

CHEM 1020 INTRODUCTORY CHEMISTRY



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CHEM 1020 Textbook Readings

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CHAPTER OVERVIEW

1: Module 1 "Start Seeing Chemistry!" Textbook Readings

- 1.1: Atoms - Ideas from the Ancient Greeks
- 1.2: The Scientific Method
- 1.3: Classification of Matter
- 1.4: Chemistry - A Study of Matter and Its Changes

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1.1: Atoms - Ideas from the Ancient Greeks

Learning Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus to atomic theory.

You learned earlier that all matter in the universe is made out of tiny building blocks called atoms; all modern scientists accept the concept of the atom. However, when the concept of the atom was first proposed about 2,500 years ago, ancient philosophers laughed at the idea. It has always been difficult to convince people of the existence of things that are too small to see. We will spend some time considering the evidence (observations) that convince scientists of the existence of atoms.

Democritus and the Greek Philosophers

About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge entity. In other words, "everything was one." They believed that all objects, all matter, and all substances were connected as a single, unchangeable "thing." One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that **atomos**, or atomon—tiny, indivisible, solid objects—make up all matter in the universe.



Figure 1.1.1 (Top) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (Bottom) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

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

Summary

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called "atomos." However, other Greek philosophers disliked Democritus' "atomos" theory because they felt it was illogical.

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1.2: The Scientific Method

Learning Objectives

- Describe the differences between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.
- Identify the components of the scientific method.

Although many have taken science classes throughout their course of studies, incorrect or misleading ideas about some of the most important and basic principles in science are still commonplace. Most students have heard of *hypotheses*, *theories*, and *laws*, but what do these terms really mean? Before you read this section, consider what you have learned about these terms previously, and what they mean to you. When reading, notice if any of the text contradicts what you previously thought. What do you read that supports what you thought?

What is a Fact?

A **fact** is a basic statement established by experiment or observation. All facts are true under the specific conditions of the observation.

What is a Hypothesis?

One of the most common terms used in science classes is a "**hypothesis**". The word can have many different definitions, dependent on the context in which it is being used:

- An educated guess: a scientific hypothesis provides a suggested solution based on evidence.
- Prediction: if you have ever carried out a science experiment, you probably made this type of hypothesis, in which you predicted the outcome of your experiment.
- Tentative or proposed explanation: hypotheses can be suggestions about why something is observed. In order for a hypothesis to be scientific, a scientist must be able to test the explanation to see if it works, and if it is able to correctly predict what will happen in a situation. For example, "if my hypothesis is correct, I should see _____ result when I perform _____ test."

A hypothesis is tentative; it can be easily changed.

What is a Theory?

The *United States National Academy of Sciences* describes a theory as:

"Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."

"A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses," but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact."

Note some key features of theories that are important to understand from this description:

- Theories are explanations of natural phenomenon. They aren't predictions (although we may use theories to make predictions). They are explanations of why something is observed.
- Theories aren't likely to change. They have a lot of support and are able to explain many observations satisfactorily. Theories can, indeed, be facts. Theories can change in some instances, but it is a long and difficult process. In order for a theory to change, there must be many observations or evidence that the theory cannot explain.

- Theories are not guesses. The phrase "just a theory" has no room in science. To be a scientific theory carries a lot of weight—it is not just one person's idea about something

Theories aren't likely to change.

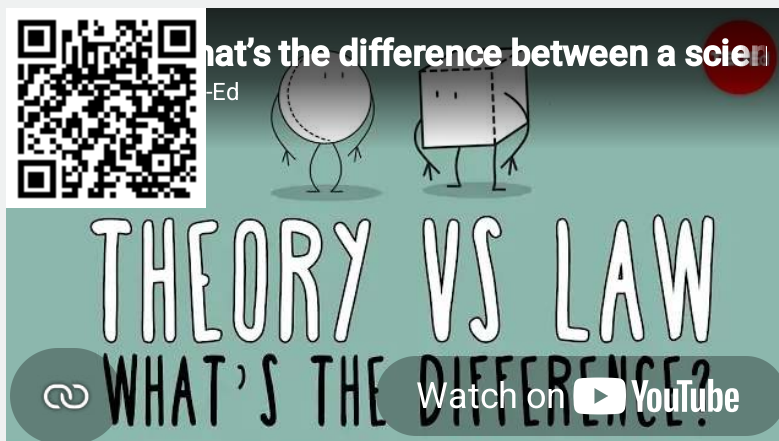
What is a Law?

Scientific **laws** are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are typically well-supported by observations and/or experimental evidence. Usually, scientific laws refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories are overarching explanations of how nature works, and why it exhibits certain characteristics. As a comparison, theories explain why we observe what we do, and laws describe what happens.

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after numerous tests, that certain patterns existed in their observations of gas behavior. If the temperature of the gas increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do not describe why the patterns exist.

📌 Laws vs Theories

A common misconception is that scientific theories are rudimentary ideas that will eventually graduate into scientific laws when enough data and evidence has been accumulated. A theory does not change into a scientific law with the accumulation of new or better evidence. Remember, theories are explanations; laws are patterns seen in large amounts of data, frequently written as an equation. A theory will always remain a theory, a law will always remain a law.



Video 1.2.1 What is the difference between scientific law and theory?

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 1.2.1).

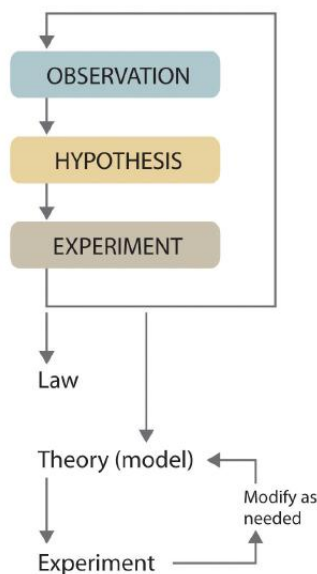


Figure 1.2.1 *The steps in the scientific method.*

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—the chemical name of which 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 which correspond to observed movements of the sun, moon, clouds, and shadows, is consistent with either of two hypotheses:

1. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun.
2. The sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. In the case of 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 supports or refutes 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 the case of validity, the scientist can proceed to step 5. In other cases, experiments may demonstrate that the hypothesis is incorrect or that it

must be modified, thus requiring further experimentation.

Step 5: Development of 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 states what happens; it does not address the question of why.

One example of a law, the law of definite proportions (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. 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. A theory, in contrast, is incomplete and imperfect; it evolves with time to explain new facts as they are discovered.

Because scientists can enter the cycle shown in Figure 1.2.1 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.

✓ Example 1.2.1

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

1. Ice always floats on liquid water.
2. Birds evolved from dinosaurs.
3. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
4. When 10 g of ice was added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
5. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution

1. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
2. This is a possible explanation for the origin of birds, so it is a hypothesis.
3. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.
4. The temperature is measured before and after a change is made in a system, so these are observations.
5. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

? Exercise 1.2.1

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

1. Measured amounts of acid were added to a Roloids tablet to see whether it really “consumes 47 times its weight in excess stomach acid.”
2. Heat always flows from hot objects to cooler ones, not in the opposite direction.
3. The universe was formed by a massive explosion that propelled matter into a vacuum.
4. Michael Jordan is the greatest pure shooter ever to play professional basketball.
5. Limestone is relatively insoluble in water, but dissolves readily in dilute acid with the evolution of a gas.

Answer 1:

experiment

Answer 2:

law

Answer 3:

theory

Answer 4:

hypothesis

Answer 5:

observation

Summary

- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is a well-supported explanation of observations.
- A scientific law is a statement that summarizes the relationship between variables.
- An experiment is a controlled method of testing a hypothesis.
- 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 are:
 - 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 of a law and/or theory.

Contributors and Attributions

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1.3: Classification of Matter

Learning Objectives

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

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 **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small or large amount of salt 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 substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

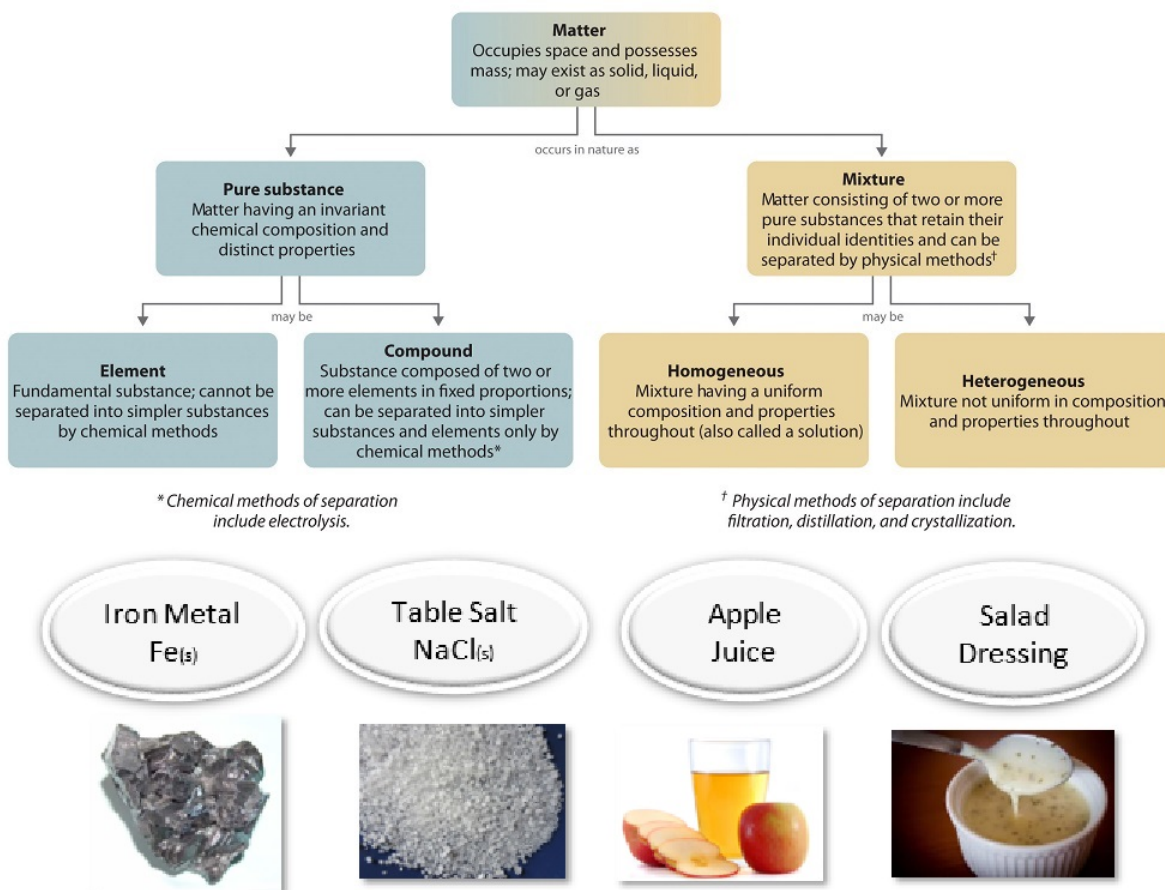


Figure 1.3.3: Relationships between the types of matter and the methods used to separate mixtures.

Atoms, Elements, and Compounds

Elements, Atoms, Molecules, Ions, Ions

Organic Chemistry Tutor

Watch on YouTube

Atom

Molecule

Compound

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

Phase

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

✓ Example 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:

homogeneous mixture (solution)

Answer b:

element

Answer c:

heterogeneous mixture

Answer d:

compound

✓ 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 substance; more specifically, because water is composed of hydrogen and oxygen, it is a compound.
- d. Oxygen, a substance, is an element.

? Exercise 1.3.3

How would a chemist categorize each example of matter?

- a. filtered coffee
- b. hydrogen
- c. an egg

Answer a:

a homogeneous mixture (solution)

Answer b:

element

Answer c:

heterogeneous mixture.

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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1.4: Chemistry - A Study of Matter and Its Changes

Learning Objectives

- Separate physical from chemical properties.
- Label a change as chemical or physical.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—the screen on which you are reading this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter. Because of this, chemistry is a science that has its fingers in just about everything. Even a description of the ingredients in a cake, and how those ingredients change when the cake is baked, is chemistry!

Physical and Chemical Properties

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

Physical Property

A **physical property** is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets—a property called malleability. Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point. In Table 1.4.1, notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will “float.”

Table 1.4.1 Densities of Common Substances

Substance	Density at 25°C (g/cm ³)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

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



Figure 1.4.1 A pencil (top) and a diamond ring (bottom). Both are a form of carbon, but exhibit very different physical properties.

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

Chemical Property

A **chemical property** of matter describes its "potential" to undergo some chemical change or reaction by virtue of its composition; as well as what elements, electrons, and bonding are present to give the potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, students of chemistry should be able to look at the formula of a compound and state a chemical property. For example, hydrogen has the potential to ignite and explode, given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with acids. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.

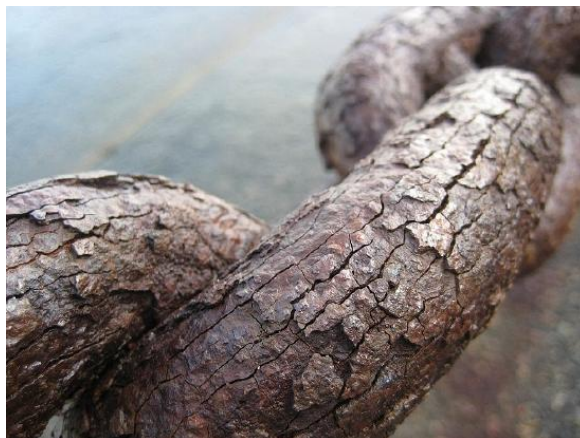


Figure 1.4.2 Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0;).

A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust (Figure 1.4.2). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are *burn*, *rot*, *explode*, *decompose*, and *ferment*. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 1.4.2: Contrasting Physical and Chemical Properties

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

✓ Example 1.4.1

Which of the following is a chemical property of iron?

1. Iron corrodes in moist air.
2. Density = 7.874 g/cm^3
3. Iron is soft when pure.
4. Iron melts at 1808 K.

Solution

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

? Exercise 1.4.1A

Which of the following is a physical property of matter?

1. corrosiveness
2. pH (acidity)
3. density
4. flammability

Answer

3.

? Exercise 1.4.1B

Which of the following is a chemical property?

1. flammability
2. melting point
3. boiling point
4. density

Answer

1.

Physical and Chemical Changes

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

Physical Change

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

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

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



Figure 1.4.3: Ice melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; However, this change is only physical, as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

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

Chemical Change

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

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



Figure 1.4.4: The burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

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

Observations that help to indicate chemical change include:

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

✓ Example 1.4.2

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

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

Solution

1. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written as: $\text{H}_2\text{O} (l) \rightarrow \text{H}_2\text{O} (g)$
2. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
3. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been combined. No color *change*, or other evidence of chemical change, occurred.
4. Chemical: the formation of a precipitate and the color change from colorless to yellow indicates a chemical change.

? Exercise 1.4.2

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

1. A mirror is broken.
2. An iron nail corroded in moist air.

3. Copper metal is melted.
4. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer 1:

physical change

Answer 2:

chemical change

Answer 3:

physical change

Answer 4:

chemical change

Summary

- A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points.
- A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change.
- Physical changes are changes that do not alter the identity of a substance.
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes. Observations that indicate a chemical change occurred include color change, temperature change, light given off, formation of bubbles, and formation of a precipitate.

Contributors and Attributions

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CHAPTER OVERVIEW

2: Module 2 "Measurement" Textbook Readings

2.1: The Measurement of Matter

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2.1: The Measurement of Matter

Learning Objectives

- Express quantities properly, using a number and a unit.
- State the different measurement systems used in chemistry.
- Express a large number or a small number in scientific notation.
- Learn how to use SI prefixes.

A coffee maker's instructions tell you to fill the coffeepot with 4 cups of water and use 3 scoops of coffee. When you follow these instructions, you are measuring. When you visit a doctor's office, a nurse checks your temperature, height, weight, and perhaps blood pressure (Figure 2.1.1); the nurse is also measuring.



Figure 2.1.1: A nurse or a doctor measuring a patient's blood pressure is taking a measurement. Figure used with permission (GFDL; Pia von Lützu).

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

To understand chemistry, we need a clear understanding of the units chemists work with and the rules they follow for expressing numbers. The next two sections examine the rules for expressing numbers.

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. The **International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of the seven base units shown in Table 2.1.1.

Table 2.1.1 SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

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

Exponential Numbers: Powers of Ten

Chemists often work with numbers that are exceedingly large or small. For example, entering the mass in grams of a hydrogen atom into a calculator would require a display with at least 24 decimal places. A system called **scientific notation** avoids much of the tedium and awkwardness of manipulating numbers with large or small magnitudes. Furthermore, use of prefixes is another way to express measurements involving large and small numbers.

Scientific Notation

In scientific notation, numbers are expressed in the form

$$N \times 10^n$$

where N is greater than or equal to 1 and less than 10 ($1 \leq N < 10$), and n is a positive or negative integer ($10^0 = 1$). The number 10 is called the base because it is this number that is raised to the power n . Although a base number may have values other than 10, the base number in scientific notation is always 10.

A simple way to convert numbers to scientific notation is to move the decimal point as many places to the left or right as needed to give a number from 1 to 10 (N). The magnitude of n is then determined as follows:

- If the decimal point is moved to the left n places, n is positive.
- If the decimal point is moved to the right n places, n is negative.

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example 2.1.1.

✓ Example 2.1.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

- 637.8
- 0.0479
- 12,378
- 0.00032

Solution

Example 2.1.1 Solution and Explanation

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, $n = 2$.	6.378×10^2
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, $n = -2$.	4.79×10^{-2}
c	Because the decimal point was moved four places to the left, $n = 4$.	1.2378×10^4
d	Because the decimal point was moved four places to the right, $n = -4$.	3.2×10^{-4}

? Exercise 2.1.1

Convert each ordinary number to scientific notation, or vice versa.

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

Answer a

$$6.7 \times 10^{10}$$

Answer b

$$1.689 \times 10^3$$

Answer c

$$1.26 \times 10^1$$

? Exercise 2.1.2

Convert each ordinary number to scientific notation, or vice versa.

- a. 0.000006567
- b. 6.22×10^{-2}
- c. 9.9×10^{-9}

Answer a

$$6.567 \times 10^{-6}$$

Answer b

$$0.0622$$

Answer c

$$0.0000000099$$

Metric Prefixes

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word *μεγας*, meaning "great". Table 2.1.2 lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 2.1.2 SI Prefixes

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

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

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

✓ Example 2.1.2: Unit Abbreviations

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

- kiloliter
- microsecond
- decimeter
- nanogram

Solutions

Example 2.1.2 Solution and Explanation

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

? Exercise 2.1.3

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

- kilometer
- milligram
- nanosecond
- centiliter

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The mass of an object is made in comparison to the standard mass of 1 kilogram. The kilogram was originally defined as the mass of 1 L of liquid water at 4°C (volume of a liquid changes slightly with temperature). In the laboratory, mass is measured with a balance (see below), which must be calibrated with a standard mass so that its measurements are accurate.



Figure 2.1.2 (From left to right:) Triple beam balance ([source\(opens in new window\)](#)), digital top-loading balance ([source\(opens in new window\)](#)), and an analytical balance ([source](#)).

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are 1000 g in 1 kg. A milligram is 1/1000th of a gram, so there are 1000 mg in 1 g.

Mass is often confused with the term weight. **Weight** is a measure of force that is equal to the gravitational pull on an object. The weight of an object is dependent on its location. On the moon, the force due to **gravity** is about one sixth that of the gravitational force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale, by reading the distance that a certain object pulls down and stretches a spring.

Length and Volume

Length is the measurement of the extent of something along its greatest dimension. The SI basic unit of length, or linear measure, is the **meter** (m). All measurements of length may be made in meters, though the prefixes listed in various tables will often be more convenient. The width of a room may be expressed as about 5 meters (m), whereas a large distance, such as the distance between New York City and Chicago, is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 10 micrometers (μm).

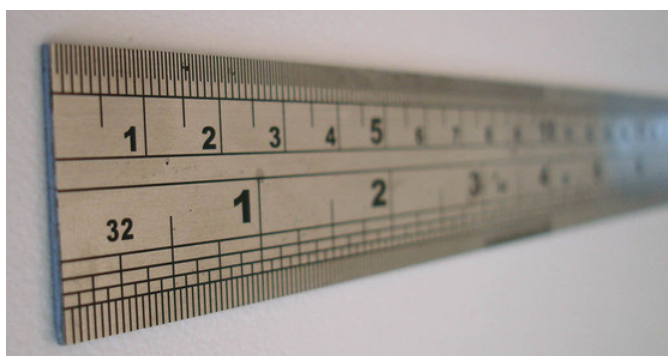


Figure 2.1.3: Metric and English measuring devices. (CC BY-SA 4.0; Ejay)

Volume is the amount of space occupied by a sample of matter. The volume of a regular object can be calculated by multiplying its length by its width and its height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter (m^3), which is the volume occupied by a cube that measures 1 m on each side. This very large volume is not very convenient for typical use in a chemistry laboratory. A liter (L) is the volume of a cube that measures 10 cm (1 dm) on each side. A liter is thus equal to both 1000 cm^3 ($10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm}$) and to 1 dm^3 .

$$1\text{L} = 1000 \text{ cm}^3 (10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm}) = 1 \text{ dm}^3$$

A smaller unit of volume that is commonly used is the milliliter (mL). A milliliter is the volume of a cube that measures 1 cm on each side. Therefore, a milliliter is equal to a cubic centimeter (cm^3).

$$1 \text{ mL} (\text{cm}^3)$$

There are 1000 mL in 1 L, which is the same as saying that there are 1000 cm^3 in 1 dm^3 .



Figure 2.1.4 A typical water bottle is 1 liter in volume.



Figure 2.1.5 This Rubik's cube is 5.7 cm on each side and has a volume of 185.2 cm^3 or 185.2 mL.

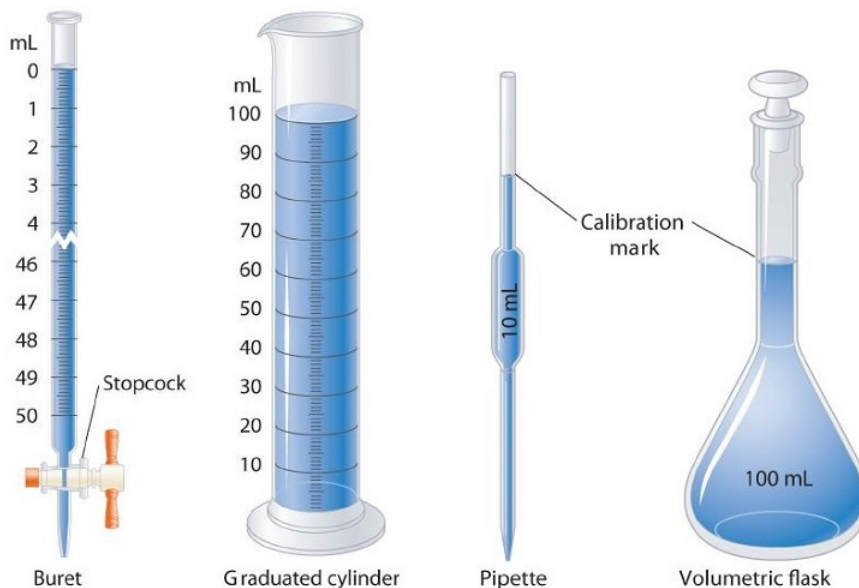


Figure 2.1.6: Buret & graduated cylinder: graduated glassware is used to deliver variable volumes of liquid. Pipette & volumetric flask: volumetric glassware is used to deliver (pipette) or contain (volumetric flask) a single volume accurately when filled to the mark. (CC BY-SA-NC; anonymous).

During your studies of chemistry (and physics also), you will note that mathematical equations are used in many different applications. Many of these equations have a number of different variables with which you will need to work. You should also note that these equations will often require you to use measurements with their units. Algebra skills become very important here!

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can be used to convert meters into centimeters, or a more complex one can be used to convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or much more complex. For instance, you already know that 12 eggs equal one dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/sec. Either one of these can be used as a conversion factor, depending on what type of calculation you might be working with (Table 2.1.1).

Table 2.1.3 Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid ounce (oz)	2.96 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume

*Pounds and ounces are technically units of force, not mass, but this fact is often ignored by the non-scientific community.

Of course, there are other ratios which are not listed in Table 2.1.1. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %).
- Conversions in the metric system, as covered earlier in this chapter.
- Common knowledge ratios (such as 60 seconds = 1 minute).

If you learned the SI units and prefixes described, then you know that 1 cm is 1/100th of a meter.

$$1 \text{ cm} = \frac{1}{100} \text{ m} = 10^{-2} \text{ m}$$

or

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

Suppose we divide both sides of the equation by 1m (both the number *and* the unit):

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1 \text{ m}}{1 \text{ m}}$$

As long as we perform the same operation on both sides of the equals sign, the expression remains an equality. Look at the right side of the equation; it now has the same quantity in the numerator (the top) as it has in the denominator (the bottom). Any fraction that has the same quantity in the numerator and the denominator has a value of 1:

$$\frac{100 \text{ cm}}{1 \text{ m}} = \frac{1000 \text{ mm}}{1 \text{ m}} = \frac{1 \times 10^6 \mu\text{m}}{1 \text{ m}} = 1$$

We know that 100 cm is 1 m, so we have the same quantity on the top and the bottom of our fraction, although it is expressed in different units.

Contributors and Attributions

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CHAPTER OVERVIEW

3: Module 3 "Calculations and Conversions" Textbook Readings

3.1: Conversions and Calculations

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3.1: Conversions and Calculations

Learning Objectives

- Perform unit conversions using conversion factors.

Performing Dimensional Analysis

Dimensional analysis is amongst the most valuable tools used by physical scientists. Simply put, it is the conversion between an amount in one unit to the corresponding amount in a desired unit, using various conversion factors. This is valuable because certain measurements are more accurate or easier to find than others. The use of units in a calculation to ensure that we obtain the final proper units is called *dimensional analysis*.

Here is a simple example: How many centimeters are there in 3.55 m? Perhaps you can determine the answer in your head. If there are 100 cm in every meter, then 3.55 m equals 355 cm. To solve the problem more formally with a conversion factor, we first write the quantity we are given, 3.55 m. Then we multiply this quantity by a conversion factor, which is the same as multiplying it by 1. We can write 1 as $\frac{100 \text{ cm}}{1 \text{ m}}$ and multiply:

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

The 3.55 m can be thought of as a fraction with a 1 in the denominator. Because m, the abbreviation for meters, occurs in both the numerator *and* the denominator of our expression, they cancel out:

$$\frac{3.55 \cancel{\text{m}}}{1} \times \frac{100 \text{ cm}}{1 \cancel{\text{m}}}$$

The final step is to perform the calculation that remains once the units have been canceled:

$$\frac{3.55}{1} \times \frac{100 \text{ cm}}{1} = 355 \text{ cm}$$

In the final answer, we omit the 1 in the denominator. Thus, by a more formal procedure, we find that 3.55 m equals 355 cm. A generalized description of this process is as follows:

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

You may be wondering why we use a seemingly complicated procedure for a straightforward conversion. In later studies, the conversion problems you will encounter *will not always be so simple*. If you can master the technique of applying conversion factors, you will be able to solve a large variety of problems.

In the previous example, we used the fraction $\frac{100 \text{ cm}}{1 \text{ m}}$ as a conversion factor. Does the conversion factor $\frac{1 \text{ m}}{100 \text{ cm}}$ also equal 1? Yes, it does; it has the same quantity in the numerator as in the denominator (except that they are expressed in different units). Why did we not use *that* conversion factor? If we had used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what we would have gotten:

$$3.55 \text{ m} \times \frac{1 \text{ m}}{100 \text{ cm}} = 0.0355 \frac{\text{m}^2}{\text{cm}}$$

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure 3.1.1 shows a concept map for constructing a proper conversion.

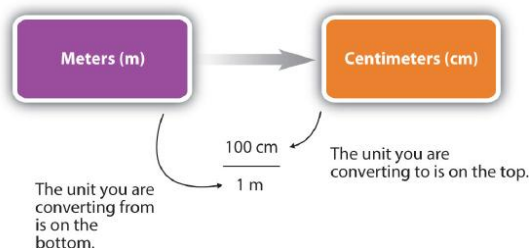


Figure 3.1.7: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another.

📌 The General Steps in Performing Dimensional Analysis

1. Identify the **"given"** information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to **"find"**? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

Steps for Performing Dimensional Analysis and Examples

	Example 3.1.2	Example 3.1.3
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in gallons?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: gal	Given: 18 ms Find: s
List other known quantities.	$1 L = 3.785 gal$	$1 ms = 10^{-3} s$
Prepare a concept map and use the proper conversion factor.		
Cancel units and calculate.	$4.7 \cancel{L} \times \frac{1 \cancel{L}}{3.785 \cancel{L}} = 1.2 \text{ gal}$	$18 \cancel{ms} \times \frac{10^{-3} s}{1 \cancel{ms}} = 0.018 s$ or $18 \cancel{ms} \times \frac{1 s}{1,000 \cancel{ms}} = 0.018 s$
Think about your result.	The amount in gal should be slightly less than 4 times smaller than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

? Exercise 3.1.4

Perform each conversion.

- 101,000 ns to seconds
- 32.08 kg to grams
- 1.53 grams to cg

Answer a:

$$1.01000 \times 10^{-4} s$$

Answer b:

$$3.208 \times 10^4 g$$

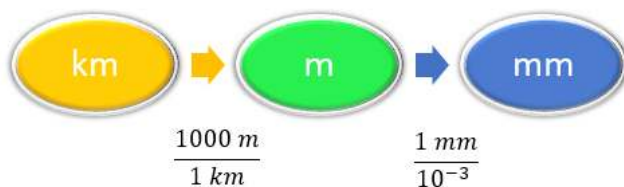
Answer c:

$$1.53 \times 10^2 g$$

Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence. We first convert the given amount in km to the base unit, which is meters. We know that 1,000 m = 1 km. Then we convert meters to mm, remembering that 1 mm = 10^{-3} m.

Concept Map



Convert kilometers to meters to millimeters: use conversion factors 1000 meters per 1 kilometer and 1 millimeter per 0.001 meter

Calculation

$$54.7 \cancel{\text{ km}} \times \frac{1,000 \cancel{\text{ m}}}{1 \cancel{\text{ km}}} \times \frac{1 \cancel{\text{ mm}}}{10^{-3} \cancel{\text{ m}}} = 54,700,000 \text{ mm} \quad (3.1.1)$$

$$= 5.47 \times 10^7 \text{ mm} \quad (3.1.2)$$

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

✓ Example 3.1.4: Unit Conversion

Convert 58.2 ms to megaseconds in one multi-step calculation.

Solution

Example 3.1.4 Steps for Problem Solving Unit Conversion

Steps for Problem Solving

Unit Conversion

Identify the "given" information and what the problem is asking you to "find."

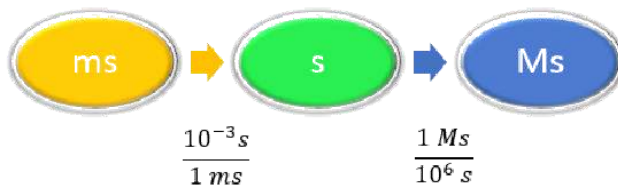
Given: 58.2 ms
Find: Ms

List other known quantities.

$$1\text{ ms} = 10^{-3}\text{ s}$$

$$1\text{ Ms} = 10^6\text{ s}$$

Prepare a concept map.



Convert milliseconds to seconds to microseconds: use conversion factors 0.001 second per millisecond and 1 microsecond per 1 million seconds

Calculate.

$$58.2 \cancel{\text{ ms}} \times \frac{10^{-3} \cancel{\text{ s}}}{1 \cancel{\text{ ms}}} \times \frac{1 \text{ Ms}}{1,000,000 \cancel{\text{ s}}} = 0.0000000582 \text{ Ms}$$

$$= 5.82 \times 10^{-8} \text{ Ms}$$

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

✓ Example 3.1.5: Unit Conversion

How many seconds are in a day?

Solution

Example 3.1.5: Steps for Problem Solving Unit Conversions

Steps for Problem Solving

Unit Conversion

Identify the "given" information and what the problem is asking you to "find."

Given: 1 day
Find: s

List other known quantities.

$$1 \text{ day} = 24 \text{ hours}$$

$$1 \text{ hour} = 60 \text{ minutes}$$

$$1 \text{ minute} = 60 \text{ seconds}$$

Prepare a concept map.



Convert day to hour to minute to second: use conversion factors 24 hours per day, 60 minutes per hour, and 60 seconds per minute

Calculate.

$$1 \text{ d} \times \frac{24 \text{ hr}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ s}}{1 \text{ min}} = 86,400 \text{ s}$$

? Exercise 3.1.5

Perform each conversion in one multistep calculation.

- a. 43.007 ng to kg
- b. 1005 in to ft
- c. 12 mi to km

Answer a:

$$4.3007 \times 10^{-14} \text{ kg}$$

Answer b:

$$83.75 \text{ ft}$$

Answer c:

$$19 \text{ km}$$

Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.
- Mass is a measure of the amount of matter that an object contains.
- Weight is a measure of force that is equal to the gravitational pull on an object.
- Mass is independent of location, while weight depends on location.
- Length is the measurement of the extent of something along its greatest dimension.
- Volume is the amount of space occupied by a sample of matter.
- Volume can be determined by knowing the length of each side of the item.
- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce desired units.

Contributors and Attributions

-
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CHAPTER OVERVIEW

4: Module 4 "Elements, Atoms, and Molecules" Textbook Readings

[4.1: The Periodic Table](#)

[4.2: Atoms and Molecules](#)

[4.3: Atoms and Subatomic Particles](#)

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4.1: The Periodic Table

 Interactive periodic table of the elements, in pictures and words

<https://elements.wlonk.com/ElementsTable.htm>

Interactive Periodic Table of the Elements, in Pictures and Words © 2005-2016 Keith Enevoldsen elements.wlonk.com

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elements.wlonk.com

Learning Objectives

- Describe the classification of the different elements.

Elements that have similar chemical properties are grouped in columns called **groups (or families)**. As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements). In and atomic , the **main group** is the group of elements whose lightest members are represented by , , , , , and as arranged in the of the elements. The main group includes the elements (except , which is sometimes not included) in groups 1 and 2, and groups 13 to 18. Main-group elements (with some of the lighter) are the most abundant elements on , in the , and in the . They are sometimes also called the **representative elements**. In older nomenclature the main-group elements are groups IA and IIA, and groups IIIB to 0 (CAS groups IIIA to VIIIA). Group 12 is labelled as group IIB in both systems. Group 3 is labelled as group IIIA in the older nomenclature (CAS group IIIB).

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

One way to categorize the elements of the periodic table is shown in Figure 4.1.1. The first two columns on the left and the last six columns on the right as mentioned earlier are the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

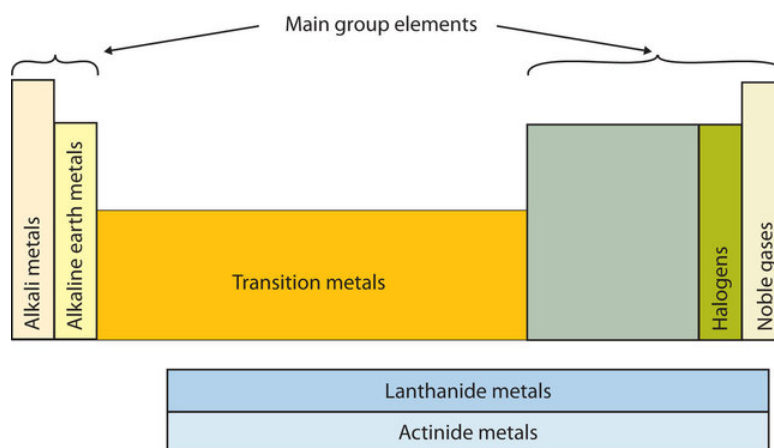


Figure 4.1.1: Special names for sections of the periodic table.

Family Groups

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical

origins.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal.

The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth's crust, respectively.

Group 2: The Alkaline Earth Metals

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

Group 17: The Halogens

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

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

Group 18: The Noble Gases

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

✓ Example 4.1.1: Groups

Provide the family or group name of each element.

- a. Li
- b. Ar
- c. Cl

Solution

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

? Exercise 4.1.1: Groups

Provide the family or group name of each element.

- a. F

- b. Ca
- c. Kr

Answer a:

Fluorine is a halogen (Group 17)

Answer b:

Calcium is a alkaline earth metal (Group 2)

Answer c:

Krypton is a noble gas (Group 18)

Metals and Nonmetals

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure 4.1.2. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 4.1.2, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

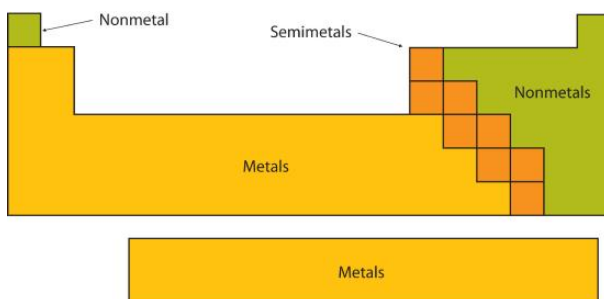


Figure 4.1.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table.

✓ Example 4.1.2

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

- a. Se
- b. Mg
- c. Ge

Solution

- a. In Figure 4.1.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
- b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
- c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

? Exercise 4.1.2

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

Answer

Indium is a metal.

✓ Example 4.1.3: Classification of Elements

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

- Li
- Ar
- Am
- Fe

Solution

- Lithium is a metal.
- Argon is a non metal
- Americium is an inner transition metal
- Iron is a transition metal.

? Exercise 4.1.3: Classification of Elements

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

- F
- U
- Cu

Answer a:

Fluorine is a nonmetal.

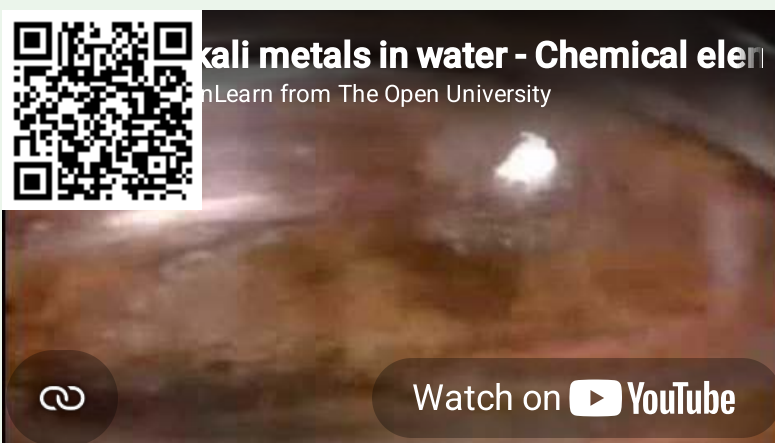
Answer b:

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

Answer c:

Copper is a metal (and a transition metal too)

Alkali Metals



Video 4.1.2 Alkali metals in water.

Summary

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

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

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Contributors and Attributions

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4.2: Atoms and Molecules

Learning Objective

- Know the difference between an atom and a molecule.

An **atom** is the smallest unit of ordinary matter that has the properties of a chemical element. Every solid, liquid, gas, and plasma is composed of either neutral (un-ionized), or ionized atoms. Atoms are extremely small; typical sizes are around 100 picometers (1×10^{-10} m, a ten-millionth of a millimeter, or 1/254,000,000 of an inch). They are small enough that attempting to predict their behavior using classical physics—as if they were billiard balls, for example—gives noticeably incorrect predictions due to quantum effects. Current atomic models have incorporated quantum mechanics to better explain and predict this behavior.

Every atom is composed of a nucleus and one or more electrons bound to the nucleus. The nucleus is made of one or more protons and a number of neutrons; only the most common variety (isotope) of hydrogen has no neutrons.

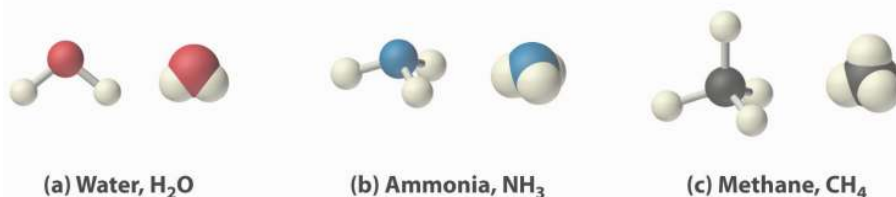
A **molecule** is an electrically neutral group of two or more atoms held together by chemical bonds. Molecules are distinguished from ions by their lack of electrical charge. However, in quantum physics, organic chemistry, and biochemistry, the term *molecule* is often used less strictly—as it is also applied to polyatomic ions.

In the kinetic theory of gases, the term *molecule* is often used for any gas, regardless of its composition. According to this definition, noble gas atoms are considered molecules because they are monatomic molecules.

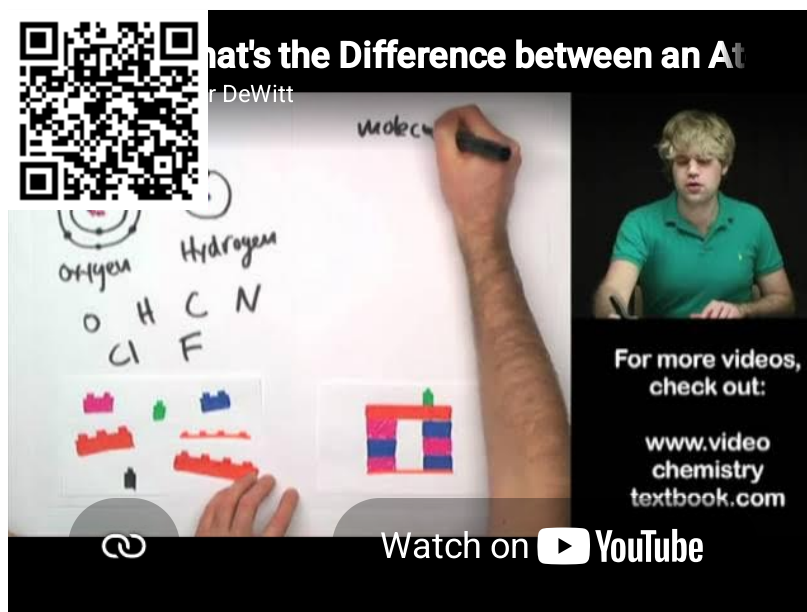
A molecule may be homonuclear—that is, it consists of atoms of one chemical element, as with oxygen (O_2). A heteronuclear molecule is a chemical compound composed of more than one element, as with water (H_2O). Atoms and complexes connected by non-covalent interactions, such as hydrogen bonds or ionic bonds, are typically not considered single molecules.

Molecules as components of matter are common in organic substances (and therefore biochemistry). They also make up most of the oceans and atmosphere. However, the majority of familiar solid substances on Earth—including most of the minerals that make up the crust, mantle, and core of the Earth—contain many chemical bonds, but are *not* made of identifiable molecules. Also, no typical molecule can be defined for ionic crystals (salts) and covalent crystals (network solids), although these are often composed of repeating unit cells that extend either in a plane (such as in graphene) or three-dimensionally (such as in diamond, quartz, or sodium chloride). The theme of repeated unit-cellular-structure also holds for most condensed phases with metallic bonding, which means that solid metals are also not made of molecules. In glasses (solids that exist in a vitreous disordered state), atoms may also be held together by chemical bonds with no presence of any definable molecule, nor any of the regularity of repeating units that characterizes crystals.

Three common molecules shown below.



From left to right: 3D model of water molecule, 3D model of ammonia molecule, and 3D model of methane molecule



Video 4.2.1: The difference between an atom and a molecule.

Summary

- An atom is the smallest unit of matter that has the properties of a chemical element
- A molecule is an electrically neutral group of two or more atoms held together by chemical bonds.

Contributors and Attributions

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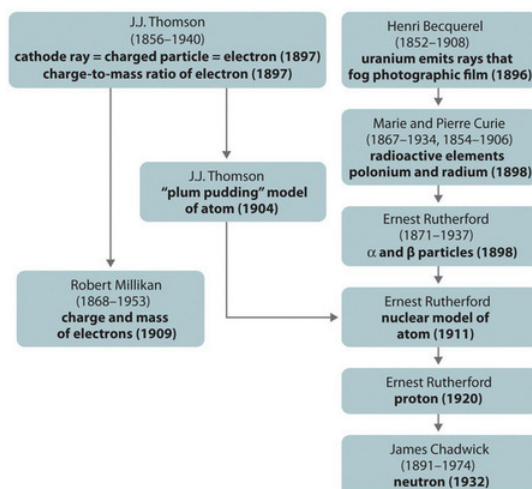
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4.3: Atoms and Subatomic Particles

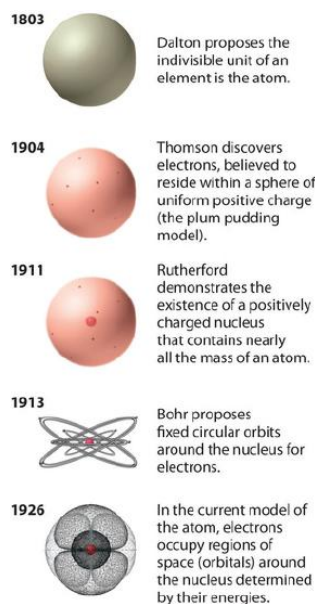
Learning Objectives

- List the properties of the three main subatomic particles.
- Define atomic mass unit (amu).
- Define atomic number and mass number.
- Define isotopes.
- Determine the number of protons, neutrons, and electrons in an atom with a given mass number.



The historical development of the different models of the atom's structure is summarized in Figure 4.3.1. J.J. Thomson and Robert Millikan conducted experiments to study the properties of **electrons**. Rutherford established that the nucleus of the hydrogen atom was a positively charged particle, for which he coined the name **proton** in 1920. He also suggested that the nuclei of elements other than hydrogen must contain electrically neutral particles with approximately the same mass as the proton. The **neutron**, however, was not discovered until 1932, when James Chadwick (1891–1974, a student of Rutherford; Nobel Prize in Physics, 1935) discovered it. As a result of Rutherford's work, it became clear that an α particle contains two protons and neutrons, and is therefore the nucleus of a helium atom.

Figure 4.3.1 A summary of the historical development of models of the components and structure of the atom. The dates in parentheses are the years in which the key experiments were performed. (CC BY-SA-NC).



Rutherford's model of the atom is essentially the same as the modern model, except that it is now known that electrons are not uniformly distributed throughout an atom's volume. Instead, they are distributed according to a set of principles described by Quantum Mechanics.

Figure 4.3.2
Page Index

Figure 4.3.2 shows how the model of the atom has evolved over time from the indivisible unit of Dalton to the modern view taught today.

Figure 4.3.2 The Evolution of Atomic Theory, as Illustrated by Models of the Oxygen Atom. Bohr's model and the current model are described in Chapter 6, "The Structure of Atoms." Image used with Permission (CC BY-SA-NC).

The **nucleus** (plural, nuclei) is a positively charged region at the center of the atom. It consists of two types of subatomic particles packed tightly together. The particles are protons, which have a positive electric charge, and neutrons, which are neutral in electric charge. Outside of the nucleus, an atom is mostly empty space, with orbiting negative particles called electrons whizzing through it. Figure 4.3.3 below shows these parts of the atom.

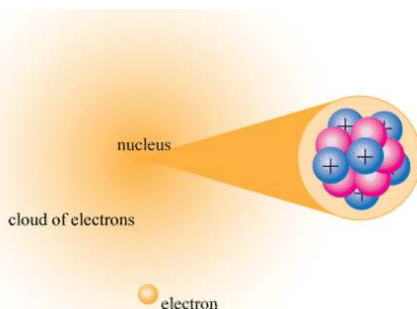


Figure 4.3.3 The nuclear atom

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

📌 Holding It All Together

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

Table 4.3.1 gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit** (amu) is defined as one-twelfth the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

Table 4.3.1: Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p ⁺	1	1	+1	inside the nucleus
electron	e ⁻	5.45 × 10 ⁻⁴	0.00055	-1	outside nucleus
neutron	n ⁰	1	1	0	inside the nucleus

📌 amu in gram and kilogram

$$1 \text{ amu} = 1.6605 \times 10^{-24} \text{ g} = 1.6605 \times 10^{-27} \text{ kg}$$

Atomic Number

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table 4.3.2). If an atom has only one proton, we know it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. The number of protons in an atom is called its **atomic number** (*Z*). This number is very important because it is unique for atoms of a given element. All atoms of an element have the same number of protons, and every element has a different number of protons in its atoms. For example, all helium atoms have two protons, and no other elements have atoms with two protons.

Table 4.3.2: Atoms of the First Six Elements

Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, 79 electrons as well.

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table 4.3.2). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

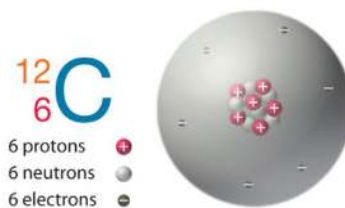
The **mass number** (*A*) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table 4.3.2).

Consider helium again. Most helium atoms have two neutrons in addition to two protons. Therefore the mass of most helium atoms is 4 atomic mass units (2 amu for the protons + 2 amu for the neutrons). However, some helium atoms have more or less than two neutrons. Atoms with the same number of protons but different numbers of neutrons are called isotopes. Because the number of neutrons can vary for a given element, the mass numbers of different atoms of an element may also vary. For example, some helium atoms have three neutrons instead of two (these are called isotopes and are discussed in detail later on)

Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). Counting the number of protons and neutrons tells scientists about the total mass of an atom. An atom's mass number is very easy to calculate provided you know the number of protons and neutrons in an atom. The mass number of a carbon atom with 6 protons and 7 neutrons is calculated and shown as follows:

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

$$\text{mass number} = 6 + 6 = 12$$



✓ Example 4.3.1

What is the mass number of an atom of helium that contains 2 neutrons?

Solution

(number of protons) = 2 (Remember that an atom of helium always has 2 protons.)

(number of neutrons) = 2

mass number = (number of protons) + (number of neutrons)

mass number = 2 + 2 = 4

Isotopes

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

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure 4.3.4. Most hydrogen atoms have just one proton and one electron and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

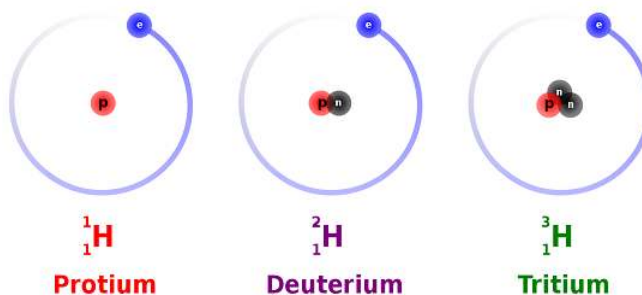


Figure 4.3.4: The three most stable isotopes of hydrogen: protium ($A = 1$), deuterium ($A = 2$), and tritium ($A = 3$). (CC SA-BY 3.0; Balajijagadesh).

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

✓ Example 4.3.2: Lithium Isotopes

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

Solution

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

a.

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

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

b.

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

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

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

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$$\begin{aligned}[x \mapsto s]x &= s \\ [x \mapsto s]y &= y && \text{als } y \neq x \\ [x \mapsto s](\lambda(y)t_1) &= \lambda(y)[x \mapsto s]t_1 && \text{als } y \neq x \text{ en } y \notin FV(s) \\ [x \mapsto s](t_1 t_2) &= ([x \mapsto s]t_1) ([x \mapsto s]t_2)\end{aligned}$$

? Exercise 4.3.1

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

Symbols for Isotopes

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

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



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

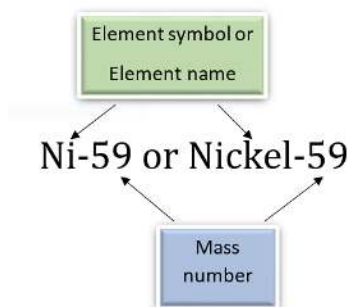
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.





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

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



One way to represent isotopes: The element symbol or name comes first, then a hyphen, then the mass number

✓ Example 4.3.3: POTASSIUM-40

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

Solution

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

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

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

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

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

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

✓ Example 4.3.4: Zinc-65

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

Solution

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

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

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

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

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

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

? Exercise 4.3.2

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

- ${}^{60}_{27}\text{Co}$
- Na-24
- ${}^{45}_{20}\text{Ca}$
- Sr-90

Answer a:

27 protons, 27 electrons, 33 neutrons

Answer b:

11 protons, 11 electrons, 13 neutrons

Answer c:

20 protons, 20 electrons, 25 neutrons

Answer d:

38 protons, 38 electrons, 52 neutrons

Table 4.3.3: Properties of Isotopes of the First Six Elements.

Element	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass (amu)	% Natural Abundance
hydrogen	${}^1_1\text{H}$ (protium)	1	1	0	1.0078	99.989
	${}^2_1\text{H}$ (deuterium)	1	1	1	2.0141	0.0115
	${}^3_1\text{H}$ (tritium)	1	1	2	3.01605	— (trace)
helium	${}^3_2\text{He}$	2	2	1	3.01603	0.00013
	${}^4_2\text{He}$	2	2	2	4.0026	100
lithium	${}^6_3\text{Li}$	3	3	3	6.0151	7.59
	${}^7_3\text{Li}$	3	3	4	7.0160	92.41
beryllium	${}^9_4\text{Be}$	4	4	5	9.0122	100
boron	${}^{10}_5\text{B}$	5	5	5	10.0129	19.9
	${}^{11}_5\text{B}$	5	5	6	11.0093	80.1
carbon	${}^{12}_6\text{C}$	6	6	6	12.0000	98.89
	${}^{13}_6\text{C}$	6	6	7	13.0034	1.11
	${}^{14}_6\text{C}$	6	6	8	14.0032	— (trace)

Summary

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

- Each isotope of a given element has the same atomic number but a different mass number (A), which is the sum of the numbers of protons and neutrons.
- The relative masses of atoms are reported using the atomic mass unit (amu) which is defined as one-twelfth of the mass of one atom of carbon-12, with 6 protons, 6 neutrons, and 6 electrons. The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

Contributors and Attributions

-
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- TextMap: Chemistry-The Central Science (Brown et al.)
- TextMap: Fundamentals of General, Organic, and Biological Chemistry (McMurry et al.)
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CHAPTER OVERVIEW

5: Module 5 "Atomic Structure" Textbook Readings

[5.1: The Bohr Model of the Atom](#)

[5.2: Arrangement of Electrons in an Atom](#)

[5.3: Valence Electrons and the Periodic Table](#)

[5.4: The Octet Rule](#)

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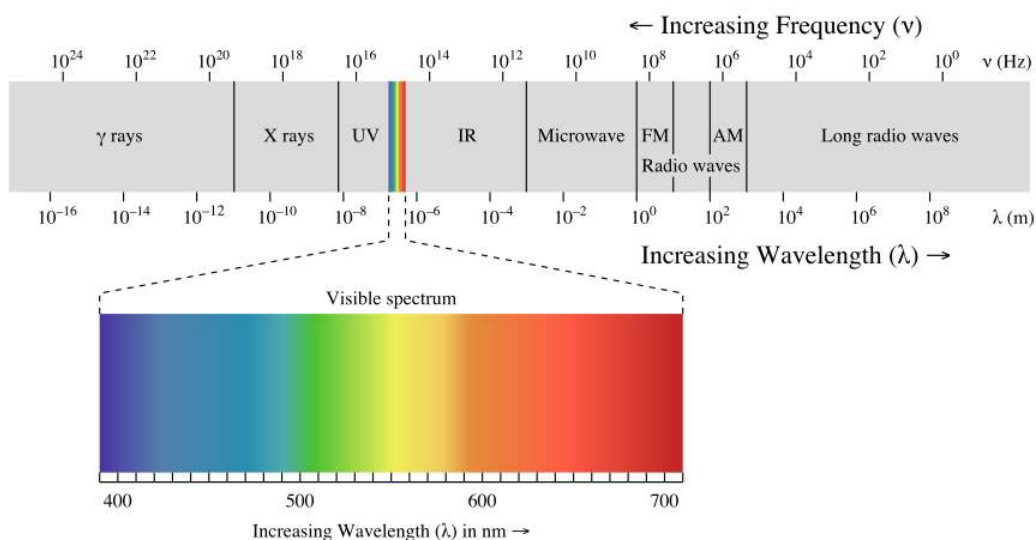
5.1: The Bohr Model of the Atom

Learning Objectives

- Summarize the properties of different types of electromagnetic radiation.
- Define an energy level in terms of the Bohr model.
- Describe the arrangement of electrons using atomic diagrams.

Electromagnetic waves have an extremely wide range of wavelengths, frequencies, and energies. The highest energy form of electromagnetic waves are gamma (γ) rays and the lowest energy form are radio waves.

The figure below shows the **electromagnetic spectrum**, which is all forms of electromagnetic radiation. On the far left of Figure 5.1.1 are the highest energy electromagnetic waves. These are called **gamma rays** and can be quite dangerous, in large numbers, to living systems. The next lower energy form of electromagnetic waves are called **x-rays**. Most of you are familiar with the penetration abilities of these waves. They can also be dangerous to living systems. Humans are advised to limit as much as possible the number of medical x-rays they have per year. Next lower, in energy, are **ultraviolet rays**. These rays are part of sunlight and the upper end of the ultraviolet range can cause sunburn and perhaps skin cancer. The tiny section next in the spectrum is the **visible range of light** ... this section has been greatly expanded in the bottom half of the figure so it can be discussed in more detail. The visible range of electromagnetic radiation are the frequencies to which the human eye responds. Lower in the spectrum are infrared rays and radio waves.



The electromagnetic spectrum, with its various regions labeled. The borders of each region are approximate. As frequency increases wavelength decreases. Gamma rays: frequency between 10^{20} and 10^{24} , wavelength between 10^{-16} and 10^{-12} ; X-Rays: frequency between 10^{16} and 10^{20} , wavelength between 10^{-12} and 10^{-8} ; UV: frequencies between 10^{15} and 10^{17} , wavelengths between 10^{-8} and 10^{-6} ; visible spectrum of light: wavelength to between 400 and 700 nm; IR: frequencies between 10^{11} and 10^{15} , wavelengths between 10^{-6} and 10^{-3} ; Microwave: frequencies between 10^8 and 10^{11} , wavelengths between 10^{-3} and 10^0 ; radio waves: frequencies between 10^5 and 10^8 , wavelengths between 10^0 and 10^3 ; long radio waves: the frequencies between 10^0 and 10^5 .

The light energies that are in the visible range are electromagnetic waves that cause the human eye to respond when those frequencies enter the eye. The eye sends a signal to the brain and the individual "sees" various colors. The highest energy waves in the visible region cause the brain to see violet and as the energy decreases, the colors change to blue, green, yellow, orange, and red. When the energy of the wave is above or below the visible range, the eye does not respond to them. When the eye receives several different frequencies at the same time, the colors are blended by the brain. If all frequencies of light strike the eye together, the brain sees white and if there are no visible frequencies striking the eye, the brain sees black. The objects that you see around you are light absorbers - that is, the chemicals on the surface of the object will absorb certain frequencies and not others. Your eyes detect the frequencies that strike your eye. Therefore, if your friend is wearing a red shirt, it means the dye in that shirt absorbs

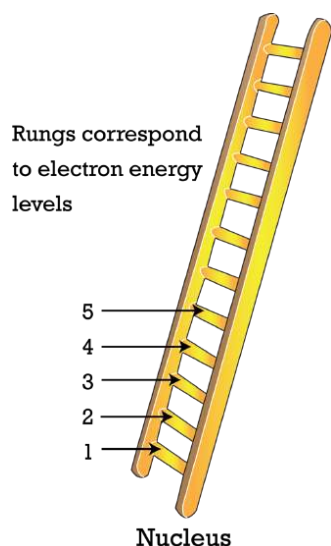
every frequency except red and the red frequencies are reflected. If your only light source was one exact frequency of blue light and you shined it on a shirt that was red in sunlight, the shirt would appear black because no light would be reflected. The light from fluorescent types of lights do not contain all the frequencies of sunlight and so clothes inside a store may appear to be a slightly different color than when you get them home.

Bohr's Explanation of Atomic Structure

In 1913, the Danish physicist Niels Bohr proposed a model of the atom in which electrons orbit the nucleus. Understanding Bohr's model requires some knowledge of electromagnetic radiation (or light). Bohr's key idea in his model of the atom is that electrons occupy definite orbitals that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in

one of the allowable orbitals and it must have the precise energy required for that orbit. Orbits closer to the nucleus would require smaller amounts of energy for an electron and orbits farther from the nucleus would require the electrons to have a greater amount of energy. The possible orbits are known as **energy levels (n)**. One of the weaknesses of Bohr's model was that he could not offer a reason why only certain energy levels or orbits were allowed.

Figure 5.1.5: Niels Bohr with Albert Einstein at Paul Ehrenfest's home in Leiden (December 1925).



Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized**, meaning that only specific amounts are possible. It would be like a ladder that had rungs only at certain heights. The only way you can be on that ladder is to be on one of the rungs and the only way you could move up or down would be to move to one of the other rungs. Suppose we had such a ladder with 10 rungs. Other rules for the ladder are that only one person can be on a rung in normal state and the ladder occupants must be on the lowest rung available. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all the lower rungs are full. Bohr worked out rules for the maximum number of electrons that could be in each energy level in his model and required that an atom in its normal state (ground state) had all electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, Bohr's model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.

Figure 5.1.6 The energy levels ($n = 1, 2, 3, \dots$) of the electrons can be viewed as rungs on a ladder.

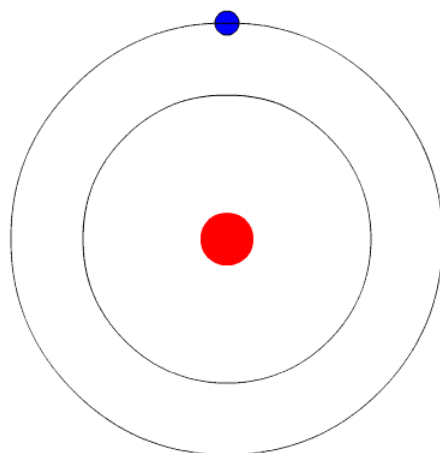


Figure 5.1.7 In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. (CC BY-SA 3.0; [Kurzban](#)).

The evidence used to support Bohr's model came from spectroscopy experiments.

Building Atoms: Main Shells

An electron "shell" is the outside part of an atom around the atomic nucleus. It is a group of atomic orbitals with the same value of the principal quantum number n . Electron shells have one or more electron subshells, or sublevels. The name for electron shells comes from the Bohr model, in which groups of electrons were believed to go around the nucleus at certain distances, so that their orbits formed "shells".

An electron shell may be thought of as an orbit followed by electrons around an atom nucleus. Because each shell can contain only a fixed number of electrons, each shell is associated with a particular range of electron energy, and thus each shell must fill completely before electrons can be added to an outer shell. The electrons in the outermost shell determine the chemical properties of the atom (see Valence shell). For an explanation of why electrons exist in these shells see electron configuration.

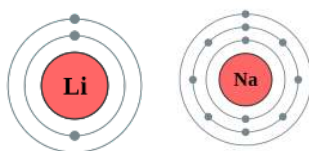


Figure 5.1.10 A shell diagram of lithium (left) and Sodium (right)

The electron shells are labeled K, L, M, N, O, P, and Q; or 1, 2, 3, 4, 5, 6, and 7; going from innermost shell outwards. Electrons in outer shells have higher average energy and travel farther from the nucleus than those in inner shells. This makes them more important in determining how the atom reacts chemically and behaves as a conductor, because the pull of the atom's nucleus upon them is weaker and more easily broken. In this way, a given element's reactivity is highly dependent upon its electronic configuration.

Drawing Shell Models



Shell Models

raptor
 shell model for Sulfur



$1^4 \rightarrow 2e^-$
 $2,3 \rightarrow 8e^-$
 $4,5 \rightarrow 18e^-$
 $6,7 \rightarrow 32e^-$

RECORDED WITH
 SCREENCASTOMATIC

Watch on YouTube

Video 5.1.1 How to draw the shell model for sulfur.

Note: The number of electrons that can occupy each energy level are 2 (first level), 8 (2nd level), 18 (3rd level), and 32 (4th level) based on the formula: # of electrons = $2(n)^2$, wherein n = principle energy level.

Table 5.1.1 shows the number of electrons that fill each shell for neutral atoms of several elements. As mentioned earlier, the innermost shell (corresponding to lowest energy) is filled first and only a fixed number of electrons is allowed in each shell. The

only electron in hydrogen ($Z=1$) goes to the first shell. In lithium atom ($Z=3$), the two electrons fill the first shell, and the third electron goes to the second shell. An argon atom ($Z=18$) has 18 electrons. The 10 electrons fill the first and second shells, and the remaining 8 electrons go to the third shell. The electron configuration for elements past argon are covered in more detail in section 3.7.

Table 5.1.1 *Electron Configuration (Arrangement) of Several Elements.*

Element Symbol	Atomic Number* (Z)	First Shell n=1 (2 electrons allowed)	Second Shell n=2 (8 electrons allowed)	Third Shell n=3 (18 electrons allowed)
H	1	1		
He	2	2		
C	6	2	4	
N	7	2	5	
Na	11	2	8	1
Mg	12	2	8	2
Cl	17	2	8	7
Ar	18	2	8	8

Note:* In a neutral atom the number of protons is equal to the number of electrons.

Summary

- Electromagnetic radiation has a wide spectrum, including gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves, and radio waves.
- The different colors of light differ in their frequencies (or wavelengths).
- Bohr's model suggests each atom has a set of unchangeable energy levels and electrons in the electron cloud of that atom must be in one of those energy levels.
- Bohr's model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for Bohr's model of the atom.
- Bohr's model was only successful in calculating energy levels for the hydrogen atom.
- The shell model is a good representation of electron arrangement only for elements 1-18.

Contributors and Attributions

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- 3.6: Electron Arrangement- The Bohr Model (Orbits) is licensed [CC BY-NC-SA 4.0](#).

5.2: Arrangement of Electrons in an Atom

Learning Objectives

- Represent the organization of electrons in an atom by writing an electron configuration.

Quantum mechanics is the study of the motion of objects that are atomic or subatomic in size and thus demonstrate wave-particle duality. One of the fundamental (and hardest to understand) principles of quantum mechanics is that the electron is both a particles and a wave. In the everyday macroscopic world of things we can see, something cannot be both. But this duality can exist in the quantum world of the submicroscopic at the atomic scale.

At the heart of quantum mechanics is the idea that we cannot specify accurately the location of an electron. All we can say is that there is a probability that it exists within this certain volume of space. The scientist Erwin Schrödinger developed an equation that deals with these calculations, which we will not pursue at this time. Recall that in the Bohr model, the exact path of the electron was restricted to very well-defined circular **orbits** around the nucleus. An **orbital** is the quantum mechanical refinement of Bohr's orbit. In contrast to his concept of a simple circular orbit with a fixed radius, orbitals are mathematically derived regions of space with different *probabilities* of having an electron.



Figure 5.2.1 The Erwin Schrödinger.

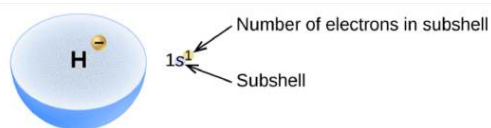
Building Atoms by Orbital Filling

In the quantum-mechanical model of an atom, electrons in the same atom that have the same principal quantum number (n) or principal energy level are said to occupy an electron **shell** of the atom. Orbitals define regions in space where you are likely to find electrons. Orbitals within a shell are divided into **subshells (sublevels)**. There are two subshells in the second shell namely the 2s subshell (with one orbital) and the 2p subshell (with three orbitals).

In contrast to the shell model discussed in section 3.6, chemists more often use a more detailed electron configuration to represent the organization of electrons in shells and subshells in an atom. An atom's electron configuration includes the shell and subshell labels, with a right superscript giving the number of electrons in that subshell. The shells and subshells are listed in the order of filling. Electrons are typically organized around an atom by starting at the lowest possible quantum numbers first, which are the shells-subshells with lower energies.

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

He has two electrons in the 1s subshell. Its electron configuration is **He: $1s^2$**



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

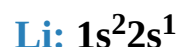


Figure 5.2.3 The diagram of an electron configuration of hydrogen right specifies the subshell and superscript number of electrons.

Table 5.2.2: Electron Configurations of First and Second-Period Elements

Element Name	Symbol	Atomic Number	Electron Configuration
Hydrogen	H	1	$1s^1$
Helium	He	2	$1s^2$
Lithium	Li	3	$1s^2 2s^1$
Beryllium	Be	4	$1s^2 2s^2$
Boron	B	5	$1s^2 2s^2 2p^1$
Carbon	C	6	$1s^2 2s^2 2p^2$
Nitrogen	N	7	$1s^2 2s^2 2p^3$
Oxygen	O	8	$1s^2 2s^2 2p^4$
Fluorine	F	9	$1s^2 2s^2 2p^5$
Neon	Ne	10	$1s^2 2s^2 2p^6$

In order to create ground state electron configurations for any element, it is necessary to know the way in which the atomic sublevels are organized in order of increasing energy. Figure 5.2.4 shows the order of increasing energy of the sublevels.

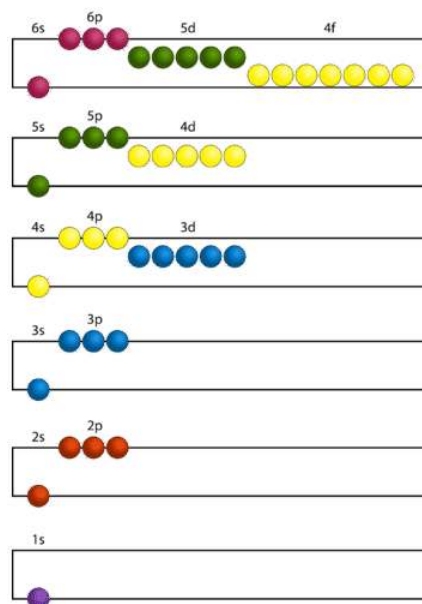


Figure 5.2.4: Electrons are added to atomic orbitals in order from low energy (bottom of the graph) to high (top of the graph) according to the Aufbau principle. Principle energy levels are color coded, while sublevels are grouped together and each circle represents an orbital capable of holding two electrons.

Figure 5.2.5 illustrates the traditional way to remember the filling order for atomic orbitals.

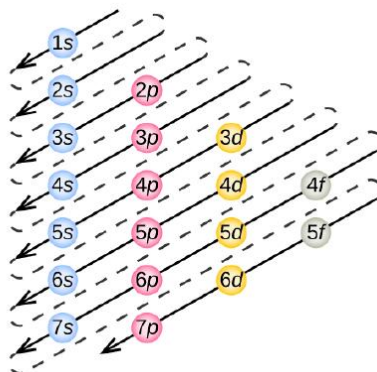


Figure 5.2.5 The arrow leads through each subshell in the appropriate filling order for electron configurations. This chart is straightforward to construct. Simply make a column for all the s orbitals with each n shell on a separate row. Repeat for p , d , and f . Be sure to only include orbitals allowed by the quantum numbers (no $1p$ or $2d$, and so forth). Finally, draw diagonal lines from top to bottom as shown.

✓ Example 5.2.1: Nitrogen Atoms

Nitrogen has 7 electrons. Write the electron configuration for nitrogen.

Solution:

Take a close look at Figure 9.6.5, and use it to figure out how many electrons go into each sublevel, and also the order in which the different sublevels get filled.

1. Begin by filling up the $1s$ sublevel. **This gives $1s^2$.** Now all of the orbitals in the red $n = 1$ block are filled.

Since we used 2 electrons, there are $7 - 2 =$ **5 electrons left**

2. Next, fill the $2s$ sublevel. This gives $1s^2 2s^2$. Now all of the orbitals in the s sublevel of the orange $n = 2$ block are filled.

Since we used another 2 electrons, there are $5 - 2 =$ **3 electrons left**

3. Notice that we haven't filled the entire $n = 2$ block yet... there are still the p orbitals!

The final 3 electrons go into the $2p$ sublevel. This gives $1s^2 2s^2 2p^3$

The overall electron configuration is: **$1s^2 2s^2 2p^3$** .

✓ Example 5.2.2: Potassium Atoms

Potassium has 19 electrons. Write the electron configuration code for potassium.

Solution

This time, take a close look at Figure 9.6.5.

1. Begin by filling up the $1s$ sublevel. **This gives $1s^2$.** Now the $n = 1$ level is filled.

Since we used 2 electrons, there are $19 - 2 =$ **17 electrons left**

2. Next, fill the $2s$ sublevel. **This gives $1s^2 2s^2$**

Since we used another 2 electrons, there are $17 - 2 =$ **15 electrons left**

3. Next, fill the $2p$ sublevel. **This gives $1s^2 2s^2 2p^6$.** Now the $n = 2$ level is filled.

Since we used another 6 electrons, there are $15 - 6 =$ **9 electrons left**

4. Next, fill the 3s sublevel. **This gives $1s^2 2s^2 2p^6 3s^2$**

Since we used another 2 electrons, there are $9 - 2 = 7$ **electrons left**

5. Next, fill the 3p sublevel. **This gives $1s^2 2s^2 2p^6 3s^2 3p^6$**

Since we used another 6 electrons, there are $7 - 6 = 1$ **electron left**

Here's where we have to be careful – right after $3p^6$!!

Remember, 4s comes before 3d!

6. The final electron goes into the 4s sublevel. **This gives $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$**

The overall electron configuration is: **$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$.**

? Exercise 5.2.1: Magnesium and Sodium Atoms

What is the electron configuration for Mg and Na?

Answer Mg

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

Answer Na

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

Summary

- Orbitals define regions in space where you are likely to find electrons.
- Electron configuration notation simplifies the indication of where electrons are located in a specific atom. Superscripts are used to indicate the number of electrons in a given sublevel.
- Electrons are added to atomic orbitals in order from low energy to high energy.

Contributors and Attributions

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- TextMap: Beginning Chemistry (Ball et al.)

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5.3: Valence Electrons and the Periodic Table

 Interactive periodic table of the elements, in pictures and words

<https://elements.wlonk.com/ElementsTable.htm>

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Learning Objectives

- Define valence electrons.
- Describe how the number of valence electrons are related the arrangement of the main group elements and its chemical behavior.

Family Features: Outer Electron Configurations

Valence Electrons

The valence shell is the outermost shell of an atom in its uncombined state, which contains the electrons most likely to account for the nature of any reactions involving the atom and of the bonding interactions it has with other atoms. **Valence electrons** are the outer-shell electrons of an atom. These are electrons that can participate in the formation of a chemical bond. The presence of valence electrons can determine the element's chemical properties and whether it may bond with other elements.

An atom with a closed shell of valence electrons tends to be chemically inert. An atom with one or two valence electrons more than a closed shell is highly reactive, because the extra valence electrons are easily removed to form a positive ion. An atom with one or two valence electrons fewer than a closed shell is also highly reactive, because of a tendency either to gain the missing valence electrons (thereby forming a negative ion), or to share valence electrons (thereby forming a covalent bond).

The number of valence electrons of an element can be determined by the periodic table group (vertical column) in which the element is categorized (Table 5.3.1). With the exception of groups 3–12 (the transition metals), the units digit of the group number identifies how many valence electrons are associated with a neutral atom of an element listed under that particular column.

Table 5.3.1: Number of Valence Electrons for Main Periodic Table Groups.

Periodic table group	Valence electrons
Group 1 (I) (alkali metals)	1
Group 2 (II) (alkaline earth metals)	2
Groups 3-12 (transition metals)	2*
Group 13 (III) (boron group)	3
Group 14 (IV) (carbon group)	4
Group 15 (V) (pnictogens)	5
Group 16 (VI) (chalcogens)	6
Group 17 (VII) (halogens)	7
Group 18 (VIII or 0) (noble gases)	8**

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

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

 Valence Electrons and the Periodic Table

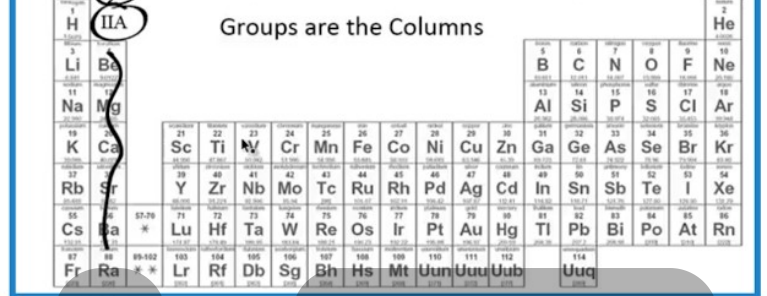


Finding the Number of Valence Electrons

Dr. B. Breslyn (Dr. B.)

Valence Electrons

Groups are the Columns



Watch on  YouTube

Video 5.3.1 Finding the valence electron for an element.

Summary

- Valence electrons are the outer-shell electrons of an atom. These are electrons that can participate in the formation of a chemical bond.

References

- Petrucchi, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
- Sisler, Harry H. Electronic structure, properties, and the periodic law. New York; Reinhold publishing corporation, 1963.
- Petrucchi, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

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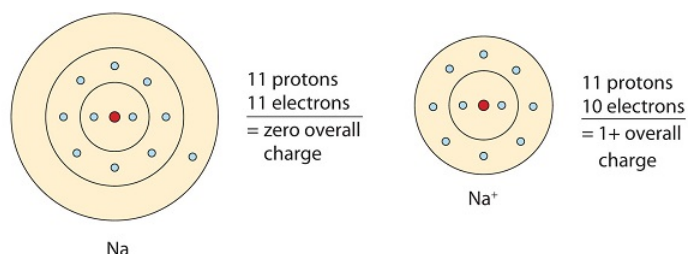
5.4: The Octet Rule

Learning Objective

- Define the octet rule.

The **octet rule** is a chemical rule of thumb that reflects observation that elements tend to bond in such a way that each atom has eight electrons in its valence shell, giving it the same electronic configuration as a noble gas.

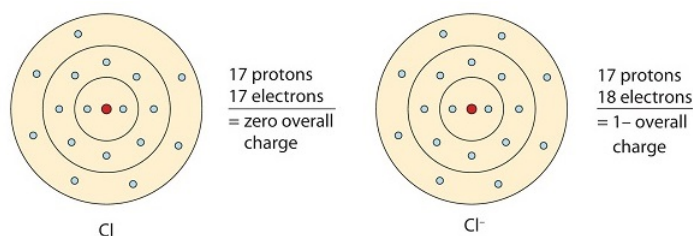
Referring to the octet rule, atoms attempt to get a noble gas electron configuration, which is eight valence electrons. Metals have few electrons in their outer-most orbitals. By losing those electrons, these metals can achieve noble gas configuration and satisfy the octet rule. Sodium has one valence electron, so giving it up would result in the same electron configuration as **neon**. Due to the loss of electron, a positively charged cation (Na^+) called sodium ion is formed.



Sodium (Na) neutral (left): 11 protons and 11 electrons equal zero overall charge; Sodium (Na) positive (right): 11 protons and 10 electrons equals one plus overall charge

Similarly, nonmetals that have close to 8 electrons in their valence shells tend to readily accept electrons to achieve noble gas configuration. Chlorine has seven valence electrons, so if it takes one it will have eight (an octet). Chlorine has the electron configuration of **argon** when it gains an electron. The resulting ion (Cl^-) is called chloride ion.

The octet rule could have been satisfied if chlorine gave up all seven of its valence electrons and sodium took them. In that case, both would have the electron configurations of noble gasses, with a full valence shell. However, their charges would be much higher. It would be Na^{7-} and Cl^{7+} , which is much less stable than Na^+ and Cl^- . Atoms are more stable when they have no charge, or a small charge.



Chlorine (Cl) neutral (left): 17 protons and 17 electrons equal zero overall charge; chlorine (Cl) negative (right): 17 protons and 18 electrons equals one minus overall charge

Summary

The octet rule refers to the tendency of atoms to prefer to have eight electrons in the valence shell.

Contributor

Organic Chemistry Supplemental_Modules_(Organic_Chemistry)/Fundamentals/Ionic_and_Covalent_Bonds

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CHAPTER OVERVIEW

6: Module 6 "Ionic Compounds" Textbook Readings

6.1: Lewis Symbols for Ionic Compounds

6.2: Formulas and Names of Ionic Compounds

6.3: Polyatomic Ions

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6.1: Lewis Symbols for Ionic Compounds

Learning Objectives

- State the octet rule.
- Define cations and anions.
- Define *ionic bond*.
- Draw Lewis structures for ionic compounds.

The Octet Rule: The Drive for Eight

The **octet rule** refers to the tendency of atoms to prefer to have eight electrons in the *valence shell*. When atoms have fewer than eight electrons, they tend to react and form more stable compounds. When discussing the octet rule, we do not consider d or f electrons. Only the s and p electrons are involved in the octet rule, making it useful for the *main group elements* (elements not in the transition metal or inner-transition metal blocks); an octet in these atoms corresponds to an electron configurations ending with s^2p^6 .

Note

The noble gases rarely form compounds. They have the most stable configuration (full octet, no charge), so they have no reason to react and change their configuration. All other elements attempt to gain, lose, or share electrons to achieve a noble gas configuration.

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

Cations

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

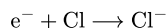


The **cation** produced in this way, Na^+ , is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. F

Anions

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

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



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

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

Note the usefulness of the periodic table in predicting likely ion formation and charge (Figure 6.1.1). Moving from the far left to the right on the periodic table, main-group elements tend to form cations with a charge equal to the group number. That is, group 1 elements form 1+ ions; group 2 elements form 2+ ions, and so on. Moving from the far right to the left on the periodic table, elements often form anions with a negative charge equal to the number of groups moved left from the noble gases. For example, group 17 elements (one group left of the noble gases) form 1− ions; group 16 elements (two groups left) form 2− ions, and so on. This trend can be used as a guide in many cases, but its predictive value decreases when moving toward the center of the periodic table. In fact, transition metals and some other metals often exhibit variable charges that are not predictable by their location in the table. For example, copper can form ions with a 1+ or 2+ charge, and iron can form ions with a 2+ or 3+ charge.

Periodic Table of the Elements

Period	Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1																			He
2		Li ⁺	Be ²⁺												C ^{4−}	N ^{3−}	O ^{2−}	F [−]	Ne
3		Na ⁺	Mg ²⁺											Al ³⁺		P ^{3−}	S ^{2−}	Cl [−]	Ar
4		K ⁺	Ca ²⁺				Cr ³⁺ Cr ⁶⁺	Mn ²⁺	Fe ²⁺ Fe ³⁺	Co ²⁺	Ni ²⁺	Cu ⁺ Cu ²⁺	Zn ²⁺			As ^{3−}	Se ^{2−}	Br [−]	Kr
5		Rb ⁺	Sr ²⁺									Ag ⁺	Cd ²⁺				Te ^{2−}	I [−]	Xe
6		Cs ⁺	Ba ²⁺								Pt ²⁺	Au ⁺ Au ³⁺	Hg ₂ ²⁺ Hg ²⁺					At [−]	Rn
7		Fr ⁺	Ra ²⁺																

*
**

Figure 6.1.1: Some elements exhibit a regular pattern of ionic charge when they form ions.

Group one of the periodic table contains L i superscript plus sign in period 2, N a superscript plus sign in period 3, K superscript plus sign in period 4, R b superscript plus sign in period 5, C s superscript plus sign in period 6, and F r superscript plus sign in period 7. Group two contains B e superscript 2 plus sign in period 2, M g superscript 2 plus sign in period 3, C a superscript 2 plus sign in period 4, S r superscript 2 plus sign in period 5, B a superscript 2 plus sign in period 6, and R a superscript 2 plus sign in period 7. Group six contains C r superscript 3 plus sign and C r superscript 6 plus sign in period 4. Group seven contains M n superscript 2 plus sign in period 4. Group eight contains F e superscript 2 plus sign and F e superscript 3 plus sign in period 4. Group nine contains C o superscript 2 plus sign in period 4. Group ten contains N i superscript 2 plus sign in period 4, and P t superscript 2 plus sign in period 6. Group 11 contains C U superscript plus sign and C U superscript 2 plus sign in period 4, A g superscript plus sign in period 5, and A u superscript plus sign and A u superscript 3 plus sign in period 6. Group 12 contains Z n superscript 2 plus sign in period 4, C d superscript 2 plus sign in period 5, and H g superscript 2 plus sign in period 6. Group 13 contains A l superscript 3 plus sign in period 3. Group 14 contains C superscript 4 negative sign in period 2. Group 15 contains N superscript 3 negative sign in period 2, P superscript 3 negative sign in period 3, and A s superscript 3 negative sign in period 4. Group 16 contains O superscript 2 negative sign in period 2, S superscript 2 negative sign in period 3, S e superscript 2 negative sign in period 4 and T e superscript 2 negative sign in period 5. Group 17 contains F superscript negative sign in period 2, C l superscript negative sign in period 3, B r superscript negative sign in period 4, I superscript negative sign in period 5, and A t superscript negative sign in period 6. Group 18 contains H e in period 1, N e in period 2, A r in period 3, K r in period 4, X e in period 5 and R n in period 6.

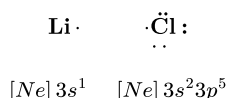
Lewis Symbols for Ionic Compounds

In the section above, we saw how ions are formed by losing electrons to make cations or by gaining electrons to form anions. The astute reader may have noticed something: Many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of

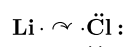
electrons. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider lithium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Li^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Li^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a $1+$ charge after losing a single electron. It turns out that the Li^+ ion has a complete octet in its new valence shell, the $n = 2$ shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

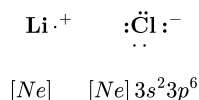
Now consider an Li atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:



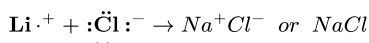
For the Li atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:



resulting in two ions—the Na^+ ion and the Cl^- ion:

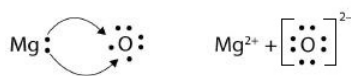


Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:



where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond**, and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:

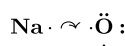


The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:



Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:



The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:



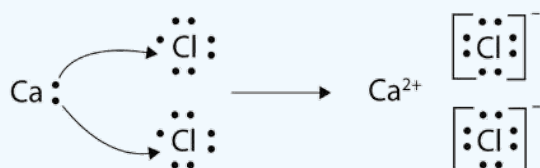
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O . The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

✓ Example 6.1.1: Synthesis of calcium chloride from Elements

With arrows, illustrate the transfer of electrons to form calcium chloride from Ca atoms and Cl atoms.

Solution

A Ca atom has two valence electrons, while a Cl atom has seven electrons. A Cl atom needs only one more to complete its octet, while Ca atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one Ca atom. The transfer process looks as follows:

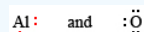


The oppositely charged ions attract each other to make CaCl_2 .

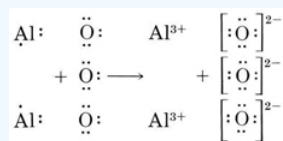
✓ Example 6.1.2: Ionic Formula

Find the formula of the ionic compound formed from O and Al .

Solution: We first write down Lewis diagrams for each atom involved:



We now see that each O atom needs 2 electrons to make up an octet, while each Al atom has 3 electrons to donate. In order that the same number of electrons would be donated as accepted, we need 2 Al atoms ($2 \times 3e^-$ donated) and 3 O atoms ($3 \times 2e^-$ accepted). The whole process is then

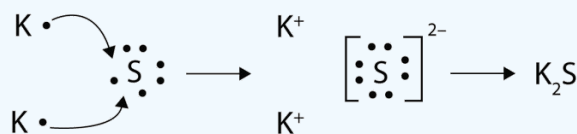


The resultant oxide consists of aluminum ions, Al^{3+} , and oxide ions, O^{2-} , in the ratio of 2:3. The formula is Al_2O_3 .

? Exercise 6.1.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from K atoms and S atoms.

Answer:



Summary

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

Contributors and Attributions

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6.2: Formulas and Names of Ionic Compounds

Learning Objectives

- Name ionic compounds.
- Write the correct formula for an ionic compound.

Names of Binary Ionic Compounds

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

Naming Ions

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

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 4.4.1). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+} . The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system** (Table 6.2.1), an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word ion. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a $1+$ ion, so there is no ambiguity about the name *sodium ion*.

Table 6.2.1: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name (Stock System)	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix

represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 6.2.1 lists the elements that use the common system, along with their respective cation names.

Table 6.2.2: Some Monatomic Anions

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

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^{-} is “chlor-” + “-ide ion,” or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 6.2.2 lists the names of some common monatomic ions.

✓ Example 6.2.1

Name each ion.

- Ca^{2+}
- S^{2-}
- Cu^{+}

Solution

- the calcium ion
- the sulfide ion
- the copper(I) ion or the cuprous ion

? Exercise 6.2.1

Name each ion.

- Fe^{2+}
- Fe^{3+}
- Ba^{2+}

Answer a:

iron(II) ion

Answer b:

iron(III) ion

Answer c:

barium ion

✓ Example 6.2.2

Write the formula for each ion.

- the bromide ion
- the cupric ion
- the magnesium ion

Solution

- Br^-
- Cu^{2+}
- Mg^{2+}

? Exercise 6.2.2

Write the formula for each ion.

- the fluoride ion
- the stannous ion
- the potassium ion

Answer a:



Answer b:



Answer c:



Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

Now that we know how to name ions, we are ready to name ionic compounds. A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 6.2.1 for the compound BaCl_2 . The word *ion* is dropped from both parts.

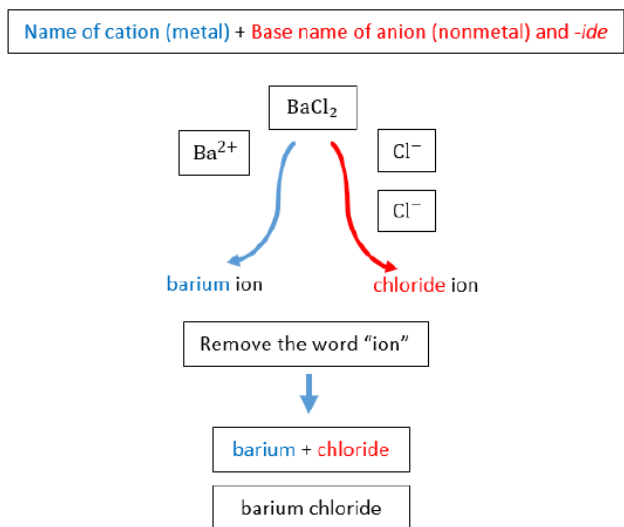


Figure 6.2.1 Naming BaCl_2

Naming formula: Name of metal cation + base name of nonmetal anion + suffix *-ide*. BaCl_2 is named as barium chloride.

Subscripts in the formula do not affect the name.

✓ Example 6.2.3: Naming Ionic Compounds

Name each ionic compound.

- a. CaCl_2
- b. AlF_3
- c. KCl

Solution

- a. Using the names of the ions, this ionic compound is named calcium chloride.
- b. The name of this ionic compound is aluminum fluoride.
- c. The name of this ionic compound is potassium chloride

? Exercise 6.2.3

Name each ionic compound.

- a. AgI
- b. MgO
- c. Ca_3P_2

Answer a:

silver iodide

Answer b:

magnesium oxide

Answer c:

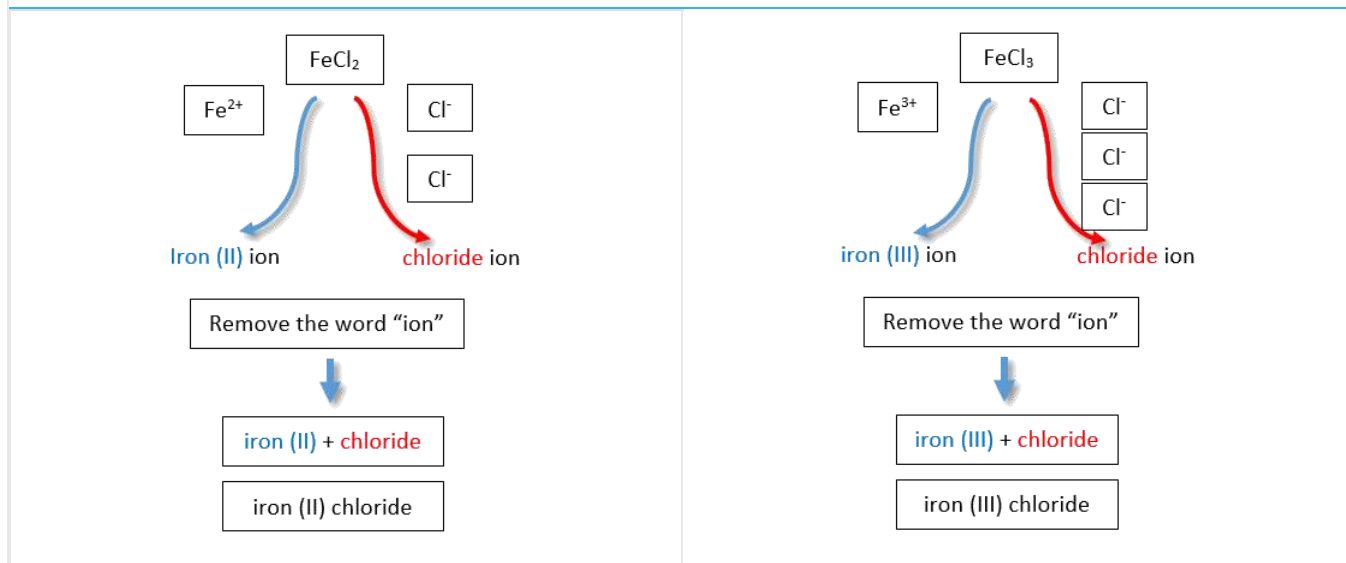
calcium phosphide

Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider FeCl_2 and FeCl_3 . In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1- charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure 6.2.2).

Figure 6.2.2 Naming the FeCl_2 and FeCl_3 compounds in the modern/stock system.

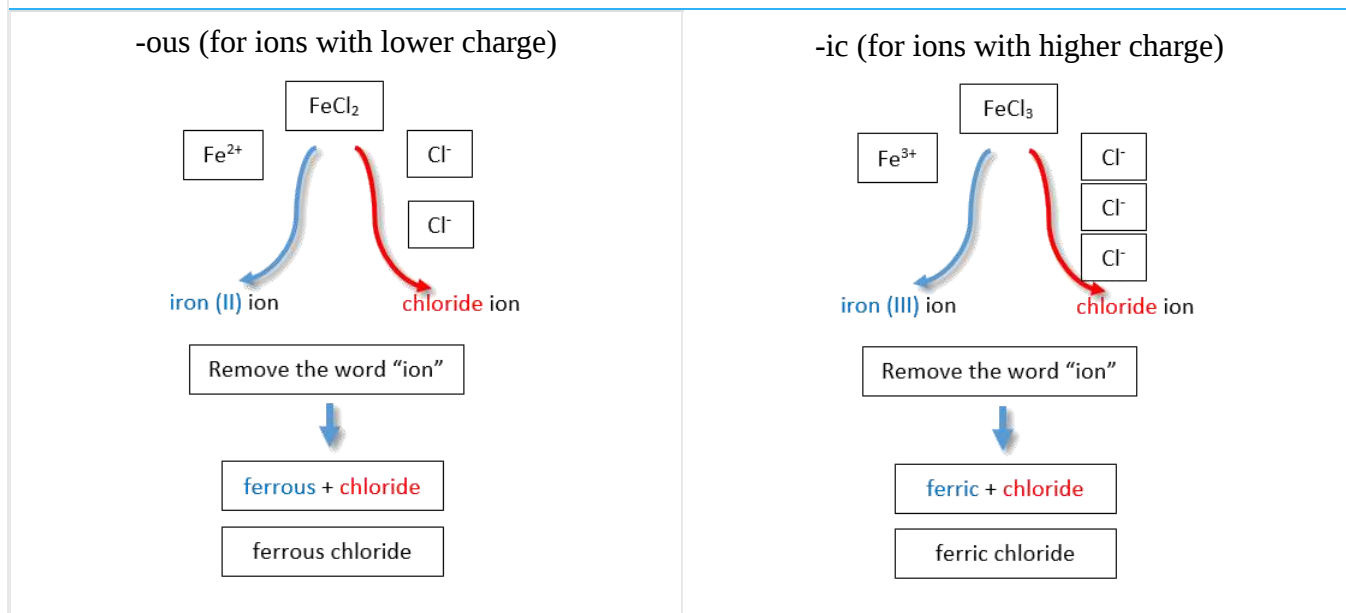
Name of cation (metal) + (Roman Numeral in parenthesis) + Base name of anion (nonmetal) and *-ide*



If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure 6.2.3) .

Figure 6.2.3 Naming the $FeCl_2$ and $FeCl_3$ compounds in the old/common System.

"Old" base name of cation (metal) and *-ic or -ous* + Base name of anion (nonmetal) and *-ide*



✓ Example 6.2.4:

Name each ionic compound.

a. Co_2O_3

b. FeCl_2

Solution

Example 6.2.4: Explanation of the names of ionic compounds

	Ionic Compound	Explanation	Answer
d	Co_2O_3	We know that cobalt can have more than one possible charge; we just need to determine what it is. - Oxide always has a 2^- charge, so with three oxide ions, we have a total negative charge of 6^-. - This means that the two cobalt ions have to contribute 6^+, which for two cobalt ions means that each one is 3^+. - Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
d	FeCl_2	Iron can also have more than one possible charge. - Chloride always has a 1^- charge, so with two chloride ions, we have a total negative charge of 2^-. - This means that the one iron ion must have a 2^+ charge. - Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

? Exercise 6.2.4

Name each ionic compound.

a. AuCl_3

b. PbO_2

c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Figure 6.2.4 is a synopsis of how to name simple ionic compounds.

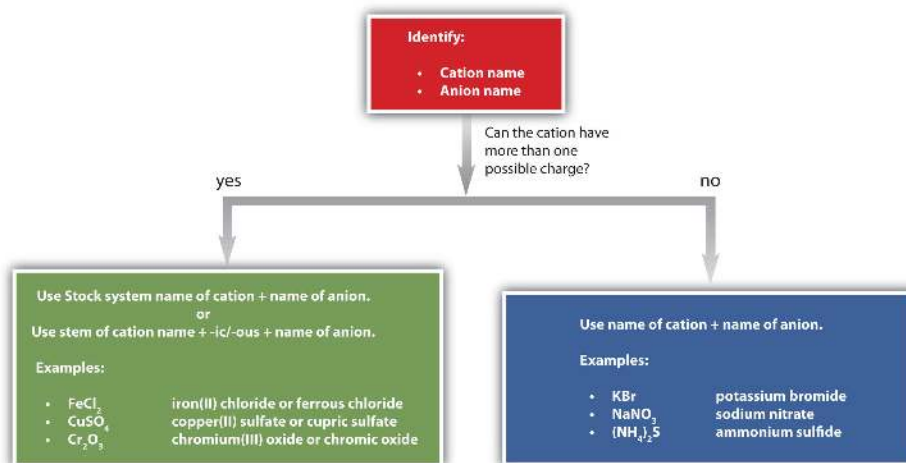


Figure 6.2.4: A Guide to Naming Simple Ionic Compounds.

? Exercise 6.2.5

Name each ionic compound.

- ZnBr₂
- Al₂O₃
- AuF₃
- AgF

Answer a:

zinc bromide

Answer b:

aluminum oxide

Answer c:

gold(III) fluoride or auric fluoride

Answer d:

silver fluoride

Writing Formulas of Ionic Compounds

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattice containing many ions each of the cation and anion. An ionic formula, like NaCl, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na₂S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charge to cancel each other out.

✓ Example 6.2.5: Aluminum Nitride and Lithium Oxide

Write the formula for aluminum nitride and lithium oxide.

Solution

Example 6.2.5: Steps for Problem Solving Writing Formulas

Steps for Problem Solving	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	$\text{Al}^{3+} \text{N}^{3-}$	$\text{Li}^{+} \text{O}^{2-}$
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions $1(3+) = 1(3-)$ $+3 = -3$	total charge of cations = total charge of anions $2(1+) = 1(2-)$ $+2 = -2$
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

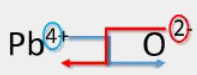
An alternative way to writing a correct formula for an ionic compound is to use the **crisscross method**. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped.

✓ Example 6.2.6: The Crisscross Method for lead (IV) oxide

Write the formula for lead (IV) oxide.

Solution

Example 6.2.6: Steps For Problem Solving the Crisscross Method for writing formulas.

Crisscross Method	Example: Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	$\text{Pb}^{4+} \text{O}^{2-}$
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all subscripts that are 1.	PbO_2

? Exercise 6.2.6

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

- the calcium ion and the oxygen ion
- the 2+ copper ion and the sulfur ion
- the 1+ copper ion and the sulfur ion

Answer a:



Answer b:



Answer c:



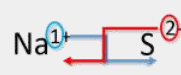
Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.

✓ Example 6.2.7: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Example 6.2.7: Steps For Problem Solving the Crisscross Method for writing formulas.

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	$\text{Na}^+ \quad \text{S}^{2-}$
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all all subscripts that are 1.	Na_2S

? Exercise 6.2.7

Write the formula for each ionic compound.

- sodium bromide
- lithium chloride
- magnesium oxide

Answer a:

NaBr

Answer b:

LiCl

Answer c:

MgO

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between metal and nonmetal elements are usually ionic. For example, CaBr_2 contains a metallic element (calcium, a group 2 (or 2A) metal) and a nonmetallic element (bromine, a group 17 (or 7A) nonmetal). Therefore, it is most likely an ionic compound. (In fact, it is ionic.) In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 15 (or 5A), and oxygen, from group 16 (or 6A)). It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $\text{Ba}(\text{NO}_3)_2$, you may recognize the “ NO_3 ” part as the nitrate ion, NO_3^- . (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2- charge from the two nitrate ions. Thus, this compound is also ionic.

✓ Example 6.2.8

Identify each compound as ionic or not ionic.

- Na_2O
- PCl_3
- OF_2

Solution

Example 6.2.8: Steps For Problem Solving Identifying Ionic Compounds

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, Na_2O is expected to be ionic.	Na_2O , ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic.	PCl_3 , not ionic
c. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.	OF_2 , ionic

? Exercise 6.2.8

Identify each compound as ionic or not ionic.

- N_2O
- FeCl_3

Answer a:

not ionic

Answer b:

ionic

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parenthesis.
- Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

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- Henry Agnew (UC Davis)

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6.3: Polyatomic Ions

Learning Objectives

- Identify polyatomic ions in chemical formulas.
- Write the correct formula for ionic compounds containing polyatomic ions.
- Name ionic compounds containing polyatomic ions.

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 6.3.1 lists the most common polyatomic ions.

Table 6.3.1 Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$\text{C}_2\text{H}_3\text{O}_2^-$ (also written CH_3CO_2^-)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$\text{Cr}_2\text{O}_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	H_2PO_4^-
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $\text{Ba}(\text{NO}_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

✓ Example 6.3.1: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Example 6.3.1: Steps For Problem Solving, Write the formula for calcium nitrate

Criss Cross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	$\text{Ca}^{2+} \quad \text{NO}_3^-$
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$\text{Ca}_1(\text{NO}_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$\text{Ca}(\text{NO}_3)_2$

✓ Example 6.3.2

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion

Solution

Example 6.3.2: Explanation for Writing the Chemical Formula for an Ionic Compound Composed of the Potassium Ion and the Sulfate Ion

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

? Exercise 6.3.1

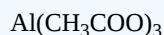
Write the chemical formula for an ionic compound composed of each pair of ions.

- the magnesium ion and the carbonate ion
- the aluminum ion and the acetate ion

Answer a:



Answer b:



Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium Sulfate compound in Figure 6.3.6.

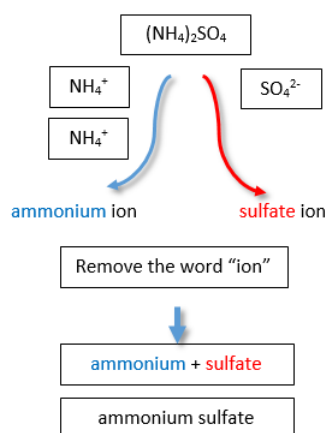


Figure 6.3.2 Naming Ionic Compounds with Polyatomic Ions

✓ Example 6.3.3: Naming Ionic Compounds

Write the proper name for each ionic compound.

- $(\text{NH}_4)_2\text{S}$
- AlPO_4
- $\text{Fe}_3(\text{PO}_4)_2$

Solution

Example 6.3.3: Explanation for Naming Ionic Compounds

Ionic Compound	Explanation	Answer
a. $(\text{NH}_4)_2\text{S}$	a. The ammonium ion has a 1+ charge and the sulfide ion has a 2- charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound's name is ammonium sulfide.	ammonium sulfide
b. AlPO_4	b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. $\text{Fe}_3(\text{PO}_4)_2$	c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6-, two phosphate ions are needed. The compound's name is iron(II) phosphate.	iron(II) phosphate

? Exercise 6.3.2

Write the proper name for each ionic compound.

- $(\text{NH}_4)_3\text{PO}_4$

b. $\text{Co}(\text{NO}_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Summary

- Formulas for ionic compounds contain the symbols and number of each atom and/or polyatomic ion present in a compound in the lowest whole number ratio.
- The process of naming ionic compounds with polyatomic ions are the same as binary ionic compounds.

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CHAPTER OVERVIEW

7: Module 7 "Covalent Compounds" Textbook Readings

7.1: Covalent Bonds- Shared Electron Pairs

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7.1: Covalent Bonds- Shared Electron Pairs

Learning Objectives

- Define covalent bond.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won't give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

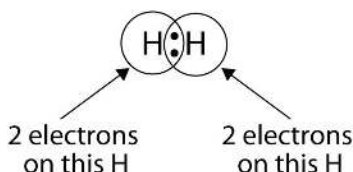
Chemists frequently use Lewis diagrams to represent covalent bonding in molecular substances. For example, the Lewis diagrams of two separate hydrogen atoms are as follows:



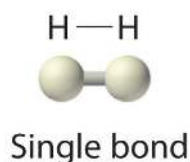
The Lewis diagram of two hydrogen atoms sharing electrons looks like this:



We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity's sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:



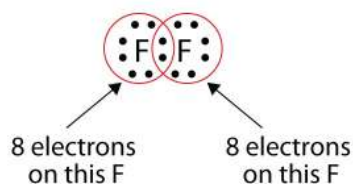
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:



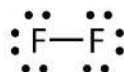
These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.



Note that each F atom has a complete octet around it now:



We can also write this using a dash to represent the shared electron pair:

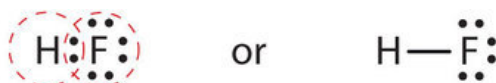


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



In this molecule, the hydrogen atom does not have nonbonding electrons, while the fluorine atom has six nonbonding electrons (three lone electron pairs). The circles show how the valence electron shells are filled for both atoms.

✓ Example 7.1.1:

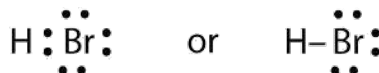
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:



The two atoms can share their unpaired electron:



? Exercise 7.1.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl₂.

Answer:

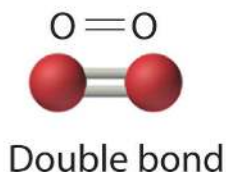


When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far and there are still some electrons hanging out there unattached. You can't just leave them there. So where do you put them?

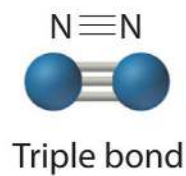
Multiple Covalent Bonds

The sharing of a pair of electrons represents a single covalent bond, usually just referred to as a **single bond**. However, in many molecules atoms attain complete octets by sharing more than one pair of electrons between them:

- Two electron pairs shared a **double bond**



- Three electron pairs shared a **triple bond**



Because each nitrogen contains 5 valence electrons, they need to share 3 pairs to each achieve a valence octet. N_2 is fairly inert, due to the strong triple bond between the two nitrogen atoms.



Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

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CHAPTER OVERVIEW

8: Module 8 "Density" Textbook Readings

8.1: Density

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8.1: Density

Learning Objectives

- Define density.
- Compare the density of various substances.

Density (ρ) is a physical property found by dividing the mass of an object by its volume. Regardless of the sample size, density is always constant. For example, the density of a pure sample of tungsten is always 19.25 grams per cubic centimeter. This means that whether you have one gram or one kilogram of the sample, the density will never vary. The equation, as we already know, is as follows:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

or just

$$\rho = \frac{m}{V}$$

Table 8.1.1 Densities of Common Substances

Substance	Density at 25°C (g/cm ³)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Notice in Table 8.1.1 that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will “float.”

Density can be measured for all substances, solids, liquids and gases. For solids and liquids, density is often reported using the units of g/cm³. Densities of gases, which are significantly lower than the densities of solids and liquids, are often given using units of g/L.

✓ Example 8.1.1: Ethyl Alcohol

Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.71002 g

Solution

$$\rho = \frac{23.71002 \text{ g}}{30.2 \text{ mL}} = 0.785 \text{ g/mL}$$

? Exercise 8.1.1

- Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.
- If you have a 2.130 mL sample of acetic acid with mass 2.234 g, what is the density in g/mL?

Answer a

$$0.274 \text{ kg/L}$$

Answer b

1.049 g/mL

Summary

- Density is defined as the mass of an object divided by its volume.

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CHAPTER OVERVIEW

9: Module 9 "Chemical Equations" Textbook Readings

9.1: Chemical Sentences- Equations

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9.1: Chemical Sentences- Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and → appropriately when writing a chemical reaction.
- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and rearrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

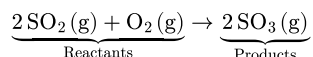
To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that, what is produced or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, some baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.



Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\text{SO}_2 + \text{O}_2$, are reactants and sulfur trioxide, SO_3 , is the product.



In chemical reactions, the reactants are found before the symbol "→" and the products are found after the symbol "→". The general equation for a reaction is:



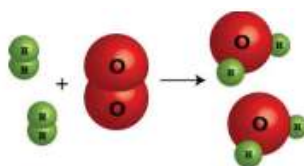
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. You will find there are others but these are the main ones that we need to know. Table 9.1.1 shows a listing of symbols used in chemical equations.

Table 9.1.1 Symbols Used in Chemical Equations

Symbol	Description
+	used to separate multiple reactants or products
→	yield sign; separates reactants from products
⇌	replaces the yield sign for reversible reactions that reach equilibrium
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated
(s)	reactant or product in the solid state
(l)	reactant or product in the liquid state
(g)	reactant or product in the gas state
(aq)	reactant or product in an aqueous solution (dissolved in water)

Chemists have a choice of methods for describing a chemical reaction.

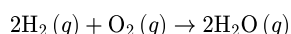
1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:

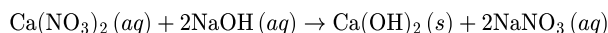
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. They could write the equation in chemical shorthand.

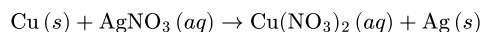


In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products and symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:



How much easier is that to read? Let's try it in reverse? Look at the following reaction in shorthand and write the word equation for the reaction:



The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver."

To turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know what symbols go on each side of the arrow and where the + signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).

3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning they are always found in pairs in nature. They include those elements listed in the table.

Table 9.1.2 Diatomic Elements

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

✓ Example 9.1.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

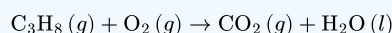
- $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
- Gaseous propane, C_3H_8 , burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
- Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

a. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

b. Reactants: propane (C_3H_8) and oxygen (O_2)

Product: carbon dioxide (CO_2) and water (H_2O)



c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

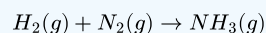


? Exercise 9.1.1

Transfer the following symbolic equations into word equations or word equations into symbolic equations.

- hydrogen gas reacts with nitrogen gas to produce gaseous ammonia
- $HCl(aq) + LiOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
- copper metal is heated with oxygen gas to produce solid copper(II) oxide.

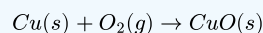
Answer a



Answer b

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

Answer c



Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products - they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance is used or produced.

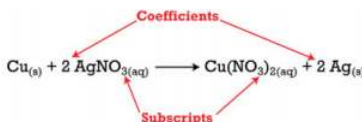


Figure 9.1.1 Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances.

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 9.1.2.

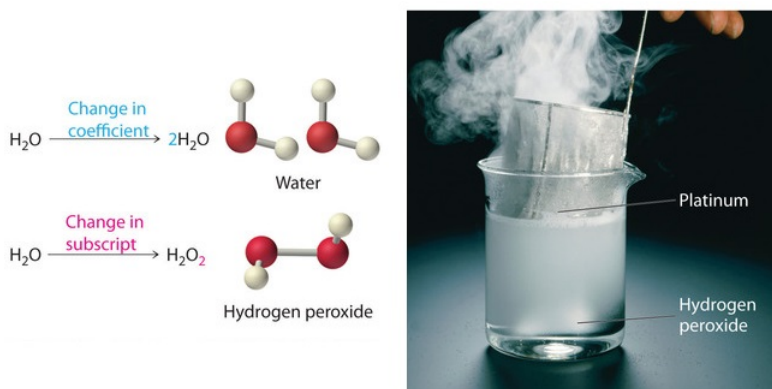


Figure 9.1.2 H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum (left), whereas no such reaction occurs between water and platinum (right).

The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance polyatomic ions (if present on both sides of the chemical equation) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

✓ Example 9.1.2: Combustion of Heptane

Balancing the chemical Equation for the combustion of Heptane (C_7H_{16})



Solution

Example 9.1.2: Steps for Problem Solving the Combustion Of Heptane

Steps	Explanation															
1. Identify the most complex substance.	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.															
2. Adjust the coefficients.	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $C_7H_{16}(l) + O_2(g) \longrightarrow \underline{7}CO_2(g) + H_2O(g)$ Reactants and Products of Elements <table><thead><tr><th>Reactants</th><th>Element/Polyatomic Ion</th><th>Products</th></tr></thead><tbody><tr><td>7</td><td>C</td><td>7</td></tr></tbody></table> b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $C_7H_{16}(l) + O_2(g) \longrightarrow 7CO_2(g) + \underline{8}H_2O(g)$ Reactants and Products of Elements <table><thead><tr><th>Reactants</th><th>Element/Polyatomic Ion</th><th>Products</th></tr></thead><tbody><tr><td>7</td><td>C</td><td>7</td></tr><tr><td>16</td><td>H</td><td>16</td></tr></tbody></table>	Reactants	Element/Polyatomic Ion	Products	7	C	7	Reactants	Element/Polyatomic Ion	Products	7	C	7	16	H	16
Reactants	Element/Polyatomic Ion	Products														
7	C	7														
Reactants	Element/Polyatomic Ion	Products														
7	C	7														
16	H	16														
3. Balance polyatomic ions as a unit.	There are no polyatomic ions to be considered in this reaction.															
4. Balance the remaining atoms.	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $C_7H_{16}(l) + \underline{11}O_2(g) \longrightarrow 7CO_2(g) + 8H_2O(g)$ Reactants and Products of Elements <table><thead><tr><th>Reactants</th><th>Element/Polyatomic Ion</th><th>Products</th></tr></thead><tbody><tr><td>7</td><td>C</td><td>7</td></tr><tr><td>16</td><td>H</td><td>16</td></tr><tr><td>22</td><td>O</td><td>22</td></tr></tbody></table>	Reactants	Element/Polyatomic Ion	Products	7	C	7	16	H	16	22	O	22			
Reactants	Element/Polyatomic Ion	Products														
7	C	7														
16	H	16														
22	O	22														
5. Check your work.	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and															

22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

✓ Example 9.1.3: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})



Solution

The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Example 9.1.3: Solution for the combustion of isooctane

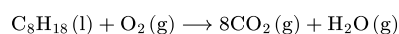
steps

1. Identify the most complex substance.

examples

The most complex substance is the one with the largest number of different atoms, which is C_8H_{18} . We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.

a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products:

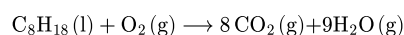


Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Products
8	C	8

2. Adjust the coefficients.

b. Eighteen hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products:



Reactants and Products of Elements

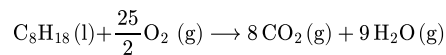
Reactants	Element/Polyatomic Ion	Products
8	C	8
18	H	18

3. Balance polyatomic ions as a unit.

There are no polyatomic ions to be considered in this reaction.

4. Balance the remaining atoms.

The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2 , but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms:

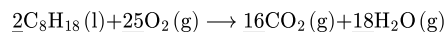


Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Products
8	C	8
18	H	18
25	O	25

The equation is now balanced, but we usually write equations with whole-number coefficients. We can eliminate the fractional

coefficient by multiplying all coefficients on both sides of the chemical equation by 2:



The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side.

Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Products
16	C	16
36	H	36
50	O	50

Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

5. Check your work.

✓ Example 9.1.4: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

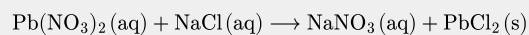
Example 9.1.4: Solution for the combustion of isooctane

Steps for Problem Solving

Example 9.1.4

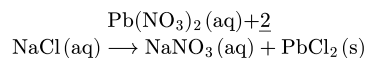
1. Identify the most complex substance.

The most complex substance is lead (II) chloride.



2. Adjust the coefficients.

There are twice as many chloride ions in the reactants than in the products. Place a 2 in front of the NaCl in order to balance the chloride ions.

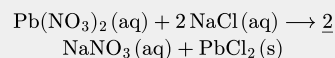


Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Products
1	Pb	1
2	Na	1
2	Cl	2

3. Balance polyatomic ions as a unit.

The nitrate ions are still unbalanced. Place a 2 in front of the NaNO₃. The result is:



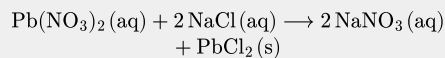
Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Products
1	Pb	1
2	Na	2
2	Cl	2
2	NO ₃ ⁻	2

4. Balance the remaining atoms.

There is no need to balance the remaining atoms because they are already balanced.

5. Check your work.



Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Products
1	Pb	1
2	Na	2
2	Cl	2
2	NO ₃ ⁻	2

? Exercise 9.1.2

Is each chemical equation balanced?

- $2 \text{Hg}(\ell) + \text{O}_2(\text{g}) \rightarrow \text{Hg}_2\text{O}_2(\text{s})$
- $\text{C}_2\text{H}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$
- $\text{Mg}(\text{NO}_3)_2(\text{s}) + 2 \text{Li}(\text{s}) \rightarrow \text{Mg}(\text{s}) + 2 \text{LiNO}_3(\text{s})$

Answer a

yes

Answer b

no

Answer c

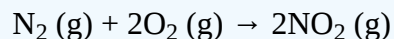
yes

? Exercise 9.1.3

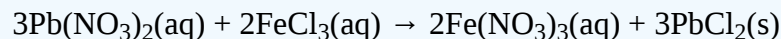
Balance the following chemical equations.

- $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$
- $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{FeCl}_3(\text{aq}) \rightarrow \text{Fe}(\text{NO}_3)_3(\text{aq}) + \text{PbCl}_2(\text{s})$
- $\text{C}_6\text{H}_{14}(\text{l}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$

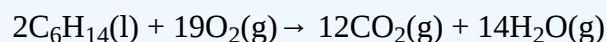
Answer a



Answer b



Answer c



Summary

- A chemical reaction is the process in which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.
- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.

- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Contributors and Attributions

-
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CHAPTER OVERVIEW

10: Module 10 "The Chemical Mole" Textbook Readings

10.1: Avogadro's Number and the Mole

10.2: Molar Mass- Mole-to-Mass and Mass-to-Mole Conversions

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10.1: Avogadro's Number and the Mole

Learning Objectives

- Use Avogadro's number to convert to moles and vice versa given the number of particles of an element.
- Describe the chemical mole.
- Determine the formula mass of an ionic or covalent compound.

When objects are very small, it is often inconvenient or inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross".

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number: Counting Atoms

Owing to their tiny size, atoms and molecules cannot be counted by direct observation. But much as we do when "counting" beans in a jar, we can estimate the number of particles in a sample of an element or compound if we have some idea of the volume occupied by each particle and the volume of the container. Once this has been done, we know the number of formula units (to use the most general term for any combination of atoms we wish to define) in any arbitrary weight of the substance. The number will of course depend both on the formula of the substance and on the weight of the sample. However, if we consider a weight of substance that is the same as its formula (molecular) weight expressed in grams, we have only one number to know: **Avogadro's number**.

Avogadro's number

Avogadro's number is known to ten significant digits:

$$N_A = 6.022141527 \times 10^{23}.$$

However, you only need to know it to three significant figures:

$$N_A \approx 6.02 \times 10^{23}. \quad (10.1.1)$$

So 6.02×10^{23} of *what*? Well, of anything you like: apples, stars in the sky, burritos. However, the only *practical* use for N_A is to have a more convenient way of expressing the huge numbers of the tiny particles such as atoms or molecules that we deal with in chemistry. Avogadro's number is a *collective number*, just like a dozen. Students can think of 6.02×10^{23} as the "chemist's dozen".

Before getting into the use of Avogadro's number in problems, take a moment to convince yourself of the reasoning embodied in the following examples.

Things to understand about Avogadro's number

- It is a *number*, just as is "dozen", and thus is *dimensionless*.
- It is a *huge* number, far greater in magnitude than we can visualize
- Its practical use is limited to counting tiny things like atoms, molecules, "formula units", electrons, or photons.
- The value of N_A can be known only to the precision that the number of atoms in a measurable weight of a substance can be estimated. Because large numbers of atoms cannot be counted directly, a variety of ingenious indirect measurements have been made involving such things as Brownian motion and X-ray scattering.

The Mole: "A Dozen Eggs and a Mole of Sugar, Please"

The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly $6.02214076 \times 10^{23}$ particles, e.g., atoms, molecules, ions or electrons. The current definition was adopted in November 2018, revising its old definition based on the number of atoms in 12 grams of carbon-12 (^{12}C) (the isotope of carbon with relative atomic mass 12 Daltons by definition).

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.02×10^{23} .



Video 10.1.1 *How big is a mole?*

Converting Between Number of Atoms to Moles and Vice Versa

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of atoms of an element X, we can convert it into moles of by using the relationship

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

An example on the use of Avogadro's number as a conversion factor is given below for carbon.

✓ Example 10.1.1: Moles of Carbon

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

Solution

Example 10.1.1: Steps for Problem Solving how many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

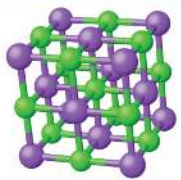
Steps for Problem Solving	Example
Identify the "given" information and what the problem is asking you to "find."	Given: 4.72×10^{24} C atoms Find: mol C
List other known quantities	$1 \text{ mol} = 6.022 \times 10^{23} \text{ C atoms}$
Prepare a concept map and use the proper conversion factor.	 $\frac{1 \text{ mol C}}{6.022 \times 10^{23} \text{ C atoms}}$
Cancel units and calculate.	$4.72 \times 10^{24} \text{ C atoms} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ C atoms}}$ $= 7.84 \text{ mol C}$
Think about your result.	The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calculation is rounded to three significant figures

Formula Mass

One skill needed in future chapters is the ability to determine the mass of the formula of various chemical substances. This quantity is called the formula mass. The **formula mass** is obtained by adding the masses of each individual atom in the formula of the substance. Because a proper formula is electrically neutral (with no net electrons gained or lost), the ions can be considered atoms for the purpose of calculating the formula mass.

Let us start by calculating the formula mass of sodium chloride (NaCl). This formula mass is the sum of the atomic masses of one sodium atom and one chlorine atom, which we find from the periodic table; here, we use the masses to two decimal places:

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
Formula mass					58.44



A table and diagram are shown. The table is made up of six columns and four rows. The header row reads: "Element," "Quantity," a blank space, "Average atomic mass (a m u)," a blank space and "Subtotal (a m u)." The first column contains the symbols "N a", "C l," and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers "1" and "1" as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers "22.99" and "35.45" as well as the merged cell. The fifth column contains the symbol "=" in each cell except for the last, merged cell. The sixth column contains the values "22.99," "35.45," and "58.44." There is a thick black line below the number "35.45." The merged cell under the first five columns reads "Formula mass." To the left of the table is a diagram of a chemical structure. The diagram shows green and purple spheres placed in an alternating pattern, making up the corners of eight stacked cubes to form one larger cube. The green spheres are slightly smaller than the purple spheres.

To two decimal places, the formula mass of NaCl is 58.44 amu.

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a **molecular mass**. Consider chloroform (CHCl_3), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the "anti-stick" polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
Molecular mass					119.37



A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space, and “Subtotal (a m u).” The first column contains the symbols “C,” “H,” “Cl” and a blank, merged cell that runs the width of the first five columns. The second column contains the numbers “1,” “1,” and “3” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “12.01,” “1.008,” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “12.01,” “1.008,” “106.35,” and “119.37.” There is a thick black line below the number 106.35. The merged cell under the first five columns reads “Molecular mass.” To the left of the table is a diagram of a molecule. Three green spheres are attached to a slightly smaller black sphere, which is also attached to a smaller white sphere. The green spheres lie beneath and to the sides of the black sphere while the white sphere is located straight up from the black sphere.

To two decimal places, the formula mass of CHCl_3 is 119.37 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. as shown in the example below for aluminum sulfate, $\text{Al}_2(\text{SO}_4)_3$.

✓ Example 10.1.2 Formula Mass for an Ionic Compound

Aluminum sulfate, $\text{Al}_2(\text{SO}_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $\text{Al}_2\text{S}_3\text{O}_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:



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src="/@api/deki/files/56171/CNX_Chem_03_01_alsulfatmass_img.jpg">

The formula mass for $\text{Al}_2(\text{SO}_4)_3$, is 342.14 amu.

? Exercise 10.1.1

Use the atomic masses (rounded to two decimal places) to determine the formula mass for each ionic compound.

- TiO_2
- AgBr
- $\text{Au}(\text{NO}_3)_3$
- $\text{Fe}_3(\text{PO}_4)_2$

Answer

- 79.87 amu
- 187.77 amu
- 383.0 amu

Summary

- The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly $6.02214076 \times 10^{23}$ particles, e.g., atoms, molecules, ions or electrons.
- Avogadro's number is related to moles of any substance X as follows:

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

- Formula masses of ionic and covalent compounds can be determined from the masses of the atoms in their formulas.

Contributors and Attributions

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- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a [Creative Commons Attribution License 4.0](https://creativecommons.org/licenses/by-nc-sa/4.0/) license. Download for free at <http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110>.
- Stephen Lower, Professor Emeritus ([Simon Fraser U.](https://www.sfu.ca/)) [Chem1 Virtual Textbook](#)
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10.2: Molar Mass- Mole-to-Mass and Mass-to-Mole Conversions

Learning Objectives

- Perform conversions between mass and moles of a substance.

Molar Mass

The **molar mass** of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. 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). 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:

$$\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \text{ and } \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \quad (10.2.1)$$

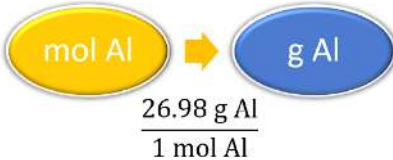
The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do "by eye."

✓ Example 10.2.1

What is the mass of 3.987 mol of Al?

Solution

Example 10.2.1: Steps for Problem Solving the mass of 3.987 moles of aluminum (Al).

Steps for Problem Solving	Example
Identify the "given" information and what the problem is asking you to "find."	Given: 3.987 mol of Al Find: g Al
List other known quantities	1 mol Al = 26.98 g Al
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$3.987 \cancel{\text{ mol Al}} \times \frac{26.98 \text{ g Al}}{1 \cancel{\text{ mol Al}}} = 107.6 \text{ g Al}$
Think about your result.	The calculated value makes sense because it is almost four times times the mass for 1 mole of aluminum. Our final answer is expressed to four significant figures.

? Exercise 10.2.1

How many moles are present in 100. g of Al? (Hint: you will have to use the other conversion factor we obtained for aluminum.)

Answer

3.71 g Al

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 10.2.1 is a chart for determining what conversion factor is needed, and Figure 10.2.2 is a flow diagram for the steps needed to perform a conversion.

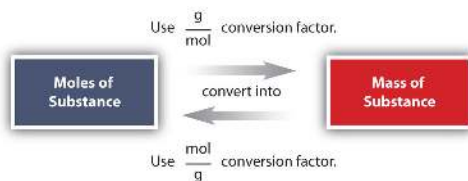


Figure 10.2.1 A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles.

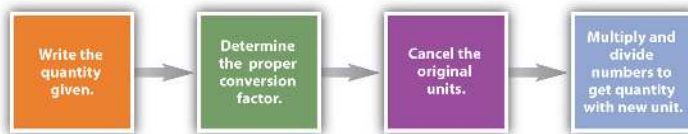


Figure 10.2.2 A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

Suppose that for a certain experiment you need 3.00 moles of calcium chloride (CaCl_2). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. Dimensional analysis will allow you to calculate the mass of CaCl_2 that you should measure as show in Example 10.2.3.

✓ Example 10.2.2: Calcium Chloride

Calculate the mass of 3.00 moles of calcium chloride (CaCl_2).



Figure 10.2.1: Calcium chloride is used as a drying agent and as a road deicer.

Solution

Example 10.2.2: Steps for Problem Solving the mass of 3.00 moles of calcium chloride (CaCl_2).

Steps for Problem Solving

Identify the "given" information and what the problem is asking you to "find."

List other known quantities

Prepare a concept map and use the proper conversion factor.

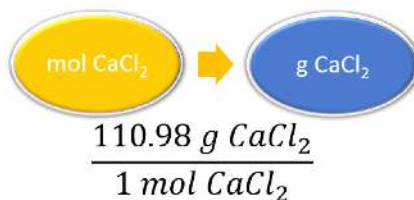
Cancel units and calculate.

Think about your result.

Example

Given: 3.00 moles of CaCl_2
Find: g CaCl_2

1 mol CaCl_2 = 110.98 g CaCl_2



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

? Exercise 10.2.2: Calcium Oxide

What is the mass of 7.50 mol of (calcium oxide) CaO ?

Answer:

420.60 g

✓ Example 10.2.3: Water

How many moles are present in 108 grams of water?

Solution

Example 10.2.3: Steps for Problem Solving moles in 108 grams of water?

Steps for Problem Solving

Identify the "given" information and what the problem is asking you to "find."

List other known quantities

Prepare a concept map and use the proper conversion factor.

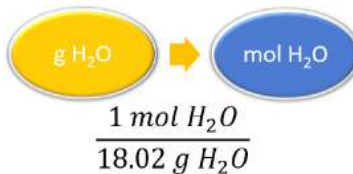
Cancel units and calculate.

Think about your result.

Example

Given: 108 g H_2O
Find: mol H_2O

1 mol H_2O = 18.02 g H_2O



$$108 \cancel{\text{ g H}_2\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \cancel{\text{ g H}_2\text{O}}} = 5.99 \text{ mol H}_2\text{O}$$

? Exercise 10.2.3: Nitrogen Gas

What is the mass of 7.50 mol of Nitrogen gas N_2 ?

Answer:

210 g

Summary

- Calculations involving conversions between moles of a substance and the mass of that substance are described.

Contributors and Attributions

- TextMap: The Basics of GOB Chemistry (Ball et al.)
- [Marisa Alviar-Agnew](#) ([Sacramento City College](#))
- Henry Agnew (UC Davis)

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CHAPTER OVERVIEW

11: Module 11 "Chemical Reactions" Textbook Readings

11.1: Mole-to-Mass and Mass-to-Mole Conversions in Chemical Equations

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11.1: Mole-to-Mass and Mass-to-Mole Conversions in Chemical Equations

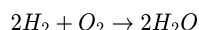
Learning Objectives

- Use a balanced chemical equation to determine molar relationships between substances.

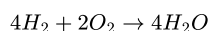
Molar Relationships in Chemical Equations

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.

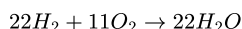
Consider the simple chemical equation



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

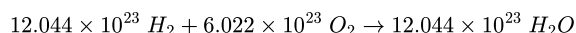


The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

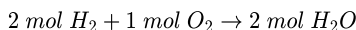


because 22:11:22 also reduces to 2:1:2.

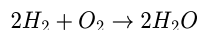
Suppose we want to use larger numbers. Consider the following coefficients:



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



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



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.”

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:

$$\begin{aligned} \frac{2 \text{ mol } H_2}{1 \text{ mol } O_2} &\text{ or } \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2} \\ \frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2} &\text{ or } \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} \\ \frac{2 \text{ mol } H_2}{2 \text{ mol } H_2O} &\text{ or } \frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2} \end{aligned}$$

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*.

✓ Example 11.1.5

How many moles of oxygen react with hydrogen to produce 27.6 mol of H₂O?

Solution

Example 11.1.5: Steps for Problem Solving how many moles of oxygen react with hydrogen to produce 27.6 mol of H₂O?

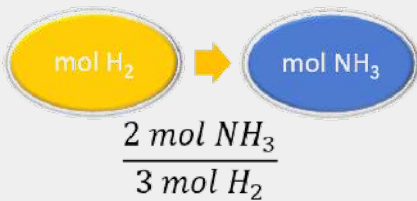
Steps for Problem Solving	Example
Find a balanced equation that describes the reaction	Unbalanced: $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$ Balanced: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
Identify the "given" information and what the problem is asking you to "find."	Given: moles H ₂ O Find: moles oxygen
List other known quantities	1 mol O ₂ = 2 mol H ₂ O
Prepare a concept map and use the proper conversion factor.	 $\frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}}$
Cancel units and calculate.	$\cancel{27.6 \text{ mol H}_2\text{O}} \times \frac{1 \text{ mol O}_2}{\cancel{2 \text{ mol H}_2\text{O}}} = 13.8 \text{ mol O}_2$ To produce 27.6 mol of H₂O, 13.8 mol of O₂ react.
Think about your result.	Since each mole of oxygen produces twice as many moles of water, it makes sense that the produced amount is greater than the reactant amount

✓ Example 11.1.6

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

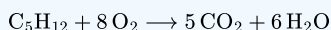
Solution

Example 11.1.6: Steps for Problem Solving how many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen

Steps for Problem Solving	Example
Find a balanced equation that describes the reaction	Unbalanced: $\text{N}_2 + \text{H}_2 \rightarrow \text{NH}_3$ Balanced: $\text{N}_2 + \underline{3}\text{H}_2 \rightarrow \underline{2}\text{NH}_3$
Identify the "given" information and what the problem is asking you to "find."	Given: $\text{H}_2 = 4.20 \text{ mol}$ Find: mol of NH_3
List other known quantities	$3 \text{ mol H}_2 = 2 \text{ mol NH}_3$
Prepare a concept map and use the proper conversion factor.	 $\frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2}$
Cancel units and calculate.	$\cancel{4.20 \text{ mol H}_2} \times \frac{2 \text{ mol NH}_3}{\cancel{3 \text{ mol H}_2}} = 2.80 \text{ mol NH}_3$ The reaction of 4.20 mol of hydrogen with excess nitrogen produces 2.80 mol of ammonia
Think about your result.	The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

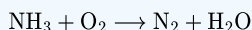
? Exercise 11.1.5

- a. Given the following balanced chemical equation,



how many moles of H_2O can be formed if 0.0652 mol of C_5H_{12} were to react?

- b. Balance the following unbalanced equation and determine how many moles of H_2O are produced when 1.65 mol of NH_3 react.



Answer a:

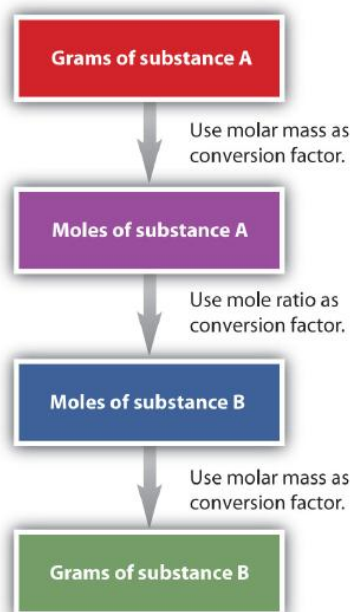
3.14 mol H_2O

Answer b:

$4\text{NH}_3 + 3\text{O}_2 \rightarrow 2\text{N}_2 + 6\text{H}_2\text{O}$; 2.48 mol H_2O

Mass Relationships in Chemical Equations

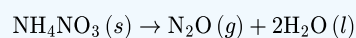
It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert 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.

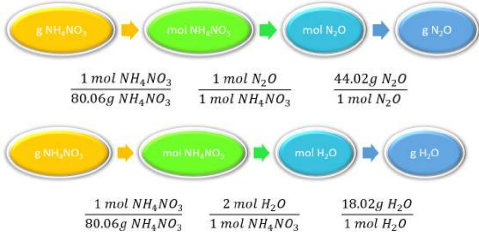
✓ Example 11.1.7: Decomposition of Ammonium Nitrate

Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.



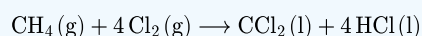
In a certain experiment, 45.7 g of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Example 11.1.7: Steps for Problem Solving the Decomposition of Ammonium Nitrate

Steps for Problem Solving	Example 11.1.2
Identify the "given" information and what the problem is asking you to "find."	Given: 45.7 g NH_4NO_3 Find: Mass N_2O = ? g Mass H_2O = ? g
List other known quantities	1 mol NH_4NO_3 = 80.06 g/mol 1 mol N_2O = 44.02 g/mol 1 mol H_2O = 18.02 g/mol 1 mol NH_4NO_3 to 1 mol N_2O to 2 mol H_2O
Prepare two concept maps and use the proper conversion factor.	
Cancel units and calculate.	$45.7 \text{ g } \text{NH}_4\text{NO}_3 \times \frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{80.06 \text{ g } \text{NH}_4\text{NO}_3} \times \frac{1 \text{ mol } \text{N}_2\text{O}}{1 \text{ mol } \text{NH}_4\text{NO}_3} \times \frac{44.02 \text{ g } \text{N}_2\text{O}}{1 \text{ mol } \text{N}_2\text{O}} = 25.1 \text{ g } \text{N}_2\text{O}$ $45.7 \text{ g } \text{NH}_4\text{NO}_3 \times \frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{80.06 \text{ g } \text{NH}_4\text{NO}_3} \times \frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{NH}_4\text{NO}_3} \times \frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} = 20.6 \text{ g } \text{H}_2\text{O}$
Think about your result.	The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

? Exercise 11.1.6: Carbon Tetrachloride

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows.



How many grams of HCl are produced by the reaction of 100.0g of CH_4 ?

Answer

908.7g HCl

Summary

- The balanced chemical reaction can be used to determine molar and mass relationships between substances.

Contributors and Attributions

- TextMap: The Basics of GOB Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

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CHAPTER OVERVIEW

12: Module 12 "Shapes of Molecules" Textbook Readings

[12.1: Molecular Shapes- VSEPR Theory](#)

[12.2: Polar and Nonpolar Molecules](#)

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12.1: Molecular Shapes- VSEPR Theory

Learning Objective

- Determine the shape of simple molecules and polyatomic ions.

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

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

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible— 180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2 :



Figure 12.1.1 Beryllium hydride and carbon dioxide bonding.

The two molecules, shown in the figure below in a "ball and stick" model.



Figure 12.1.2 Beryllium hydride and carbon dioxide models.

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle— 120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3 :

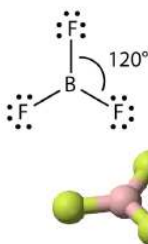


Figure 12.1.3 Boron trifluoride bonding.

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2 :

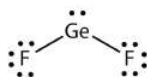


Figure 12.1.4: Germanium difluoride bonding.

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 12.1.1 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

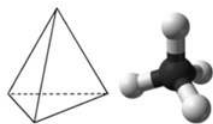


Figure 12.1.5 Tetrahedral structure of methane.

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

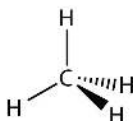


Figure 12.1.6 Methane bonding.

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.

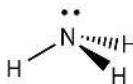


Figure 12.1.7 Ammonia bonding.

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 12.1.8 Water bonding.

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

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 12.1.9.



Figure 12.1.9 Lewis electron dot diagram of formaldehyde.

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.

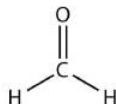


Figure 12.1.10: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule. Figure 12.1.11 illustrates several representations of the water, ammonia, and methane molecules.

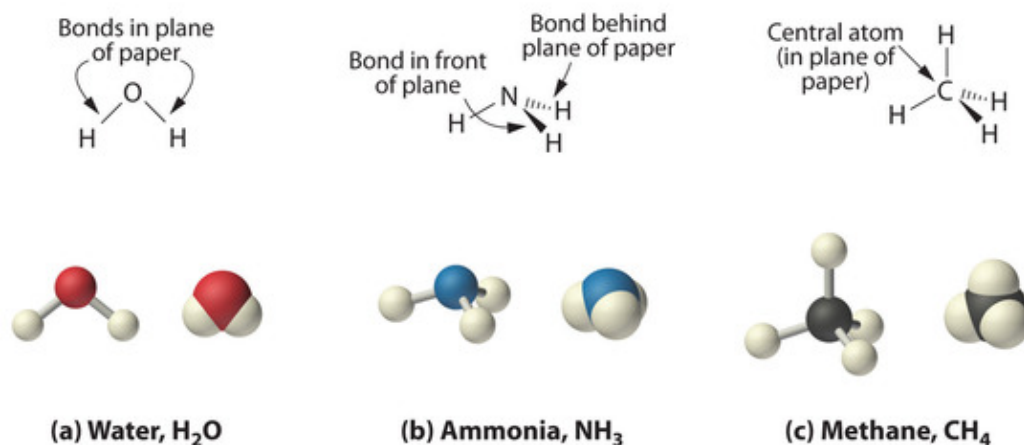


Figure 12.1.11 The three-dimensional structures, ball and stick models, and space filling models of water, ammonia, and methane. (a) Water is a V-shaped molecule, in which all three atoms lie in a plane. (b) In contrast, ammonia has a pyramidal structure, in which the three hydrogen atoms form the base of the pyramid and the nitrogen atom is at the vertex. (c) The four hydrogen atoms of methane form a tetrahedron; the carbon atom lies in the center.

Table 12.1.1 summarizes the shapes of molecules based on their number of electron groups and surrounding atoms and the common shapes are shown in figure 12.1.1.

Table 12.1.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

Figure 12.1.12: Common structures for molecules and polyatomic ions that consist of a central atom bonded to two or three other atoms.

✓ Example 12.1.1:

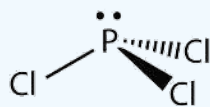
What is the approximate shape of each molecule?

- PCl_3
- NOF

Solution

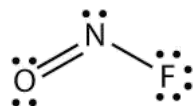
The first step is to draw the Lewis structure of the molecule.

- 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 12.1.1

What is the approximate molecular shape of CH_2Cl_2 ?

Answer

Tetrahedral

? Exercise 12.1.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. Hint, hydrogen is a terminal atom.

Answer

Trigonal planar about both central C atoms

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.)
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12.2: 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 CCl_4 is nonpolar Figure (12.2.1. Another non polar molecule shown below is boron trifluoride, BF_3 . BF_3 is a trigonal planar molecule and all three peripheral atoms are the same.

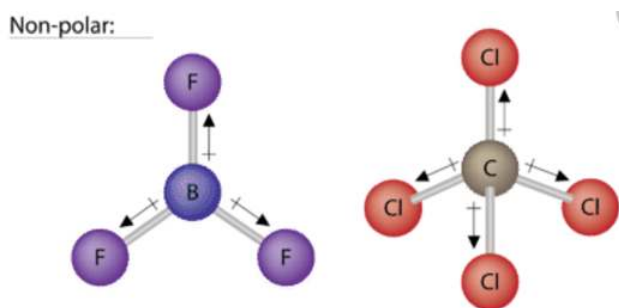
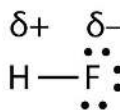


Figure 12.2.1 Some examples of nonpolar molecules based on molecular geometry (BF_3 and CCl_4).

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 $\delta-$ (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 $\delta+$. 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 12.2.2 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.

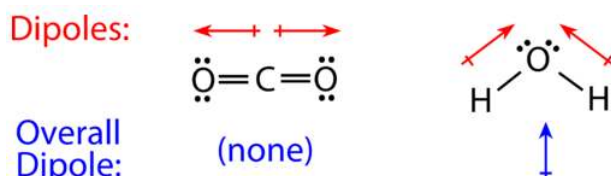


Figure 12.2.3 The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO_2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

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.

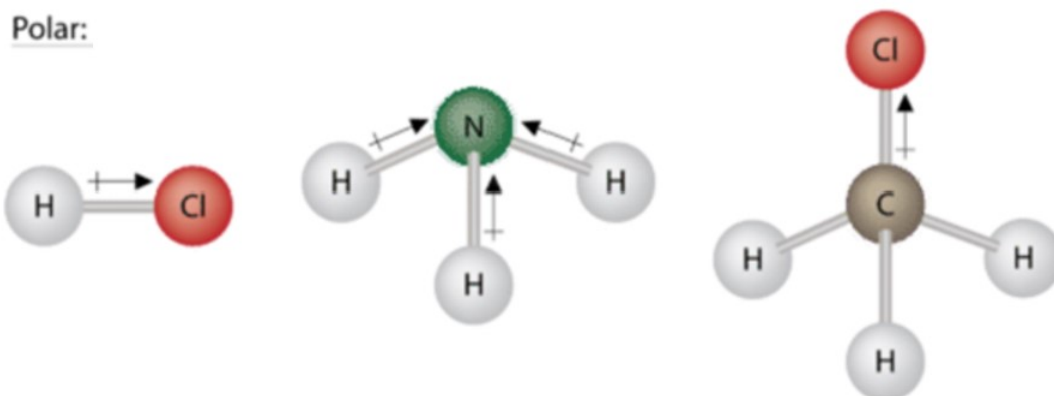


Figure 12.2.4 Some examples of polar molecules based on molecular geometry (HCl , NH_3 and CH_3Cl).

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 12.2.1:

Label each of the following as polar or nonpolar.

- a. Water, H_2O : 

b. Methanol, CH_3OH :



c. Hydrogen Cyanide, HCN : $\text{H}-\text{C}\equiv\text{N}$

d. Oxygen, O_2 : $\text{:}\ddot{\text{O}}=\ddot{\text{O}}\text{:}$

e. Propane, C_3H_8 :



Solution

- Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
- Methanol is polar. This is not a symmetric molecule. The $-\text{OH}$ side is different from the other 3 $-\text{H}$ sides.
- Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
- Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
- 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 12.2.1

Label each of the following as polar or nonpolar.

a. SO_3

b. NH_3

Answer a

non polar

Answer b

polar

Summary

- Non polar 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

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CHAPTER OVERVIEW

13: Module 13 "Solids, Liquids, and Gases" Textbook Readings

13.1: Solids, Liquids, and Gases

13.2: Intermolecular Forces

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13.1: Solids, Liquids, and Gases

Learning Objectives

- Describe the different states of matter.
- Describe the different changes that occur between solid, liquid, and gas.

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

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

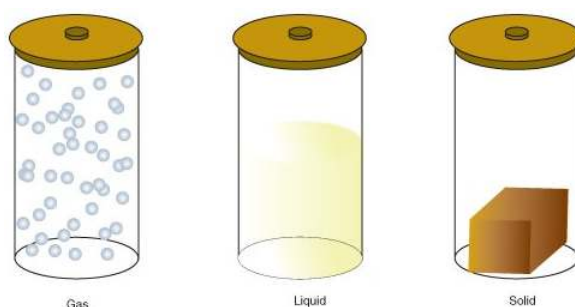


Figure 13.1.1: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

In Figure 13.1.2, all three phases are shown at the submicroscopic level in animations. Notice how the movement and freedom of molecules steadily increases as attractive forces decrease from solid to liquid to gas phase.

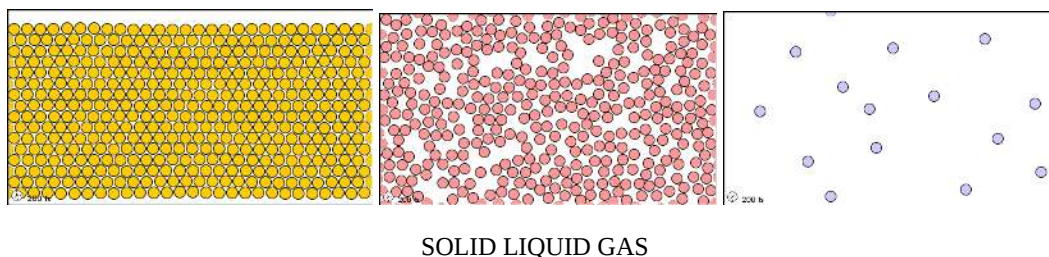


Figure 13.1.2 Animation of all three phases at the submicroscopic level.

The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Figures 13.1.3 and 13.1.4 show the differences among solids, liquids, and gases at the molecular level, while Table 13.1.1 lists the different characteristics of these states.

Table 13.1.1 Characteristics of the Three States of Matter

Characteristic	Solid	Liquid	Gas
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

✓ Example 13.1.1

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

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

Solution

- This statement describes either the liquid state or the solid state.
- This statement describes either the liquid state or the gas state.
- This statement describes the liquid state.

? Exercise 13.1.1

What state or states of matter does each statement describe?

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

Answer a

solid

Answer b

gas

Answer c

liquid or gas

Phase Changes

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

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C . At 0°C , both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O . If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O . A similar process can occur at 100°C : adding heat increases the amount of gaseous H_2O , while removing heat increases the amount of liquid H_2O (Figure 13.1.3).

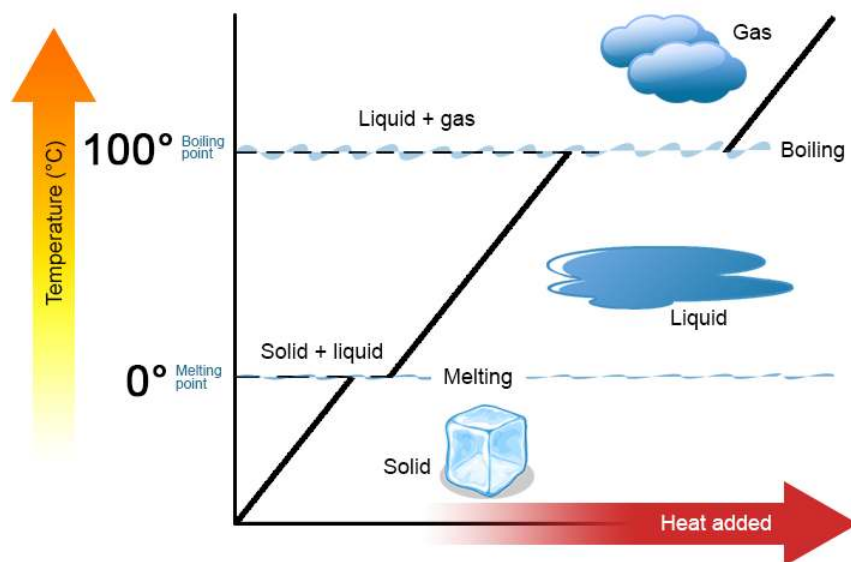


Figure 13.1.3: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0 °C, the *melting point*. At this point, the *phase change*, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).

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

Second, as shown in Figure 13.1.3, **the temperature of a substance does not change** as the substance goes **from one phase to another**. In other words, phase changes are isothermal (isothermal means “constant temperature”). Again, consider H₂O as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of H₂O can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

Summary of Phase Changes

All of the changes of state that occur between solid, liquid, and gas are summarized in the diagram in the figure below. **Freezing** is the opposite of melting, and both represent the equilibrium between the solid and liquid states. **Evaporation** occurs when a liquid turns to a gas. **Condensation** is the opposite of vaporization, and both represent the equilibrium between the liquid and gas states. **Deposition** is the opposite of sublimation, and both represent the equilibrium between the solid and gas states.

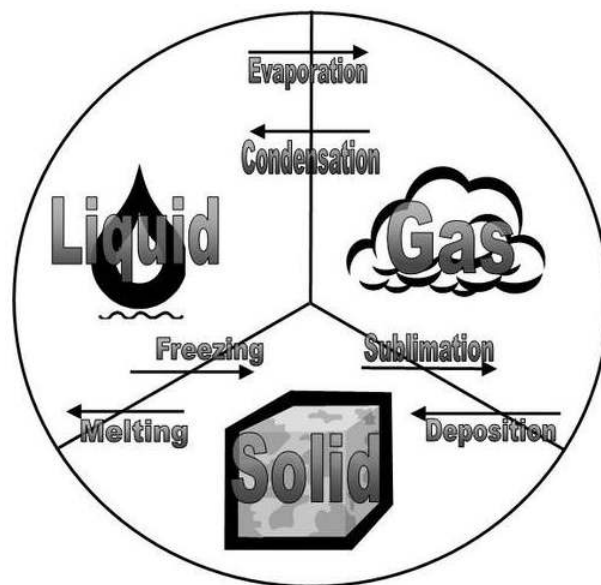


Figure 13.1.4: Solid, liquid, and gas states with the terms for each change of state that occurs between them.

Liquid evaporates into gas in freezes into a solid; gas condenses into a liquid depositions into a solid; solids melt into a liquid and sublimation into a gas

Summary

- Solids, liquids, and gases are different phases that have their own unique properties.
- A change of state (or phase change) can be brought about by putting heat into a system or removing it from the system.

Contributors and Attributions

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13.2: Intermolecular Forces

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

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not *intramolecular* forces. *Intramolecular* forces are those *within* the molecule that keep the molecule together, for example, the bonds between the atoms. *Intermolecular* forces are the attractions *between* molecules, which determine many of the physical properties of a substance. Figure 13.2.1 illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

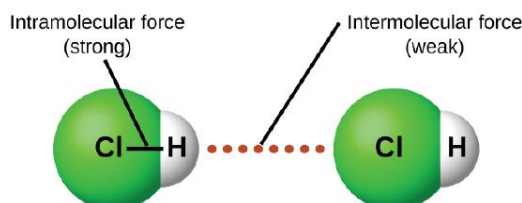


Figure 13.2.1 Intramolecular forces keep a molecule intact. Intermolecular forces hold multiple molecules together and determine many of a substance's properties.

An image is shown in which two molecules composed of a green sphere labeled “Cl” connected on the right to a white sphere labeled “H” are near one another with a dotted line labeled “Intermolecular force (weak)” drawn between them. A line connects the two spheres in each molecule and the line is labeled “Intramolecular force (strong).”

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.

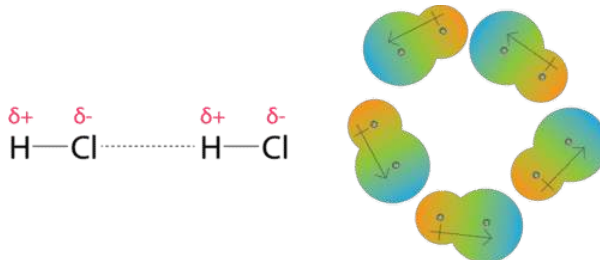


Figure 13.2.2 Dipole-dipole forces result from the attraction between the positive end of one dipole and the negative end of a neighboring dipole. Dipole-dipole forces are similar to ionic bonds, but because they involve only partial charges, they are much weaker.

Dipole-dipole forces in hydrogen chloride

The strength of the intermolecular interactions increases as the dipole moment of the molecules increases, as shown in Table 13.2.1.

Table 13.2.1: Relationships between the Dipole Moment and the Boiling Point for Organic Compounds of Similar Molar Mass

Compound	Molar Mass (g/mol)	Dipole Moment (D)	Boiling Point (K)
C ₃ H ₆ (cyclopropane)	42	0	240
CH ₃ OCH ₃ (dimethyl ether)	46	1.30	248
CH ₃ CN (acetonitrile)	41	3.9	355

The attractive energy between two ions is proportional to $1/r$, whereas the attractive energy between two dipoles is proportional to $1/r^6$.

✓ Example 13.2.1 Dipole-Dipole Forces and Their Effects

Predict which will have the higher boiling point: N₂ or CO. Explain your reasoning.

Solution

CO and N₂ are both diatomic molecules with masses of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N₂ is nonpolar, its molecules cannot exhibit dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N₂ molecules, so CO is expected to have the higher boiling point.

A common method for preparing oxygen is the decomposition

? Exercise 13.2.1

Predict which will have the higher boiling point: ICl or Br₂. Explain your reasoning.

Answer

ICl. ICl and Br₂ have similar masses (~160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br₂ is nonpolar and does not. The relatively stronger dipole-dipole attractions require more energy to overcome, so ICl will have the higher boiling point

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.

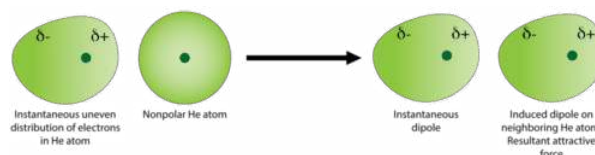


Figure 13.2.3 Random fluctuations in the electron density within the electron cloud of a helium atom results in a short-lived ("instantaneous") dipole. The attractive force between instantaneous dipoles and the resulting induced dipoles in neighboring molecules is called the London dispersion force.

From left to right: instantaneous uneven distribution of electrons in helium atom; nonpolar helium atom; instantaneous dipole; induce dipole on neighboring helium atom press resulting in an attractive 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. Larger and heavier atoms and molecules

exhibit stronger dispersion forces than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in Table 13.2.1.

Table 13.2.2: Melting and Boiling Points of the Halogens

Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K
chlorine, Cl_2	71 g/mol	99 pm	172 K	238 K
bromine, Br_2	160 g/mol	114 pm	266 K	332 K
iodine, I_2	254 g/mol	133 pm	387 K	457 K
astatine, At_2	420 g/mol	150 pm	575 K	610 K

The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther from the nuclei than in a smaller atom. Thus, they are less tightly held and can more easily form the temporary dipoles that produce the attraction. The measure of how easy or difficult it is for another electrostatic charge (for example, a nearby ion or polar molecule) to distort a molecule's charge distribution (its electron cloud) is known as polarizability. A molecule that has a charge cloud that is easily distorted is said to be very polarizable and will have large dispersion forces; one with a charge cloud that is difficult to distort is not very polarizable and will have small dispersion forces.

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

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 13.2.4 shows 2,2-dimethylpropane (neopentane) and n -pentane, both of which have the empirical 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 ($9.5^\circ C$) is more than $25^\circ C$ lower than the boiling point of n -pentane ($36.1^\circ C$).

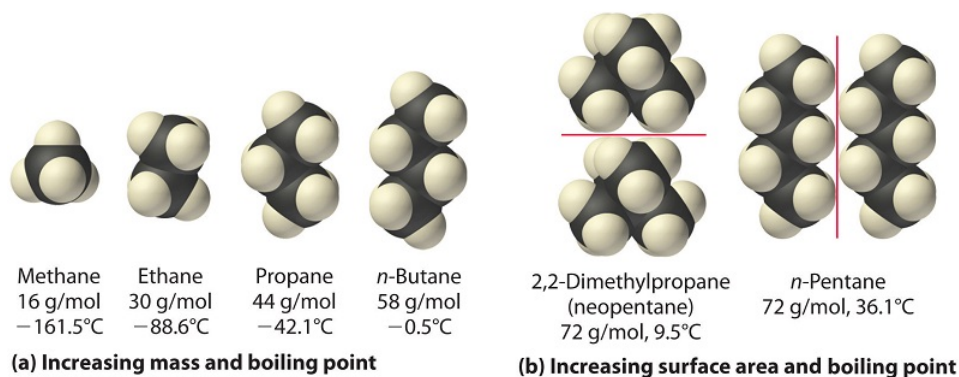


Figure 13.2.4 Mass and Surface Area Affect the Strength of London Dispersion Forces. (a) In this series of four simple alkanes, larger molecules have stronger London forces between them than smaller molecules and consequently higher boiling points. (b) Linear n -pentane molecules have a larger surface 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 13.2.2 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: compounds

Asked for: order of increasing boiling points

Strategy:

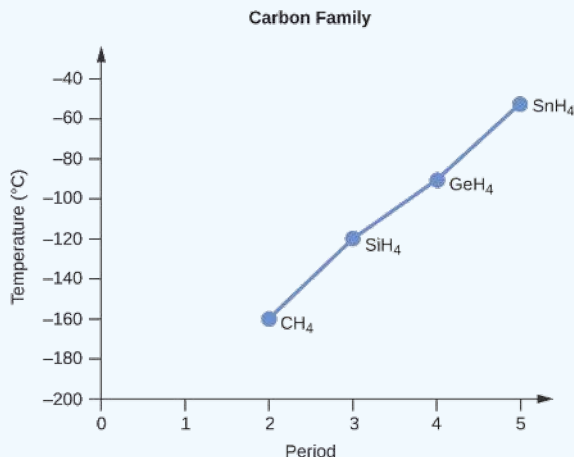
Determine the intermolecular forces in the compound. Also, compare the molar masses and surface area of compounds with similar types of intermolecular force. The substance with the weakest forces and the lowest molar masses and least surface area have the lowest boiling point.

Solution

Applying the skills acquired in the chapter on chemical bonding and molecular geometry, all of these compounds are predicted to be nonpolar, so they may experience only dispersion forces: the smaller the molecule, the less polarizable and the weaker the dispersion forces; the larger the molecule, the larger the dispersion forces. The molar masses of CH_4 , SiH_4 , GeH_4 , and SnH_4 are approximately 16 g/mol, 32 g/mol, 77 g/mol, and 123 g/mol, respectively. Therefore, CH_4 is expected to have the lowest boiling point and SnH_4 the highest boiling point. The ordering from lowest to highest boiling point is expected to be



A graph of the actual boiling points of these compounds versus the period of the group 14 elements shows this prediction to be correct:



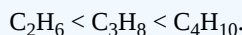
A line graph, titled "Carbon Family," is shown where the y-axis is labeled "Temperature, (degree sign C)" and has values of "negative 200" to "negative 40" from bottom to top in increments of 20. The x-axis is labeled "Period" and has values of "0" to "5" in increments of 1. The first point on the graph is labeled "C H subscript 4" and is at point "2, negative 160." The second point on the graph is labeled "S i H subscript 4" and is at point "3, negative 120" while the third point on the graph is labeled "G e H subscript 4" and is at point "4, negative 100." The fourth point on the graph is labeled "S n H subscript 4" and is at point "5, negative 60."

? Exercise 13.2.2

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

Answer

All of these compounds are nonpolar and only have London dispersion forces: the larger the molecule, the larger the dispersion forces and the higher the boiling point. The ordering from lowest to highest boiling point is therefore



Hydrogen Bonds

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

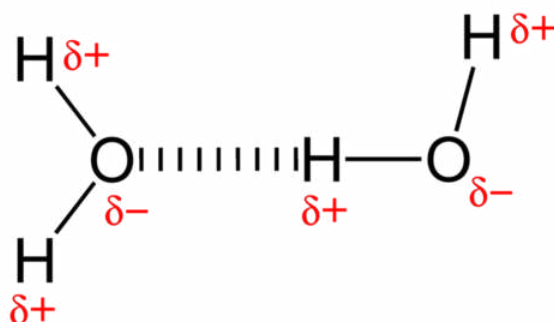


Figure 13.2.5 A hydrogen bond in water occurs between the hydrogen atom of one water molecule and the lone pair of electrons on the oxygen atom of a neighboring water molecule.

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

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

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

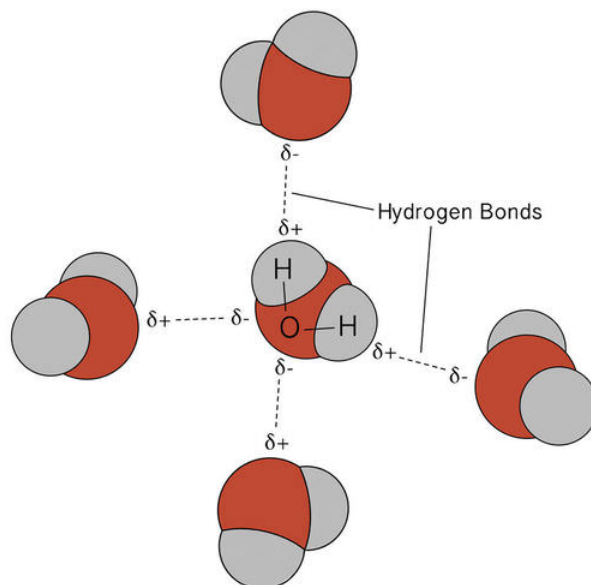


Figure 13.2.6 Multiple hydrogen bonds occur simultaneously in water because of its bent shape and the presence of two hydrogen atoms per molecule.

In the liquid state, the hydrogen bonds of water can break and reform as the molecules flow from one place to another. When water is cooled, the molecules begin to slow down. Eventually, when water is frozen to ice, the hydrogen bonds become more rigid and form a well-defined network (Figure 13.2.7). 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.

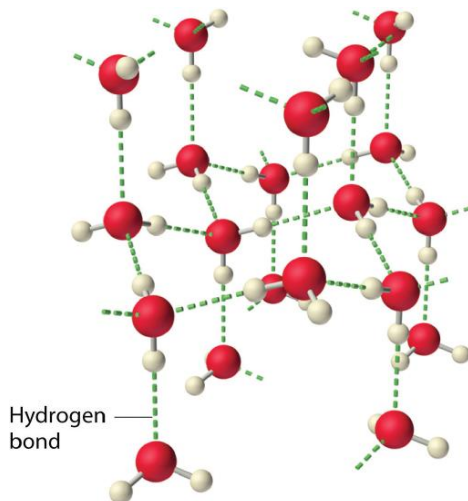


Figure 13.2.7 The Hydrogen-Bonded Structure of Ice.

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.

Strength of Intermolecular Forces

Molecules with hydrogen atoms bonded to electronegative atoms such as O, N, and F tend to exhibit unusually strong intermolecular interactions due to hydrogen bonds as illustrated for the covalent hydrides of elements of groups 14–17 in Figure 13.2.8. These result in much higher boiling points than are observed for substances in which dipole-dipole forces or London dispersion forces dominate. All the polar molecules (with predominantly dipole-dipole forces) in group 16 (H_2S to H_2Te) have lower boiling points than H_2O (with hydrogen bonds). Methane and its heavier congeners in group 14 form a series whose boiling points increase smoothly with increasing molar mass. This is the expected trend in nonpolar molecules, for which London dispersion forces are the exclusive intermolecular forces. In contrast, the hydrides of the lightest members of groups 15–17 have boiling points that are more than 100°C greater than predicted on the basis of their molar masses. The effect is most dramatic for water: if we extend the straight line connecting the points for H_2Te and H_2Se to the line for period 2, we obtain an estimated boiling point of -130°C for water! Imagine the implications for life on Earth if water boiled at -130°C rather than 100°C .

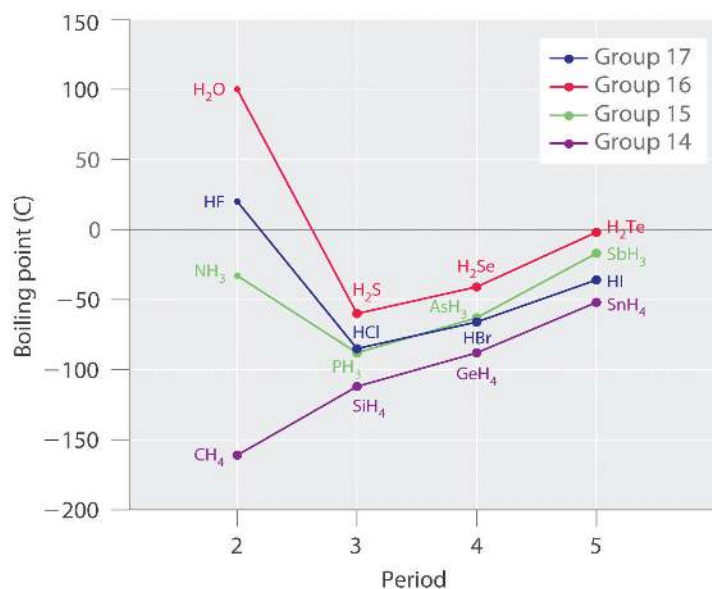


Figure 13.2.8 The Effects of Hydrogen Bonding on Boiling Points. These plots of the boiling points of the covalent hydrides of the elements of groups 14–17 show that the boiling points of the lightest members of each series for which hydrogen bonding is possible (HF, NH₃, and H₂O) are anomalously high for compounds with such low molecular masses.

✓ Example 13.2.3

Considering CH₃OH, C₂H₆, Xe, and (CH₃)₃N, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Given: compounds

Asked for: formation of hydrogen bonds and structure

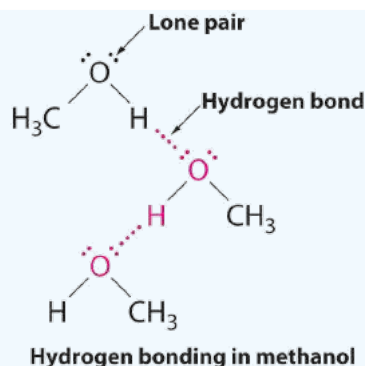
Strategy:

- 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.
- 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 (C₂H₆), and trimethylamine [(CH₃)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 (CH₃OH), 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:

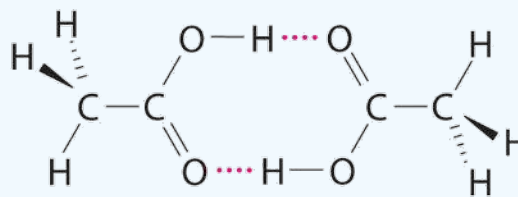
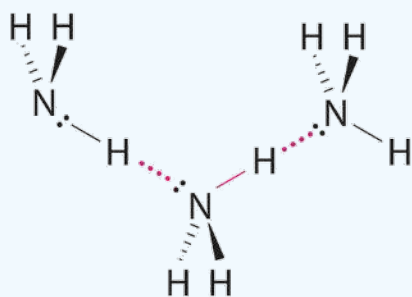


? Exercise 13.2.3

Considering $\text{CH}_3\text{CO}_2\text{H}$, $(\text{CH}_3)_3\text{N}$, NH_3 , and CH_3F , which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Answer

$\text{CH}_3\text{CO}_2\text{H}$ and NH_3 ;



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.

✓ Example 13.2.4: Buckyballs

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

Given: compounds

Asked for: order of increasing boiling points

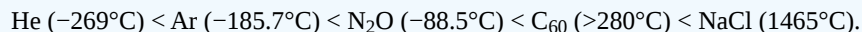
Strategy:

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:

Electrostatic interactions are strongest for an ionic compound, 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 molar mass (for London dispersion forces). Helium is nonpolar and by far the lightest, so it should have the lowest boiling point. Argon and N_2O have very similar molar masses (40 and 44 g/mol, 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 its molar mass is 720 g/mol, much greater than that of Ar or N_2O . Because the boiling points of nonpolar substances increase rapidly with molecular mass, 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:



? Exercise 13.2.4

Arrange 2,4-dimethylheptane, Ne, CS_2 , Cl_2 , and KBr in order of decreasing boiling points.

Answer

$\text{KBr} (1435^\circ\text{C}) > \text{2,4-dimethylheptane} (132.9^\circ\text{C}) > \text{CS}_2 (46.6^\circ\text{C}) > \text{Cl}_2 (-34.6^\circ\text{C}) > \text{Ne} (-246^\circ\text{C})$

✓ Example 13.2.5:

Identify the most significant intermolecular force in each substance.

- C_3H_8
- CH_3OH
- H_2S

Solution

- Although C–H bonds are polar, they are only minimally polar. The most significant intermolecular force for this substance would be dispersion forces.
- This molecule has an H atom bonded to an O atom, so it will experience hydrogen bonding.
- 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 13.2.5

Identify the most significant intermolecular force in each substance.

- HF
- HCl

Answer a

hydrogen bonding

Answer b

dipole-dipole interactions

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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- [Marisa Alviar-Agnew](#) ([Sacramento City College](#))
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CHAPTER OVERVIEW

14: Module 14 "Solutions" Textbook Readings

[14.1: Solutions](#)

[14.2: Properties of Solutions](#)

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14.1: Solutions

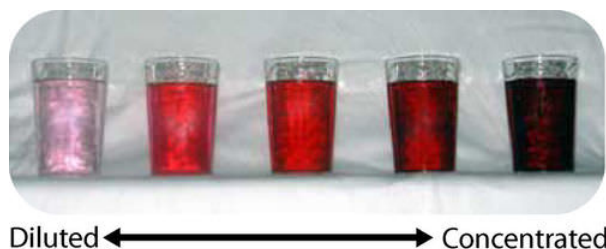
Learning Objectives

- Define terms involving solutions.
- Describe the amount of solute in a solution in various concentration units.

When one substance dissolves into another, a solution is formed. A **solution** is a homogenous mixture consisting of a solute dissolved into a solvent. The **solute** is the substance that is being dissolved, while the **solvent** is the dissolving medium. Solutions can be formed with many different types and forms of solutes and solvents. In this chapter, we will focus on solution where the solvent is water. An **aqueous solution** is water that contains one or more dissolved substance. The dissolved substances in an aqueous solution may be solids, gases, or other liquids.

In order to be a true solution, a mixture must be stable. When sugar is fully dissolved into water, it can stand for an indefinite amount of time, and the sugar will not settle out of the solution. Further, if the sugar-water solution is passed through a filter, it will remain with the water. This is because the dissolved particles in a solution are very small, usually less than 1 nm in diameter. Solute particles can be atoms, ions, or molecules, depending on the type of substance that has been dissolved. Water typically dissolves most ionic compounds and polar molecules. Nonpolar molecules, such as those found in grease or oil, do not dissolve in water.

Solution Concentrations



Solutions of a red dye in water from the most dilute (on the left) to the most concentrated (on the right).

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. What you have to do is take the frozen juice from inside these containers and usually empty it into 3 or 4 times the container size full of water to mix with the juice concentrate and make your container of juice. Therefore, you are diluting the concentrated juice. When we talk about solute and solvent, the concentrated solution

has a lot of solute versus the dilute solution that would have a smaller amount of solute.

The terms "concentrated" and "dilute" provide qualitative methods of describing concentration. Although qualitative observations are necessary and have their place in every part of science, including chemistry, we have seen throughout our study of science that there is a definite need for quantitative measurements in science. This is particularly true in solution chemistry. In this section, we will explore some quantitative methods of expressing solution concentration.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms whose meanings depend on various factors. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution and we know exactly by how much.

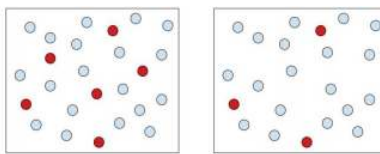


Figure 14.1.2 The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).

Molarity

One 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.

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \quad (14.1.1)$$

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 $[\text{Ag}^+]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to calculate 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 therefore

$$\frac{1.5 \text{ mol NaCl}}{0.500 \text{ L solution}} = 3.0 \text{ M NaCl}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” You would 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?

First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} = 0.614 \text{ mol HCl}$$

Now we can use the definition of molarity to determine a concentration:

$$M = \frac{0.614 \text{ mol HCl}}{1.56 \text{ L solution}} = 0.394 \text{ M HCl}$$

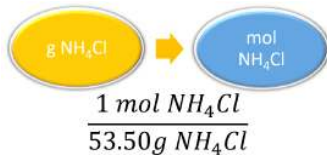
Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

✓ Example 14.1.1

A solution is prepared by dissolving 42.23 g of NH_4Cl into enough water to make 500.0 mL of solution. Calculate its molarity.

Solution

Example 14.1.1: Steps for Problem Solving Molarity

Steps for Problem Solving	Example
Identify the "given" information and what the problem is asking you to "find."	Given: Mass = 42.23 g NH_4Cl Volume solution = 500.0 mL = 0.5000 L Find: Molarity = ? M
List other known quantities.	Molar mass NH_4Cl = 53.50 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. $M = \frac{\text{mol } \text{NH}_4\text{Cl}}{\text{L solution}}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23 \text{ g } \cancel{\text{NH}_4\text{Cl}} \times \frac{1 \text{ mol } \cancel{\text{NH}_4\text{Cl}}}{53.50 \text{ g } \cancel{\text{NH}_4\text{Cl}}} = 0.7893 \text{ mol } \text{NH}_4\text{Cl} \quad (14.1.2)$ $\frac{0.7893 \text{ mol } \text{NH}_4\text{Cl}}{0.5000 \text{ L solution}} = 1.579 \text{ M} \quad (14.1.3)$
Think about your result.	The molarity is 1.579 M, meaning that a liter of the solution would contain 1.579 mol NH_4Cl . Four significant figures are appropriate.

? Exercise 14.1.1

What is the molarity of a solution made when 66.2 g of $\text{C}_6\text{H}_{12}\text{O}_6$ are dissolved to make 235 mL of solution?

Answer

1.57 M $\text{C}_6\text{H}_{12}\text{O}_6$

? Exercise 14.1.2

What is the concentration, in mol/L, where 137 g of NaCl has been dissolved in enough water to make 500 mL of solution?

Answer

4.69 M NaCl

Figure 14.1.1 illustrates how a solution is prepared with the measured amount of a solid solute and the desired final volume of the solution.

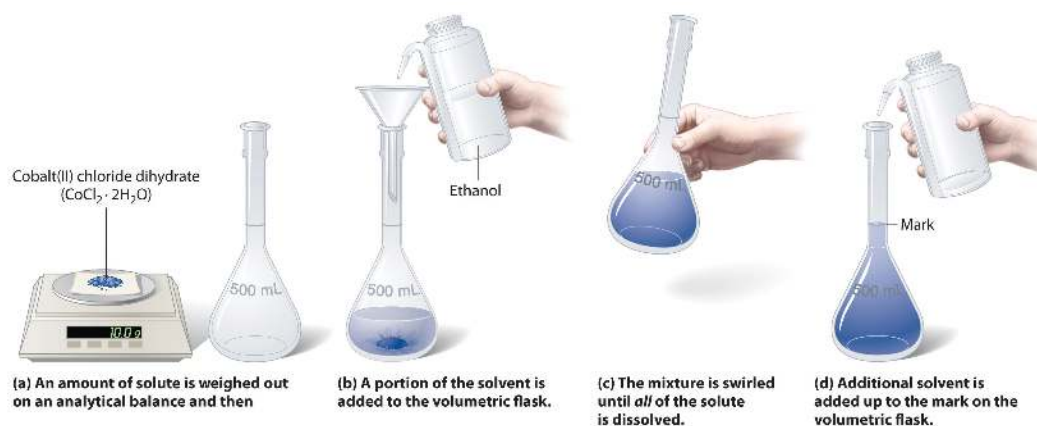


Figure 14.1.3: Preparation of a solution of known concentration using a solid solute.

(a) An amount of solute is weighed out on an analytical balance, (b) a portion of the solvent is added to the volumetric flask. (c) the mixture is swirled until all of the solute is dissolved. (d) additional solvent is added up to the mark on the volumetric flask.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Concentration of solutions can be more precisely expressed in terms of molarity.

Contributors and Attributions

-
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14.2: Properties of Solutions

Learning Objectives

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

A **solution** is another name for a homogeneous mixture. A **mixture** as a material composed of two or more substances. In a solution, the combination is so intimate that the



Figure 14.2.1 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).

different substances cannot be differentiated by sight, even with a microscope. Compare, for example, a mixture of salt and pepper and another mixture consisting of salt and water. In the first mixture, we can readily see individual grains of salt and the flecks of pepper. A mixture of salt and pepper is not a solution. However, in the second mixture, no matter how carefully we look, we cannot see two different substances. Salt dissolved in water is a solution.

The major component of a solution, called the **solvent**, is typically the same phase as the solution itself. Each minor component of a solution (and there may be more than one) is called the **solute**. In most of the solutions we will describe in this textbook, there will be no ambiguity about whether a component is the solvent or the solute.) For example, in a solution of salt in water, the solute is salt, and solvent is water Figure 14.2.1.

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table 14.2.1 lists some common types of solutions, with examples of each.

Table 14.2.1 Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol ($\text{C}_2\text{H}_5\text{OH}$) in H_2O (alcoholic beverages)
liquid	solid	salt water
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	$\text{Hg}(\ell)$ in dental fillings
solid	solid	steel alloys

✓ Example 14.2.1: 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 14.2.1

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

Answer

solute: $HCl(g)$; solvent: CH_3OH

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, an nonpolar molecules, does dissolve in water and it is this oxygen that the fish take in through their gills. Or, one more example of a nonpolar compound that dissolves in water is the reason we can enjoy carbonated sodas. 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 14.2.2 Summary of Solubilities

Solute	Solvent	Is Solution Formed?
Polar Covalent	Polar	yes
Non-polar Covalent	Non-polar	yes
Polar Covalent	Non-polar	no
Non-polar Covalent	Polar	no
Ionic	Polar	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



Figure 14.2.1 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).

the previous section). 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. *Liquids that dissolve in one another in all proportions are said to be **miscible**. Liquids that do not dissolve in one another are called **immiscible**.*

Figure 14.2.2 Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

✓ Example 14.2.2: Polar and Nonpolar Solvents

Would I_2 be more soluble in CCl_4 or H_2O ? Explain your answer.

Solution

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

? Exercise 14.2.2

Would C_3H_7OH be more soluble in CCl_4 or H_2O ? Explain your answer.

Answer

H_2O because both experience hydrogen bonding

✓ Example 14.2.3

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

- methanol (CH_3OH)
- sodium sulfate (Na_2SO_4)
- octane (C_8H_{18})

Solution

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

- 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*.
- 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 14.2.3

Toluene ($\text{C}_6\text{H}_5\text{CH}_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

- water (H_2O)
- sodium sulfate (Na_2SO_4)
- octane (C_8H_{18})

Answer

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

The Dissolving Process

Water typically dissolves most ionic compounds and polar molecules. Nonpolar molecules, such as those found in grease or oil, do not dissolve in water. We will first examine the process that occurs when an ionic compound, such as table salt (sodium chloride), dissolves in water.

Water molecules move about continuously due to their kinetic energy. When a crystal of sodium chloride is placed into water, the water's molecules collide with the crystal lattice. Recall that the crystal lattice is composed of alternating positive and negative ions. Water is attracted to the sodium chloride crystal because water is polar; it has both a positive and a negative end. The positively charged sodium ions in the crystal attract the oxygen end of the water molecules because they are partially negative. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules because they are partially positive. The action of the polar water molecules takes the crystal lattice apart (see figure below).

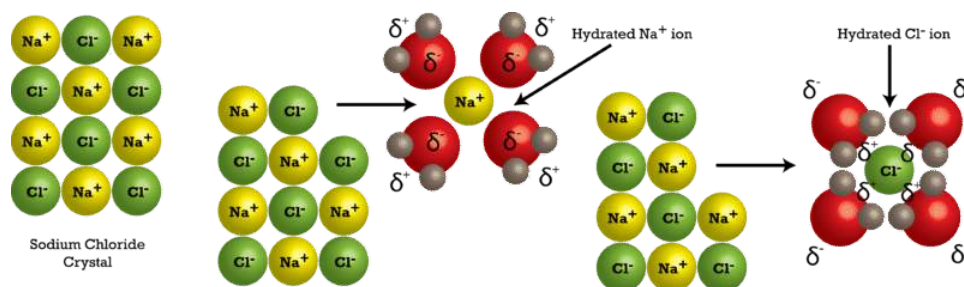


Figure 14.2.3: The process of an ionic sodium chloride crystal being dissolved and hydrated by water.

After coming apart from the crystal, the individual ions are then surrounded by solvent particles in a process called **solvation**. Note in the figure above that the individual Na^+ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. Thus, numerous **ion-dipole** interactions are formed. **Hydration** is the process of solute particles being surrounded by water molecules arranged in a specific manner. Hydration helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and forming a precipitate.

Table sugar is made of the molecular compound sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$). Solid sugar consists of individual sugar molecules held together by intermolecular attractive forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the attractive forces, but it *does not break the covalent bonds* between the carbon, hydrogen, and oxygen atoms. Dissolved sugar molecules are also hydrated. The hydration shell around a molecule of sucrose is arranged so that its partially negative oxygen atoms are near the partially positive hydrogen atoms in the solvent, and vice versa.

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.
- Liquids that dissolve in one another in all proportions are said to be **miscible**.
- Liquids that do not dissolve in one another are called **immiscible**.

- **Ion-dipole** interactions are formed when ionic compounds dissolve in water.

Contributors and Attributions

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- Henry Agnew (UC Davis)
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CHAPTER OVERVIEW

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[15.1: The pH Scale](#)

[15.2: Acid-Base Neutralization Reactions](#)

[15.3: Acids and Bases](#)

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15.1: The pH Scale

Learning Objectives

- Define pH .
- Interpret the pH of solutions.

Molar concentration values of hydrogen $[H^+]$ can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions. This is known as the pH scale. The range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to make a quick determination whether a given aqueous solution is acidic, basic, or neutral.

pH is a *logarithmic* scale. A solution that has a pH of 1.0 has 10 times the $[H^+]$ as a solution with a pH of 2.0, which in turn has 10 times the $[H^+]$ as a solution with a pH of 3.0 and so forth.

Warning: The pH scale has no limits

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H^+]$, we can summarize as follows:

- If $pH < 7$, then the solution is acidic.
- If $pH = 7$, then the solution is neutral.
- If $pH > 7$, then the solution is basic.

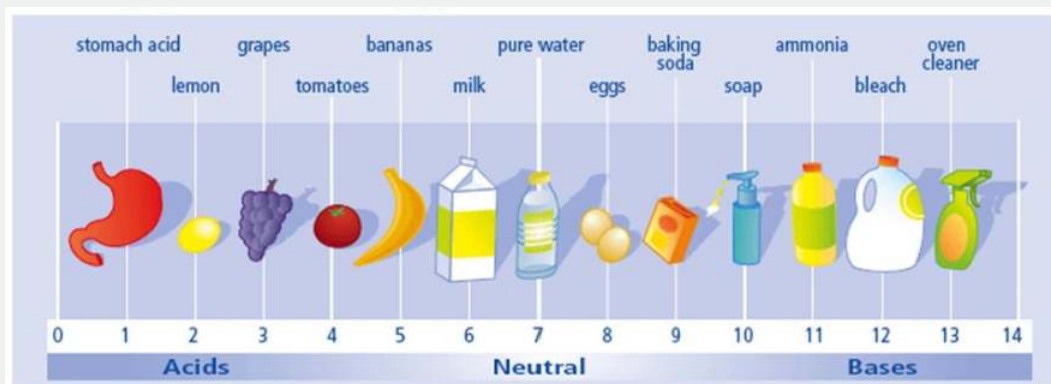


Figure 15.1.1 The pH scale.

Figure 15.1.2 illustrates the relationship between pH and the hydrogen ion concentration, 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 $[H^+]$** , which will give a positive value for pH .

The general formula for determining $[H^+]$ from pH is as follows:

$$[H^+] = 10^{-pH}$$

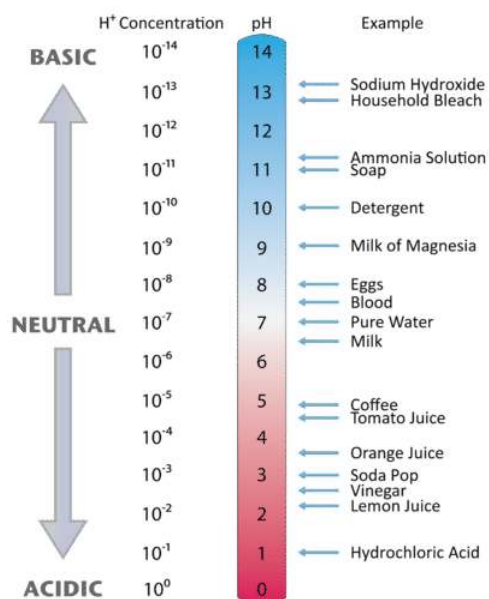


Figure 15.1.2 The pH values for several common materials with corresponding hydrogen ion concentration, $[H^+]$.

✓ Example 15.1.1

Label each solution as acidic, basic, or neutral based only on the stated pH .

- milk of magnesia, $pH = 10.5$
- pure water, $pH = 7$
- wine, $pH = 3.0$

Answer

- With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
- Pure water, with a pH of 7, is neutral.
- With a pH of less than 7, wine is acidic.

? Exercise 15.1.1

Identify each substance as acidic, basic, or neutral based only on the stated pH .

- human blood with $pH = 7.4$
- household ammonia with $pH = 11.0$
- cherries with $pH = 3.6$

Answer a

basic

Answer b

basic

Answer c

acidic

Table 15.1.1: gives the typical pH values of some common substances. Note that several food items are on the list, and most of them are acidic.

Table 15.1.1 Typical pH Values of Various Substances*

Substance	pH
stomach acid	1.7
lemon juice	2.2
vinegar	2.9
soda	3.0
wine	3.5
coffee, black	5.0
milk	6.9
pure water	7.0
blood	7.4
seawater	8.5
milk of magnesia	10.5
ammonia solution	12.5
1.0 M NaOH	14.0

*Actual values may vary depending on conditions

Summary

pH is usually (but not always) between 0 and 14.

- If $\text{pH} < 7$, then the solution is acidic.
- If $\text{pH} = 7$, then the solution is neutral.
- If $\text{pH} > 7$, then the solution is basic.

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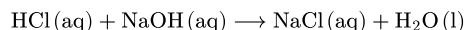
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15.2: Acid-Base Neutralization Reactions

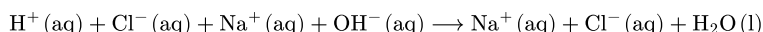
Learning Objective

- Write equations for acid-base neutralization reactions.

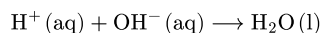
What happens when an acid such as HCl is mixed with a base such as NaOH?



When an acid and a base are combined, water and a salt are the products. Double displacement reactions of this type are called **neutralization reactions**. We can write an expanded version of this equation, with aqueous substances written in their longer form:



Removing the spectator ions we get the net ionic equation:



When a strong acid and a strong base are combined in the proper amounts - when $[\text{H}^+]$ equals $[\text{OH}^-]$ - a neutral solution results in which $\text{pH} = 7$. The acid and base have neutralized each other, and the acidic and basic properties are no longer present.

Salt solutions do not always have a pH of 7, however. Through a process known as **hydrolysis**, the ions produced when an acid and base combine may react with the water molecules to produce a solution that is slightly acidic or basic. We will not go into details here, but generally if a strong acid is mixed with a weak base there the resulting solution will be slightly acidic; if a strong base is mixed with a weak acid the solution will be slightly basic.

 Neutralization



Video 15.2.1 Acid-base neutralization reaction.

✓ Example 15.2.1: propionic acid + calcium hydroxide

Calcium propionate is used to inhibit the growth of molds in foods, tobacco, and some medicines. Write a balanced chemical equation for the reaction of aqueous propionic acid ($\text{CH}_3\text{CH}_2\text{CO}_2\text{H}$) with aqueous calcium hydroxide $[\text{Ca}(\text{OH})_2]$.

Solution

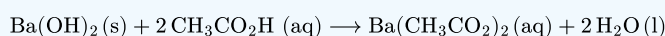
Example 15.2.1: steps for problem solving the reaction of propionic acid + calcium hydroxide

Steps	Reaction
Write the unbalanced equation. This is a double displacement reaction, so the cations and anions swap to create the water and the salt.	$\text{CH}_3\text{CH}_2\text{CO}_2\text{H}(\text{aq}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow (\text{CH}_3\text{CH}_2\text{CO}_2)_2\text{Ca}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
Balance the equation. Because there are two OH^- ions in the formula for $\text{Ca}(\text{OH})_2$, we need two moles of propionic acid, $\text{CH}_3\text{CH}_2\text{CO}_2\text{H}$ to provide H^+ ions.	$2\text{CH}_3\text{CH}_2\text{CO}_2\text{H}(\text{aq}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow (\text{CH}_3\text{CH}_2\text{CO}_2)_2\text{Ca}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$

? Exercise 15.2.1

Write a balanced chemical equation for the reaction of solid barium hydroxide with dilute acetic acid.

Answer



Applications

In analytical laboratories, chemical methods are used for determining the amounts of acids or bases in foods and household and industrial chemicals.

In , chemical neutralization methods are often applied to reduce the damage that an effluent may cause upon release to the environment.

There are many uses of neutralization reactions that are acid-alkali reactions. A very common use is antacid tablets (section 7.8).

In chemical synthesis of nanomaterials, the heat of neutralization reaction can be used to facilitate the chemical reduction of metal precursors.

Also in the digestive tract, neutralization reactions are used when food is moved from the stomach to the intestines. In order for the nutrients to be absorbed through the intestinal wall, an alkaline environment is needed, so the pancreas produce an antacid bicarbonate to cause this transformation to occur.

Another common use, though perhaps not as widely known, is in fertilizers and control of soil pH. Slaked lime (calcium hydroxide) or limestone (calcium carbonate) may be worked into soil that is too acidic for plant growth. Fertilizers that improve plant growth are made by neutralizing sulfuric acid (H_2SO_4) or nitric acid (HNO_3) with ammonia gas (NH_3), making or . These are salts utilized in the fertilizer.

Industrially, a by-product of the burning of coal, gas, may combine with water vapor in the air to eventually produce sulfuric acid, which falls as acid rain. To prevent the sulfur dioxide from being released, a device known as a scrubber gleans the gas from smoke stacks. This device first blows calcium carbonate into the combustion chamber where it decomposes into calcium oxide (lime) and carbon dioxide. This lime then reacts with the sulfur dioxide produced forming calcium sulfite. A suspension of lime is then injected into the mixture to produce a slurry, which removes the and any remaining unreacted sulfur dioxide.

Fluffy Cakes

Q: How do acid-base reactions make cake fluffy?



Figure 15.2.1 Cake with buttercream frosting. Image use with permission (Public domain; Scheinwerfermann).png

A: Usually cakes include an acidic ingredient (this varies) and sodium bicarbonate, a base. When they react, the proton from the acid is transferred to the bicarbonate, making the weak acid carbonic acid. Carbonic acid is the product of an acid anhydride reaction between carbon dioxide and water. This reaction can be reversed, or carbonic acid can decompose into water and carbon dioxide. Especially at the high temperatures inside a baking cake, this decomposition will happen, and produce carbon dioxide gas. The pressure of the hot gas will form bubbles inside the cake, making it fluffy.

Summary

- Acids and bases combine and produce salt and water.

Contributors and Attributions

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15.3: Acids and Bases

Learning Objectives

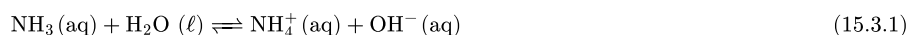
- Identify acids and bases.

The Brønsted-Lowry Theory of Acids and Bases

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, a **Brønsted-Lowry acid** is a **proton donor (PD)**, while a **Brønsted-Lowry base** is a **proton acceptor (PA)**.

A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:



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 15.3.2):

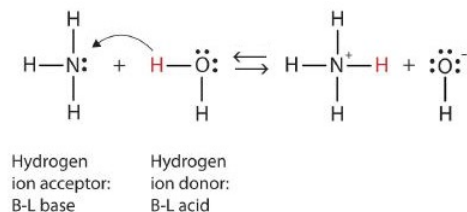


Figure 15.3.2 reaction of ammonia and water.

Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+ , which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

The Hydronium Ion

A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like H_5O_2^+ or H_9O_4^+ rather than H_3O^+ . It is simpler, however, to use H_3O^+ to represent the [hydronium ion](#).

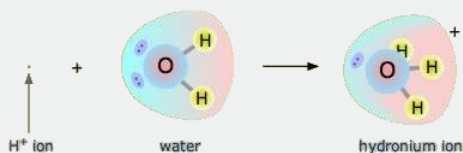
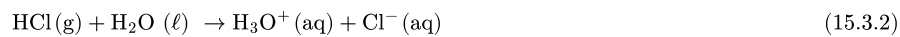


Figure 15.3.3 reaction of hydrogen and water to form hydronium. (CC BY-NC 3.0)

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 H₂O:



We can depict this process using Lewis electron dot diagrams (Figure 15.3.4):

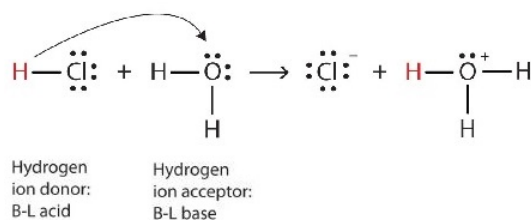


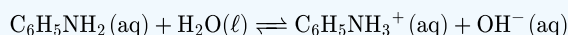
Figure 15.3.4 Lewis electron dot diagram of HCl dissolved in H₂O

✓ Example 15.3.1

Aniline (C₆H₅NH₂) 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

C₆H₅NH₂ and H₂O are the reactants. When C₆H₅NH₂ accepts a proton from H₂O, it gains an extra H and a positive charge and leaves an OH⁻ ion behind. The reaction is as follows:



Because C₆H₅NH₂ accepts a proton, it is the Brønsted-Lowry base. The H₂O molecule, because it donates a proton, is the Brønsted-Lowry acid.

? Exercise 15.3.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.



Answer

Brønsted-Lowry acid: H₂PO₄⁻; Brønsted-Lowry base: H₂O

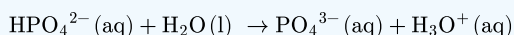
? Exercise 15.3.2

Which of the following compounds is a Brønsted-Lowry base?

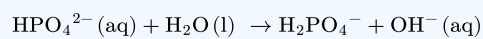
- HCl
- HPO₄²⁻
- H₃PO₄
- NH₄⁺
- CH₃NH₃⁺

Answer:

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H⁺. This eliminates HCl, H₃PO₄, NH₄⁺ and CH₃NH₃⁺ because they are Brønsted-Lowry acids. They all give away protons. In the case of HPO₄²⁻, consider the following equation:



Here, it is clear that HPO_4^{2-} is the acid since it donates a proton to water to make H_3O^+ and PO_4^{3-} . Now consider the following equation:



In this case, HPO_4^{2-} is the base since it accepts a proton from water to form $\text{H}_2\text{PO}_4^{-}$ and OH^{-} . Thus, HPO_4^{2-} is an acid and base together, making it amphoteric.

Since HPO_4^{2-} is the only compound from the options that can act as a base, the answer is **(b) HPO_4^{2-}** .

Summary

- A Brønsted-Lowry acid is a proton donor; a Brønsted-Lowry base is a proton acceptor.

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