   4. strong
4. 1. strong
   2. weak
   3. weak
   4. strong

5. 1. weak
   2. strong
   3. strong
   4. weak
6. 1. strong
   2. weak
   3. weak
   4. weak

7. HF(aq) ⇌ H^+(aq) + F^-(aq); 3b: $HC_2H_3O_2$(aq) ⇌ H^+(aq) + $C_2H_3O_2^-$(aq)

8. HSO_4^-(aq) ⇌ H^+(aq) + SO_4^{2-}(aq); 4c: HPO_4^{2-}(aq) ⇌ H^+(aq) + PO_4^{3-}(aq)

9. NH_3(aq) + H_2O ⇌ NH_4^+(aq) + OH^-(aq); 5d: $Cu(OH)_2$(aq) ⇌ Cu^{2+}(aq) + $2OH^-$(aq)

10. H_2O + H_2O ⇌ H_3O^+(aq) + OH^-(aq); 6c: $Fe(OH)_2$(aq) ⇌ Fe^{2+}(aq) + $2OH^-$(aq); 6d: Fe(OH)3(aq) ⇌ Fe^{3+}(aq) + $3OH^-$(aq)

Answers

Exercise 11.4.1

a. HF(aq) + $H_2O(\ell)$ ⇌ F^-(aq) + H_3O^+(aq)

b. AgOH(aq) ⇌ Ag^+(aq) + OH^-(aq)

c. CH_3NH_2(aq) + $H_2O(\ell)$ ⇌ $CH_3NH_3^+$(aq) + OH^-(aq)

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11.5: The pH Scale

Learning Objectives

- Define the pH scale and use it to describe acids and bases.
- Calculate pH, pOH, [H_3O^+], and [OH^-]

The Ion-Product 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:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{11.5.1}$$

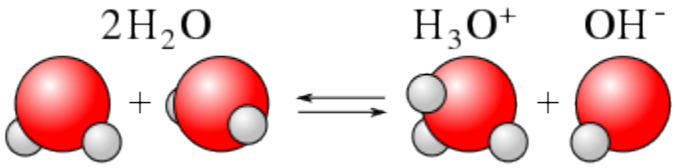


Figure 11.5.2: Autoionization of water, giving hydroxide and hydronium ions.

We often use the simplified form of the reaction:

$$H_2O\,(l) \rightleftharpoons H^+\,(aq) + OH^-\,(aq) \tag{11.5.2}$$

No matter what the pH, there is a relationship between the concentrations of hydronium ions and hydroxide ions. When those concentrations are multiplied together they give a constant called the ion-product of water K_w.

$$K_w = [H_3O^+][OH^-] \tag{11.5.3}$$

The variables with square brackets around chemical formulas, $[H_3O^+]$ and $[OH^-]$, indicate the concentration of each substance in a unit of molarity. The value of K_w is very small because the concentrations of hydronium and hydroxide ions are small. At 25°C, the experimentally determined value of K_w in pure water is 1.0×10^{-14}. The unit for K_w is customarily omitted.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{11.5.4}$$

Although the concentrations are small, they can be detected and measured experimentally. 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.**

$$[H_3O^+] = [OH^-] = 1.0\times10^{-7} \text{ at pH 7}$$

The product of these two concentrations is 1.0×10^{-14}

$$[H^+]\times[OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14} \tag{11.5.5}$$

- For acids, the concentration of H^+ or [H_3O^+]) is greater than 1.0×10^{-7} M
- For bases, the concentration of OH^- or [OH^-] is greater than 1.0×10^{-7} M.

Aqueous HCl is an example of **acidic solution**. Hydrogen chloride (HCl) ionizes to produce H^+ and Cl^- ions upon dissolving in water.

$$\mathrm{HCl\,(g) \longrightarrow H^+\,(aq) + Cl^-\,(aq)}$$

This increases the concentration of H_3O^+ ions in the solution. According to Le Chatelier's principle, the equilibrium represented by

$$\mathrm{H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + OH^-\,(aq)}$$

is forced to the left, towards the reactant. As a result, the concentration of the hydroxide ion decreases.

Now, consider KOH (aq), a **basic solution**. Solid potassium hydroxide (KOH) dissociates in water to yield potassium ions and hydroxide ions.

$$\mathrm{KOH(s) \rightarrow K^+(aq) + OH^-(aq)}$$

The increase in concentration of the OH^- ions will cause a decrease in the concentration of the H^+ ions because they react to form water.

No matter whether the aqueous solution is an acid, a base, or neutral:and the ion-product of $[H^+][OH^-]$ 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

This means that if you know $[H_3O^+]$ for a solution, you can calculate what $[OH^-]$ has to be for the product to equal 1.0×10^{-14}, or if you know $[OH^-]$), you can calculate $[H_3O^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_w.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{11.5.6}$$

Example 11.5.2

Hydrochloric acid (HCl) is a strong acid, meaning it is 100% ionized in solution. What is the $[H^+]$ and the $[OH^-]$ in a solution of 2.0×10^{-3} M HCl?

Solution

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

Known

- $[HCl] = 2.0\times10^{-3}$ M
- $Kw = 1.0\times10^{-14}$

Unknown

- $[H^+]$=?M
- $[OH^-]$=?M

Because HCl is 100% ionized, the concentration of H+ ions in solution will be equal to the original concentration of HCl. Each HCl molecule that was originally present ionizes into one H+ ion and one Cl– ion. The concentration of OH– can then be determined from the [H+] and Kw.

Step 2: Solve.

$$[H^+]=2.0\times10^{-3}\ \mathrm{M}$$

$$Kw = [H^+][OH^-] = 1.0\times10^{-14}$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} = 5.0 \times 10^{-12} M \tag{11.5.7}$$

Step 3: Think about your result.

The $[H_3O^+]$ is much higher than the $[OH^-]$ because the solution is acidic.

Exercise 11.5.2

Sodium hydroxide (NaOH) is a strong base. What is the $[H^+]$ and the $[OH^-]$ in a 0.001 M NaOH solution at 25 °C?

Answer available at the end of this section.

The pH Scale

One measure of the strength of an acid or a base solution is the pH scale, which is based on the concentration of the hydronium (or hydrogen) ion in aqueous solution.

$$pH = -\log[H^+] \tag{11.5.8}$$

or

$$pH = -\log[H_3O^+] \tag{11.5.9}$$

Figure 11.5.3 illustrates this relationship, along with some examples of various solutions. Because hydrogen ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$), the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the ***negative* log of** $[H_3O^+]$, which will give a positive value for pH.

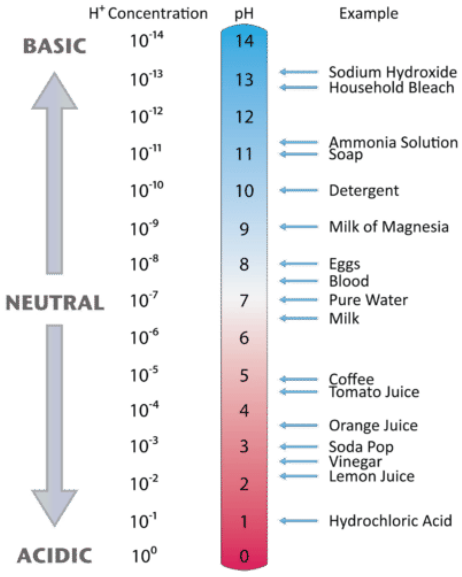


Figure 11.5.3: The relationship between $[H^+]$ and pH *values for several common materials.*

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 11.5.4.

Table 11.5.4: Acidic, Basic and Neutral pH Values

Classification	Relative Ion Concentrations	pH at 25 °C
acidic	$[H^+] > [OH^-]$	pH < 7
neutral	$[H^+] = [OH^-]$	pH = 7
basic	$[H^+] < [OH^-]$	pH > 7

Example 11.5.3

Find the pH, given the $[H^+]$ 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. $pH = - \log [1 \times 10^{-3}] = 3.0$ (1 decimal place since 1 has 1 significant figure)
b. $pH = - \log [2.5 \times 10^{-11}] = 10.60$ (2 decimal places since 2.5 has 2 significant figures)
c. $pH = - \log [4.7 \times 10^{-9}] = 8.33$ (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.

Exercise 11.5.3

Find the pH, given $[H^+]$ of the following:

a. 5.8×10^{-4} M
b. 1.0×10^{-7} M

Answer available at the end of this section.

Table 11.5.5 lists the pH of several common solutions. The most acidic among the listed solutions is battery acid with the lowest pH value (0.3). The most basic is 1M NaOH solution with the highest pH value of 14.0. Notice that some biological fluids (stomach acid and urine) are not neutral. You may also notice that many food products are slightly acidic because they contain solutions of weak acids. pH depends on both the identity of the substance and its concentration, so concentrations are listed for pure substances.

Table 11.5.5: The pH Values of Some Common Solutions

Solution	pH
battery acid	0.3
stomach acid	1–2

Solution	pH
lemon or lime juice	2.1
vinegar	2.8–3.0
Coca-Cola	3
wine	2.8–3.8
beer	4–5
coffee	5
milk	6
urine	6
pure H_2O	7
(human) blood	7.3–7.5
sea water	8
antacid (milk of magnesia)	10.5
NH_3 (1 M)	11.6
bleach	12.6
NaOH (1 M)	14.0

Example 11.5.4

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.

Exercise 11.5.4

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 available at the end of this section.

Calculating $[H_3O^+]$ and $[OH^-]$

pH can be determined experimentally using chemical indicator or a pH meter. The pH can then be used to calculate $[H_3O^+]$ by rearranging the pH equation.

$$pH = -log(H_3O^+) \tag{11.5.10}$$

rearranges to

$$[H_3O^+] = 10^{-pH} \tag{11.5.11}$$

in order to get $[H_3O^+]$ by itself. If the pH of human blood is 7.4, then the concentration of $[H_3O^+]$ is

$$[H_3O^+] = 10^{-pH} = 10^{-7.4} = 4x10^{-8}M \tag{11.5.12}$$

A concentration of 4 x 10^{-8} M makes sense for a pH of 7.4 because it is close to 1 x 10^{-7} M which we would expect for a pH of 7.

If the pH is known and the concentration of hydroxide ions is desired, then a two step calculation is needed. You could use

$$[H_3O^+] = 10^{-pH} \tag{11.5.13}$$

to use the pH to calculate $[H_3O^+]$ followed by

$$K_w = [\mathrm{H_3O^+}][\mathrm{OH^-}] \tag{11.5.14}$$

in which the $[H_3O^+]$ is used to calculate \([OH^-]).

Another option is to calculate pOH which is the -log[OH-]. Just as pH is a more convenient way to express $[H_3O^+]$, the pOH is a more convenient way to express \([OH^-]). While there are not usually instruments available that measure pOH directly, there is a simple relationship between pH and pOH.

$$14 = pH + pOH \tag{11.5.15}$$

For example, the antacid milk of magnesia has a pH of 10.5. We can use that to calculate pOH

$$14 = pH + pOH \tag{11.5.16}$$

$$pOH = 14 - pH = 14 - 10.5 = 3.5 \tag{11.5.17}$$

Then the concentration of $[OH^-]$ can be calculated using an equation analogous to $[H_3O^+] = 10^{-pH}$

$$[OH^-] = 10^{-pOH} = 10^{-3.5} = 3x1 -^{-4} M \tag{11.5.18}$$

A concentration of hydroxide ions greater than $1x10^{-7}$ M makes sense since milk of magnesia is basic. (Remember that -4 is greater than -7).

Acid Rain

Normal rainwater has a pH between 5 and 6 due to the presence of dissolved CO_2 which forms carbonic acid:

$$\mathrm{H_2O\,(l) + CO_2\,(g) \longrightarrow H_2CO_3\,(aq)} \tag{11.5.19}$$

$$\mathrm{H_2CO_3\,(aq) \rightleftharpoons H^+\,(aq) + HCO_3{}^-\,(aq)} \tag{11.5.20}$$

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:

$$\mathrm{H_2O\,(l) + SO_3\,(g) \longrightarrow H_2SO_4\,(aq)} \tag{11.5.21}$$

$$\mathrm{H_2SO_4\,(aq) \longrightarrow H^+\,(aq) + HSO_4{}^-\,(aq)} \tag{11.5.22}$$

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 11.5.4). 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 11.5.4: (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)

Key Takeaways

- The ion-product of $[H^+][OH^-]$ in an aqueous solution remains constant.
- A ***p*H** value is simply the *negative* of the logarithm of the H^+ ion concentration ($-\log[H^+]$).
- The pH scale is used to succinctly communicate the acidity or basicity of a solution.
- A solution is **acidic** if *p*H < 7.
- A solution is **basic** if *p*H > 7.
- A solution is **neutral** if *p*H = 7.

Concept Review Exercises

1. Define *pH*.

Answers

1. pH is a measure of the hydrogen ion concentration.

Exercises

Indicate whether solutions with the following pH values are acidic, basic, or neutral:

1. pH = 9.4
2. pH = 7.0
3. pH = 1.2
4. pH = 6.5

Answers

1. basic

2. neutral
3. acidic (strongly)
4. acidic (mildly)

Answers

Exercise 11.5.2

$[OH^-]$ = 0.001M or 1 x 10^{-3}M; $[H^+]$=1×10^{-11}M.

Exercise 11.5.3

a. 3.24

b. 7.00

Exercise 11.5.4

a. slightly basic

b. basic

c. acidic

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11.6: Buffers

Learning Objectives

- To define *buffer* and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have 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—a 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 change in its pH when H^+ or OH^- ions are added or removed owing to some other reaction taking place in the same solution. Buffers do so by being composed of certain pairs of solutes: 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 acetate 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^+ or OH^- it is expected to buffer against. 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:

$$HC_2H_3O_{2(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(\ell)} + C_2H_3O^-_{2(aq)} \tag{11.6.1}$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

> Many people are aware of the concept of buffers from *buffered aspirin*, which is aspirin that also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt acts like a base, while aspirin is itself a weak acid.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$H^+_{(aq)} + C_2H_3O^-_{2(aq)} \rightarrow HC_2H_3O_{2(aq)} \tag{11.6.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid. Figure 11.6.1 illustrates both actions of a buffer.

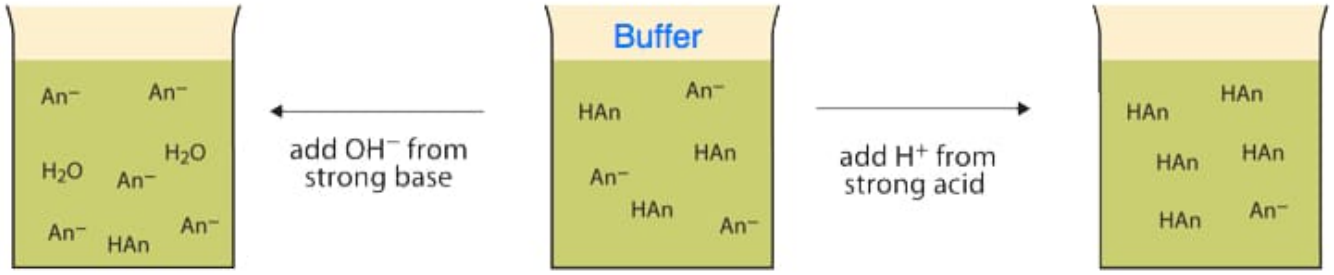


Figure 11.6.1: The Action of Buffers. Buffers can react with both strong acids (top) and strong bases (bottom) to minimize large changes in pH.

A simple buffer system might be a 0.2 *M* solution of sodium acetate; the conjugate pair here is 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+ or OH^- that may result from other processes going on in the solution.

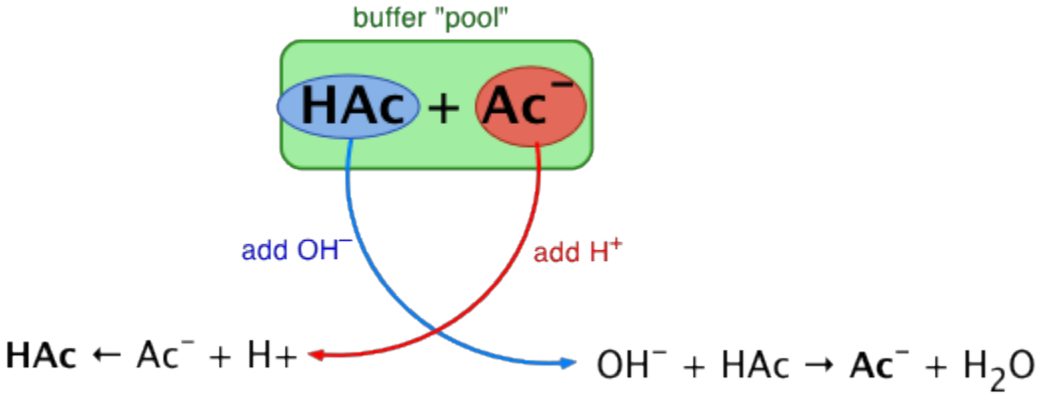


Figure 11.6.2: How HAc/Ac buffer works

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain capacity. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

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:

$$NH_{3(aq)} + H^+_{(aq)} \rightarrow NH^+_{4(aq)} \tag{11.6.3}$$

while the ammonium ion [NH_4^+(aq)] can react with any hydroxide ions introduced by strong bases:

$$NH^+_{4(aq)} + OH^-_{(aq)} \rightarrow NH_{3(aq)} + H_2O_{(\ell)} \tag{11.6.4}$$

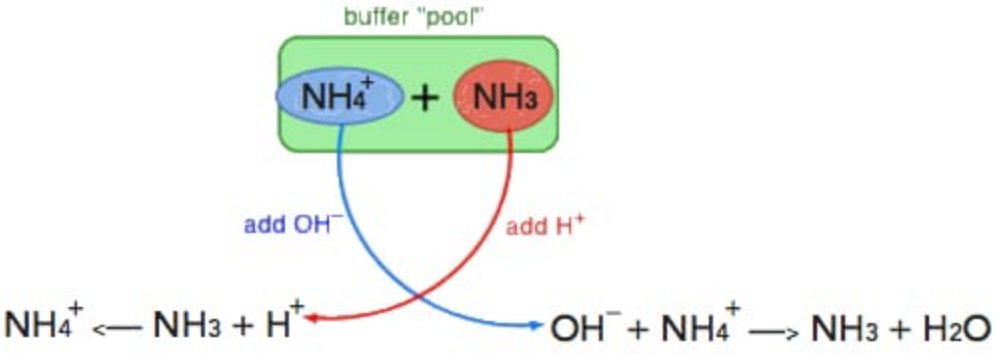


Figure 11.6.3: How NH_4^+/NH_3 buffer works

Example 11.6.1

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

a. $HCHO_2$ and $NaCHO_2$
b. HCl and NaCl
c. CH_3NH_2 and CH_3NH_3Cl
d. NH_3 and NaOH

Solution

a. Formic acid ($HCHO_2$) is a weak acid, while $NaCHO_2$ is a salt supplying —formate ion (CHO_2^-), the conjugate base of $HCHO_2$. The combination of these two solutes would make a buffer solution.
b. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
c. Methylamine (CH_3NH_2) is like ammonia, a weak base. The compound CH_3NH_3Cl is a salt supplying $CH_3NH_3^+$, the conjugate acid of CH_3NH_2. The combination of these two solutes would make a buffer solution.
d. Ammonia (NH_3) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

Exercise 11.6.1

Which solute combinations can make a buffer solution? Assume all are aqueous solutions.

a. $NaHCO_3$ and NaCl
b. H_3PO_4 and NaH_2PO_4
c. NH_3 and $(NH_4)_3PO_4$
d. NaOH and NaCl

Answer available at the end of this section.

Application: 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 and antipyretic (fever reducer).

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. Because it is not given in Table 10.5.1, acetylsalicylic acid is a weak acid. However, it is still an acid, and given that 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.

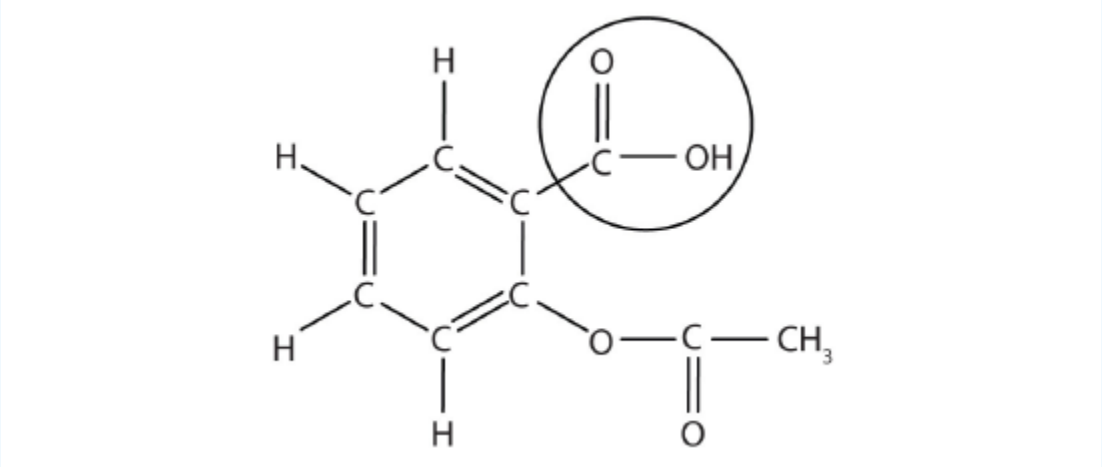


Figure 11.6.4 The Molecular Structure of Aspirin. The circled atoms are the acid part of the molecule.

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.

Buffer solutions are essential components of all living organisms.

- Our blood is buffered to maintain a pH of 7.4 that must remain unchanged as metabolically-generated CO_2 (carbonic acid) 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 (Figure \({11{.}6{.}\eatSpaces{1} \endSpaces}) is maintained by buffer systems.

Table 7.3.2: 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

Compartment	pH
Cerebrospinal Fluid	7.3
Blood	7.43-7.45
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 $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Medicine: The Buffer System in Blood

The normal pH of human blood is about 7.4. The carbonate buffer system in the blood uses the following equilibrium reaction:

$$\mathrm{CO_2}(g) + 2\,\mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_2CO_3}(aq) \rightleftharpoons \mathrm{HCO_3}^-(aq) + \mathrm{H_3O}^+(aq)$$

The concentration of carbonic acid, H_2CO_3 is approximately 0.0012 *M*, and the concentration of the hydrogen carbonate ion, $\mathrm{HCO_3}^-$, is around 0.024 *M*. Using the Henderson-Hasselbalch equation and the pK_a of carbonic acid at body temperature, we can calculate the pH of blood:

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} + \log\frac{[\mathrm{base}]}{[\mathrm{acid}]} = 6.1 + \log\frac{0.024}{0.0012} = 7.4$$

The fact that the H_2CO_3 concentration is significantly lower than that of the $\mathrm{HCO_3}^-$ ion may seem unusual, but this imbalance is due to the fact that most of the by-products of our metabolism that enter our bloodstream are acidic. Therefore, there must be a larger proportion of base than acid, so that the capacity of the buffer will not be exceeded.

Lactic acid is produced in our muscles when we exercise. As the lactic acid enters the bloodstream, it is neutralized by the $\mathrm{HCO_3}^-$ 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^+] 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.

Key Takeaway

- A buffer is a solution that resists sudden changes in pH.

Concept Review Exercise

1. Explain how a buffer prevents large changes in pH.

Answer

1. 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. 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. 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.

Answers

Exercise 11.6.1

b. H_3PO_4 (weak acid) and $H_2PO_4^-$ (conjugate base of H_3PO_4) make a buffer.

c. NH_3 (weak base) and NH_4^+ (conjugate acid of NH_3) make a buffer

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12: GASES

Much of our focus so far has been on condensed phases: solids and liquids; now we turn our attention to gases. The main reason for the difference in behaviour between gases and the other phases of matter is the lack of intermolecular forces between the particles of gases.

12.1: Gases and Pressure

12.2: Ideal Gas Law

12.3: The Combined Gas Law

12.4: Dalton's Law

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12.1: Gases and Pressure

Learning Objectives

- Describe the gas phase using kinetic molecular theory
- Convert among pressure units

The gas phase is unique among the three states of matter in that there are some simple models we can use to predict the physical behavior of all gases—independent of their identities. We cannot do this for the solid and liquid states.

How is it that we can model all gases independent of their chemical identity? The answer is in a group of statements called the kinetic theory of gases:

- Gases are composed of tiny particles that are separated by large distances.
- Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container.
- The velocity of gas particles is related to the temperature of a gas.
- Gas particles do not experience any force of attraction or repulsion with each other.

Did you notice that none of these statements relates to the identity of the gas? This means that all gases should behave similarly regardless of whether they are composed of an element (e.g. helium), a compound (e.g. carbon dioxide), or a mixture (e.g. air). A gas that follows these statements perfectly is called an *ideal gas*. Most gases show slight deviations from these statements and are called *real gases*. However, the existence of real gases does not diminish the importance of the kinetic theory of gases.

One of the statements of the kinetic theory mentions collisions. As gas particles are constantly moving, they are also constantly colliding with each other and with the walls of their container. There are forces involved as gas particles bounce off the container walls (Figure 12.1.1). The force generated by gas particles divided by the area of the container walls yields pressure. Pressure is a property we can measure for a gas, but we typically do not consider pressure for solids or liquids.

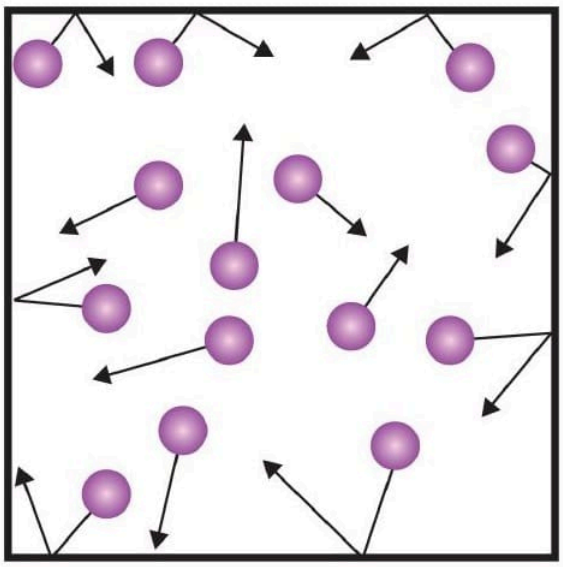


Figure 12.1.1: Gas Pressure. Pressure is what results when gas particles rebound off the walls of their container.

Common pressure units used in science include pascals (Pa), bars (bar), atmospheres (atm), and millimeters of mercury (mm Hg). One pascal is not a very large amount of pressure. A more useful unit of pressure is the bar, which is 100,000 Pa (1 bar = 100,000 Pa). The units of pressure that we will use most often in this course are the atmosphere (atm), which was originally defined as the average pressure of Earth's atmosphere at sea level; and mm Hg (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. There are exactly 760 mm Hg (or torr) in 1 atm. A bar equals 1.01325

atm. The ability to convert from one pressure unit to another is a useful skill because different instruments measure in different units.

Example 12.1.1: Converting Pressures

How many atmospheres are in 1,547 mm Hg?

Solution

Because there are 760 mm Hg in 1 atm, we can use this conversion factor to do the mathematical conversion:

$$1,547 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 2.035 \text{ atm}$$

Note how the mm Hg units cancel algebraically.

Exercise 12.1.1: Converting Pressures

How many millimeters of mercury are in 9.65 atm?

Answer available at the end of this section.

The kinetic theory also states that there is no interaction between individual gas particles. Although we know that there are, in fact, intermolecular interactions in real gases, the kinetic theory assumes that gas particles are so far apart that the individual particles don't "feel" each other. Thus, we can treat gas particles as tiny bits of matter whose identity isn't important to certain physical properties.

Concept Review Exercise

1. What is pressure, and what units do we use to express it?

Answer

1. Pressure is the force per unit area; common units include pascals, torr, millimeters of mercury, or atmospheres.

Key Takeaway

- The gas phase has certain general properties characteristic of that phase.

Exercises

1. What is the kinetic theory of gases?
2. According to the kinetic theory of gases, the individual gas particles are (always, frequently, never) moving.
3. Why does a gas exert pressure?
4. Why does the kinetic theory of gases allow us to presume that all gases will show similar behavior?
5. Arrange the following pressure quantities in order from smallest to largest: 1 mm Hg, 1 Pa, and 1 atm.
6. Which unit of pressure is larger—the torr or the atmosphere?

7. How many mm Hg are there in 1.56 atm?

8. Blood pressures are expressed in millimeters of mercury. What would be the blood pressure in atmospheres if a patient's systolic blood pressure is 120 mm Hg and the diastolic blood pressure is 82 mm Hg? (In medicine, such a blood pressure would be reported as "120/82," spoken as "one hundred twenty over eighty-two.")

9. In weather forecasting, barometric pressure is expressed in inches of mercury (in. Hg), where there are exactly 25.4 mm Hg in every 1 in. Hg. What is the barometric pressure in millimeters of mercury if the barometric pressure is reported as 30.21 in. Hg?

Answers

1. Gases are composed of tiny particles that are separated by large distances. Gas particles are constantly moving, experiencing collisions with other gas particles and the walls of their container. The velocity of gas particles is related to the temperature of a gas. Gas particles do not experience any force of attraction or repulsion with each other.

2. always

3. A gas exerts pressure as its particles rebound off the walls of its container.

4. Because the molecules are far apart and don't have attractive forces (intermolecular forces) between them

5. 1 Pa < 1 mm Hg < 1 atm

6. atm

7. 1,190 mm Hg

8. 0.158 atm; 0.108 atm

9. 767.3 mm Hg

Answers

Exercise 12.1.1**: Converting Pressures**

$$9.65\ \text{atm} \times \frac{760\ \text{mm Hg}}{1\ \text{atm}} = 7,330\ \text{mm Hg}.$$

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12.2: Ideal Gas Law

Experimentation has shown that several 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*). What we find is that 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 Ideal Gas Law

The ideal gas law relates all the independent properties of a gas under any particular condition. This law can be represented mathematically as follows:

$$PV = nRT \tag{12.2.1}$$

In this equation, *P* is pressure, *V* is volume, *n* is amount of moles, and *T* is temperature. *R* is called the ideal gas law constant and is a proportionality constant that relates the values of pressure, volume, amount, and temperature of a gas sample. If three of the properties (plus the constant) are known, then the other property can be determined.

The value of *R* depends on what units are used to express the other quantities. If volume is expressed in liters and pressure in atmospheres, then the proper value of *R* is as follows:

$$\mathrm{R} = 0.08206\ \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}} \tag{12.2.2}$$

This may seem like a strange unit, but that is what is required for the units to work out algebraically.

At this point, the concept of temperature must be clarified. Although the Kelvin scale is the preferred temperature scale, the Celsius scale is also a common temperature scale used in science. The Celsius scale is based on the melting and boiling points of water and is actually the common temperature scale used by most countries around the world (except for the United States, which still uses the Fahrenheit scale). The value of a Celsius temperature is directly related to its Kelvin value by a simple expression:

$$\text{Kelvin temperature} = \text{Celsius temperature} + 273.15$$

$$T_K = T_{°C} + 273.15$$

Thus, it is easy to convert from one temperature scale to another.

> The Kelvin scale is sometimes referred to as the absolute scale because the zero point on the Kelvin scale is at absolute zero, the coldest possible temperature. On the other temperature scales, absolute zero is –260 °C or –459 °F.

Example 12.2.8

What is the volume in liters of 1.45 mol of N_2 gas at 298 K and 3.995 atm?

Solution

Using the ideal gas law where *P* = 3.995 atm, *n* = 1.45 mol, and *T* = 298 K,

$$PV = nRT$$

$$(3.995\ \mathrm{atm}) \times \mathrm{V} = (1.45\ \mathrm{mol}) \left(0.08205\ \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}\right) (298\ \mathrm{K})$$

On the right side, the moles and kelvins cancel. Also, because atmospheres appear in the numerator on both sides of the equation, they also cancel. The only remaining unit is liters, a unit of volume. So

$$3.995 \times V = (1.45)(0.08205)(298)\ \text{L}$$

Dividing both sides of the equation by 3.995 and evaluating, we get V = 8.87 L. Note that the conditions of the gas are not changing. Rather, the ideal gas law allows us to determine what the fourth property of a gas (e.g. volume) *must* be if three other properties (e.g. amount, pressure, and temperature) are known.

Exercise 12.2.8

What is the pressure of a sample of CO_2 gas if 0.557 mol is held in a 20.0 L container at 451 K?

Answer available at the end of this section.

Example 12.2.9

Cyclopropane (C_3H_6) is a gas that formerly was used as an anesthetic. How many moles of gas are there in a 100.0 L sample if the gas is at 0 °C and 1.00 atm?

Solution

Remember that °C can never be used as a unit when solving gas law problems. Therefore, we can start by converting the temperature to Kelvin.

$$0\ ^\circ\text{C} + 273.15 = 273.15\ \text{K}$$

Now we can set up the ideal gas law problem.

$$PV = nRT$$

$$(1.00\ \text{atm}) \times (100.0\ \text{L}) = \text{n} \left(0.08206\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (273.15\ \text{K})$$

There are almost 4.5 mol of gas in 100.0 L at this temperature and pressure.

Note: Because of its flammability, cyclopropane is no longer used as an anesthetic gas.

Exercise 12.2.9

Freon is a trade name for a series of fluorine- and chlorine-containing gases that formerly were used in refrigeration systems. What volume does 8.75 mol of Freon have at 0 °C and 1.00 atm?

Note: Many gases known as Freon are no longer used because their presence in the atmosphere destroys the ozone layer, which protects us from ultraviolet light from the sun.

Answer available at the end of this section.

Airbags (Figure 12.2.3) are a safety feature provided in most automobiles since the 1990s. The effective operation of an airbag requires that it be rapidly inflated with an appropriate amount (volume) of gas when the vehicle is involved in a collision. This requirement is satisfied in many automotive airbag systems through use of explosive chemical reactions, one common choice being the decomposition of sodium azide, NaN_3. When sensors in the vehicle detect a collision, an electrical current is passed through a carefully measured amount of NaN_3 to initiate its decomposition:

$$2\,\text{NaN}_3(s) \rightarrow 3\,\text{N}_2(g) + 2\,\text{Na}(s) \tag{12.2.3}$$

This reaction is very rapid, generating gaseous nitrogen that can deploy and fully inflate a typical airbag in a fraction of a second (~0.03–0.1 s). Among many engineering considerations, the amount of sodium azide used must be appropriate for generating enough nitrogen gas to fully inflate the air bag and ensure its proper function. For example, a small mass (~100 g) of NaN_3 will generate approximately 50 L of N_2.



Figure 12.2.3: Airbags deploy upon impact to minimize serious injuries to passengers. (credit: Jon Seidman)

Answers

Exercise 12.2.8

1.03 atm

Exercise 12.2.9

196 L

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12.3: The Combined Gas Law

Learning Objectives

- To predict the properties of gases using the gas laws.

In the previous section you read about the ideal gas law which can be used to determine a missing property of a gas when the other three properties are known. This is helpful when one variable cannot be easily measured, such as the pressure inside a container that does not have a pressure gauge.

In this section we will turn our focus to making predictions about gases that are changing. For instance, what is the new pressure of a gas when the temperature is increased? What is the new volume of a gas when the pressure is decreased?

A systematic way to learn relationships between properties of gases is to control as many variables as possible and to study how one variable is affected when another changes. For instance, study how the pressure of a gas changes when the temperature is increased or decreased (and other variables such as the amount of gas and volume are kept constant). With four different possible variables (pressure, volume, temperature, and amount of gas) there are several possible relationships to study. Each can be represented by a mathematical equation.

A more realistic scenario in which multiple properties are changing simultaneously can be represented by a mathematical equation that includes all of the variables that describe a gas. Such an equation is called the combined gas law because it is constructed from a combination of the individual gas laws (mathematical equations) that describe pairs of properties.

In this section you will learn to use the combined gas law equation, including how to use it as a replacement for all of the individual gas laws that describe how different pairs of properties change individually.

If the units of similar quantities are not the same, one of them must be converted to the other quantity's units for the calculation to work out properly. It does not matter which quantity is converted to a different unit; the only thing that matters is that the conversion and subsequent algebra are performed properly. The following example illustrates this process.

Combined Gas Law

Other gas laws can be constructed, but we will focus on one in this section, the combined gas law.

$$\frac{\mathrm{P_1V_1}}{\mathrm{n_1T_1}} = \frac{\mathrm{P_2V_2}}{\mathrm{n_2T_2}} \tag{12.3.1}$$

Notice that this equation contains copies of each variable (i.e two pressures, two volumes, etc.). The subscript 1 indicates an initial value (before a change); a subscript 2 indicates a final value (after a change). As with the ideal gas law, the temperature must be expressed in Kelvin. For the other variables any relevant unit may be used, but the units must be paired. In other words, both pressures must have the same unit, both volumes must have the same unit, etc.

Example 12.3.6

A sample of gas has P_1 = 1.50 atm, V_1 = 10.5 L, and T_1 = 300 K. What is the final volume if P_2 = 0.750 atm and T_2 = 350 K?

Solution

Using the combined gas law, substitute for five of the quantities. Since the amount of gas is not changing, n_1 and n_2 are the same and cancel out.

$$\frac{(1.50\ \mathrm{atm})(10.5\ \mathrm{L})}{300\ \mathrm{K}} = \frac{(0.750\ \mathrm{atm})(\mathrm{V_2})}{350\ \mathrm{K}}$$

We algebraically rearrange this expression to isolate V_2 on one side of the equation:

$$\mathrm{V_2} = \frac{(1.50\ \mathrm{atm})(10.5\ \mathrm{L})(350\ \mathrm{K})}{(300\ \mathrm{K})(0.750\ \mathrm{atm})} = 24.5\ \mathrm{L}$$

Note how all the units cancel except the unit for volume.

Exercise 12.3.6

A sample of gas has P_1 = 0.768 atm, V_1 = 10.5 L, and T_1 = 300 K. What is the final pressure if V_2 = 7.85 L and T_2 = 250 K?

Answer available at the end of this section.

Example 12.3.7

A balloon containing a sample of gas has a temperature of 22°C and a pressure of 1.09 atm in an airport in Cleveland. The balloon has a volume of 1,070 mL. The balloon is transported by plane to Denver, where the temperature is 11°C and the pressure is 655 torr. What is the new volume of the balloon?

Solution

The first task is to convert all quantities to the proper and consistent units. The temperatures must be expressed in kelvins, and the pressure units are different so one of the quantities must be converted. Let us convert the atmospheres to torr:

22°C + 273 = 295 K = T_1

11°C + 273 = 284 K = T_2

$$1.09\ \mathrm{atm} \times \frac{760\ \mathrm{torr}}{1\ \mathrm{atm}} = 828\ \mathrm{torr} = \mathrm{P_1}$$

Now we can substitute the quantities into the combined has law:

$$\frac{(828\ \mathrm{torr})(1,070\ \mathrm{mL})}{295\ \mathrm{K}} = \frac{(655\ \mathrm{torr}) \times \mathrm{V_f}}{284\ \mathrm{K}}$$

To solve for V_2, we multiply the 284 K in the denominator of the right side into the numerator on the left, and we divide 655 torr in the numerator of the right side into the denominator on the left:

$$\frac{(828\ \mathrm{torr})(1,070\ \mathrm{mL})(284\ \mathrm{K})}{(295\ \mathrm{K})(655\ \mathrm{torr})} = \mathrm{V_2}$$

Notice that torr and kelvins cancel, as they are found in both the numerator and denominator. The only unit that remains is milliliters, which is a unit of volume. So V_2 = 1,300 mL. The overall change is that the volume of the balloon has increased by 230 mL.

Exercise 12.3.7

A balloon used to lift weather instruments into the atmosphere contains gas having a volume of 1,150 L on the ground, where the pressure is 0.977 atm and the temperature is 18°C. Aloft, this gas has a pressure of 6.88 torr and a temperature of –15°C. What is the new volume of the gas?

Answer available at the end of this section.

Direct vs. Inverse Relationships

If only two properties of a gas are changing, it is usually possible to make a qualitative prediction about the size of the unknown variable. This can be very helpful when checking your work!

Direct Relationships

Experiments indicate that as the temperature of a gas sample is increased, its volume increases as long as the pressure and the amount of gas remain constant. Because volume increases as temperature increases, these two properties are directly proportional. If you are solving a problem in which only temperature and volume are changing, then you can check your work to make sure that volume increases when temperature increases and vice versa.

Example 12.3.4: Increasing Temperature

What happens to the volume of a gas if its temperature is decreased? Assume that all other conditions remain constant.

Solution

If the temperature of a gas sample is decreased, the volume decreases as well.

Exercise 12.3.4

What happens to the temperature of a gas if its volume is increased? Assume that all other conditions remain constant.

Answer available at the end of this section.

Example 12.3.5: Checking Your Work

A gas sample at 20°C has an initial volume of 20.0 L. What is its volume if the temperature is changed to 60°C? Does the answer make sense? Assume that the pressure and the amount of the gas remain constant.

Solution

Although the temperatures are given in degrees Celsius, we must convert them to the kelvins before we can use Charles's law. Thus,

$$20°C + 273 = 293\ K = T_1\ 60°C + 273 = 333\ K = T_2$$

Now we can substitute these values into Charles's law, along with the initial volume of 20.0 L:

$$\frac{20.0\ \mathrm{L}}{293\ \mathrm{K}} = \frac{\mathrm{V}_2}{333\ \mathrm{K}}$$

Multiplying the 333 K to the other side of the equation, we see that our temperature units will cancel:

$$\frac{(333\ \mathrm{K})(20.0\ \mathrm{L})}{293\ \mathrm{K}} = \mathrm{V}_2$$

Solving for the final volume, V_2 = 22.7 L. So, as the temperature is increased, the volume increases. This makes sense because volume is directly proportional to the absolute temperature (as long as the pressure and the amount of the remain constant).

Exercise 12.3.5: Making a Prediction

A gas sample at 35°C has an initial volume of 5.06 L. What is its volume if the temperature is changed to –35°C? First make a prediction: will the new volume be higher or lower than 5.06 L? Then calculate the answer. Assume that the pressure and the amount of the gas remain constant.

Answer available at the end of this section.

Inverse Relationships

If the amount of gas in a sample and its temperature are kept constant, then as the pressure of a gas is increased, the volume of the gas decreases proportionately. Because volume decreases as pressure increases, these two properties are inversely proportional. If you are solving a problem in which only pressure and volume are changing, then you can check your work to make sure that volume decreases when pressure increases and vice versa.

Example 12.3.1: Increasing Pressure in a Gas

What happens to the volume of a gas if its pressure is increased? Assume all other conditions remain the same.

Solution

If the pressure of a gas is increased, the volume decreases in response.

Exercise 12.3.1: Increasing Volume in a Gas

What happens to the pressure of a gas if its volume is increased? Assume all other conditions remain the same.

Answer available at the end of this section.

Example 12.3.2: Making a Prediction

If a sample of gas has an initial pressure of 1.56 atm and an initial volume of 7.02 L, what is the final volume if the pressure is reduced to 0.987 atm? Assume that the amount and the temperature of the gas remain constant. First predict whether the answer will be larger or smaller than 7.02 L, then perform the calculation to see if you were correct.

Solution

The key in problems like this is to be able to identify which quantities represent which variables from the relevant equation. The way the question is worded, you should be able to tell that 1.56 atm is P_1, 7.02 L is V_1, and 0.987 atm is P_2. What we are looking for is the final volume—V_2. Therefore, substituting these values into $P_1V_1 = P_2V_2$: (Because the amount of gas and temperature are not changing n_1 and n_2 cancel out and T_1 and T_2 cancel out)

$$(1.56\text{ atm})(7.02\text{ L}) = (0.987\text{ atm}) \times V_2$$

The expression has atmospheres on both sides of the equation, so they cancel algebraically:

$$(1.56)(7.02\text{ L}) = (0.987) \times V_2$$

Now we divide both sides of the expression by 0.987 to isolate V_2, the quantity we are seeking:

$$\frac{(1.56)(7.02\text{ L})}{0.987} = \text{V}_2$$

Performing the multiplication and division, we get the value of V_2, which is 11.1 L. The volume increases. This should make sense because the pressure decreases, so pressure and volume are inversely related.

Exercise 12.3.2

If a sample of gas has an initial pressure of 3.66 atm and an initial volume of 11.8 L, what is the final pressure if the volume is reduced to 5.09 L? Assume that the amount and the temperature of the gas remain constant.

Answer available at the end of this section.

Example 12.3.3: Checking Your Work

If a sample of gas has an initial pressure of 1.56 atm and an initial volume of 7.02 L, what is the final volume if the pressure is changed to 1,775 mm Hg? Does the answer make sense? Assume that the amount and the temperature of the gas remain constant.

Solution

This example is similar to Example 12.3.2, except now the final pressure is expressed in mm Hg. For the math to work out properly, one of the pressure values must be converted to the other unit. Let us change the initial pressure to mm Hg:

$$1.56\ \text{atm} \times \frac{760\ \text{mm Hg}}{1\ \text{atm}} = 1,190\ \text{mm Hg}$$

Now we can use Boyle's law:

$$(1{,}190\ \text{mm Hg})(7.02\ \text{L}) = (1{,}775\ \text{mm Hg}) \times V_2$$

Torr cancels algebraically from both sides of the equation, leaving

$$(1{,}190)(7.02\ \text{L}) = (1{,}775) \times V_2$$

Now we divide both sides of the equation by 1,775 to isolate V_2 on one side. Solving for the final volume,

$$\text{V}_2 = \frac{(1,190)(7.02\ \text{L})}{1,775} = 4.71\ \text{L}$$

Because the pressure increases, it makes sense that the volume decreases.

The answer for the final volume is essentially the same if we converted the 1,775 mm Hg to atmospheres: $1,775\ \text{mm} \times \frac{1\ \text{atm}}{760\ \text{mm Hg}} = 2.336\ \text{atm}$. Using Boyle's law: (1.56 atm)(7.02 L) = (2.335 atm) × V_2; $\text{V}_\text{f} = \frac{(1.56\ \text{atm})(7.02\ \text{L})}{2.336\ \text{atm}} = 4.69\ \text{L}$.

Exercise 12.3.3

If a sample of gas has an initial pressure of 375 torr and an initial volume of 7.02 L, what is the final pressure if the volume is changed to 4,577 mL? Does the answer make sense? Assume that amount and the temperature of the gas remain constant.

Answer available at the end of this section.

To Your Health: Breathing

Breathing certainly is a major contribution to your health! Without breathing, we could not survive. Curiously, the act of breathing itself is little more than an application of the inverse relationship between volume and pressure (called Boyle's law).

The lungs are a series of ever-narrowing tubes that end in a myriad of tiny sacs called alveoli. It is in the alveoli that oxygen from the air transfers to the bloodstream and carbon dioxide from the bloodstream transfers to the lungs for exhalation. For air to move in and out of the lungs, the pressure inside the lungs must change, forcing the lungs to change volume—just as predicted by Boyle's law.

The pressure change is caused by the diaphragm, a muscle that covers the bottom of the lungs. When the diaphragm moves down, it expands the size of our lungs. When this happens, the air pressure inside our lungs decreases slightly. This causes new air to rush in, and we inhale. The pressure decrease is slight—only 3 torr, or about 0.4% of an atmosphere. We inhale only 0.5–1.0 L of air per normal breath.

Exhaling air requires that we relax the diaphragm, which pushes against the lungs and slightly decreases the volume of the lungs. This slightly increases the pressure of the air in the lungs, and air is forced out; we exhale. Only 1–2 torr of extra pressure is needed to exhale. So with every breath, our own bodies are performing an experimental test of Boyle's law.

Determining Relationships

Think about the two relationships that were just discussed:

- Volume and temperature are directly related: when one increases, the other also increases
- Pressure and volume are inversely related: when one increases, the other decreases

Now take a look at the combined gas law and see if you can figure out whether relationships between other variables (e.g. pressure and temperature) will be direct or inverse.

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} \tag{12.3.2}$$

You can determine whether a pair of properties are directly or inversely related by looking at their relative positions in the mathematical equation. If two variables must be multiplied to obtain a constant, like P_1 and V_1, then they are inversely proportional. If two variables must be divided to obtain a constant, like V_1 and T_1, then they are directly proportional.

Concept Review Exercises

1. What properties do the gas laws help us predict?
2. What makes the ideal gas law different from the other gas laws?

Answers

1. Gas laws relate four properties: pressure, volume, temperature, and number of moles.
2. The ideal gas law does not require that the properties of a gas change.

Key Takeaway

- The physical properties of gases are predictable using mathematical formulas known as gas laws.

Exercises

1. What conditions of a gas sample should remain constant for Boyle's law to be used?

2. What conditions of a gas sample should remain constant for Charles's law to be used?

3. Does the identity of a gas matter when using Boyle's law? Why or why not?

4. Does the identity of a gas matter when using Charles's law? Why or why not?

5. A sample of nitrogen gas is confined to a balloon that has a volume of 1.88 L and a pressure of 1.334 atm. What will be the volume of the balloon if the pressure is changed to 0.662 atm? Assume that the temperature and the amount of the gas remain constant.

6. A sample of helium gas in a piston has a volume of 86.4 mL under a pressure of 447 torr. What will be the volume of the helium if the pressure on the piston is increased to 1,240 torr? Assume that the temperature and the amount of the gas remain constant.

7. If a gas has an initial pressure of 24,650 Pa and an initial volume of 376 mL, what is the final volume if the pressure of the gas is changed to 775 torr? Assume that the amount and the temperature of the gas remain constant.

8. A gas sample has an initial volume of 0.9550 L and an initial pressure of 564.5 torr. What would the final pressure of the gas be if the volume is changed to 587.0 mL? Assume that the amount and the temperature of the gas remain constant.

9. A person draws a normal breath of about 1.00 L. If the initial temperature of the air is 18°C and the air warms to 37°C, what is the new volume of the air? Assume that the pressure and amount of the gas remain constant.

10. A person draws a normal breath of about 1.00 L. If the initial temperature of the air is –10°C and the air warms to 37°C, what is the new volume of the air? Assume that the pressure and the amount of the gas remain constant.

11. An air/gas vapor mix in an automobile cylinder has an initial temperature of 450 K and a volume of 12.7 cm^3. The gas mix is heated to 565°C. If pressure and amount are held constant, what is the final volume of the gas in cubic centimeters?

12. Given the following conditions for a gas: V_i = 0.665 L, T_i = 23.6°C, V_f = 1.034 L. What is T_f in degrees Celsius and kelvins?

13. Assuming the amount remains the same, what must be the final volume of a gas that has an initial volume of 387 mL, an initial pressure of 456 torr, an initial temperature of 65.0°C, a final pressure of 1.00 atm, and a final temperature of 300 K?

14. When the nozzle of a spray can is depressed, 0.15 mL of gas expands to 0.44 mL, and its pressure drops from 788 torr to 1.00 atm. If the initial temperature of the gas is 22.0°C, what is the final temperature of the gas?

15. Use the ideal gas law to show that 1 mol of a gas at STP has a volume of about 22.4 L.

16. Use a standard conversion factor to determine a value of the ideal gas law constant R that has units of L•torr/mol•K.

17. How many moles of gas are there in a 27.6 L sample at 298 K and a pressure of 1.44 atm?

18. How many moles of gas are there in a 0.066 L sample at 298 K and a pressure of 0.154 atm?

19. A 0.334 mol sample of carbon dioxide gas is confined to a volume of 20.0 L and has a pressure of 0.555 atm. What is the temperature of the carbon dioxide in kelvins and degrees Celsius?

20. What must V be for a gas sample if n = 4.55 mol, P = 7.32 atm, and T = 285 K?

21. What is the pressure of 0.0456 mol of Ne gas contained in a 7.50 L volume at 29°C?

22. What is the pressure of 1.00 mol of Ar gas that has a volume of 843.0 mL and a temperature of –86.0°C?

Answers

1. temperature and amount of the gas

2. pressure and amount of the gas

3. The identity does not matter because the variables of Boyle's law do not identify the gas.
4. The identity does not matter because the variables of Charles law do not identify the gas.
5. 3.79 L

6. 31.1 mL

7. 92.1 mL

8. 918.4 torr

9. 1.07 L

10. 1.18 L

11. 23.7 cm^3

12. 461 K; 188°C

13. 206 mL

14. 835 K; 562°C

15. The ideal gas law confirms that 22.4 L equals 1 mol.

16. $\frac{760\ torr}{1\ atm}$

17. 1.63 mol

18. 4.2 x 10^{-4} mol

19. 405 K; 132°C

20. 14.5 L

21. 0.151 atm

22. 18.2 atm

Answers

Exercise 12.3.6
0.856 atm

Exercise 12.3.7
110,038 L

Exercise 12.3.4
The temperature increases.

Exercise 12.3.5**: Making a Prediction**
The final volume should be lower than 5.06 L because when temperature decreases volume also decreases. The final volume is 3.91 L.

Exercise 12.3.1**: Increasing Volume in a Gas**
If the volume of a gas is increased, the pressure decreases.

Exercise 12.3.2
8.48 atm

Exercise 12.3.3
575 torr. Since the volume decreased, it makes sense that the pressure increased because pressure and volume are inversely related.

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12.4: Dalton's Law

Dalton's Law of Partial Pressures

The ideal gas equation of state applies to mixtures just as to pure gases. It was in fact with a gas mixture, ordinary air, that Boyle, Gay-Lussac and Charles did their early experiments. The only new concept we need in order to deal with gas mixtures is the *partial pressure*, a concept invented by the famous English chemist John Dalton (1766-1844). Dalton reasoned that the low density and high compressibility of gases indicates that they consist mostly of empty space; from this it follows that when two or more different gases occupy the same volume, they behave entirely independently. The contribution that each component of a gaseous mixture makes to the total pressure of the gas is known as the *partial pressure* of that gas.

The definition of Dalton's Law of Partial Pressures that address this is:

The total pressure of a gas is the sum of the partial pressures of its components

which is expressed algebraically as

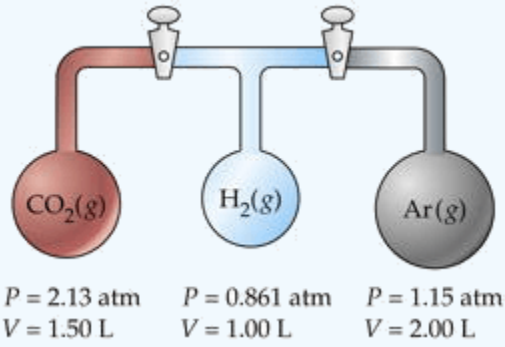
$$P_{total} = P_1 + P_2 + P_3... \tag{12.4.1}$$

Example 12.4.10

Three flasks having different volumes and containing different gases at various pressures are connected by stopcocks as shown. When the stopcocks are opened,

a. What will be the pressure in the system?
b. Which gas will be most abundant in the mixture?

Assume that the temperature is uniform and that the volume of the connecting tubes is negligible.



Solution

The trick here is to note that the total number of moles n_T and the temperature remain unchanged, so we can make use of Boyle's law PV = constant. We will work out the details for CO_2 only, denoted by subscripts a.

For CO_2,

$$P_aV_a = (2.13\ atm)(1.50\ L) = 3.19\ L \cdot atm \tag{12.4.2}$$

Adding the PV products for each separate container, we obtain

$$\sum_i P_iV_i = 6.36\ L \cdot atm = n_TRT \tag{12.4.3}$$

We will call this sum P_1V_1. After the stopcocks have been opened and the gases mix, the new conditions are denoted by P_2V_2.

From Boyle's law,

$$P_1V_1 = P_2V_2 = 6.36\ L \cdot atm \tag{12.4.4}$$

$$V_2 = \sum_i V_i = 4.50\ L \tag{12.4.5}$$

Solving for the final pressure P_2 we obtain (6.36 L-atm)/(4.50 L) = **1.41 atm**.

For part *(b)*, note that the number of moles of each gas is $n = PV/RT$. The mole fraction of any one gas is $X_i = n_i / n_T$. For CO_2, this works out to (3.19/*RT*) / (6.36/*RT*) = 0.501. Because this exceeds 0.5, we know that this is the most abundant gas in the final mixture.

Dalton's law states that in a gas mixture (P_{total}) each gas will exert a pressure independent of the other gases (P_n) and each gas will behave as if it alone occupies the total volume. By extension, the partial pressure of each gas can be calculated by multiplying the total pressure (P_{total}) by the gas percentage (%).

$$P_{Total} = P_1 + P_2 + P_3 + P_4 + ... + P_n \tag{12.4.6}$$

or

$$P_n = \frac{n}{P_{Total}} \tag{12.4.7}$$

Table 12.4.1: *Partial Pressures for the gases in air on a typical day*

Gas	Partial Pressure (mm Hg)	Percentage (%)
Nitrogen, N_2	P_{N_2} = 594	78
Oxygen, O_2	P_{O_2}= 160	21
Carbon Dioxide, CO_2	P_{CO_2} = 0.25	0.033
Water Vapor, H_2O	P_{H_2O} = 5.7	0.75
Other trace gases	P_{Other} = 0.05	0.22
Total air	P_{Total} = 760	1

Application of Dalton's Law: Collecting Gases over Water

A common laboratory method of collecting the gaseous product of a chemical reaction is to conduct it into an inverted tube or bottle filled with water, the opening of which is immersed in a larger container of water. This arrangement is called a *pneumatic trough*, and was widely used in the early days of chemistry. As the gas enters the bottle it displaces the water and becomes trapped in the upper part.

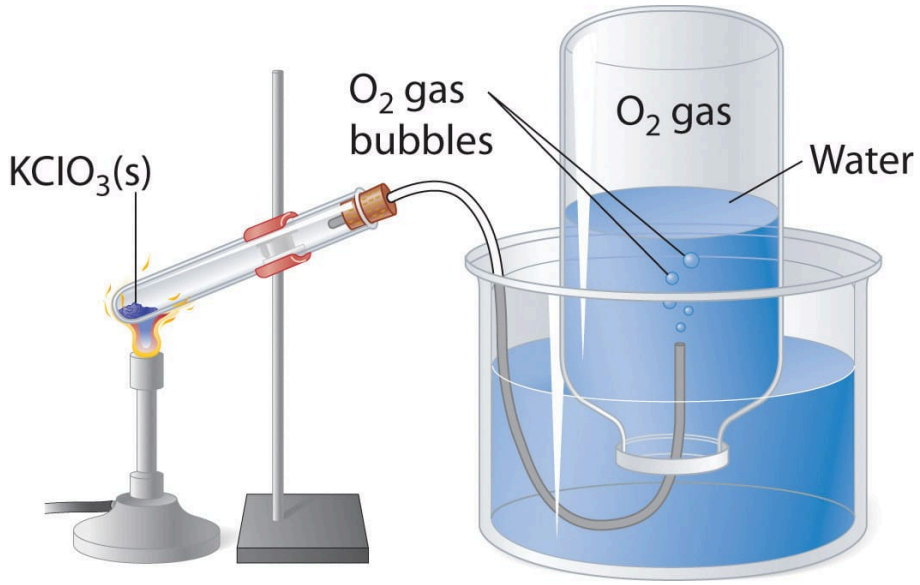


Figure 12.4.1:An Apparatus for Collecting Gases by the Displacement of Water

The volume of the gas can be observed by means of a calibrated scale on the bottle, but what about its pressure? The total pressure confining the gas is just that of the atmosphere transmitting its force through the water. (An exact calculation would also have to take into account the height of the water column in the inverted tube.) But liquid water itself is always in equilibrium with its vapor, so the space in the top of the tube is a mixture of two gases: the gas being collected, and gaseous H_2O. The partial pressure of H_2O is known as the vapor pressure of water and it depends on the temperature. In order to determine the quantity of gas we have collected, we must use Dalton's Law to find the partial pressure of that gas.

Example 12.4.11

Oxygen gas was collected over water as shown above. The atmospheric pressure was 754 torr, the temperature was 22°C, and the volume of the gas was 155 mL. The vapor pressure of water at 22°C is 19.8 torr. Use this information to estimate the number of moles of O_2 produced.

Solution

From Dalton's law,

$$P_{O_2} = P_{total} – P_{H_2O} = 754 – 19.8 = 734\ torr = 0.966\ atm$$

Now use the Ideal Gas Law to convert to moles

$$n = \frac{PV}{RT} = \frac{(0.966\ atm)(0.155\ L)}{(0.082\ Latmmol^{-1}K^{-1})(295\ K)} = 0.00619\ mol$$

Career Focus: Respiratory Therapist

Certain diseases—such as emphysema, lung cancer, and severe asthma—primarily affect the lungs. Respiratory therapists help patients with breathing-related problems. They can evaluate, help diagnose, and treat breathing disorders and even help provide emergency assistance in acute illness where breathing is compromised.

Most respiratory therapists must complete at least two years of college and earn an associate's degree, although therapists can assume more responsibility if they have a college degree. Therapists must also pass state or national certification exams. Once certified, respiratory therapists can work in hospitals, doctor's offices, nursing homes, or patient's homes. Therapists work with equipment such as oxygen tanks and respirators, may sometimes dispense medication to aid in breathing, perform tests, and educate patients in breathing exercises and other therapy.

Because respiratory therapists work directly with patients, the ability to work well with others is a must for this career. It is an important job because it deals with one of the most crucial functions of the body.

Exercises

1. A mixture of the gases N_2, O_2, and Ar has a total pressure of 760 mm Hg. If the partial pressure of N_2 is 220 mm Hg and of O_2 is 470 mm Hg, What is the partial pressure of Ar?
2. What percent of the gas above is Ar?

Answers

1. 70 mm Hg
2. 9.2%

Summary

P_{gas} is the partial pressure of the gas (often in units of Atm)

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13: APPENDIX- ALGEBRA REVIEW

13.1: Solving Equations- Addition and Subtraction
13.2: Solving Equations- Multiplication and Division

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13.1: Solving Equations- Addition and Subtraction

Learning Objectives

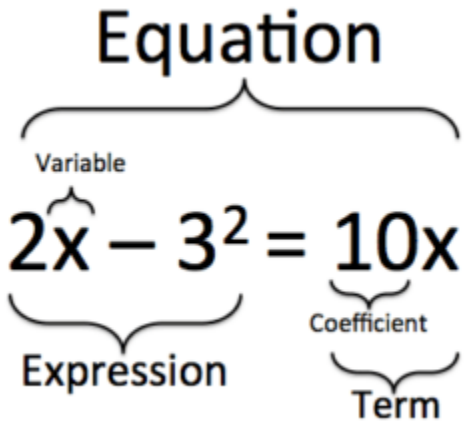
- Solve algebraic equations using the addition property of equality
- Solve algebraic equations using the multiplication property of equality

Terminology

First, let's define some important terminology:

- **variables:** variables are symbols that stand for an unknown quantity, they are often represented with letters, like *x*, *y*, or *z*.
- **coefficient:** Sometimes a variable is multiplied by a number. This number is called the coefficient of the variable. For example, in 3*x* the coefficient is 3.
- **expression:** groups of terms connected by addition and subtraction. $2x^2 - 5$ is an expression
- **equation:** an equation is a mathematical statement that two expressions are equal. An equation will always contain an equal sign with an expression on each side. Think of an equal sign as meaning "the same as." Some examples of equations are $\frac{3}{4}r = v^3 - r$, and $2(6 - d) + f(3 + k) = \frac{1}{4}d$

The following figure shows how coefficients, variables, terms, and expressions all come together to make equations. In the equation x, a coefficient is $10x$, an expression is $2x - 3^2$.



Equation made of coefficients, variables, terms and expressions.

Addition and Subtraction

An important property of equations is that you can add the same quantity to both sides of an equation and still maintain an equivalent equation. Sometimes people refer to this as keeping the equation "balanced." If you think of an equation as being like a balance scale, the quantities on each side of the equation are equal, or balanced.

Let's look at a simple numeric equation, $3 + 7 = 10$, to explore the idea of an equation as being balanced.



The expressions on each side of the equal sign are equal, so you can add the same value to each side and maintain the equality. Let's see what happens when 5 is added to each side.

$$3 + 7 + 5 = 10 + 5$$

Since each expression is equal to 15, you can see that adding 5 to each side of the original equation resulted in a true equation. The equation is still "balanced."

On the other hand, let's look at what would happen if you added 5 to only one side of the equation.

$$\begin{aligned} 3 + 7 &= 10 \\ 3 + 7 + 5 &= 10 \\ 15 &\neq 10 \end{aligned}$$

Adding 5 to only one side of the equation resulted in an equation that is false. The equation is no longer "balanced," and it is no longer a true equation!

Addition Property of Equality

For all real numbers *a*, *b*, and *c*: If $a + c = b + c$.

If two expressions are equal to each other, and you add the same value to both sides of the equation, the equation will remain equal.

Solve algebraic equations using the addition property of equality

When you solve an equation, you find the value of the variable that makes the equation true. In order to solve the equation, you **isolate the variable**. Isolating the variable means rewriting an equivalent equation in which the variable is on one side of the equation and everything else is on the other side of the equation.

When the equation involves addition or subtraction, use the inverse operation to "undo" the operation in order to isolate the variable. For addition and subtraction, your goal is to change any value being added or subtracted to 0, the additive identity.

In the following simulation, you can adjust the quantity being added or subtracted to each side of an equation to see how important it is to perform the same operation on both sides of an equation when you are solving.

Example 13.1.1

Solve $x - 6 = 8$.

Solution

This equation means that if you begin with some unknown number, *x*, and subtract 6, you will end up with 8. You are trying to figure out the value of the variable *x*.

Using the Addition Property of Equality, add 6 to both sides of the equation to isolate the variable. You choose to add 6 because 6 is being subtracted from the variable.

$$\begin{array}{rcr} x-6 & = & 8 \\ +6 & & +6 \\ \hline x & = & 14 \end{array}$$

Since subtraction can be written as addition (adding the opposite), the **addition property of equality** can be used for subtraction as well. So just as you can add the same value to each side of an equation without changing the meaning of the equation, you can subtract the same value from each side of an equation.

Example 13.1.2

Solve $x+10=-65$. Check your solution.

Solution

$$x+10=-65$$

Since 10 is being added to the variable, subtract 10 from both sides. Note that subtracting 10 is the same as adding -10.

$$\begin{array}{rcr} x+10 & = & -65 \\ -10 & & -10 \\ \hline x & = & -75 \end{array}$$

To check, substitute the solution, -75 for *x* in the original equation, then simplify.

$$\begin{aligned} x+10 &= -65 \\ -75+10 &= -65 \\ -65 &= -65 \end{aligned}$$

This equation is true, so the solution is correct.

Exercise 13.1.1

Solve $x+5=27$.

Answer available at the end of this section.

Exercise 13.1.2

Solve $x-4=-32$. Check your solution.

Answer available at the end of this section.

Exercise 13.1.3

Solve $12.5+t=-7.5$.

Answer available at the end of this section.

The examples above are sometimes called **one-step equations** because they require only one step to solve. In these examples, you either added or subtracted a **constant** from both sides of the equation to isolate the variable and solve the equation.

With any equation, you can check your solution by substituting the value for the variable in the original equation. In other words, you evaluate the original equation using your solution. If you get a true statement, then your solution is correct.

Writing and solving algebraic equations is an important part of mathematics. Equations can be used to describe economic, cultural, physical, and biological processes. They help business people make decisions, and help doctors and scientists find ways to heal and help people. Without mathematical equations we would not have the physical infrastructure that rely on every day for transportation and clean water.

Equations can help you model situations and solve problems in which quantities are unknown (like how long Joan should wait before she drives home). The simplest type of algebraic equation is a linear equation that has just one variable.

When you follow the steps to solve an equation, you try to isolate the variable. The variable is a quantity we don't know yet. You have a solution when you get the equation *x* = some value.

Exercise 13.1.4: Chemistry connection

In chemistry you may use the equation $T_K = T_{^\circ C} + 273.15$. If $T_K = 293.65$, what is the value of $T_{^\circ C}$?

Answer available at the end of this section.

Exercise 13.1.5: Chemistry connection

Another equation you may use in chemistry is 14 = pH + pOH. Rearrange this equation to solve for pOH.

Answer available at the end of this section.

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Answers

Exercise 13.1.1

This equation means that if you begin with some unknown number, *x*, and add 5, you will end up with 27. You are trying to figure out the value of the variable *x*.

Using the Addition Property of Equality, subtract 5 from both sides of the equation to isolate the variable. You choose to subtract 5, as 5 is being added from the variable.

$$\begin{aligned} x+5 &= 27 \\ \underline{-5} &\quad \underline{-5} \\ x &= 22 \end{aligned}$$

Exercise 13.1.2

$$x-4=-32$$

Since 4 is being subtracted from the variable, add 4 to both sides.

$$\begin{aligned} x-4 &= -32 \\ \underline{+4} &\quad \underline{+4} \\ x &= -28 \end{aligned}$$

Check:

To check, substitute the solution, -28 for *x* in the original equation, then simplify.

$$\begin{aligned} x-4 &= -32 \\ -28-4 &= -32 \\ -32 &= -32 \end{aligned}$$

This equation is true, so the solution is correct. It is always a good idea to check your answer whether you are requested to or not.

Exercise 13.1.3

To solve this equation you need to remember how to add or subtract decimal numbers. You also need to remember that when you subtract a number from a negative number, your result will be negative.

Using the Addition Property of Equality, subtract 12.5 from both sides of the equation to isolate the variable, *t*. You choose to subtract 12.5 because 12.5 is being added to the variable, *t*.

$$\begin{aligned} 12.5+t &= -7.5 \\ \underline{-12.5} &\quad \underline{-12.5} \\ t &= -20 \end{aligned}$$

Exercise 13.1.4**: Chemistry connection**

$20.5\ ^{\circ}C$

Exercise 13.1.5**: Chemistry connection**

When asked to solve for pOH, that means to get it by itself. You need to move the pH away from pOH by doing the opposite of the current operation. In the given equation pH is being added to pOH. The opposite of addition is subtraction. Remember to do the same thing to both sides of the equation; subtract pH from both sides.

$$14-pH=pH+pOH-pH$$

The pH's on the right cancel leaving

$$14-pH=pOH$$

If you were given a value for pH you could plug it into the answer above to solve for pOH.

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13.2: Solving Equations- Multiplication and Division

Just as you can add or subtract the same exact quantity on both sides of an equation, you can also multiply or divide both sides of an equation by the same quantity to write an equivalent equation. Let's look at a numeric equation, $5 \cdot 3 = 15$, to start. If you multiply both sides of this equation by 2, you will still have a true equation.

$$\begin{aligned} 5 \cdot 3 &= 15 \\ 2 \cdot 5 \cdot 3 &= 15 \cdot 2 \\ 30 &= 30 \end{aligned}$$

This characteristic of equations is generalized in the **multiplication property of equality**.

Multiplication Property of Equality

For all real numbers *a*, *b*, and *c*: If *a* = *b*, then $a \cdot c = b \cdot c$ (or *ac* = *bc*).

If two expressions are equal to each other and you multiply both sides by the same number, the resulting expressions will also be equivalent.

When the equation involves multiplication or division, you can "undo" these operations by using the inverse operation to isolate the variable. When the operation is multiplication or division, your goal is to change the coefficient to 1, the multiplicative identity.

Example 13.2.1

Solve $3x = 24$. When you are done, check your solution.

Solution

Divide both sides of the equation by 3 to isolate the variable (have a coefficient of 1). Dividing by 3 is the same as having multiplied by $\frac{1}{3}$.

$$\begin{aligned} \frac{3x}{3} &= \frac{24}{3} \\ x &= 8 \end{aligned}$$

Check by substituting your solution, 8, for the variable in the original equation.

$$\begin{aligned} 3x &= 24 \\ 3 \cdot 8 &= 24 \\ 24 &= 24 \end{aligned}$$

The solution is correct!

You can also multiply the coefficient by the multiplicative inverse (reciprocal) in order to change the coefficient to 1.

Example 13.2.2

Solve $\frac{1}{2}x = 8$

Solution

Multiply both sides by 2 to get x by itself:

$$(2)\tfrac{1}{2}x = 8(2)$$

On the left side the 2's cancel leaving x by itself. On the right side (8)(2) = 16

$$x = 16$$

In the next example, we will solve a one-step equation using the multiplication property of equality. You will see that the variable is part of a fraction in the given equation, and using the multiplication property of equality allows us to remove the variable from the fraction. Remember that fractions imply division, so you can think of this as the variable *k* is being divided by 10. To "undo" the division, you can use multiplication to isolate *k*. Lastly, note that there is a negative term in the equation, so it will be important to think about the sign of each term as you work through the problem. Stop after each step you take to make sure all the terms have the correct sign.

Example 13.2.3

Solve $\frac{7}{2} = \frac{k}{10}$ for *k*.

Solution

We want to isolate the *k*, which is being divided by 10. The first thing we should do is multiply both sides by 10.

$$\frac{10 \cdot 7}{2} = \frac{k \cdot 10}{10}$$

On the right side, the 10's divide out. Solve the left side to get the value for k.

$$\frac{10 \cdot 7}{2} = k$$

$$k = 35$$

Exercise 13.2.1: Chemistry Connection

A common equation in an introductory chemistry course is $d = \frac{m}{V}$. Rearrange this equation to solve for m.

Answer available at the end of this section.

Exercise 13.2.2: Chemistry Connection

Students often find it more difficult to solve for a value in the denominator. Rearrange the density equation from the previous question to solve for V.

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

Tip: You must ensure that the variable you are solving for is in the numerator, not the denominator, so start by multiplying both sides by V. Then rearrange the equation to get V by itself.

Answer available at the end of this section.

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Answers

Exercise 13.2.1**: Chemistry Connection**

"Solve for m" means to get the variable m by itself. To do so we need to move V to the other side of the equation using the opposite of the current operation. In the equation above m is divided by V so we will multiply both sides by V.

$$V \cdot d = \frac{m \cdot V}{V}$$

On the right side the V's cancel out leaving

$$V \cdot d = m$$

If you had values for the density and volume you could plug them into the answer to solve for the mass of a substance.

Exercise 13.2.2**: Chemistry Connection**

Since the requested variable is in the denominator, start by multiplying both sides by V to get it into the numerator.

$$V \cdot d = \frac{m \cdot V}{V}$$

Now that V is in the numerator we can proceed as usual to get it by itself. V is being multiplied by d. The opposite of multiplying by d is dividing by d, so divide both sides by d in order to get V by itself.

$$\frac{V \cdot d}{d} = \frac{m}{d}$$
$$V = \frac{m}{d}$$

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GLOSSARY

A

Alkali metal | A metal in Group 1 of the periodic table

Alkaline earth metal | A metal in Group 2 of the periodic table

Amino acid | A biomolecule consisting of an amino group (amine) and a carboxylic acid. The building block of proteins.

Anion | A negative ion

Aqueous | Dissolved in water

Atom | The smallest particle of an element that retains the properties of the element

Atomic mass | The average mass of all isotopes of an element. Note: The atomic mass is usually listed on the periodic table.

Atomic number | The number of protons in an atom. Note: The atomic number is usually listed on the periodic table.

C

Cation | A positive ion

Chemical change | A change in which the identity of the substance changes due to chemical bonds breaking and/or forming. Also called a chemical reaction.

Chemical equation | A written description of a chemical reaction using symbols and formulas.

Chemical formula | A description of a compound using chemical symbols and subscripts

Coefficient | (in chemistry) A number written before a chemical formula to indicate the number of copies needed to balance a chemical equation.

Compound | A pure substance composed of two or more elements bonded together in a fixed ratio.

Covalent bond | An attraction between atoms due to the sharing of valence electrons.

E

Electron | A small, negatively charged particle found around the nucleus of an atom.

Element | A substance made of only one type of atom. Note: Elements are listed on the periodic table.

F

Fatty acid | A biomolecule consisting of a carboxylic acid functional group with a long (~10 - 20 C atom) chain.

Formula unit | The lowest ratio of elements in an ionic compound.

H

Halogen | A nonmetal in Group 17 of the periodic table

Heterogeneous mixture | A mixture in which different components are visible, e.g. in chunks or layers

Homogeneous mixture | A mixture in which different components are not visible. The mixture looks the same throughout. Also called a solution.

I

Inner transition metal | Metals found in the two rows at the bottom of the periodic table. Also called lanthanides (top row) and actinides (bottom row)

Inorganic compound | A compound that does not contain any C-H bonds.

Ion | A charged particle

Ionic compound | A compound formed by the attraction between cations and anions (ionic bond).

Isotope | A form of an element with a specific number of neutrons.

L

Law | Something that has been observed over and over again in experiments. Helpful in making predictions.

M

Mass number | The total number of protons and neutrons in a particular isotope of an element. Note: The mass number is NOT listed on the periodic table.

Molecular compound | A compound consisting of molecules in which the atoms are held together by covalent bonds.

Molecule | Two or more atoms covalently bonded together.

Monomer | A small molecule that bonds to other small molecules to make a polymer.

Monosaccharide | The building block (monomer) of carbohydrates.

N

Neutron | An uncharged particle found in the nucleus of an atom.

Noble gas | A nonmetal in Group 18 of the periodic table

Nucleic acid | A type of biomolecule that stores and processes genetic information. RNA and DNA are nucleic acids.

Nucleobase | The bases in nucleic acids. Their structure consists of 1 - 2 rings with several nitrogen atoms.

Nucleoside | A component of nucleic acids. Consists of a nucleobase bonded to a monosaccharide.

Nucleotide | The building block of nucleic acids. Consists of a nucleobase bonded to a monosaccharide bonded to a phosphate group.

O

Octet rule | Atoms tend to gain, lose, or share valence electrons in order to get to a total of eight valence electrons.

Organic compound | A compound containing carbon and hydrogen bonded together. It may also contain other elements.

P

Peptide | A biomolecule consisting of several amino acids (~2 - 100), but shorter than a protein

Periodic table | A list of all known elements

Phospholipid | A biomolecule with a polar head and two long nonpolar tails.

Physical change | A change in which the identity of the substance stays constant, i.e. it can still be represented by the same symbol or formula after the change.

Polymer | A large molecule with a repeating structure. May be natural, e.g. protein and DNA, or synthetic, e.g. plastics.

Precipitate | A solid that forms from solutions.

Precipitation reaction | A reaction between two aqueous solutions that results in the formation of a precipitate

Product | The result of a chemical reaction. Written after the arrow in a chemical equation.

Proton | A positively charged particle in the nucleus of an atom.

R

Reactant | A substance present before a chemical reaction occurs. Written before the arrow in a chemical equation.

S

SI | A set of units that scientists and many countries use. Can be used with metric prefixes.

Subscript | (in chemistry) A number written lower than the normal text indicating the number of copies of an atom or ion in a molecule or formula unit.

T

Theory | An explanation developed after a lot of experimentation.

Transition metal | A metal located in Groups 3 - 12 of the periodic table

Triacylglycerol | A biomolecule formed from glycerol and three fatty acids. Also known as a triglyceride.

V

Valence electron | An electron in the outermost shell (layer) of an atom

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