

CHE 103: PRINCIPLES OF CHEMISTRY I



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CHE 103: Principles of Chemistry I

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CHAPTER OVERVIEW

1: Matter, Measurements, and Calculations

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- 1.2: Chemical Change vs. Physical Change
- 1.3: Indivisible - The Atomic Theory
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1.1: What Is Chemistry?

Learning Objectives

1. Define chemistry in relation to other sciences.
2. Identify the general steps in the scientific method.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Being able to describe the ingredients in a cake and how they change when the cake is baked is called chemistry. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter.

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure 1.1.1). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

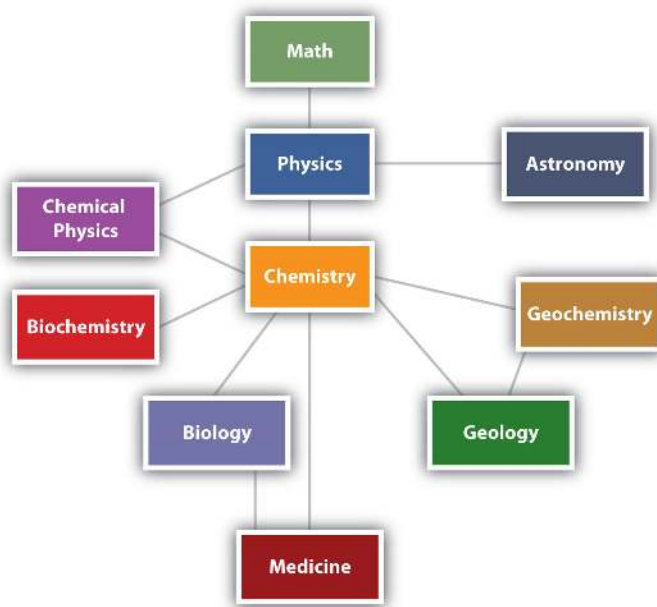


Figure 1.1.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure 1.1.1 shows how many of the individual fields of science are related.

There are many other fields of science in addition to the ones listed here.

Alchemy

As our understanding of the universe has changed over time, so has the practice of science. Chemistry in its modern form, based on principles that we consider valid today, was developed in the 1600s and 1700s. Before that, the study of matter was known as *alchemy* and was practiced mainly in China, Arabia, Egypt, and Europe.

Alchemy was a somewhat mystical and secretive approach to learning how to manipulate matter. Practitioners, called alchemists, thought that all matter was composed of different proportions of the four basic elements—fire, water, earth, and air—and believed that if you changed the relative proportions of these elements in a substance, you could change the substance. The long-standing attempts to “transmute” common metals into gold represented one goal of alchemy. Alchemy’s other major

[illegible]

Esprits acides.	Terre absorbante.	Q Cuivre.	♂ Soufre minéral. [Principe.
Acide du sel marin.	^{SN} Substances métalliques.	♂ Fer.	♂ Principe huileux ou Soufre
Acide nitreux.	☿ Mercure.	♂ Plomb.	♂ Esprit de vinaigre.
Acide vitriolique.	☿ Règle d'Antimoine	♂ Etain.	✓ Eau.
Sel alcali fixe.	☿ Or.	♂ Zinc	☉ Sel [dents
Sel alcali volatil.	☿ Argent.	♂ Pierre Calaminaire.	☉ Esprit de vin et Esprits ar-

In spite of this secrecy, in its time alchemy was respected as a serious, scholarly endeavor. Isaac Newton, the great mathematician and physicist, was also an alchemist.

- a. sculpture
- b. astronomy

Sculpture is not considered a science because it is not a study of some aspect of the natural universe.

Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

- a. physiology (the study of the function of an animal's or a plant's body)
- b. geophysics
- c. agriculture
- d. politics

Answer

a, b and c only

How do scientists work? Generally, they follow a process called the scientific method. The scientific method is an organized procedure for learning answers to questions and making explanations for observations. To find the answer to a question (for example, “Why do birds fly toward Earth’s equator during the cold months?”), a scientist goes through the following steps, which are also illustrated in Figure 1.1.2:

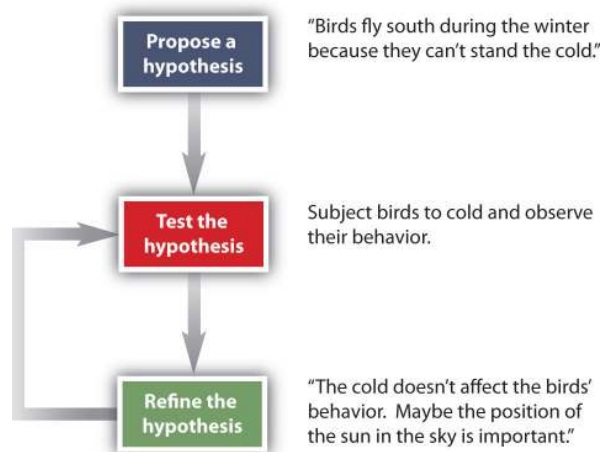


Figure 1.1.2: The General Steps of the Scientific Method. After an observation is made or a question is identified, a hypothesis is made and experiments are designed to test the hypothesis.

The steps may not be as clear-cut in real life as described here, but most scientific work follows this general outline.

- 1. Propose a hypothesis.** A scientist generates a testable idea, or hypothesis, to try to answer a question or explain an observation about how the natural universe works. Some people use the word *theory* in place of hypothesis, but the word hypothesis is the proper word in science. For scientific applications, the word theory is a general statement that describes a large set of observations and data. A theory represents the highest level of scientific understanding.
- 2. Test the hypothesis.** A scientist evaluates the hypothesis by devising and carrying out experiments to test it. If the hypothesis passes the test, it may be a proper answer to the question. If the hypothesis does not pass the test, it may not be a good answer.
- 3. Refine the hypothesis if necessary.** Depending on the results of experiments, a scientist may want to modify the hypothesis and then test it again. Sometimes the results show the original hypothesis to be completely wrong, in which case a scientist will have to devise a new hypothesis.

Not all scientific investigations are simple enough to be separated into these three discrete steps. But these steps represent the general method by which scientists learn about our natural universe.

1.1.3Exercise

Define science and chemistry.

Answer

Science is a process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Chemistry is the study of matter.

1.1.4Exercise

Name the steps of the scientific method.

Answer

After identifying the problem or making an observation, propose a hypothesis, test the hypothesis, and refine the hypothesis, if necessary

Key Takeaways

- Chemistry is the study of matter and how it behaves.
 - The scientific method is the general process by which we learn about the natural universe.
1. Based on what you know, which fields are branches of science?
 - a. meteorology (the study of weather)
 - b. astrophysics (the physics of planets and stars)
 - c. economics (the study of money and monetary systems)
 - d. astrology (the prediction of human events based on planetary and star positions)
 - e. political science (the study of politics)
 2. Based on what you know, which fields are a branches of science?
 - a. history (the study of past events)
 - b. ornithology (the study of birds)
 - c. paleontology (the study of fossils)
 - d. zoology (the study of animals)
 - e. phrenology (using the shape of the head to determine personal characteristics)
 3. Which of the following are examples of matter?
 - a. a baby
 - b. an idea
 - c. the Empire State Building
 - d. an emotion
 - e. the air
 - f. Alpha Centauri, the closest known star (excluding the sun) to our solar system
 4. Which of the following are examples of matter?
 - a. your textbook
 - b. brain cells
 - c. love
 - d. a can of soda
 - e. breakfast cereal
 5. Suggest a name for the science that studies the physics of rocks and the earth.
 6. Suggest a name for the study of the physics of living organisms.
 7. Engineering is the practical application of scientific principles and discoveries to develop things that make our lives easier. Is medicine science or engineering? Justify your answer.
 8. Based on the definition of engineering in Exercise 7, would building a bridge over a river or road be considered science or engineering? Justify your answer.
 9. When someone says, "I have a theory that excess salt causes high blood pressure," does that person really have a theory? If it is not a theory, what is it?
 10. When a person says, "My hypothesis is that excess calcium in the diet causes kidney stones," what does the person need to do to determine if the hypothesis is correct?
 11. Some people argue that many scientists accept many scientific principles on faith. Using what you know about the scientific method, how might you argue against that assertion?
 12. Most students take multiple English classes in school. Does the study of English use the scientific method?

Answers

1.
 - a. science
 - b. science
 - c. not science
 - d. not science
 - e. not science
2.
 - a. not science
 - b. science
 - c. science
 - d. science
 - e. not science
3.
 - a. matter
 - b. not matter
 - c. matter
 - d. not matter
 - e. matter
 - f. matter
4.
 - a. matter
 - b. matter
 - c. not matter
 - d. matter
 - e. matter
5. geophysics
6. biophysics
7. Medicine is probably closer to a field of engineering than a field of science, but this may be arguable. Ask your doctor.
8. Engineering
9. In scientific terms, this person has a hypothesis.
10. Conduct experiments to determine if kidney stones contain calcium.
11. Science is based on reproducible facts, not blind belief.
12. No.

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1.2: Chemical Change vs. Physical Change

The difference between a physical change and a chemical change (or reaction) is **composition**. In a chemical change, there is a change in the composition of the substances in question; in a physical change there is a difference in the appearance, smell, or simple display of a sample of matter without a change in composition. Although we call them physical "reactions," no reaction is actually occurring. In order for a reaction to take place, there must be a change in the elemental composition of the substance in question. Thus, we shall simply refer to physical "reactions" as physical changes from now on.

Introduction

Physical changes are limited to changes that result in a difference in display without changing the composition. Some common changes (but not limited to) are:

- Texture
- Color
- Temperature
- Shape
- Change of State (Boiling Point and Melting Point are significant factors in determining this change.)

Physical properties include many other aspects of a substance. The following are (but not limited to) physical properties.

- Luster
- Malleability
- Ability to be drawn into a thin wire
- Density
- Viscosity
- Solubility
- Mass
- Volume

Any change in these physical properties is referred to as a physical change.

Chemical changes, on the other hand, are quite different. A chemical change occurs when the substance's composition is changed. When bonds are broken and new ones are formed a chemical change occurs. The following are indicators of chemical changes:

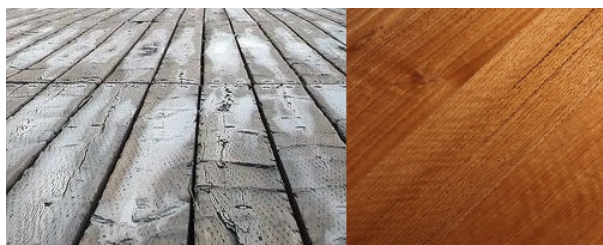
- Change in Temperature
- Change in Color
- Noticeable Odor (after reaction has begun)
- Formation of a Precipitate
- Formation of Bubbles

Note: When two or more reactants are mixed and a change in temperature, color, etc. is noticed, a chemical reaction is probably occurring. These are not definite indicators; a chemical reaction may not be occurring. A change in color is not always a chemical change. If one were to change the color of a substance in a non-chemical reaction scenario, such as painting a car, the change is physical and not chemical. This is because the composition of the car has not changed. Also, changes in temperature or color can be a consequence of both: physical and chemical changes. In order to decide, we need to understand the *cause* of the change.

Common Physical Changes

Texture

The **texture** of a substance can differ with a physical change. For example, if a piece of wood was sanded, waxed, and polished, it would have a very different texture than it initially had as a rough piece of wood.



(left) Rough plank boardwalk, Quebec City, Canada (right) Finished mountain ash floor. (CC BY-SA 4.0; [WikiPedant](#) and CC BY-SA 2.5; [MarkAnthonyBoyle](#), respectively).

As you can see, the texture of the finished wood is much smoother than the initial grainy wood.

Color

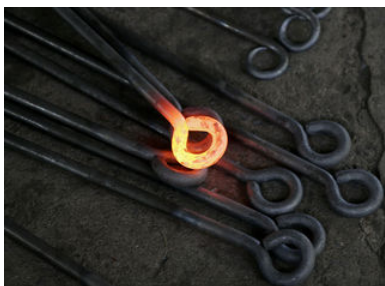
The changing of **color** of a substance is not necessarily an indicator of a chemical change. For example, changing the color of a metal does not change its physical properties. However, in a chemical reaction, a color change is usually an indicator that a reaction is occurring. Painting the metal car does not changing the composition of the metallic substance.



Robotic arm applying paint on car parts. Image use with permission (CC BY-SA 4.0 [RoboGuru](#)).

Temperature

Although we cannot see **temperature** change, unless if a change of state is occurring, it is a physical change.



Hot metalwork. (CC BY-SA-NC 2.0; [flagstaffotos.com.au](#))

One cannot see the pan physically changing shape, color, texture, or any of the other physical properties. However, if one were to touch the pan, it would be incredibly hot and could cause a burn. Sitting idle in a cupboard, this pan would be cold. One cannot assess this change only through visual exposure; the use of a thermometer or other instrument is necessary.

Shape

The **shape** of an object can be changed and the object will still remain true to its chemical composition. For example, if one were to fold money, as shown by the figure below, the money is still chemically the same.

Origami Money



Change of State

The **change of state** is likewise a physical change. In this scenario, one can observe a number of physical properties changing, such as viscosity and shape. As ice turns into water, it does not retain a solid shape and now becomes a viscous fluid. The physical "reaction" for the change of ice into liquid water is:



Figure: Ice Melting. used with permission from Wikipedia

The following are the changes of state:

Solid → Liquid	Melting
Liquid → Gas	Vaporization
Liquid → Solid	Freezing
Gas → Liquid	Condensation
Solid → Gas	Sublimation

- If heat is added to a substance, such as in melting, vaporization, and sublimation, the process is **endothermic**. In this instance, heat is increasing the speed of the molecules causing them to move faster.
- If heat is removed from a substance, such as in freezing and condensation, then the process is **exothermic**. In this instance, heat is decreasing the speed of the molecules causing them to move slower.

Physical Properties

Luster

The **luster** of an element is defined as the way it reacts to light. Luster is a quality of a metal. Almost all of the metals, transition metals, and metalloids are lustrous. The non-metals and gases are not lustrous. For example, oxygen and bromine are not lustrous. Shown below is are lustrous paper clips.

Lustrous Paperclips



Malleability

Malleability is also a quality of metals. Metals are said to be malleable. This means that the metals can deform under an amount of stress. For example, if you can hit a metal with a mallet and it deforms, it is malleable. Also, a paperclip can be shaped with bare hands.

Bent Paperclip

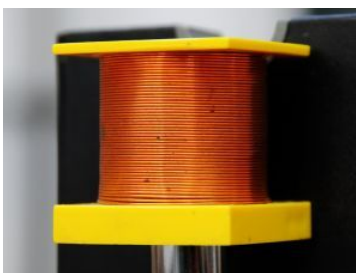


The image shows the malleability of a certain metal as stress is applied to it.

Ability to be drawn into a thin wire

In materials science, this property is called **ductility**. For example, raw copper can be obtained and it can be purified and wrapped into a cord. Once again, this property is characteristic of mainly metals, nonmetals do not possess this quality.

Copper Wire



Density

The **density** of an object is its mass divided by its volume ($d=m/v$). A substance will have a higher density if it has more mass in a fixed amount of volume. For example, take a ball of metal, roughly the size of a baseball, compressed from raw metal. Compare this to a baseball made of paper. The baseball made of metal has a much greater weight to it in the same amount of volume. Therefore the baseball made out of metal has a much higher density. The density of an object will also determine whether it will sink or float in a particular chemical. Water for example has a density of 1g/cm^3 . Any substance with a density lower than that will float, while any substance with a density above that will sink.

Oil Sinking in a Glass of Water



Viscosity

Viscosity is defined to be the resistance to deformation of a particular chemical substance when a force is applied to it. In the example below, one can see two cubes falling into two different test tubes. The upper substance shows a violent reaction to the dropping of the cube. The lower substance simply engulfs it slowly without much reaction. The upper substance has a lower viscosity relative to the lower substance, which has very high viscosity. One may even think of viscosity in terms of thickness. The substance with more thickness has higher viscosity than a substance that is deemed "thin." Water has a lower viscosity than honey or magma, which have relatively high viscosities.

Viscosity of Fluids



Figure 1.6.21.6.2: Viscosity demonstration. The fluid on the left has a lower viscosity than the fluid on the right. (CC SA-BY 4.0; Synapticrelay).

Common Chemical Changes

The follow are all indicators of chemical reactions.

Change in Temperature

A **change in temperature** is characteristic of a chemical change. During an experiment, one could dip a thermometer into a beaker or Erlenmeyer Flask to verify a temperature change. If temperature increases, as it does in most reactions, a chemical change is likely to be occurring. This is different from the physical temperature change. During a physical temperature change, one substance, such as water is being heated. However, in this case, one compound is mixed in with another, and these reactants produce a product. When the reactants are mixed, the temperature change caused by the reaction is an indicator of a chemical change.



Figure: Violent Reaction (Fireworks) with heat as a product

As an example of a exothermic reaction, if Fe_2O_3 is mixed with Al and ignited (often with burning Mg), then the thermite reaction is initiated



This reaction generates heat as a product and is (very) exothermic.

However, physical changes can be exothermic or endothermic. The melting of an ice cube, which is endothermic, is a change in a physical property and not composition. Thus, it is a physical change.

Change in Color

A **change in color** is also another characteristic of a chemical reaction taking place. For example, if one were to observe the rusting of metal over time, one would realized that the metal has changed color and turned orange. This change in color is evidence of a chemical reaction. However, one must be careful; sometimes a change in color is simply the mixing of two colors, but no real change in the composition of the substances in question.

Metal Rusting



The reaction above is that of the rusting of iron.



Noticeable Odor

When two or more compounds or elements are mixed and a **scent or odor** is present, a chemical reaction has taken place. For example, when an egg begins to smell, (a rotten egg) a chemical reaction has taken place. This is the result of a chemical decomposition.

Spoiled Egg



Formation of a Precipitate

The **formation of a precipitate** may be one of the most common signs of a chemical reaction taking place. A precipitate is defined to be a solid that forms inside of a solution or another solid. Precipitates should not be confused with suspensions, which are solutions that are homogeneous fluids with particles floating about in them. For instance, when a soluble carbonate reacts with Barium, a Barium Carbonate precipitate can be observed.

Test Tube



Reaction:



For further information, please refer to [Classification of Matter](#).

Formation of Bubbles

The **formation of bubbles**, or rather **a gas**, is another indicator of a chemical reaction taking place. When bubbles form, a temperature change could also be taking place. Temperature change and formation of bubbles often occur together. For example, in the following image, one can see a gas spewing. This is the formation of a gas.

Gas Formation



However, most reactions are much more subtle. For instance, if the following reaction occurs, one may notice Carbon Dioxide bubbles forming. If there is enough Hydrochloric Acid, bubbles are visible. If there isn't, one can't readily notice the change:



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Outside Links

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Problems

1. Which of the following is a chemical reaction?
 - a. Freezing liquid Mercury
 - b. Adding yellow to blue to make green
 - c. Cutting a piece of paper into two pieces
 - d. Dropping a sliced orange into a vat of Sodium Hydroxide
 - e. Filling a balloon with natural air
2. Which of the following is a physical reaction?
 - a. Shattering Glass with a baseball
 - b. Corroding Metal
 - c. Fireworks Exploding
 - d. Lighting a match
 - e. Baking a cake
3. Which of the following is a chemical reaction?
 - a. Painting a wall blue
 - b. A bicycle rusting
 - c. Ice cream melting
 - d. Scratching a key across a desk
 - e. Making a sand castle
4. Which of the following is a physical reaction?
 - a. Frying an egg
 - b. Digesting carrots
 - c. A Macbook falling out of a window
 - d. Creating ATP in the human body
 - e. Dropping a fizzy tablet into a glass of water
5. Write C for Chemical Reaction or P for Physical Reaction.
 - a. Burning Leaves
 - b. Cutting Diamonds
 - c. Crushing a pencil
 - d. The salivary amylase enzyme that breaks down food in the mouth
 - e. Salt mixing in with water

Answers

1. D

2. A

3. B

4. C

5. a) C

b) P

c) P

d) C

e) Neither. This is one of the gray areas of chemical change and physical change. Although the salt has dissociated into Sodium and Chloride ions, it is still salt in water. Salt, initially is actually just a conglomerate of sodium and chloride ions and by dissociating them, just the arrangement of the ions has changed. Please [click here](#) for more information.

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1.3: Indivisible - The Atomic Theory

Learning Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus and Dalton to atomic theory.
- Summarize Dalton's atomic theory and explain its historical development.

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, but 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, big, 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 1.3.1: (left) 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. (right) 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 are known today) was fully appreciated.

📌 Greeks: "All Thought and No Action"

Greek philosophers were "all thought and no action" and did not feel the need to test their theories with reality. In contrast, Dalton's efforts were based on experimentation and testing ideas against reality.

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British

scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 1.3.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

Dalton studied the weights of various elements and compounds. He noticed that matter always combined in fixed ratios based on weight, or volume in the case of gases. Chemical compounds always contain the same proportion of elements by mass, regardless of amount, which provided further support for Proust's law of definite proportions. Dalton also observed that there could be more than one combination of two elements.

From his experiments and observations, as well as the work from peers of his time, Dalton proposed a new theory of the **atom**. This later became known as Dalton's **atomic theory**. The general tenets of this theory were as follows:

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We know now that (1) an atom can be further subdivided, (2) all atoms of an element are not identical in mass, and (3) using nuclear fission and fusion techniques, we can create or destroy atoms by changing them into other atoms.

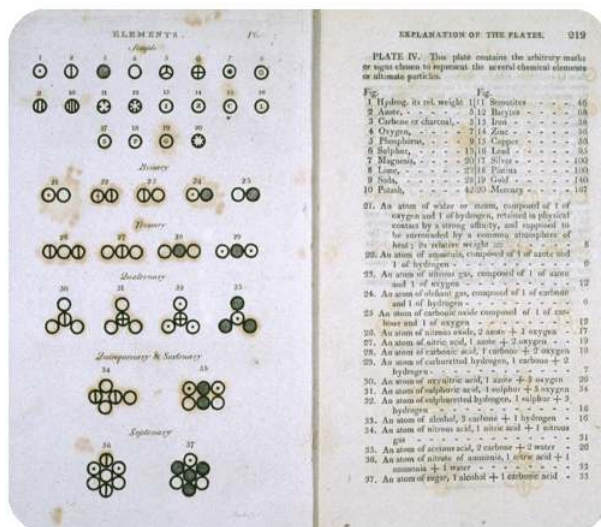


Figure 1.3.2: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the-art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video 1.3.1, control of individual atoms can be used to create animations.



Video 1.3.1: A Boy And His Atom: The World's Smallest Movie.

A Boy and His Atom is a 2012 stop-motion animated short film released by IBM Research. The movie tells the story of a boy and a wayward atom who meet and become friends. It depicts a boy playing with an atom that takes various forms. It was made by moving carbon monoxide molecules viewed with a scanning tunneling microscope, a device that magnifies them 100 million times. These molecules were moved to create images, which were then saved as individual frames to make the film.

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.
- Dalton's Atomic Theory is the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom. The general tenets of this theory are:
 - All matter is composed of extremely small particles called atoms.
 - Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
 - Atoms cannot be subdivided, created, or destroyed.
 - Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
 - In chemical reactions, atoms are combined, separated, or rearranged.

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1.4: The Classification of Matter

Learning Objectives

- Use physical and chemical properties, including phase, to describe matter.
- Identify a sample of matter as an element, a compound, or a mixture.

Part of understanding matter is being able to describe it. One way chemists describe matter is to assign different kinds of properties to different categories.

Physical and Chemical Properties

The properties that chemists use to describe matter fall into two general categories. **Physical properties** are characteristics that describe matter. They include characteristics such as size, shape, color, and mass. These characteristics can be observed or measured without changing the *identity* of the matter in question. **Chemical properties** are characteristics that describe how matter changes its chemical structure or composition. An example of a chemical property is flammability—a material's ability to burn—because burning (also known as combustion) changes the chemical composition of a material. The observation of chemical properties involves a *chemical change* of the matter in question, resulting in matter with a different *identity* and different physical and chemical properties.



Figure 1.4.1: 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 the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass.



Figure 1.4.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. Image used with permission (CC-SA-BY-3.0; Andrikkos)

Elements and Compounds

Any sample of matter that has the *same physical and chemical properties throughout* the sample is called a **substance**. There are two types of substances. A substance that cannot be broken down into chemically simpler components is called an **element**. Aluminum, which is used in soda cans and is represented by the symbol Al, is an element. A substance that can be broken down

into chemically simpler components (because it consists of more than one element) is called a **compound**. Water is a compound composed of the elements hydrogen and oxygen and is described by the chemical formula, H_2O . Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

Sometimes the word *pure* is used to describe a substance, but this is not absolutely necessary. By definition, any single substance, element or compound is *pure*.

The smallest part of an element that maintains the identity of that element is called an **atom**. Atoms are extremely tiny; to make a line of iron atoms that is 1 inch long, you would need approximately 217 million iron atoms. The smallest part of a compound that maintains the identity of that compound is called a **molecule**. Molecules are composed of two or more different atoms that are attached together and behave as a unit. Scientists usually work with millions and millions of atoms and molecules at a time. When a scientist is working with large numbers of atoms or molecules at a time, the scientist is studying the *macroscopic viewpoint* of the universe. However, scientists can also describe chemical events on the level of individual atoms or molecules, which is referred to as the *microscopic viewpoint*. We will see examples of both macroscopic and microscopic viewpoints throughout this book (Figure 1.4.1).

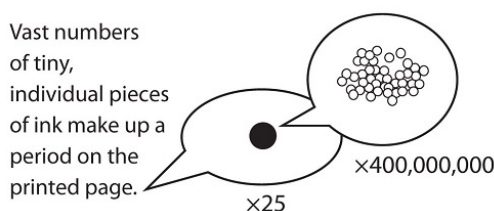


Figure 1.4.3: How Many Particles Are Needed for a Period in a Sentence? Although we do not notice it from a macroscopic perspective, matter is composed of microscopic particles so tiny that billions of them are needed to make a speck we can see with the naked eye. The $\times 25$ and $\times 400,000,000$ indicate the number of times the image is magnified.

Mixtures

A material composed of two or more substances is a **mixture**. In a mixture, the individual substances maintain their chemical identities. Many mixtures are obvious combinations of two or more substances, such as a mixture of sand and water. Such mixtures are called **heterogeneous mixtures**. In some mixtures, the components are so intimately combined that they act like a single substance (even though they are not). Mixtures with a consistent or uniform composition throughout are called **homogeneous mixtures**. Homogeneous mixtures are also known as **solutions**. For example, when sugar is dissolved in water to form a liquid solution, the individual properties of the components cannot be distinguished. Other examples of homogeneous mixtures include solid solutions, like the metal alloy steel, and gaseous solutions, like air which is a mixture of mainly nitrogen and oxygen.

Example 1.4.1

How would a chemist categorize each example of matter?

- saltwater
- soil
- water
- oxygen

Answer 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.

Answer b

Soil is composed of small pieces of a variety of materials, so it is a *heterogeneous mixture*.

Answer c

Water is a *substance*; more specifically, because water is composed of hydrogen and oxygen, it is a *compound*.

Answer d

Oxygen, a *substance*, is an *element*.

Exercise 1.4.1

How would a chemist categorize each example of matter?

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

Answer a

homogeneous mixture or solution

Answer b

element

Answer c

heterogeneous mixture

Phases or Physical States of Matter

All matter can be further classified by one of three physical **states** or **phases**, solid, liquid or gas. These three descriptions each imply that the matter has certain physical properties when in these states. A solid has a definite shape and a definite volume. Liquids ordinarily have a definite volume but not a definite shape; they take the shape of their containers. Gases have neither a definite shape nor a definite volume, and they expand to fill their containers.

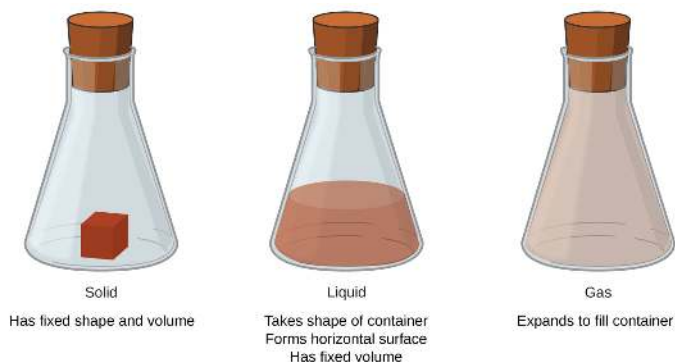


Figure 1.4.4: The three most common states or phases of matter are solid, liquid, and gas.

We encounter matter in each phase every day; in fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas) (Figure 1.4.2).



Figure 1.4.5: Boiling Water. When liquid water boils to make gaseous water, it undergoes a phase change. (CC BY-SA 3.0 Unported; Markus Schweiss via Wikipedia)

We know from our experience with water that substances can change from one phase to another if the conditions are right. Typically, varying the temperature of a substance (and, less commonly, the pressure exerted on it) can cause a **phase change**, a physical process in which a substance changes from one phase to another (Figure 1.4.2). Phase changes are identified by particular names depending on what phases are involved, as summarized in Table 1.4.1.

Table 1.4.1: Phase Changes

Change	Name
solid to liquid	melting, fusion
solid to gas	sublimation
liquid to gas	boiling, evaporation
liquid to solid	solidification, freezing
gas to liquid	condensation
gas to solid	deposition

Figure 1.4.6 illustrates the relationships between the different ways matter can be classified.

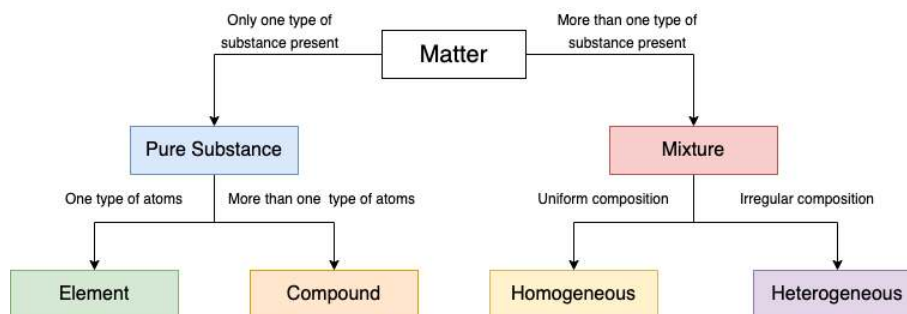


Figure 1.4.6: The Classification of Matter. Matter can be classified in a variety of ways, depending on its properties.

Concept Review Exercises

1. Explain the differences between the physical properties of matter and the chemical properties of matter.
2. What is the difference between a heterogeneous mixture and a homogeneous mixture? Give an example of each.
3. Give at least two examples of a phase change and state the phases involved in each.

Answers

1. Physical properties describe the existence of matter, and chemical properties describe how substances change into other substances.
2. A heterogeneous mixture is obviously a mixture, such as dirt; a homogeneous mixture behaves like a single substance, such as saltwater.
3. solid to liquid (melting) and liquid to gas (boiling) (answers will vary)

Key Takeaways

- Matter can be described with both physical properties and chemical properties.
- Matter can be identified as an element, a compound, or a mixture.

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1.5: Measurements

Learning Objectives

- Express quantities properly using a number and a unit.

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 1.5.1); the nurse is also measuring.



Figure 1.5.1: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GNU Free Documentation License; Pia von Lützau via Wikipedia).

Chemists measure the properties of matter using a variety of devices or measuring tools, many of which are similar to those used in everyday life. Rulers are used to measure length, balances (scales) are used to measure mass (weight), and graduated cylinders or pipettes are used to measure volume. Measurements made using these devices are expressed 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 “5.2 kilometers,” we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5.2.

5.2 kilometers
number unit

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—for example, 12 miles, 12 kilometers, 12 furlongs, or 12 yards.

Without units, a number can be meaningless, confusing, or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient's seizures and states a dosage of “100” without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

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.

✓ Exercise 1.5.1

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

- one dozen eggs
- 2.54 centimeters
- a box of pencils
- 88 meters per second

Answer a

The number is one, and the unit is dozen.

Answer b

The number is 2.54, and the unit is centimeter.

Answer c

The number 1 is implied because the quantity is only *a* box. The unit is box of pencils.

Answer d

The number is 88, and the unit is meters per second. Note that in this case the unit is actually a combination of two units: meters and seconds.

? Exercise 1.5.2

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

- a. 99 bottles of soda
- b. 60 miles per hour
- c. 32 fluid ounces
- d. 98.6 degrees Fahrenheit

Answer a

The number is 99, and the unit is bottles of soda.

Answer b

The number is 60, and the unit is miles per hour.

Answer c

The number 32, and the unit is fluid ounces

Answer d

The number is 98.6, and the unit is degrees Fahrenheit

? Exercise 1.5.2

What are the two necessary parts of a quantity?

Answer

The two necessary parts are the number and the unit.

Key Takeaway

- Identify a quantity properly with a number and a unit.

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1.6: The International System of Units

Learning Objectives

- Recognize the SI base units and explain the system of prefixes used with them.
- Define and calculate density.

People who live in the United States measure weight in pounds, height in feet and inches, and a car's speed in miles per hour. In contrast, chemistry and other branches of science use the International System of Units (also known as **SI** after *Système Internationale d'Unités*), which was established so that scientists around the world could communicate efficiently with each other. Many countries have also adopted SI units for everyday use as well. The United States is one of the few countries that has not.

Base SI Units

Base (or basic) units, are the fundamental units of SI. There are seven base units, which are listed in Table 1.6.1, Chemistry uses five of the base units: the mole for amount, the kilogram for mass, the meter for length, the second for time, and the kelvin for temperature. The degree Celsius ($^{\circ}\text{C}$) is also commonly used for temperature. The numerical relationship between kelvins and degrees Celsius is as follows:

$$K = ^{\circ}\text{C} + 273 \quad (1.6.1)$$

The United States uses the English (sometimes called Imperial) system of units for many quantities. Inches, feet, miles, gallons, pounds, and so forth, are all units connected with the English system of units. There have been [many mistakes](#) due to the improper conversion of units between the SI and English systems.

Table 1.6.1: The Seven Base SI Units

Property	Unit	Abbreviation
length	meter	m
mass	kilogram	kg
time	second	s
amount	mole	mol
temperature	kelvin	K
electrical current	ampere	amp
luminous intensity	candela	cd

The size of each base unit is defined by international convention. For example, the *kilogram* is defined as the quantity of mass of a special metal cylinder kept in a vault in France (Figure 1.6.1). The other base units have similar definitions and standards. The sizes of the base units are not always convenient for all measurements. For example, a meter is a rather large unit for describing the width of something as narrow as human hair. Instead of reporting the diameter of hair as 0.00012 m or as 1.2×10^{-4} m using scientific notation as discussed in section 1.4, SI also provides a series of **prefixes** that can be attached to the units, creating units that are larger or smaller by powers of 10.

Common prefixes and their multiplicative factors are listed in Table 1.6.2. (Perhaps you have already noticed that the base unit *kilogram* is a combination of a prefix, kilo- meaning $1,000 \times$, and a unit of mass, the gram.) Some prefixes create a multiple of the original unit: 1 kilogram equals 1,000 grams, and 1 megameter equals 1,000,000 meters. Other prefixes create a fraction of the original unit. Thus, 1 centimeter equals 1/100 of a meter, 1 millimeter equals 1/1,000 of a meter, 1 microgram equals 1/1,000,000 of a gram, and so forth.

Table 1.6.2: Prefixes Used with SI Units

Prefix	Abbreviation	Multiplicative Factor	Multiplicative Factor in Scientific Notation
giga-	G	1,000,000,000 ×	$10^9 \times$
mega-	M	1,000,000 ×	$10^6 \times$
kilo-	k	1,000 ×	$10^3 \times$
deca-	D	10 ×	$10^1 \times$
deci-	d	1/10 ×	$10^{-1} \times$
centi-	c	1/100 ×	$10^{-2} \times$
milli-	m	1/1,000 ×	$10^{-3} \times$
micro-	μ^*	1/1,000,000 ×	$10^{-6} \times$
nano-	n	1/1,000,000,000 ×	$10^{-9} \times$

*The letter μ is the Greek lowercase letter for *m* and is called “mu,” which is pronounced “myoo.”

Both SI units and prefixes have abbreviations, and the combination of a prefix abbreviation with a base unit abbreviation gives the abbreviation for the modified unit. For example, kg is the abbreviation for kilogram. We will be using these abbreviations throughout this book.



Figure 1.6.1: The Kilogram. The standard for the kilogram is a platinum-iridium cylinder kept in a special vault in France. Source: Photo reproduced by permission of the Bureau International des Poids et Mesures, who retain full internationally protected copyright.

The Difference Between Mass and Weight

The mass of a body is a measure of its inertial property or how much matter it contains. The weight of a body is a measure of the force exerted on it by gravity or the force needed to support it. Gravity on earth gives a body a downward acceleration of about 9.8 m/s^2 . In common parlance, weight is often used as a synonym for mass in weights and measures. For instance, the verb “to weigh” means “to determine the mass of” or “to have a mass of.” The incorrect use of weight in place of mass should be phased out, and the term mass used when mass is meant. The SI unit of mass is the kilogram (kg). In science and technology, the weight of a body in a particular reference frame is defined as the force that gives the body an acceleration equal to the local acceleration of free fall in that reference frame. Thus, the SI unit of the quantity weight defined in this way (force) is the newton (N).

Derived SI Units

Derived units are combinations of SI base units. Units can be multiplied and divided, just as numbers can be multiplied and divided. For example, the area of a square having a side of 2 cm is $2\text{ cm} \times 2\text{ cm}$, or 4 cm^2 (read as “four centimeters squared” or “four square centimeters”). Notice that we have squared a length unit, the centimeter, to get a derived unit for area, the square centimeter.

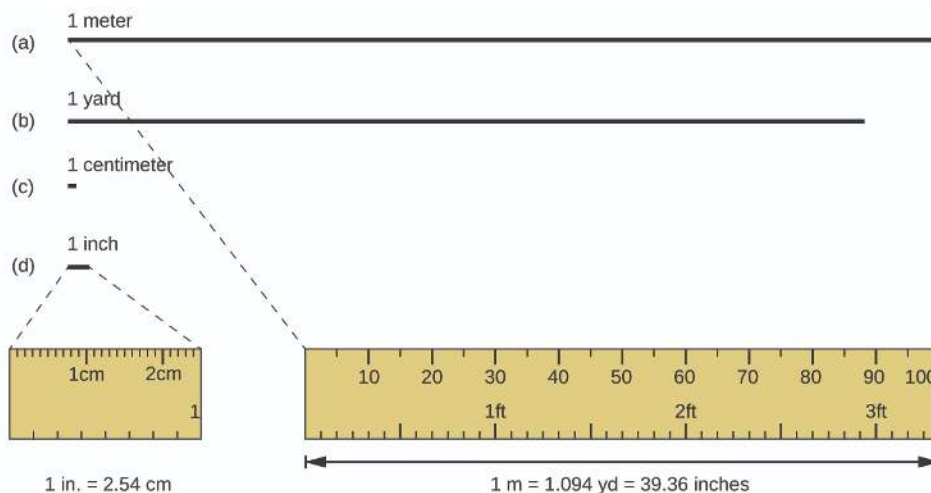


Figure 1.6.2: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd. (CC BY 4.0; OpenStax)

Volume is an important quantity that uses a derived unit. **Volume** is the amount of space that a given substance occupies and is defined geometrically as length $\times$ width $\times$ height. Each distance can be expressed using the meter unit, so volume has the derived unit $\text{m} \times \text{m} \times \text{m}$, or m^3 (read as “meters cubed” or “cubic meters”). A cubic meter is a rather large volume, so scientists typically express volumes in terms of 1/1,000 of a cubic meter. This unit has its own name—the liter (L). A liter is a little larger than 1 US quart in volume. Below are approximate equivalents for some of the units used in chemistry.

Approximate Equivalents to Some SI Units

- $1\text{ m} \approx 39.36\text{ in.} \approx 3.28\text{ ft} \approx 1.09\text{ yd}$
- $1\text{ in.} \approx 2.54\text{ cm}$
- $1\text{ km} \approx 0.62\text{ mi}$
- $1\text{ kg} \approx 2.20\text{ lb}$
- $1\text{ lb} \approx 454\text{ g}$
- $1\text{ L} \approx 1.06\text{ qt}$
- $1\text{ qt} \approx 0.946\text{ L}$

As shown in Figure 1.6.3, a liter is also $1,000\text{ cm}^3$. By definition, there are 1,000 mL in 1 L, so 1 milliliter and 1 cubic centimeter represent the same volume.

$$1\text{ mL} = 1\text{ cm}^3 \quad (1.6.2)$$

✓ Example 1.6.1

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

- kiloliter
- microsecond
- decimeter
- nanogram

Answer a

The abbreviation for a kiloliter is kL. Because kilo means “1,000 ×,” 1 kL equals 1,000 L.

Answer b

The abbreviation for microsecond is μs . Micro implies 1/1,000,000th of a unit, so 1 μs equals 0.000001 s.

Answer c

The abbreviation for decimeter is dm. Deci means 1/10th, so 1 dm equals 0.1 m.

Answer d

The abbreviation for nanogram is ng and equals 0.000000001 g.

? Exercise 1.6.1

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

- kilometer
- milligram
- nanosecond
- centiliter

Answer a

km (1,000 m)

Answer b

mg (0.001 g)

Answer c

ns (0.000000001 s)

Answer d

cL (0.01L)

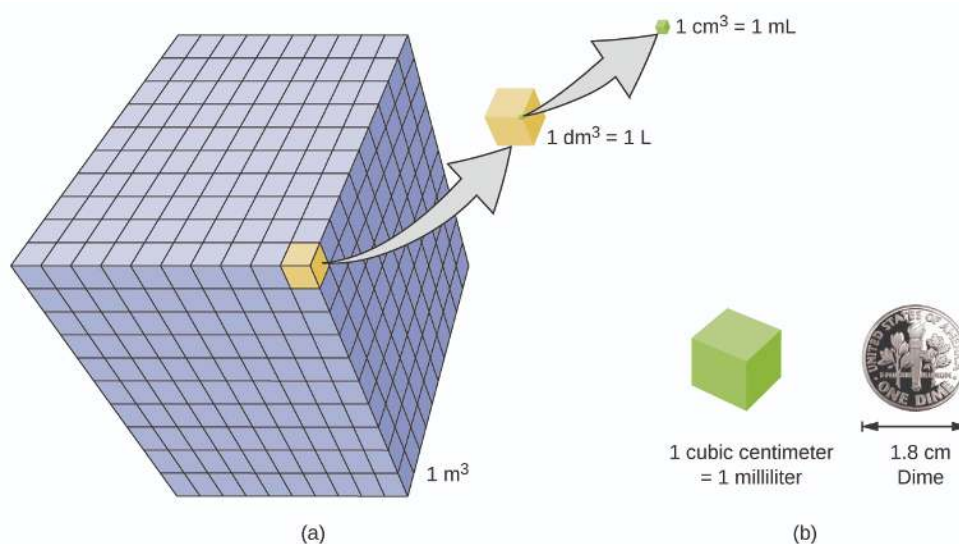


Figure 1.6.3: Units of Volume. (a) The relative volumes are shown for cubes of 1 m³, 1 dm³ (1 L), and 1 cm³ (1 mL) (not to scale). A liter (L) is defined as a cube 1 dm (1/10th of a meter) on a side. A milliliter (mL), 1/1,000th of a liter, is equal to 1 cubic centimeter. (b) The diameter of a dime is compared relative to the edge length of a 1-cm³ (1-mL) cube. (CC BY 4.0; OpenStax)

Energy, another important quantity in chemistry, is the ability to perform work, such as moving a box of books from one side of a room to the other side. It has a derived unit of $\text{kg}\cdot\text{m}^2/\text{s}^2$. (The dot between the kg and m units implies the units are multiplied together.) Because this combination is cumbersome, this collection of units is redefined as a **joule (J)**. An older unit of energy, but

likely more familiar to you, the calorie (cal), is also widely used. There are 4.184 J in 1 cal. Energy changes occur during all chemical processes and will be discussed in a later chapter.

To Your Health: Energy and Food

The food in our diet provides the energy our bodies need to function properly. The energy contained in food could be expressed in joules or calories, which are the conventional units for energy, but the food industry prefers to use the kilocalorie and refers to it as the Calorie (with a capital C). The average daily energy requirement of an adult is about 2,000–2,500 Calories, which is 2,000,000–2,500,000 calories (with a lowercase c).

If we expend the same amount of energy that our food provides, our body weight remains stable. If we ingest more Calories from food than we expend, however, our bodies store the extra energy in high-energy-density compounds, such as fat, and we gain weight. On the other hand, if we expend more energy than we ingest, we lose weight. Other factors affect our weight as well—genetic, metabolic, behavioral, environmental, cultural factors—but dietary habits are among the most important.

In 2008 the US Centers for Disease Control and Prevention issued a report stating that 73% of Americans were either overweight or obese. More alarmingly, the report also noted that 19% of children aged 6–11 and 18% of adolescents aged 12–19 were overweight—numbers that had tripled over the preceding two decades. Two major reasons for this increase are excessive calorie consumption (especially in the form of high-fat foods) and reduced physical activity. Partly because of that report, many restaurants and food companies are working to reduce the amounts of fat in foods and provide consumers with more healthy food options.

Density is defined as the mass of an object divided by its volume; it describes the amount of matter contained in a given amount of space.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (1.6.3)$$

Thus, the units of density are the units of mass divided by the units of volume: g/cm³ or g/mL (for solids and liquids), g/L (for gases), kg/m³, and so forth. For example, the density of water is about 1.00 g/cm³, while the density of mercury is 13.6 g/mL. (Remember that 1 mL equals 1 cm³.) Mercury is over 13 times as dense as water, meaning that it contains over 13 times the amount of matter in the same amount of space. The density of air at room temperature is about 1.3 g/L. Table 1.6.3 shows the densities of some common substances.

Table 1.6.2: Densities of Common Substances

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

✓ Example 1.6.2: Density of Bone

What is the density of a section of bone if a 25.3 cm³ sample has a mass of 27.8 g?

Solution

Because density is defined as the mass of an object divided by its volume, we can set up the following relationship:

$$\begin{aligned}\text{density} &= \frac{\text{mass}}{\text{volume}} \\ &= \frac{27.8 \text{ g}}{25.3 \text{ cm}^3} \\ &= 1.10 \text{ g/cm}^3\end{aligned}$$

Note that we have limited our final answer to three significant figures.

? Exercise 1.6.2: Density of Oxygen

What is the density of oxygen gas if a 15.0 L sample has a mass of 21.7 g?

Answer

1.45 g/L

Density can be used to convert between the mass and the volume of a substance. This will be discussed in the next section.

Concept Review Exercises

1. What is the difference between a base unit and a derived unit? Give two examples of each type of unit.
2. Do units follow the same mathematical rules as numbers do? Give an example to support your answer.
3. What is density?

Answers

1. Base units are the seven fundamental units of SI; derived units are constructed by making combinations of the base units; Two examples of base units: kilograms and meters (answers will vary); Two examples of derived units: grams per milliliter and joules (answers will vary).
2. yes; $\text{mL} \times \frac{\text{g}}{\text{mL}} = \text{g}$ (answers will vary)
3. Density is defined as the mass of an object divided by its volume

Key Takeaways

- Recognize the SI base units and derived units.
- Combining prefixes with base units creates new units of larger or smaller sizes.

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1.7: Expressing Numbers - Scientific Notation

Learning Objectives

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

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

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

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

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

10^6

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

Recall that:

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

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

$$\underbrace{1,500,000}_{\text{move decimal 6 places left}} = 1.5 \times 10^6$$

✓ Example 1.7.1: Scientific Notation

Express each number in scientific notation.

- 67,000,000,000
- 1,689
- 12.6

Answer a

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

Answer b

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

Answer c

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

? Exercise 1.7.1

Express each number in scientific notation.

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

Answer a

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

Answer b

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

Answer c

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

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

$$1.5 \times 10^6 = 1.\underbrace{500000}_6 = 1,500,000$$

✓ Example 1.7.2

Express each number in standard, or conventional notation.

- a. 5.27×10^4
- b. 1.0008×10^6

Answer a

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

Answer b

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

✓ Exercise 1.7.2

Express each number in standard, or conventional notation.

- a. 6.98×10^8
- b. 1.005×10^2

Answer a

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

Answer b

Moving the decimal point two places to the right gives 100.5

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

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

$$\frac{1}{10} \times \frac{1}{10} \times \frac{1}{10}$$

$$10^{-3}$$

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

Recall that:

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

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

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

Note:

In writing scientific notations, the convention is to have only one digit before the decimal point.

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

✓ Example 1.7.3

Express each number in scientific notation.

- 0.000006567
- 0.0004004
- 0.000000000000123

Answer a

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

Answer b

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

Answer c

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

? Exercise 1.7.3

Express each number in scientific notation.

- a. 0.000355
- b. 0.314159
- c. -0.051204

Answer a

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

Answer b

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

Answer c

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

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

$$6 \times 10^{-3} = \underbrace{006}_{\text{three zeros}} = 0.006$$

📌 Note

Changing a number in scientific notation to standard form:

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

✓ Example 1.7.4

Change the number in scientific notation to standard form.

- a. 6.22×10^{-2}
- b. 9.9×10^{-9}

Answer a

0.0622

Answer b

0.0000000099

? Exercise 1.7.4

Change the number in scientific notation to standard form.

- a. 9.98×10^{-5}
- b. 5.109×10^{-8}

Answer a

0.0000998

Answer b

0.00000005109

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

? Exercise 1.7.1

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

Answer

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

? Exercise 1.7.2

What is the relationship between how many places a decimal point moves and the power of 10 used in changing a conventional number into scientific notation?

Answer

The number of places the decimal point moves equals the power of 10. The power of 10 is positive if the decimal point moves to the left and negative if the decimal point moves to the right.

Key Takeaway

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

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1.8: Expressing Numbers - Significant Figures

Learning Objectives

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

Scientists have established certain conventions for communicating the degree of **precision** of a measurement, which is dependent on the measuring device used. Imagine, for example, that you are using a meterstick to measure the width of a table. The centimeters (cm) marked on the meterstick, tell you how many *centimeters* wide the table is. Many metersticks also have markings for millimeters (mm), so we can measure the table to the nearest *millimeter*. Most metersticks do not have any smaller (or more precise) markings indicated, so you cannot report the measured width of the table any more precise than to the nearest millimeter. However, you can *estimate* one past the smallest marking, in this case the millimeter, to the next decimal place in the measurement (Figure 1.8.1).

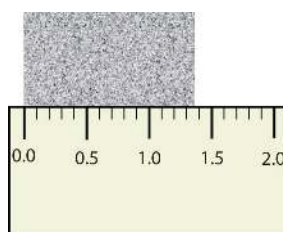


Figure 1.8.1: Measuring an Object to the Correct Number of Digits. How many digits should be reported for the length of this object?

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

“Sig figs” is a common abbreviation for significant figures.

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

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

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

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

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

5. Exact numbers can be considered to have an unlimited number of significant figures, as such calculations are not subject to errors in measurement. This may occur:

- By definition (1 minute = 60 seconds, 1 inch = 2.54 cm, 12 inches = 1 foot, etc.)
- As a result of counting (6 faces on a cube or dice, two hydrogen atoms in a water molecule, 3 peas in a pod, etc.)

Example 1.8.1

How many significant figures does each number have?

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

Answer a

four (by rules 1 and 4)

Answer b

seven (by rules 1 and 2)

Answer c

nine (by rules 1, 2, and 4)

Answer d

four (by rules 1 and 3)

Exercise 1.8.1

How many significant figures does each number have?

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

Answer a

five

Answer b

four

Answer c

four

Answer d

one

Answer e

one

Rounding off numbers

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know

that, round to that many digits, starting from the left. **If the number immediately to the right of the last significant digit is less than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.**

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

Table 1.5.1: Rounding examples

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

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

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

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

Rules for Combining Numbers

For **addition or subtraction**, the rule requires that the **answer should be rounded to the same number of decimal places** as the number with the least decimal places used in the calculation. For example, consider the following addition:

$$\begin{array}{r}
 56.789 \\
 + 102.2 \\
 + 1,300.099 \\
 \hline
 = 1,459.088
 \end{array}$$

↑ Limit to this column

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

For **multiplication or division**, the rule states that the answer should have the **same number of significant figures** as the number with the least significant figure used in the calculation, in other words, we must count the number of significant figures in each number being multiplied or divided and then limit the significant figures in the answer to the lowest count. An example is as follows:

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

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

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

Example 1.8.2

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

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

Answer a

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

Answer b

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

Answer c

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

Exercise 1.8.2

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

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

Answer a

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

Answer b

$$1437 = 1.437 \times 10^3$$

Answer c

$$156 = 1.56 \times 10^2$$

Answer d

$$2.14 = 2.14 \times 10^0$$

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

Concept Review Exercises

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

Answers

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

Key Takeaways

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

Citations

- Significant Digits. (2020, August 15). Retrieved May 18, 2021, from <https://chem.libretexts.org/@go/page/352>
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1.9: Converting Units

Learning Objectives

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

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

Conversion Factors

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

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

or

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

Suppose we divide both sides of the equation by 1 m (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:

$$\text{same quantity} \leftarrow \frac{100 \text{ cm}}{1 \text{ m}} = 1$$

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

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:

$$\frac{3.55}{1} \times \frac{100 \text{ cm}}{1} = 355 \text{ cm} \quad (1.9.1)$$

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 (Equation 1.9.1), 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 1.9.1 shows a **concept map** for constructing a proper conversion.

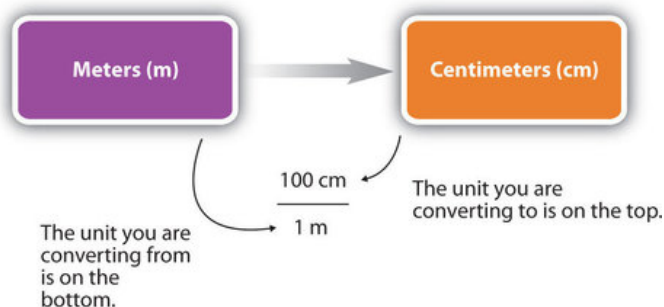


Figure 1.9.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Meters converted to centimeters. A fraction showing 100 cm over 1 m. The unit you convert to is on top, and the unit you convert from is on bottom.

Significant Figures in Conversions

How do conversion factors affect the determination of significant figures? Numbers in conversion factors based on prefix changes, such as kilograms to grams, are *not* considered in the determination of significant figures in a calculation because the numbers in such conversion factors are exact. **Exact numbers** are *defined* or *counted* numbers, not measured numbers, and can be considered as having an infinite number of significant figures. (In other words, 1 kg is exactly 1,000 g, by the definition of kilo-.) Counted numbers are also exact. If there are 16 students in a classroom, the number 16 is exact. In contrast, conversion factors that come from measurements (such as density, as we will see shortly) or are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

✓ Example 1.9.1

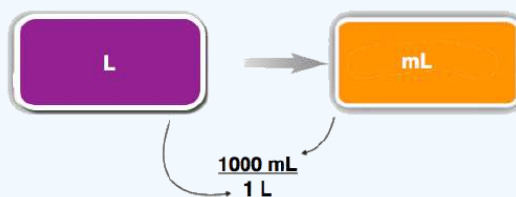
- The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?
- A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?

Solution

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

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

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



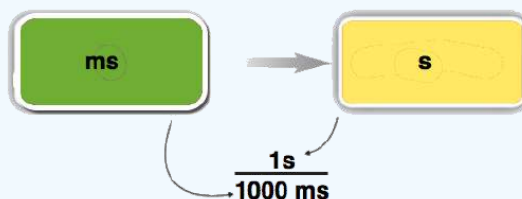
$$4.7 \cancel{\text{L}} \times \frac{1,000 \text{ mL}}{1 \cancel{\text{L}}} = 4,700 \text{ mL}$$

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

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

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

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



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

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

? Exercise 1.9.1

Perform each conversion.

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

Answer a

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

Answer b

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

Conversion Factors From Different Units

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

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

This relationship can be used to construct two conversion factors:

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

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

$$16 \cancel{\text{ mL}} \times \frac{13.6 \text{ g}}{1 \cancel{\text{ mL}}} = 217.6 \text{ g} \\ \approx 220 \text{ g}$$

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

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

✓ **Example 1.9.2: Mercury Thermometer**

A mercury thermometer for measuring a patient's temperature contains 0.750 g of mercury. What is the volume of this mass of mercury?

Solution

Because we are starting with grams, we want to use the conversion factor that has grams in the denominator. The gram unit will cancel algebraically, and milliliters will be introduced in the numerator.

$$0.750 \cancel{\text{ g}} \times \frac{1 \text{ mL}}{13.6 \cancel{\text{ g}}} = 0.055147 \dots \text{ mL} \\ \approx 0.0551 \text{ mL}$$

We have limited the final answer to three significant figures.

? **Exercise 1.9.2**

What is the volume of 100.0 g of air if its density is 1.3 g/L?

Answer

$$100.0 \cancel{\text{ g}} \times \frac{1 \text{ L}}{1.3 \cancel{\text{ g}}} = 76.92307692 \text{ L} \approx 77 \text{ L}$$

Because the density (1.3 g/L) has only 2 significant figures, we are rounding off the final answer to 2 significant figures.

📌 **Looking Closer: Density and the Body**

The densities of many components and products of the body have a bearing on our health.

Bones. Bone density is important because bone tissue of lower-than-normal density is mechanically weaker and susceptible to breaking. The density of bone is, in part, related to the amount of calcium in one's diet; people who have a diet deficient in calcium, which is an important component of bones, tend to have weaker bones. Dietary supplements or adding dairy products to the diet seems to help strengthen bones. As a group, women experience a decrease in bone density as they age. It has been estimated that fully half of women over age 50 suffer from excessive bone loss, a condition known as osteoporosis. Exact bone densities vary within the body, but for a healthy 30-year-old female, it is about 0.95–1.05 g/cm³. Osteoporosis is diagnosed if the bone density is below 0.6–0.7 g/cm³.

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

$$\frac{\text{density of some material}}{\text{density of water}}$$

Normal values for the specific gravity of urine range from 1.002 to 1.028.

Body Fat. The overall density of the body is one indicator of a person's total body fat. Fat is less dense than muscle and other tissues, so as it accumulates, the overall density of the body decreases. Measurements of a person's weight and volume provide the overall body density, which can then be correlated to the percentage of body fat. (The body's volume can be measured by immersion in a large tank of water. The amount of water displaced is equal to the volume of the body.)

Problem Solving With Multiple Conversions

Sometimes you will have to perform more than one conversion to obtain the desired unit. For example, suppose you want to convert 54.7 km into millimeters. You can either memorize the relationship between kilometers and millimeters, or you can do the conversion in two steps. Most people prefer to convert in steps.

To do a stepwise conversion, we first convert the given amount to the base unit. In this example, the base unit is meters. We know that there are 1,000 m in 1 km:

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

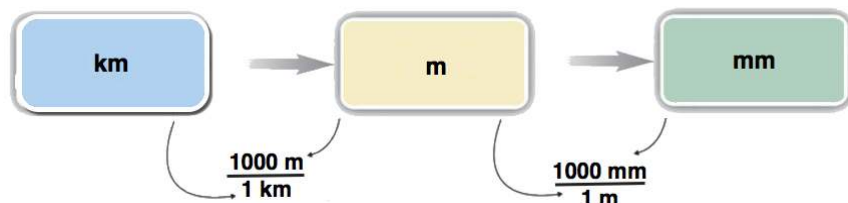
Then we take the result (54,700 m) and convert it to millimeters, remembering that there are 1,000 mm for every 1 m:

$$\begin{aligned} 54,700 \cancel{\text{ m}} \times \frac{1,000 \text{ mm}}{1 \cancel{\text{ m}}} &= 54,700,000 \text{ mm} \\ &= 5.47 \times 10^7 \text{ mm} \end{aligned}$$

We have expressed the final answer in scientific notation.

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

Concept Map



Calculation

$$\begin{aligned} 54.7 \cancel{\text{ km}} \times \frac{1,000 \cancel{\text{ m}}}{1 \cancel{\text{ km}}} \times \frac{1,000 \text{ mm}}{1 \cancel{\text{ m}}} &= 54,700,000 \text{ mm} \\ &= 5.47 \times 10^7 \text{ mm} \end{aligned}$$

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

Either method—one step at a time or all the steps together—is acceptable. If you do all the steps together, the restriction for the proper number of significant figures should be done after the last step. As long as the math is performed correctly, you should get

the same answer no matter which method you use.

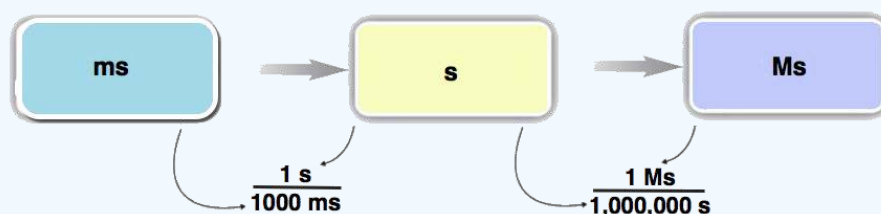
✓ Example 1.9.3

Convert 58.2 ms to megaseconds in one multistep calculation.

Solution

First, convert the given unit (ms) to the base unit—in this case, seconds—and then convert seconds to the final unit, megaseconds:

Concept Map



Calculation

$$58.2 \cancel{\text{ms}} \times \frac{1 \cancel{\text{s}}}{1,000 \cancel{\text{ms}}} \times \frac{1 \text{ Ms}}{1,000,000 \cancel{\text{s}}} = 0.000000582 \text{ Ms}$$

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

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

? Exercise 1.9.3

Convert 43.007 mg to kilograms in one multistep calculation.

Answer

$$43.007 \cancel{\text{mg}} \times \frac{1 \cancel{\text{g}}}{1,000 \cancel{\text{mg}}} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 0.000043007 \text{ kg}$$

$$= 4.3007 \times 10^{-5} \text{ kg}$$

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

📌 Career Focus: Pharmacist

A pharmacist dispenses drugs that have been prescribed by a doctor. Although that may sound straightforward, pharmacists in the United States must hold a doctorate in pharmacy and be licensed by the state in which they work. Most pharmacy programs require four years of education in a specialty pharmacy school.

Pharmacists must know a lot of chemistry and biology so they can understand the effects that drugs (which are chemicals, after all) have on the body. Pharmacists can advise physicians on the selection, dosage, interactions, and side effects of drugs. They can also advise patients on the proper use of their medications, including when and how to take specific drugs properly. Pharmacists can be found in drugstores, hospitals, and other medical facilities.

Curiously, an outdated name for pharmacist is *chemist*, which was used when pharmacists formerly did a lot of drug preparation, or *compounding*. In modern times, pharmacists rarely compound their own drugs, but their knowledge of the sciences, including chemistry, helps them provide valuable services in support of everyone's health.

Key Takeaway

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

Concept Review Exercises

- How do you determine which quantity in a conversion factor goes in the denominator of the fraction?
- State the guidelines for determining significant figures when using a conversion factor.
- Write a concept map (a plan) for how you would convert 1.0×10^{12} nanoliters (nL) to kiloliters (kL).

Answers

- The unit you want to cancel from the numerator goes in the denominator of the conversion factor.
- Exact numbers that appear in many conversion factors do not affect the number of significant figures; otherwise, the normal rules of multiplication and division for significant figures apply.
- Concept Map: Convert the given (nanoliters, nL) to liters; then convert liters to kiloliters.



1.0 times 10 to the 12th nL converts to 1000 Liters converts to 1 kiloliter. One Liter is 1.0 times 10 to the 9th nL. One kiloliter is 1000 Liters.

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1.10: Dosage Calculations

Learning Objectives

- Calculate drug dosages using conversion factors.

To Your Health: Dosages

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

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

Effects are dose-dependent

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

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

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

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

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

Calculation of dose

Calculating drug dosages for humans based on the doses used in animal studies can be based on weight (e.g., mg/kg) or surface area (e.g., mg/m²) based on weight. Drug dosage calculation is required if the physician's order is different from what is available.

Example 1.10.1

- a. A physician ordered 100 mg of Demerol. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?
- b. The doctor's order is 1.2 g of Folic Acid. Folic Acid is available as 800 mg per tablet. How many tablets should be taken?

Solution

- a. We start with the given, 100 mg. We want to change the unit from *mg* to *tablets*. There are 50 mg in 1 tablet (Remember that *per tablet* means *one tablet*. From this relationship, we can construct the conversion factor. We use the conversion factor that

will cancel out the original unit, *mg*, and introduce the unit we are converting to, which is *tablet*.

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

Hence, the nurse should administer 2 tablets.

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

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

Hence, 1.5 tablets should be taken.

? Exercise 1.10.1

Calculate each of the following.

- The physician ordered 20 mg of Valium. Valium is available as 10 mg per tablet. How many tablets should the nurse administer?
- The doctor's order is 1 g of Calcium. What is on hand is Calcium as 500 mg per tablet. How many tablets should be taken?

Answer a

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

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

Answer b

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

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

Drug Dosage Calculation based on Body Weight. Many drugs (especially in children) are dosed according to body weight (mg/kg). These calculations are carried in 3-step conversions. The first step is to convert the body weight from *pounds (lbs)* to *kg*. The second step is to convert *kg* to *mg* (the total *mg* dose calculated based on body weight). And, finally, the *mg* dose is converted to the number of *tablets*. (as in Example 1.8.1).

✓ Example 1.10.2

- Demerol is ordered 1.5mg/kg for a patient that is 220 lbs. Demerol is available as 50 mg per tablet. How many tablets should the nurse administer?
- A doctor prescribes amoxicillin 30mg/kg to a child weighing 73.5 lbs. Amoxicillin is available as 500 mg tablets. How many tablets should the nurse administer?

Solution

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

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

Hence, the nurse administers 3 tablets.

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

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

Hence, the nurse administers 2 tablets.

? Exercise 1.10.2

Calculate each of the following.

- Vancomycin is ordered 15mg/kg for a patient that is 110 lbs. Vancomycin is available as 250 mg per capsule. How many capsules should the nurse administer? (ans. 3 capsules)
- A doctor prescribes ampicillin 40mg/kg to a patient weighing 55 lbs. Ampicillin is available as 500 mg tablets. How many tablets should the nurse administer? (ans. 2 tablets)

Answer a

$$110 \cancel{\text{lbs}} \times \frac{1 \cancel{\text{kg}}}{2.2 \cancel{\text{lbs}}} \times \frac{15 \cancel{\text{mg}}}{1 \cancel{\text{kg}}} \times \frac{1 \text{ tablet}}{250 \cancel{\text{mg}}} = 3 \text{ tablets}$$

Answer b

$$55 \cancel{\text{lbs}} \times \frac{1 \cancel{\text{kg}}}{2.2 \cancel{\text{lbs}}} \times \frac{40 \cancel{\text{mg}}}{1 \cancel{\text{kg}}} \times \frac{1 \text{ tablet}}{500 \cancel{\text{mg}}} = 2 \text{ tablets}$$

References

- [The Use of Body Surface Area as a Criterion of Drug Dosage in Cancer Chemotherapy](https://cancerres.aacrjournals.org)(opens in new window) [cancerres.aacrjournals.org] "The Use of Body Surface Area as a Criterion of Drug Dosage in Cancer Chemotherapy" D Pinkel. Cancer Research 1958

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1.11: Percentages

Learning Outcomes

- Calculate percentages and use in calculations.

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

Percent

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

$$\% = \frac{\text{part}}{\text{whole}} \times 100 \quad (1.11.1)$$

Example 1.11.1

What percent of 70 is 14?

Solution

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

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

$$\% = \frac{14}{70} \times 100$$

$$\% = 0.20 \times 100$$

$$\% = 20$$

Therefore, 14 is **20%** of 70.

Example 1.11.2

What is **35%** of 80?

Solution

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

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

$$35 = \frac{\text{part}}{80} \times 100$$

$$0.35 = \frac{\text{part}}{80}$$

$$\text{part} = 28$$

So 28 is **35%** of 80.

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

Example 1.11.3

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

Solution

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

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

$$20 = \frac{\text{part}}{150} \times 100$$

$$0.20 = \frac{\text{part}}{150}$$

$$\text{part} = 30$$

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

$$\text{original} - \text{loss} = \text{final} \quad (1.11.2)$$

$$150 \frac{\text{mg}}{\text{dL}} - 30 \frac{\text{mg}}{\text{dL}} = 120 \frac{\text{mg}}{\text{dL}} \quad (1.11.3)$$

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

Supplemental Resources

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

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CHAPTER OVERVIEW

2: Atoms and Molecules

- 2.1: Chemical Symbols and Formulas
- 2.2: Atomic Structure
- 2.3: Isotopes - When the Number of Neutrons Varies
- 2.4: Masses of Atoms and Molecules
- 2.5: Avogadro's Number
- 2.6: The Mole
- 2.7: Percentage Composition

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2.1: Chemical Symbols and Formulas

Chemical Symbols and Formulas

To illustrate chemical reactions and the elements and compounds involved in them, chemists use symbols and formulas. A chemical symbol is a one- or two-letter designation of an element. Some examples of chemical symbols are **O** for oxygen, **Zn** for zinc, and **Fe** for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown below.

Table 1: Symbols and Latin Names for Elements

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

CHEMICAL FORMULAS FOR COMPOUNDS

Compounds are combinations of two or more elements. A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a two to one ratio. The chemical formula for water is **H₂O**. Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is **H₂SO₄**.

Classification of compounds according to their molecular structure

Based on the number of atoms, molecules can be classified as:

- **Monoatomic molecule:** contains only one atom. Example: Ar
- **Diatomic molecules:** two atoms. Example: Cl₂
- **Triatomic molecules:** three atoms. Example: CO₂
- **Polyatomic molecules:** four or more types of atoms. Example: H₂SO₄

Summary

A chemical symbol is a one- or two-letter designation of an element. Compounds are combinations of two or more elements. A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements. Some elements have symbols that derive from the Latin name for the element.

Explore More

Use the link below to answer the following questions:

- **Chemical Formulas** at <http://www.elementalmatter.info/chemical-formulas.htm>.

1. What is a chemical formula?

2. What is the chemical formula for glucose? What does this formula signify?
3. What is the chemical formula for caffeine?
4. What is the chemical formula for sugar?

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2.2: Atomic Structure

Learning Objectives

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

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

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

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

These concepts form the basis of chemistry. Although the word *atom* comes from a Greek word that means "indivisible," we understand now that atoms themselves are composed of smaller parts called *subatomic particles*. The first part to be discovered was the electron, a tiny subatomic particle with a negative charge. It is often represented as e^- , with the right superscript showing the negative charge. Later, two larger particles were discovered. The proton is a more massive (but still tiny) subatomic particle with a positive charge, represented as p^+ . The neutron is a subatomic particle with about the same mass as a proton, but no charge. It is represented as either n or n^0 . We now know that all atoms of all elements are composed of electrons, protons, and (with one exception) neutrons. Table 2.2.1 summarizes the properties of these three subatomic particles.

Table 2.2.1: Properties of the Three Subatomic Particles

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

How are these particles arranged in atoms? They are not arranged at random. Experiments by Ernest Rutherford in England in the 1910s pointed to a nuclear model with atoms that has the protons and neutrons in a central nucleus with the electrons in orbit about the nucleus. The relatively massive protons and neutrons are collected in the center of an atom, in a region called the nucleus of the atom (plural *nuclei*). The electrons are outside the nucleus and spend their time orbiting in space about the nucleus. (Figure 2.2.1).

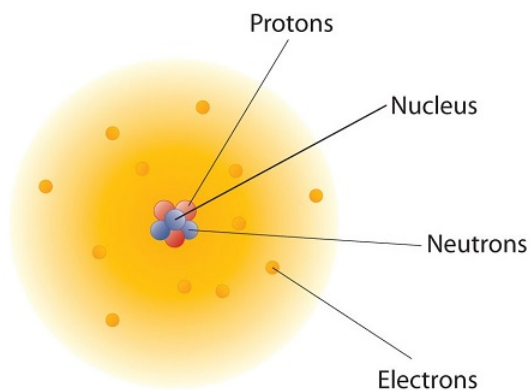


Figure 2.2.1: The Structure of the Atom. Atoms have protons and neutrons in the center, making the nucleus, while the electrons orbit the nucleus.

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

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

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

Example 2.2.1:

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

Solution

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

Exercise 2.2.1

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

When referring to an atom, we simply use the element's name: the term *sodium* refers to the element as well as an atom of sodium. But it can be unwieldy to use the name of elements all the time. Instead, chemistry defines a symbol for each element. The atomic symbol is a one- or two-letter representation of the name of an element. By convention, the first letter of an element's symbol is always capitalized, while the second letter (if present) is lowercase. Thus, the symbol for hydrogen is H, the symbol for sodium is Na, and the symbol for nickel is Ni. Most symbols come from the English name of the element, although some symbols come from an element's Latin name. (The symbol for sodium, Na, comes from its Latin name, *natrium*.) Table 2.2.2 lists some common elements and their symbols. You should memorize the symbols in Table 2.2.2, as this is how we will be representing elements throughout chemistry.

Table 2.2.2: Names and Symbols of Common Elements

Element Name	Symbol	Element Name	Symbol
Aluminum	Al	Mercury	Hg
Argon	Ar	Molybdenum	Mo
Arsenic	As	Neon	Ne
Barium	Ba	Nickel	Ni
Beryllium	Be	Nitrogen	N
Bismuth	Bi	Oxygen	O
Boron	B	Palladium	Pd
Bromine	Br	Phosphorus	P
Calcium	Ca	Platinum	Pt
Carbon	C	Potassium	K
Chlorine	Cl	Radium	Ra
Chromium	Cr	Radon	Rn
Cobalt	Co	Rubidium	Rb
Copper	Cu	Scandium	Sc
Fluorine	F	Selenium	Se
Gallium	Ga	Silicon	Si
Germanium	Ge	Silver	Ag
Gold	Au	Sodium	Na
Helium	He	Strontium	Sr
Hydrogen	H	Sulfur	S
Iodine	I	Tantalum	Ta
Iridium	Ir	Tin	Sn
Iron	Fe	Titanium	Ti
Krypton	Kr	Tungsten	W
Lead	Pb	Uranium	U
Lithium	Li	Xenon	Xe
Magnesium	Mg	Zinc	Zn
Manganese	Mn	Zirconium	Zr

The elements are grouped together in a special chart called the periodic table of all the elements. A simple periodic table is shown in Figure 2.2.2, while one may view a more extensive periodic table from another source. The elements on the periodic table are listed in order of ascending atomic number. The periodic table has a special shape that will become important to us when we consider the organization of electrons in atoms (Chapter 8). One immediate use of the periodic table helps us identify metals and nonmetals. Nonmetals are in the upper right-hand corner of the periodic table, on one side of the heavy line splitting the right-side part of the chart. All other elements are metals.

<table><tr><td>1 H 1.00794</td><td colspan="10"></td><td>1 H 1.00794</td><td>2 He 4.00260</td></tr><tr><td>3 Li 6.941</td><td>4 Be 9.012182</td><td colspan="10"></td><td>5 B 10.811</td><td>6 C 12.0107</td><td>7 N 14.00674</td><td>8 O 15.9994</td><td>9 F 18.9984032</td><td>10 Ne 20.179</td></tr><tr><td>11 Na 22.989770</td><td>12 Mg 24.3050</td><td colspan="10"></td><td>13 Al 26.981538</td><td>14 Si 28.0855</td><td>15 P 30.973761</td><td>16 S 32.066</td><td>17 Cl 35.4527</td><td>18 Ar 39.948</td></tr></table>												1 H 1.00794											1 H 1.00794	2 He 4.00260	3 Li 6.941	4 Be 9.012182											5 B 10.811	6 C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.179	11 Na 22.989770	12 Mg 24.3050											13 Al 26.981538	14 Si 28.0855	15 P 30.973761	16 S 32.066	17 Cl 35.4527	18 Ar 39.948	<table><tr><td>19 K 39.0983</td><td>20 Ca 40.078</td><td>21 Sc 44.955910</td><td>22 Ti 47.867</td><td>23 V 50.9415</td><td>24 Cr 51.9961</td><td>25 Mn 54.938049</td><td>26 Fe 55.845</td><td>27 Co 58.933200</td><td>28 Ni 58.6534</td><td>29 Cu 63.545</td><td>30 Zn 65.39</td><td>31 Ga 69.723</td><td>32 Ge 72.61</td><td>33 As 74.92160</td><td>34 Se 78.96</td><td>35 Br 79.504</td><td>36 Kr 83.80</td></tr><tr><td>37 Rb 85.4678</td><td>38 Sr 87.62</td><td>39 Y 88.90585</td><td>40 Zr 91.224</td><td>41 Nb 92.90638</td><td>42 Mo 95.94</td><td>43 Tc (98)</td><td>44 Ru 101.07</td><td>45 Rh 102.90550</td><td>46 Pd 106.42</td><td>47 Ag 196.56655</td><td>48 Cd 112.411</td><td>49 In 114.818</td><td>50 Sn 118.710</td><td>51 Sb 121.760</td><td>52 Te 127.60</td><td>53 I 126.90447</td><td>54 Xe 131.29</td></tr><tr><td>55 Cs 132.90545</td><td>56 Ba 137.327</td><td>57 La 138.9055</td><td>72 Hf 178.49</td><td>73 Ta 180.94.79</td><td>74 W 183.84</td><td>75 Re 186.207</td><td>76 Os 190.23</td><td>77 Ir 192.217</td><td>78 Pt 195.078</td><td>79 Au 196.56655</td><td>80 Hg 200.59</td><td>81 Tl 204.3833</td><td>82 Pb 207.2</td><td>83 Bi 208.58038</td><td>84 Po (209)</td><td>85 At (210)</td><td>86 Rn (222)</td></tr><tr><td>87 Fr (223)</td><td>88 Ra (226)</td><td>89 Ac (227)</td><td>104 Rf (261)</td><td>105 Db (262)</td><td>106 Sg (263)</td><td>107 Bh (262)</td><td>108 Hs (265)</td><td>109 Mt (266)</td><td>110 (269)</td><td>111 (272)</td><td>112 (277)</td><td></td><td>114 (289) (287)</td><td></td><td>116 (289)</td><td></td><td>118 (293)</td></tr></table>												19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6534	29 Cu 63.545	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.504	36 Kr 83.80	37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 196.56655	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29	55 Cs 132.90545	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.94.79	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.56655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.58038	84 Po (209)	85 At (210)	86 Rn (222)	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (277)		114 (289) (287)		116 (289)		118 (293)
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58 Ce 140.116	59 Pr 140.90765	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.92534	66 Dy 162.50	67 Ho 164.93032	68 Er 167.26	69 Tm 168.93421	70 Yb 173.04	71 Lu 174.967
90 Th 232.0381	91 Pa 231.035888	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Figure 2.2.1: A Simple Periodic Table

There is an easy way to represent isotopes using the atomic symbols. We use the construction:



where X is the symbol of the element, A is the mass number, and Z is the atomic number. Thus, for the isotope of carbon that has 6 protons and 6 neutrons, the symbol is:



where C is the symbol for the element, 6 represents the atomic number, and 12 represents the mass number.

Example 2.2.2:

- What is the symbol for an isotope of uranium that has an atomic number of 92 and a mass number of 235?
- How many protons and neutrons are in ${}^{56}_{26}\text{Fe}$?

Solution

- The symbol for this isotope is ${}^{235}_{92}\text{U}$
- This iron atom has 26 protons and $56 - 26 = 30$ neutrons.

Exercise 2.2.2

How many protons are in ${}^{23}_{11}\text{Na}$?

Answer

11 protons

It is also common to state the mass number after the name of an element to indicate a particular isotope. *Carbon-12* represents an isotope of carbon with 6 protons and 6 neutrons, while *uranium-238* is an isotope of uranium that has 146 neutrons.

A Short History of the Atomic Structure

The basic idea that matter is made up of tiny indivisible particles is very old, appearing in many ancient cultures such as Greece and India. The word *atom* is derived from the ancient Greek word *atomos*, which means "indivisible". This ancient idea was based on philosophical reasoning rather than scientific reasoning, and modern atomic theory was developed throughout a few centuries of research and experimentation.

The first atomic theory based on experimentation was stated by John Dalton, an English scientist, in 1808. By studying the chemical composition of different oxides, Dalton noticed that elements combined in very specific patterns given by small whole numbers. This observation prompted him to support that all matter is made of atoms, although Dalton had no idea about the actual structure of the atoms.

In 1897, J. J. Thomson discovered that the existence of the first *subatomic* particle. He renamed these new negatively charged particles as *electrons*. To explain the overall neutral charge of the atom, Thomson concluded that these electrons must be embedded in an uniform sea of positive charge. In this "plum pudding atomic model", the electrons were seen as embedded in the positive charge like raisins in a plum pudding.

Between 1908 and 1913, Ernest Rutherford and his colleagues Hans Geiger and Ernest Marsden came to have doubts about the Thomson model. Rutherford performed a series of experiments in which they bombarded thin foils of metal with positively charged alpha particles. They spotted alpha particles being deflected by angles greater than 90°. To explain this, Rutherford proposed that the positive charge of the atom is not distributed throughout the atom's volume as Thomson believed, but is concentrated in a tiny nucleus at the center. Rutherford's atomic model introduced the existence of the positively charged nucleus surrounded by the negatively charged electrons.

In 1913, after studying atomic spectra produced by elements, the physicist Niels Bohr proposed a model in which the electrons of an atom were assumed to orbit the nucleus but could only do so in a finite set of orbits, and could jump between these orbits only in discrete changes of energy corresponding to absorption or radiation of a photon. This quantized atomic model, also known as the planetary model of the atom, was used to explain why the electrons' orbits are stable and why elements absorb and emit electromagnetic radiation in discrete lines. The Bohr model of the atom was the first complete physical model of the atom. It described the overall structure of the atom and how atoms bond to each other. Bohr's planetary atom model was not perfect and was soon superseded by the more accurate Schrödinger model, but it is sufficient to understand most chemical and physical properties discussed in this course.

In 1924, Louis de Broglie had proposed that all particles behave like waves to some extent, including the electron. In 1926 Erwin Schrödinger used this idea to develop the Schrödinger model of the atom. In this model, electrons are considered electromagnetic waves mechanics rather than particles. According to this model, it is mathematically impossible to obtain precise values for both the position and the energy (momentum) of an electron at a given point in time. This idea became the uncertainty principle, formulated by Werner Heisenberg in 1927. Because it is impossible to determine the position and the energy of an electron in an atom, Schrödinger's atomic model is based on the "probability" of finding an electron in a certain region around the nucleus. Thus, the orbits in the planetary model of the atom were discarded in favor of **atomic orbitals**, which are zones around the nucleus where a given electron is most likely to be observed.

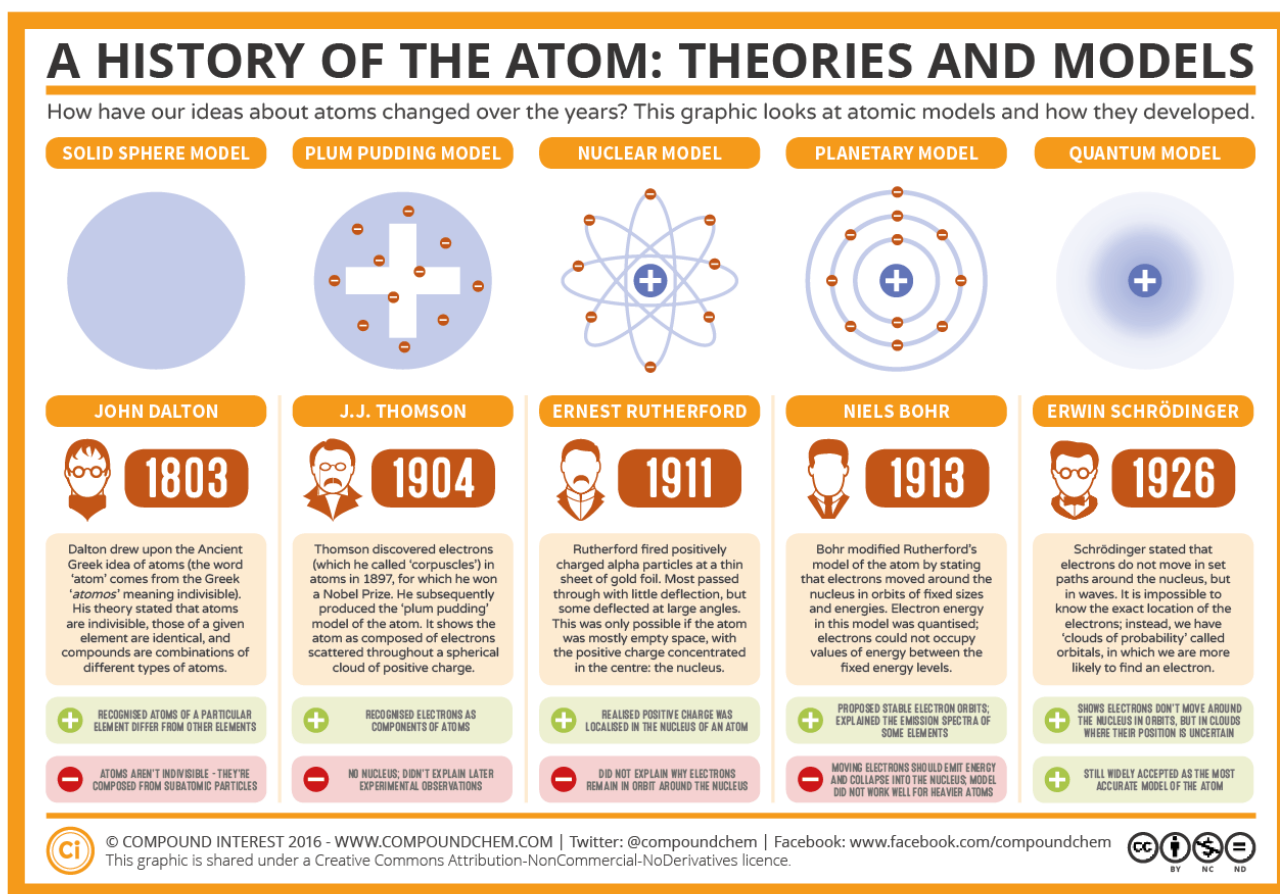


Figure 2.2.2:A summary of atomic models. Image by Compound Interest (2016) <https://www.compoundchem.com/2016/10/13/atomicmodels/>. Shared under CC BY-NC-ND 4.0 license

Key Takeaways

- Chemistry is based on the modern atomic theory, which states that all matter is composed of atoms.
- Atoms themselves are composed of protons, neutrons, and electrons.
- Each element has its own atomic number, which is equal to the number of protons in its nucleus.
- Isotopes of an element contain different numbers of neutrons.
- Elements are represented by an atomic symbol.
- The periodic table is a chart that organizes all the elements.

Citations and attributions

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2.3: Isotopes - When the Number of Neutrons Varies

Learning Objectives

- Explain what isotopes are and how an isotope affects an element's atomic mass.
- Determine the number of protons, electrons, and neutrons of an element with a given mass number.

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 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 2.3.1. Most hydrogen atoms have just one proton, 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.

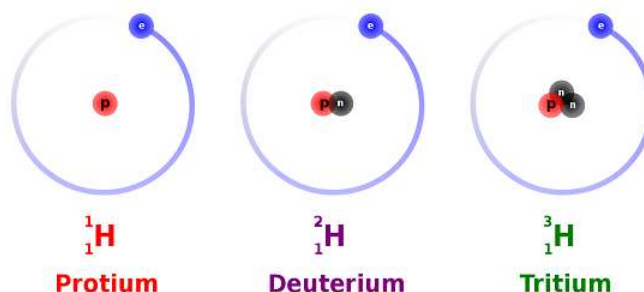


Figure 2.3.1: The three most stable isotopes of hydrogen: protium ($A = 1$), deuterium ($A = 2$), and tritium ($A = 3$). (CC SA-BY 3.0; Balajijagadesh via Wikipedia).

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 an atomic mass of 13 (6 protons + 7 neutrons = 13). These atoms are the isotope called carbon-13.

✓ Example 2.3.1: 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.

Stability of Isotopes

Atoms need a certain ratio of neutrons to protons to have a stable nucleus. Having too many or too few neutrons relative to protons results in an unstable, or radioactive, nucleus that will sooner or later break down to a more stable form. This process is called radioactive decay. Many isotopes have radioactive nuclei, and these isotopes are referred to as radioisotopes. When they decay, they release particles that may be harmful. This is why radioactive isotopes are dangerous and why working with them requires special suits for protection. The isotope of carbon known as carbon-14 is an example of a radioisotope. In contrast, the carbon isotopes called carbon-12 and carbon-13 are stable.

This whole discussion of isotopes brings us back to Dalton's Atomic Theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well! How did Dalton miss this? It turns out that elements found in nature exist as constant uniform mixtures of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundance"). In a chunk of lithium, **93%** will always be lithium with 4 neutrons, while the remaining **7%** will always be lithium with 3 neutrons.

Dalton always experimented with large chunks of an element—chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he was actually observing the averaged properties of all the different isotopes in the sample. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

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:



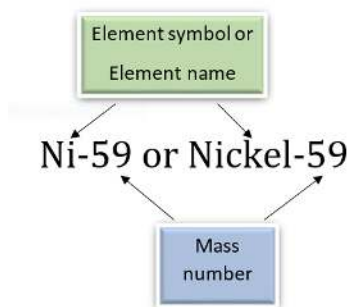
Nuclear symbol for helium-4: The element symbol is He, the mass number to the top left is 4, and the atomic number to the bottom left is 2

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 that 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 all uranium nuclei do; 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.



✓ Example 2.3.2: Potassium-40

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

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 2.3.3: 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 2.3.3

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

Summary

- The number of protons is always the same in atoms of the same element.
- The number of neutrons can be different, even in atoms of the same element.
- Atoms of the same element that contain the same number of protons, but different numbers of neutrons, are known as **isotopes**.
- Isotopes of any given element all contain the same number of protons, so they have the same atomic number (for example, the atomic number of helium is always 2).
- Isotopes of a given element contain different numbers of neutrons, therefore, different isotopes have different mass numbers.

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2.4: Masses of Atoms and Molecules

Learning Objective

- Express the masses of atoms and molecules.

Because matter is defined as anything that has mass and takes up space, it should not be surprising to learn that atoms and molecules have mass.

Individual atoms and molecules, however, are very small, and the masses of individual atoms and molecules are also very small. For macroscopic objects, we use units such as grams and kilograms to state their masses, but these units are much too big to comfortably describe the masses of individual atoms and molecules. Another scale is needed.

Atomic Mass Unit

The atomic mass unit (u; some texts use amu, but this older style is no longer accepted) is defined as one-twelfth of the mass of a carbon-12 atom, an isotope of carbon that has six protons and six neutrons in its nucleus. By this scale, the mass of a proton is 1.00728 u, the mass of a neutron is 1.00866 u, and the mass of an electron is 0.000549 u. There will not be much error if you estimate the mass of an atom by simply counting the total number of protons and neutrons in the nucleus (i.e., identify its mass number) and ignore the electrons. Thus, the mass of carbon-12 is about 12 u, the mass of oxygen-16 is about 16 u, and the mass of uranium-238 is about 238 u. More exact masses are found in scientific references—for example, the exact mass of uranium-238 is 238.050788 u, so you can see that we are not far off by using the whole-number value as the mass of the atom.

What is the mass of an element? This is somewhat more complicated because most elements exist as a mixture of isotopes, each of which has its own mass. Thus, although it is easy to speak of the mass of an atom, when talking about the mass of an element, we must take the isotopic mixture into account.

Average Atomic Mass

The average atomic mass of an element is a weighted average of the masses of the isotopes that compose an element. What do we mean by a weighted average? Well, consider an element that consists of two isotopes, 50% with mass 10 u and 50% with mass 11 u. A weighted average is found by multiplying each mass by its fractional occurrence (in decimal form) and then adding all the products. The sum is the weighted average and serves as the formal atomic mass of the element. In this example, we have the following:

$0.50 \times 10 \text{ u}$	$= 5.0 \text{ u}$
$0.50 \times 11 \text{ u}$	$= 5.5 \text{ u}$
Sum	$= 10.5 \text{ u} = \text{the atomic mass of our element}$

Note that no atom in our hypothetical element has a mass of 10.5 u; rather, that is the average mass of the atoms, weighted by their percent occurrence.

This example is similar to a real element. Boron exists as about 20% boron-10 (five protons and five neutrons in the nuclei) and about 80% boron-11 (five protons and six neutrons in the nuclei). The atomic mass of boron is calculated similarly to what we did for our hypothetical example, but the percentages are different:

$0.20 \times 10 \text{ u}$	$= 2.0 \text{ u}$
$0.80 \times 11 \text{ u}$	$= 8.8 \text{ u}$
Sum	$= 10.8 \text{ u} = \text{the atomic mass of boron}$

Thus, we use 10.8 u for the atomic mass of boron.

Virtually all elements exist as mixtures of isotopes, so atomic masses may vary significantly from whole numbers. Table **2.4.1** lists the atomic masses of some elements. The atomic masses in Table **2.4.1** are listed to three decimal places where possible, but in most cases, only one or two decimal places are needed. Note that many of the atomic masses, especially the larger ones, are not very close to whole numbers. This is, in part, the effect of an increasing number of isotopes as the atoms increase in size. (The record number is 10 isotopes for tin.)

Table **2.4.1**: Selected Atomic Masses of Some Elements

Element Name	Atomic Mass (u)	Element Name	Atomic Mass (u)
Aluminum	26.981	Molybdenum	95.94
Argon	39.948	Neon	20.180
Arsenic	74.922	Nickel	58.693
Barium	137.327	Nitrogen	14.007
Beryllium	9.012	Oxygen	15.999
Bismuth	208.980	Palladium	106.42
Boron	10.811	Phosphorus	30.974
Bromine	79.904	Platinum	195.084
Calcium	40.078	Potassium	39.098
Carbon	12.011	Radium	n/a
Chlorine	35.453	Radon	n/a
Cobalt	58.933	Rubidium	85.468
Copper	63.546	Scandium	44.956
Fluorine	18.998	Selenium	78.96
Gallium	69.723	Silicon	28.086
Germanium	72.64	Silver	107.868
Gold	196.967	Sodium	22.990
Helium	4.003	Strontium	87.62
Hydrogen	1.008	Sulfur	32.065
Iodine	126.904	Tantalum	180.948
Iridium	192.217	Tin	118.710
Iron	55.845	Titanium	47.867
Krypton	83.798	Tungsten	183.84
Lead	207.2	Uranium	238.029
Lithium	6.941	Xenon	131.293
Magnesium	24.305	Zinc	65.409
Manganese	54.938	Zirconium	91.224
Mercury	200.59	Molybdenum	95.94

Note: Atomic mass is given to three decimal places, if known.

Average Molecular Mass

Now that we understand that atoms have mass, it is easy to extend the concept to the mass of molecules. The average molecular mass is the sum of the average atomic masses of the atoms in a molecule. This may seem like a trivial extension of the concept, but it is important to count the number of each type of atom in the molecular formula. Also, although each atom in a molecule is a particular isotope, we use the weighted average, or atomic mass, for each atom in the molecule.

For example, if we were to determine the molecular mass of dinitrogen trioxide, N_2O_3 , we would need to add the atomic mass of nitrogen two times with the atomic mass of oxygen three times:

2 N masses = $2 \times 14.007 \text{ u}$	= 28.014 u
3 O masses = $3 \times 15.999 \text{ u}$	= 47.997 u
Total	= 76.011 u = the molecular mass of N_2O_3

We would not be far off if we limited our numbers to one or even two decimal places.

Example 2.4.1:

What is the molecular mass of each substance?

1. NBr_3
2. C_2H_6

Solution

1. Add one atomic mass of nitrogen and three atomic masses of bromine:

1 N mass	= 14.007 u
3 Br masses = $3 \times 79.904 \text{ u}$	= 239.712 u
Total	= 253.719 u = the molecular mass of NBr_3

2. Add two atomic masses of carbon and six atomic masses of hydrogen:

2 C masses = $2 \times 12.011 \text{ u}$	= 24.022 u
6 H masses = $6 \times 1.008 \text{ u}$	= 6.048 u
Total	= 30.070 u = the molecular mass of C_2H_6

The compound C_2H_6 also has a common name—ethane.

Exercise 2.4.1

What is the molecular mass of each substance?

- a. SO_2
- b. PF_3

Answer a

64.063 u

Answer b

87.968 u

Chemistry is Everywhere: Sulfur Hexafluoride

On March 20, 1995, the Japanese terrorist group Aum Shinrikyo (Sanskrit for "Supreme Truth") released some sarin gas in the Tokyo subway system; twelve people were killed, and thousands were injured (part (a) in the accompanying figure). Sarin (molecular formula $\text{C}_4\text{H}_{10}\text{FPO}_2$) is a nerve toxin that was first synthesized in 1938. It is regarded as one of the most deadly toxins known, estimated to be about 500 times more potent than cyanide. Scientists and engineers who study the spread of chemical weapons such as sarin (yes, there are such scientists) would like to have a similar less dangerous chemical to study, indeed one that is nontoxic, so they are not at risk themselves.

Sulfur hexafluoride is used as a model compound for sarin. SF_6 (a molecular model of which is shown in part (b) in the accompanying figure) has a similar molecular mass (about 146 u) as sarin (about 140 u), so it has similar physical properties in the vapor phase. Sulfur hexafluoride is also very easy to accurately detect, even at low levels, and it is not a normal part of the atmosphere, so there is little potential for contamination from natural sources. Consequently, SF_6 is also used as an aerial tracer for ventilation systems in buildings. It is nontoxic and very chemically inert, so workers do not have to take special precautions other than watching for asphyxiation.

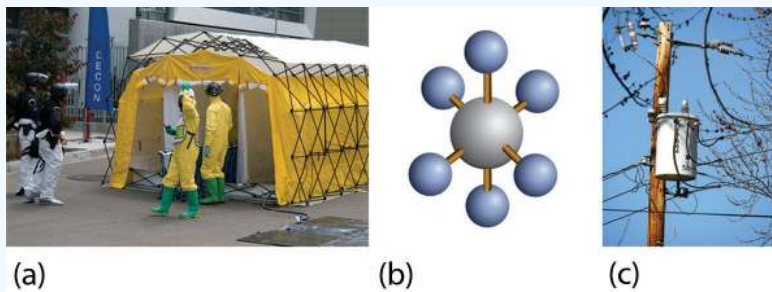


Figure 2.4.2: Sarin and Sulfur Hexafluoride © Thinkstock. (a) Properly protected workers clear out the Tokyo subway after the nerve toxin sarin was released. (b) A molecular model of SF_6 . (c) A high-voltage electrical switchgear assembly that would be filled with SF_6 as a spark suppressant.

Sulfur hexafluoride also has another interesting use: a spark suppressant in high-voltage electrical equipment. High-pressure SF_6 gas is used in place of older oils that may have contaminants that are not environmentally friendly (part (c) in the accompanying figure).

Key Takeaways

- The atomic mass unit (u) is a unit that describes the masses of individual atoms and molecules.
- The atomic mass is the weighted average of the masses of all isotopes of an element.
- The molecular mass is the sum of the masses of the atoms in a molecule.

Citations

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2.5: Avogadro's Number



Figure 2.5.1 (Public Domain; User:OSX/Wikimedia Commons via [Wikipedia](#))

Is there an easier way to load this truck?

When the weather is nice, many people begin to work on their yards and homes. For many projects, sand is needed as a foundation for a walk or to add to other materials. You could order up twenty million grains of sand and have people really stare at you. You could order by the pound, but that takes a lot of time weighing out. The best bet is to order by the yard, meaning a cubic yard. The loader can easily scoop up what you need and put it directly in your truck.

Avogadro's Number

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

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



Figure 2.5.2: Italian scientist Amadeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry. (Public Domain; C. Sentier via [Wikipedia](#))

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



Figure 2.5.3: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year. ((left; Left: Michael David Hill, 2005 (Mikiwikipikidikipedia); Right: chrisbb@prodigy.net;(left) Michael David Hill, 2005 (Mikiwikipikidikipedia); (right) chrisbb@prodigy.net via Left: http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; Right: <http://www.flickr.com/photos/chrisbrenschmidt/436990097>;(left) http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; (right) <http://www.flickr.com/photos/chrisbrenschmidt/436990097>) CC-BY 2.5; (right) CC-BY 2.0)



Summary

- A mole of any substance contains Avogadro's number (6.02×10^{23}) of representative particles.

Review

1. What is the SI unit for amount of a substance?
2. What is the representative particle for an element?
3. The formula unit is the representative particle for what?

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2.6: The Mole

Learning Objectives

- Describe the unit *mole*.
- Relate the mole quantity of substance to its mass.

So far, we have been talking about chemical substances in terms of individual atoms and molecules. Yet we do not typically deal with substances an atom or a molecule at a time; we work with millions, billions, and trillions of atoms and molecules at a time. What we need is a way to deal with macroscopic, rather than microscopic, amounts of matter. We need a unit of amount that relates quantities of substances on a scale that we can interact with.

Chemistry uses a unit called mole. A **mole** (mol) is a number of things equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \times 10^{23} \text{ things}$$

Understand that a mole means a number of things, just like a dozen means a certain number of things-twelve, in the case of a dozen. But a mole is a much larger number of things. These things can be atoms, or molecules, or eggs; however, in chemistry, we usually use the mole to refer to the amounts of atoms or molecules. Although the number of things in a mole is known to eight decimal places, it is usually fine to use only two or three decimal places in calculations. The numerical value of things in a mole is often called *Avogadro's number* (N_A), which is also known as the *Avogadro constant*, after Amadeo Avogadro, an Italian chemist who first proposed its importance.

Example 2.6.1:

How many molecules are present in 2.76 mol of H_2O ? How many molecules is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O , we can also determine the number of atoms in the sample.

$$2.76 \cancel{\text{mol H}_2\text{O}} \times \frac{6.022 \times 10^{23} \text{ molecules H}_2\text{O}}{\cancel{\text{mol H}_2\text{O}}} = 1.66 \times 10^{24} \text{ molecules H}_2\text{O} \quad (2.6.1)$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24} \cancel{\text{molecules H}_2\text{O}} \times \frac{3 \text{ atoms}}{1 \cancel{\text{molecule}}} = 4.99 \times 10^{24} \text{ atoms} \quad (2.6.2)$$

Exercise 2.6.1

How many molecules are present in 4.61×10^{-2} mol of O_2 ?

Answer

$$2.78 \times 10^{22} \text{ molecules}$$

How big is a mole? It is very large. Suppose you had a mole of dollar bills that need to be counted. If everyone on earth (about 6 billion people) counted one bill per second, it would take about 3.2 million years to count all the bills. A mole of sand would fill a cube about 32 km on a side. A mole of pennies stacked on top of each other would have about the same diameter as our galaxy, the Milky Way. A mole is a lot of things-but atoms and molecules are very tiny. One mole of carbon atoms would make a cube that is 1.74 cm on a side, small enough to carry in your pocket.

Why is the mole unit so important? It represents the link between the microscopic and the macroscopic, especially in terms of mass. *A mole of a substance has the same mass in grams as one unit (atom or molecules) has in atomic mass units.* The mole unit allows us to express amounts of atoms and molecules in visible amounts that we can understand.

For example, we already know that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022×10^{23} atoms:

$$12 \text{ g C} = 6.022 \times 10^{23} \text{ atoms C}$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50×10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25} \text{ atoms C} \times \frac{12.0000 \text{ g C}}{6.022 \times 10^{23} \text{ atoms C}} = 299 \text{ g C} \quad (2.6.3)$$

But it also goes beyond carbon. Previously we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the number of grams in 1 mol of atoms of that element, while the molecular mass of a compound is the number of grams in 1 mol of molecules of that compound. Sometimes these masses are called **molar masses** to emphasize the fact that they are the mass for 1 mol of things. (The term *molar* is the adjective form of mole and has nothing to do with teeth.)

Here are some examples. The mass of a hydrogen atom is 1.0079 u; the mass of 1 mol of hydrogen atoms is 1.0079 g. Elemental hydrogen exists as a diatomic molecule, H_2 . One molecule has a mass of $1.0079 + 1.0079 = 2.0158$ u, while 1 mol H_2 has a mass of 2.0158 g. A molecule of H_2O has a mass of about 18.01 u; 1 mol H_2O has a mass of 18.01 g. A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g. In each of these moles of substances, there are 6.022×10^{23} units: 6.022×10^{23} atoms of H, 6.022×10^{23} molecules of H_2 and H_2O , 6.022×10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

Example 2.6.2

What is the molar mass of $\text{C}_6\text{H}_{12}\text{O}_6$?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table.

$6 \text{ C} = 6 \times 12.011$	$= 72.066$
$12 \text{ H} = 12 \times 1.0079$	$= 12.0948$
$6 \text{ O} = 6 \times 15.999$	$= 95.994$
TOTAL	$= 180.155 \text{ g/mol}$

Per convention, the unit *grams per mole* is written as a fraction.

Exercise 2.6.2

What is the molar mass of AgNO_3 ?

Answer

169.87 g/mol

Knowing the molar mass of a substance, we can calculate the number of moles in a certain mass of a substance and vice versa, as these examples illustrate. The molar mass is used as the conversion factor.

Example 2.6.3

What is the mass of 3.56 mol of HgCl_2 ? The molar mass of HgCl_2 is 271.49 g/mol.

Solution

Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$3.56 \text{ mol HgCl}_2 \times \frac{271.49 \text{ g HgCl}_2}{\text{mol HgCl}_2} = 967 \text{ g HgCl}_2 \quad (2.6.4)$$

Exercise 2.6.3

What is the mass of 33.7 mol of H₂O?

Answer

607 g

Example 2.6.4

How many moles of H₂O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H₂O as a conversion factor from mass to moles. The molar mass of water is (1.0079 + 1.0079 + 15.999) = 18.015 g/mol. However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0 \cancel{\text{ g H}_2\text{O}} \times \frac{1 \text{ mol H}_2\text{O}}{18.015 \cancel{\text{ g H}_2\text{O}}} = 13.32 \text{ mol H}_2\text{O} \quad (2.6.5)$$

Exercise 2.6.4

How many moles are present in 35.6 g of H₂SO₄ (molar mass = 98.08 g/mol)?

Answer

0.363 mol

Other conversion factors can be combined with the definition of mole-density, for example.

Example 2.6.5

The density of ethanol is 0.789 g/mL. How many moles are in 100.0 mL of ethanol? The molar mass of ethanol is 46.08 g/mol.

Solution

Here, we use density to convert from volume to mass and then use the molar mass to determine the number of moles.

$$100 \cancel{\text{ mL ethanol}} \times \frac{0.789 \text{ g}}{\cancel{\text{ mL}}} \times \frac{1 \text{ mol}}{46.08 \cancel{\text{ g}}} = 1.71 \text{ mol ethanol} \quad (2.6.6)$$

Exercise 2.6.5

If the density of benzene, C₆H₆, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer

0.201 mol

Summary

The mole is a key unit in chemistry. The molar mass of a substance, in grams, is numerically equal to one atom's or molecule's mass in atomic mass units.

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2.7: Percentage Composition

We are all familiar with the term *percentage*. We take an exam that is worth 100 points, we get 96 correct and we describe our score as 96%. We arrive at 96% by first taking our score and dividing it by the total number of points to get the *fraction* that we got correct. To convert the fraction to a *percentage*, we multiply the fraction by 100.

$$\left(\frac{96 \text{ points}}{100 \text{ points total}} \right) \times 100 = 96\%$$

Applying this concept to molecules, we could describe HCl as consisting of one atom of hydrogen and one atom of chlorine. Likewise, we could use *mole* nomenclature and say that one *mole* of the molecule consists of one *mole* of hydrogen and one *mole* of chlorine, and that one *mole* of HCl has a mass of 36.46 grams. These descriptions, however, tell us nothing about how much of this mass is attributable to the hydrogen and how much comes from the chlorine. In order to do this, we need to speak of the **percentage composition** of a molecule, that is, what percent of the total mass arises from each element. These calculations are simple and involve taking the atomic mass of the element in question and dividing by the molar mass of the molecule.

For HCl, the *fraction* of hydrogen in HCl is given by the molar mass of hydrogen divided by the molar mass of HCl:

$$\frac{1.008 \frac{\text{g}}{\text{mol}}}{36.46 \frac{\text{g}}{\text{mol}}}$$

and the *percentage* of hydrogen in HCl is obtained by multiplying the *fraction* by 100:

$$0.02765 \times 100 = 2.765\%$$

Combining the steps, the *percentage* of chlorine in HCl can be calculated by dividing the molar mass of chlorine by the molar mass of HCl and multiplying by 100:

$$\frac{1.008 \frac{\text{g}}{\text{mol}}}{36.46 \frac{\text{g}}{\text{mol}}} = 0.02765$$

? Exercise 2.7.1

- Find the percentage of fluorine in calcium fluoride (CaF₂).

These types of problems can also be presented as mass calculations. For example, determine the mass of calcium in 423.6 grams of CaF₂. Collecting our known, given and find values:

The known molar mass of Ca is

$$\frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}}$$

The known molar mass of CaF₂ is

$$\frac{78.08 \text{ g CaF}_2}{1 \text{ mol CaF}_2}$$

We are given that we have a sample of CaF₂ with a mass of 423.6 grams and we want to find the mass of Ca in this sample. We could find the mass of Ca if we knew the fraction of Ca in CaF₂. We could then multiply this fraction by the known mass of CaF₂ to obtain the mass of calcium in the sample. The fraction of Ca in CaF₂ is the ratio of the two known molar masses (the percentage, before you multiply by 100). As always, to perform the dimensional analysis, we arrange the known and the given so that the units cancel, leaving only the units of the item we want to find.

$$(423.6 \text{ g CaF}_2) \times \left(\frac{1 \text{ mol}}{78.08 \text{ g CaF}_2} \right) \times \left(\frac{40.08 \text{ g Ca}}{1 \text{ mol Ca}} \right) = x \text{ g Ca} = 217.4 \text{ g Ca}$$

Note that in this example, you want to use the fraction of Ca in CaF₂, not the percentage; if you used percentage, you would have to divide your answer by 100 to get the proper number of grams of Ca in the sample.

? Exercise 2.7.1

- A sample of CaF_2 is known to contain 18.00 grams of calcium; what mass of fluorine is contained in this sample?
- Carbon dioxide is a green house gas produced in combustion. What is the percentage of oxygen in CO_2 ?
- Barium sulfate is used when x-raying the gastrointestinal track. Determine the mass of barium in 523 grams of BaSO_4 .
- Sulfuric acid, H_2SO_4 is the most used chemical in industrial processes. If a sample of sulfuric acid contained 3.67 g of hydrogen how many grams of sulfur would it contain?

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CHAPTER OVERVIEW

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3.1: Periodic Law



Figure 3.1.1: Copy and Paste Caption here. (CC-BY 2.0; Christopher Fynn via [Flickr, Mysore - Fruit vendor](#) [www.flickr.com])

How are these items related to one another?

We have all enjoyed looking around a market for delicious foods to eat later at home. When you get to the market you know you need to get fruits, vegetables and grains to ensure you eat a balanced diet. In the market, these items are all grouped together to make it easier to find the type of fruit, vegetable or grain you are looking for; it would be inconvenient if the different fruits were all scattered in different places. The periodic table is organized in a similar way, ensuring similar elements are found in the same group or period.

The Periodic Law

When Mendeleev put his periodic table together, nobody knew about the existence of the nucleus. It was not until 1911 that Rutherford conducted his gold foil experiment that demonstrated the presence of the nucleus in the atom. Just two years later, in 1913, English physicist Henry Moseley (1887-1915) examined x-ray spectra of a number of chemical elements. He would shoot x-rays through crystals of the element and study the wavelengths of the radiation that he detected. Moseley found that there was a relationship between wavelength and atomic number. His results led to the definition of atomic number as the number of protons contained in the nucleus of each atom. He then realized that the elements of the periodic table should be arranged in order of increasing atomic number, rather than increasing atomic mass.

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Uut	116 Uup	117 Uus	118 Uuo
Lanthanides																		
	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu			
Actinides																		
	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr			

Figure 3.1.2 (Public Domain; User:Cepheus/Wikimedia Commons via [Commons Wikimedia, Periodic Table](#) [commons.wikimedia.org])

When ordered by atomic number, the discrepancies within Mendeleev's table disappeared. Tellurium has an atomic number of 52, while iodine has an atomic number of 53. So even though tellurium does indeed have a greater atomic mass than iodine, it is properly placed before iodine in the periodic table. Mendeleev and Moseley are credited with being most responsible for the modern **periodic law**: When elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties. The result is the periodic table as we know it today. Each new horizontal row of the periodic table corresponds to the beginning of a new **period** because a new principal energy level is being filled with electrons. Elements with similar chemical properties appear at regular intervals, within the vertical columns called **groups**.



Summary

- Elements of the periodic table are arranged in order of increasing atomic number.
- The periodic law states: "When elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties."

Review

1. Did Mendeleev know about the nucleus of an atom?
2. Who discovered the relationship between wavelength of X-rays and atomic number?
3. What did Moseley conclude from his research?
4. What is the "periodic law"?
5. What do the vertical columns (groups) in the periodic table represent?

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3.2: The Bohr Model - Atoms with Orbits

Learning Objectives

- Define an energy level in terms of the Bohr model.
- Discuss how the Bohr model can be used to explain atomic spectra.

Electric light bulbs contain a very thin wire in them that emits light when heated. The wire is called a filament. The particular wire used in light bulbs is made of tungsten. A wire made of any metal would emit light under these circumstances, but tungsten was chosen because the light it emits contains virtually every frequency and therefore, the light emitted by tungsten appears white. A wire made of some other element would emit light of some color that was not convenient for our uses. Every element emits light when energized by heating or passing electric current through it. Elements in solid form begin to glow when they are heated sufficiently, and elements in gaseous form emit light when electricity passes through them. This is the source of light emitted by neon signs and is also the source of light in a fire.



Figure 3.2.1: Human/Need/Desire. Neon sculpture by Bruce Nauman (1983), who has been characterized as a conceptual artist.

Each Element Has a Unique Spectrum

The light frequencies emitted by atoms are mixed together by our eyes so that we see a blended color. Several physicists, including Angstrom in 1868 and Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out so they could see the individual frequencies that made up the light. The **emission spectrum** (or **atomic spectrum**) of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity.



Figure 3.2.2: Atomic Emission Spectrum of Hydrogen.

When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the color is spread out, we see that the hydrogen spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors. Every atom has its own characteristic spectrum; no two atomic spectra are alike. The image below shows the emission spectrum of iron. Because each element has a unique emission spectrum, elements can be defined using them.

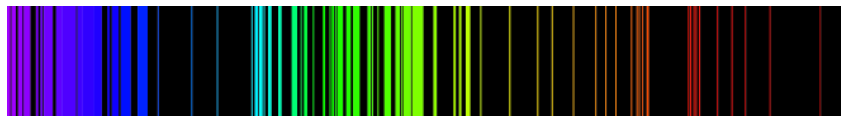


Figure 3.2.3: Atomic Emission Spectrum of Iron.

You may have heard or read about scientists discussing what elements are present in the sun or some more distant star, and after hearing that, wondered how scientists could know what elements were present in a place no one has ever been. Scientists determine what elements are present in distant stars by analyzing the light that comes from stars and finding the atomic spectrum of elements in that light. If the exact four lines that compose hydrogen's atomic spectrum are present in the light emitted from the star, that element contains hydrogen.

Bohr's Model of the Atom

By 1913, the concept of the atom had evolved from Dalton's indivisible spheres idea, to J. J. Thomson's plum pudding model, and then to Rutherford's nuclear atom theory. Rutherford, in addition to carrying out the brilliant experiment that demonstrated the presence of the atomic nucleus, also proposed that the electrons circled the nucleus in a planetary type motion. The solar system or planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system.



Figure 3.2.3: Niels Bohr with Albert Einstein at Paul Ehrenfest's home in Leiden (December 1925).

Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moves in a curved path, it gives off some form of light and loses energy in doing so. This is, after all, how we produce TV signals. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus as it loses energy, and would eventually crash into the nucleus. Furthermore, Rutherford's model was unable to describe how electrons give off light forming each element's unique atomic spectrum. These difficulties cast a shadow on the planetary model and indicated that, eventually, it would have to be replaced.

In 1913, the Danish physicist Niels Bohr proposed a model of the electron cloud of an atom in which electrons orbit the nucleus and were able to produce atomic spectra. Understanding Bohr's model requires some knowledge of electromagnetic radiation (or light).

Energy Levels

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 electron to have a greater amount of energy. The possible orbits are known as **energy levels**. 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.

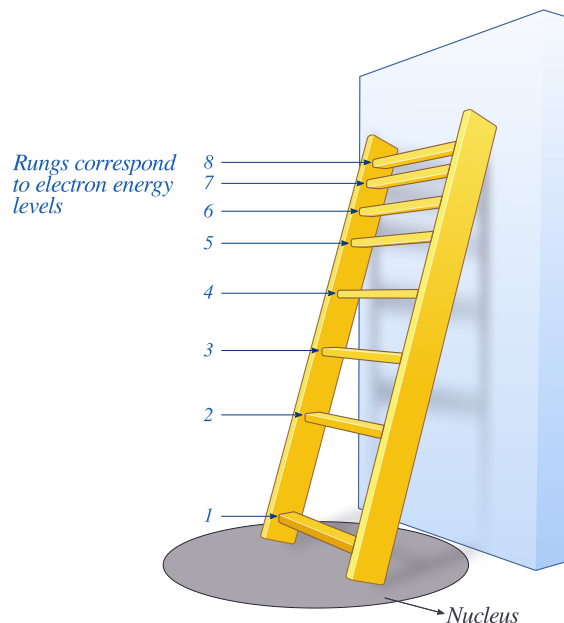


Figure 3.2.4: The energy levels of the electrons can be viewed as rungs on a ladder. Note that the spacing between rungs gets smaller at higher energies (CC BY-NC; Ümit Kaya)

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 the 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 of 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.

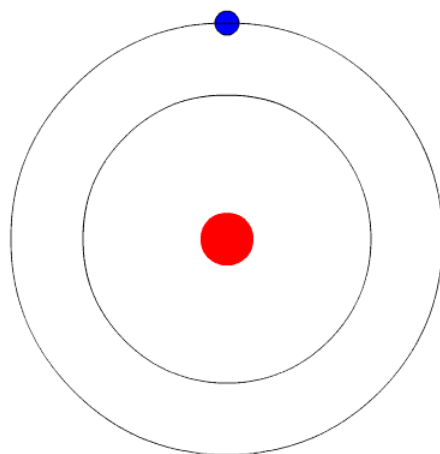


Figure 3.2.5: In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. (CC BY-SA 3.0; Kurzon).

Bohr's Model and Atomic Spectra

The evidence used to support Bohr's model came from the atomic spectra. He suggested that an atomic spectrum is made by the electrons in an atom moving energy levels. The electrons typically have the lowest energy possible, called the **ground state**. If the electrons are given energy (through heat, electricity, light, etc.) the electrons in an atom could absorb energy by jumping to a higher energy level, or **excited state**. The electrons then give off the energy in the form of a piece of light—called a **photon**—that they had absorbed, to fall back to a lower energy level. The energy emitted by electrons dropping back to lower energy levels will always be precise amounts of energy, because the differences in energy levels are precise. This explains why you see specific lines of light when looking at an atomic spectrum—each line of light matches a specific "step down" that an electron can take in that atom. This also explains why each element produces a different atomic spectrum. Because each element has different acceptable energy levels for its electrons, the possible steps each element's electrons can take differ from all other elements.

Summary

- 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.

Vocabulary

- **Emission spectrum (or atomic spectrum)** - The unique pattern of light given off by an element when it is given energy.
- **Energy levels** - Possible orbits that an electron can have in the electron cloud of an atom.
- **Ground state** - To be in the lowest energy level possible.
- **Excited state** - To be in a higher energy level.
- **Photon** - A piece of electromagnetic radiation, or light, with a specific amount of energy.
- **Quantized** - To have a specific amount.

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3.3: Quantum Mechanical Atomic Model



Figure 3.3.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Imagine that a news flash interrupts your favorite TV program... "There has been a hold-up at the First National Bank. The suspect fled in a car and is believed to be somewhere in the downtown district. Everyone is asked to be on the alert." The robber can be located only within a certain area - the police do not have an exact location, just a general idea as to the whereabouts of the thief.

Quantum Mechanical Atomic Model

In 1926, Austrian physicist Erwin Schrödinger (1887-1961) used the wave-particle duality of the electron to develop and solve a complex mathematical equation that accurately described the behavior of the electron in a hydrogen atom. **The quantum mechanical model** of the atom comes from the solution to Schrödinger's equation. Quantization of electron energies is a requirement in order to solve the equation. This is unlike the Bohr model, in which quantization was simply assumed with no mathematical basis.

Recall that in the Bohr model, the exact path of the electron was restricted to very well-defined circular orbits around the nucleus. The quantum mechanical model is a radical departure from that. Solutions to the Schrödinger wave equation, called **wave functions**, give only the probability of finding an electron at a given point around the nucleus. Electrons do not travel around the nucleus in simple circular orbits.

The location of the electrons in the quantum mechanical model of the atom is often referred to as an **electron cloud**. The electron cloud can be thought of in the following way: Imagine placing a square piece of paper on the floor with a dot in the circle representing the nucleus. Now take a marker and drop it onto the paper repeatedly, making small marks at each point the marker hits. If you drop the marker many, many times, the overall pattern of dots will be roughly circular. If you aim toward the center reasonably well, there will be more dots near the nucleus and progressively fewer dots as you move away from it. Each dot represents a location where the electron could be at any given moment. Because of the uncertainty principle, there is no way to know exactly where the electron is. An electron cloud has variable densities: a high density where the electron is most likely to be and a low density where the electron is least likely to be (see below).

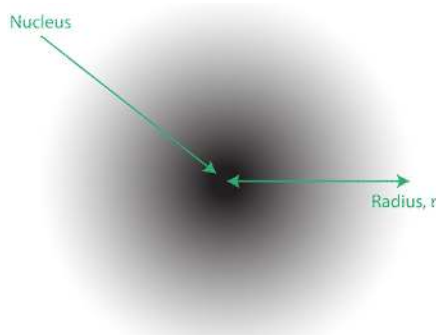


Figure 3.3.2: An electron cloud: the darker region nearer the nucleus indicates a high probability of finding the electron, while the lighter region further from the nucleus indicates a lower probability of finding the electron. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In order to specifically define the shape of the cloud, it is customary to refer to the area within which there is a **90** chance of finding the electron. This is called an **orbital**, the three-dimensional region of space that indicates where there is a high probability of finding an electron.



Summary

- The Schrödinger wave equation replaced Bohr's ideas about electron location with an uncertainty factor.
- The location of the electron can only be given as a probability that the electron is somewhere in a certain area.

Review

1. What does the quantum mechanical view of the atom require?
2. What is a wave function?
3. What does a high density electron cloud suggest?

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3.4: Quantum Numbers

If you attend a college or professional football game, you need a ticket to get in. It is very likely that your ticket may specify a gate number, section number, row, and seat number. No other ticket can have the same four parts to it. It may even have the same gate, section, and row, but it would still have to include a different seat number. Each seat is unique and allows only one occupant to fill it.

We use a series of specific numbers, called **quantum numbers**, to describe the location of an electron in an associated atom. Quantum numbers specify the properties of the atomic orbitals and the electrons in those orbitals. An electron in an atom or ion has four quantum numbers to describe its state. Think of them as important variables in an equation that describes the three-dimensional position of electrons in a given atom.

Principal Quantum Number (n)

The **principal quantum number**, signified by n , is the main **energy level (shell)** occupied by the electron. Energy levels are fixed distances from the nucleus of a given atom. They are described in whole number increments (e.g., 1, 2, 3, 4, 5, 6, ...). At location $n = 1$, an electron would be closest to the nucleus, while at $n = 2$ the electron would be farther, and at $n = 3$ farther yet. As we will see, the principal quantum number corresponds to the row number for an atom on the periodic table.

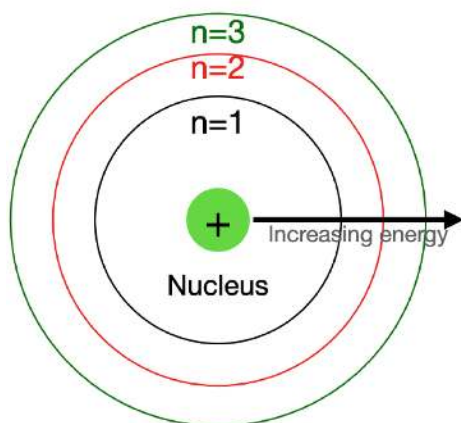


Figure 3.4.1: In Bohr's atomic model, Principal Quantum Number (n) can be interpreted as the different "energy levels" around the nucleus.

Angular Momentum Quantum Number (l)

The **angular momentum quantum number**, signified by l , describes the **energy sublevel (subshell)** and **general shape** or region an electron occupies—its orbital shape. The value of l depends on the value of the principal quantum number, n . The angular momentum quantum number can have positive values of zero to $(n - 1)$. If $n = 2$, l could be either 0 or 1.

Instead of using numbers to designate $l = 0$, we use **letters**:

If $l = 0$, then it's called **s**

If or $l = 1$, then it's called **p**

If $l = 2$ then it's called **d**

If $l = 3$, then it's called **f**

Each type of subshell (or letter) has a very specific shape, as described on the orbital shapes page.

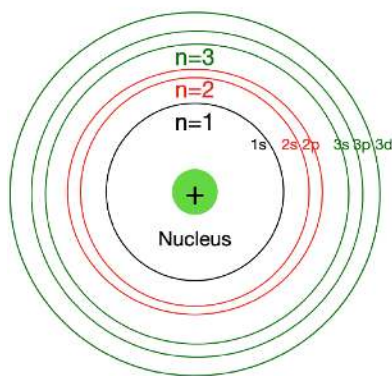


Figure 3.4.2: In Bohr's atomic model, Angular Momentum Quantum Number (l) can be interpreted as the different "energy sublevels" within each energy level

Magnetic Quantum Number (m_l)

The **magnetic quantum number**, signified as m_l , describes a particular **orbital (box)** and the **orbital orientation in space**. Electrons can be situated in one of three planes in three dimensional space around a given nucleus (x , y , and z). For a given value of the angular momentum quantum number, l , there can be $(2l + 1)$ values for m_l . As an example:

$$n = 2$$

$$l = 0 \text{ or } 1$$

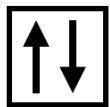
$$\text{for } l = 0, m_l = 0$$

$$\text{for } l = 1, m_l = -1, 0, +1$$

Table 3.4.1: Principal Energy Levels and Sublevels

Principal Energy Level	Number of Possible Sublevels	Possible Angular Momentum Quantum Numbers	Orbital Designation by Principal Energy Level and Sublevel
$n = 1$	1	$l = 0 \text{ or } l = s$	$1s$
$n = 2$	2	$l = 0 \text{ or } l = s$	$2s$
		$l = 1 \text{ or } l = p$	$2p$
$n = 3$	3	$l = 0 \text{ or } l = s$	$3s$
		$l = 1 \text{ or } l = p$	$3p$
		$l = 2 \text{ or } l = d$	$3d$
$n = 4$	4	$l = 0 \text{ or } l = s$	$4s$
		$l = 1 \text{ or } l = p$	$4p$
		$l = 2 \text{ or } l = d$	$4d$
		$l = 3 \text{ or } l = f$	$4f$

The table above shows the possible angular momentum quantum number values (l) for the corresponding principal quantum numbers (n) of ($n = 1, n = 2, n = 3$, and $n = 4$). A single orbital can hold a maximum of two electrons, and it can be represented as a single box. Although each one of these boxes has a specific shape given by its m_l number (check next section)



The total number of orbital per energy sublevel can be represented as:

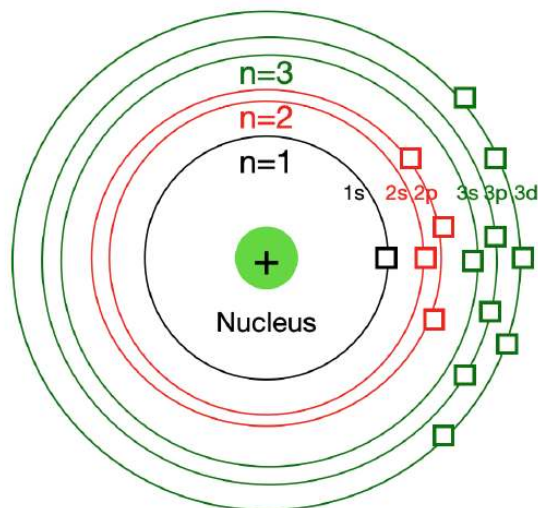


Figure 3.4.3: In Bohr's atomic model, Magnetic Quantum Number (m_l) can be considered as a "box" or a seat within each energy subshell. The number of "seats" depends on the type of subshell (s, p, d, or f).

Spin Quantum Number (m_s)

The **spin quantum number** describes the spin for a given electron. An electron can have one of two associated spins, $(+\frac{1}{2})$ spin, or $(-\frac{1}{2})$ spin. An electron cannot have zero spin. We also represent spin with arrows $\uparrow$ or $\downarrow$. A single orbital can hold a maximum of two electrons, and each must have opposite spin.

The electron spin describes the "rotation" or "spinning" of each electron around its own axis. Each electron acts as a tiny magnet that can rotate to the left $(+\frac{1}{2})$ or to the right $(-\frac{1}{2})$.

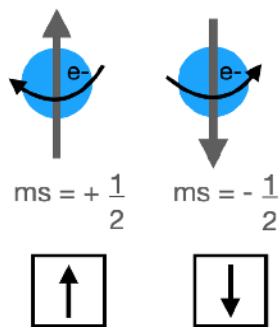


Figure 3.4.1: In Bohr's atomic model, Spin Quantum Number (m_s) can be interpreted as two possible direction in which the electron can rotate around its axis.

Summary

- Quantum numbers specify the arrangements of electrons in orbitals.
- There are four quantum numbers that provide information about various aspects of electron behavior.

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3.5: Orbital shapes

The flight path of a commercial airliner is carefully regulated by the Federal Aviation Administration. Each airplane must maintain a distance of five miles from another plane flying at the same altitude, and be 2,000 feet above and below another aircraft (1,000 feet if the altitude is less than 29,000 feet). So, each aircraft only has certain positions it is allowed to maintain while it flies. Quantum mechanics demonstrates that electrons have similar restrictions on their locations.

Knowledge of quantum numbers can be applied to describe the arrangement of electrons for a given atom. This application is done by using **electron configurations**. Electron configurations are effectively a map of the electrons for a given atom. We look at the four quantum numbers for a given electron, and then assign that electron to a specific orbital in the next Module.

s Orbitals

For any value of n , a value of $l = 0$ places that electron in an **s orbital**. This orbital is spherical in shape:

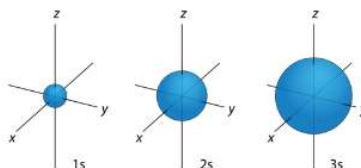


Figure 3.5.1: s orbitals have no orientational preference and resemble spheres.

p Orbitals

From the table below, it is evident that there are three possible orbitals when $l = 1$. These are designated as **p orbitals** and have dumbbell shapes. Each of the **p** orbitals has a different orientation in three-dimensional space.

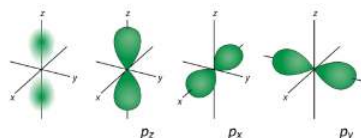


Figure 3.5.2: p orbitals have an orientational preference and resemble dumbbells.

d Orbitals

When $l = 2$, m_l values can be $-2, -1, 0, +1, +2$ for a total of five **d orbitals**. Note that all five of the orbitals have specific three-dimensional orientations.

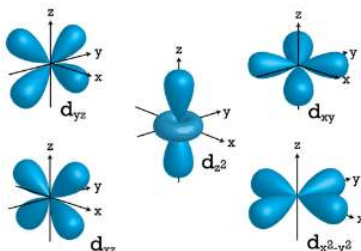


Figure 3.5.3: d orbitals have an orientational preference and exhibit complex structures.

f Orbitals

The most complex set of orbitals are the **f orbitals**. When $l = 3$, m_l values can be $-3, -2, -1, 0, +1, +2, +3$ for a total of seven different orbital shapes. Again, note the specific orientations of the different **f** orbitals.

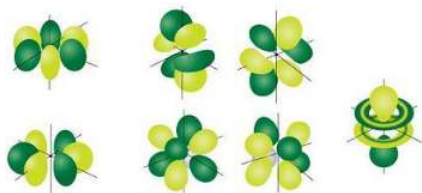


Figure 3.5.4: *d* orbitals have an orientational preference and exhibit quite complex structures.

Orbitals that have the same value of the principal quantum number form a **shell**. Orbitals within a shell are divided into **subshells** that have the same value of the angular quantum number. Some of the allowed combinations of quantum numbers are compared in Table 3.5.1.

Table 3.5.1: Electron Arrangement Within Energy Levels

Principal Quantum Number (n)	Allowable Sublevels	Number of Orbitals per Sublevel	Number of Orbitals per Principal Energy Level	Number of Electrons per Sublevel	Number of Electrons per Principal Energy Level
1	<i>s</i>	1	1	2	2
2	<i>s</i>	1	4	2	8
	<i>p</i>	3		6	
3	<i>s</i>	1	9	2	18
	<i>p</i>	3		6	
	<i>d</i>	5		10	
4	<i>s</i>	1	16	2	32
	<i>p</i>	3		6	
	<i>d</i>	5		10	
	<i>f</i>	7		14	

Summary

There are four different classes of electron orbitals. These orbitals are determined by the value of the angular momentum quantum number l .

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3.6: Aufbau Principle

Construction of a building begins at the bottom. The foundation is laid and the building goes up step by step. Construction obviously cannot start with the roof, since there is no place to hang it. The building goes from the lowest level to the highest level in a systematic way.

Aufbau Principle

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. In principle, the electron will fill available orbitals from the lowest energy towards the higher energy. However, as the energy shell increases (quantum number n), the number of energy subshells also increases (quantum number m_l). At some point, some of the subshells overlap, and the actual energy order of orbitals is not as expected. For example, the orbitals in the subshell $4s$ have lower energy than the orbitals in the $3d$ subshell. Figure 3.6.1 represents this phenomenon:

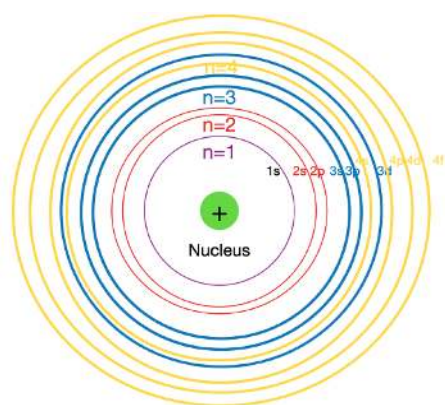


Figure 3.6.2: Overlap in the energy sublevels. Principle energy levels are color coded, while sublevels are grouped together, and each color presents an energy level, while each line represents an energy sublevel.

Figure 3.6.2 shows the order of increasing energy of the sublevels.

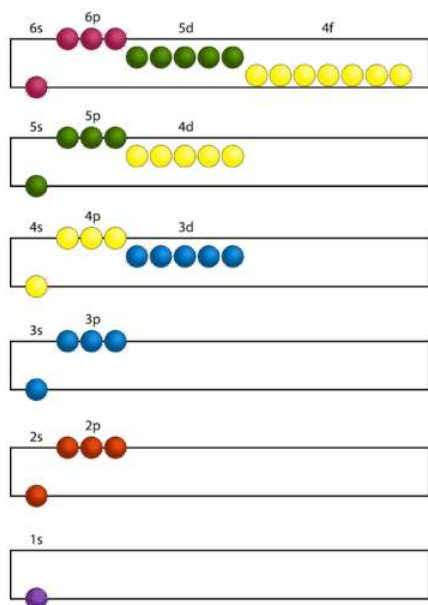


Figure 3.6.2: Electrons are added to atomic orbitals in order from low energy (bottom of the graph) to high (top of the graph), according to the Aufbau principle. Principle energy levels are color coded, while sublevels are grouped together, and each circle represents an orbital capable of holding two electrons.

The lowest energy sublevel is always the **1s** sublevel, which consists of one orbital. The single electron of the hydrogen atom will occupy the **1s** orbital when the atom is in its ground state. As we proceed to atoms with multiple electrons, those electrons are added to the next lowest sublevel: **2s**, **2p**, **3s**, and so on. The **Aufbau principle** states that an electron occupies orbitals in order from lowest energy to highest. The Aufbau (German for *building up, construction*) principle is sometimes referred to as the "building up" principle. It is worth noting that in reality, atoms are not built by adding protons and electrons one at a time, and that this method is merely an aid to understand the end result.

As seen in the figure above, the energies of the sublevels in different principal energy levels eventually begin to overlap. After the **3p** sublevel, it would seem logical that the **3d** sublevel should be the next lowest in energy. However, the **4s** sublevel is slightly lower in energy than the **3d** sublevel and thus fills first. Following the filling of the **3d** sublevel is the **4p**, then the **5s** and the **4d**. Note that the **4f** sublevel does not fill until just after the **6s** sublevel. Figure 3.6.2 is a useful and simple aid for keeping track of the order of fill of the atomic sublevels.

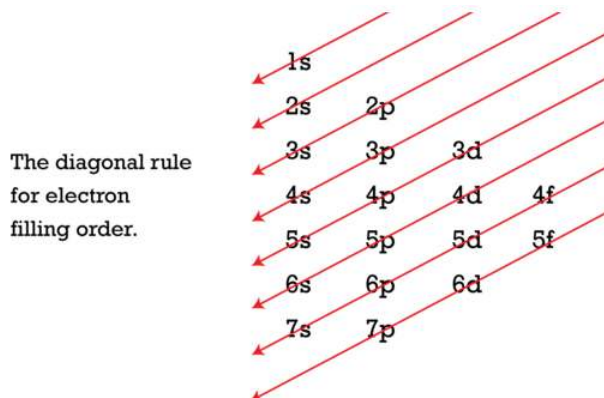


Figure 3.6.2: The Aufbau principle is illustrated in the diagram by following each red arrow in order from top to bottom: **1s**, **2s**, **2p**, **3s**, etc.

Summary

- The Aufbau principle gives the order of electron filling in an atom.
- The Aufbau principle can be used to describe the locations and energy levels of every electron in a given atom.

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3.7: Pauli Exclusion Principle



Figure 3.7.1 (Pixelbay License; Pixabay:Nemo via <http://pixabay.com/en/mail-internet-...ectronic-35636>)

Can you name one thing that easily distinguishes you from the rest of the world?

And we're not talking about DNA—that's a little expensive to sequence. For many people, it is their email address. My email address allows people all over the world to contact me. It does not belong to anyone else, but serves to identify me. Electrons also have a unique set of identifiers in the quantum numbers that describe their location and spin.

Pauli Exclusion Principle

When we look at the orbital possibilities for a given atom, we see that there are different arrangements of electrons for each different type of atom. Since each electron must maintain its unique identity, we intuitively sense that the four quantum numbers for any given electron must not match up exactly with the four quantum numbers for any other electron in that atom.

For the hydrogen atom, there is no problem since there is only one electron in the H atom. However, for helium, we see that the first three quantum numbers for the two electrons are the same: same energy level, same spherical shape. What differentiates the two helium electrons is their spin. One of the electrons has a $+\frac{1}{2}$ spin while the other electron has a $-\frac{1}{2}$ spin. So, the two electrons in the $1s$ orbital are each unique and distinct from one another because their spins are different. This observation leads to the **Pauli exclusion principle**, which states that no two electrons in an atom can have the same set of four quantum numbers. The energy of the electron is specified by the principal, angular momentum, and magnetic quantum numbers. If those three numbers are identical for two electrons, the spin numbers must be different in order for the two electrons to be differentiated from one another. The two values of the spin quantum number allow each orbital to hold two electrons. Figure 3.7.2 shows how the electrons are indicated in a diagram.

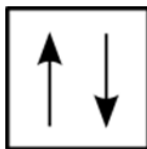


Figure 3.7.2: In an orbital filling diagram, a square represents an orbital, while arrows represent electrons. An arrow pointing upward represents one spin direction, while an arrow pointing downward represents the other spin direction. (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)



Summary

- The Pauli exclusion principle specifies limits on how identical quantum numbers can be for two electrons in the same atom.

Review

1. What is the difference between the two helium electrons?
2. What does the Pauli exclusion principle state?
3. What does the two values for the spin quantum number allow?

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3.8: Hund's Rule and Orbital Filling Diagrams

Hund's Rule

The last of the three rules for constructing electron arrangements requires electrons to be placed one at a time in a set of orbitals within the same sublevel. This minimizes the natural repulsive forces that one electron has for another. **Hund's rule** states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron and that each of the single electrons must have the same spin. The figure below shows how a set of three p orbitals is filled with one, two, three, and four electrons.

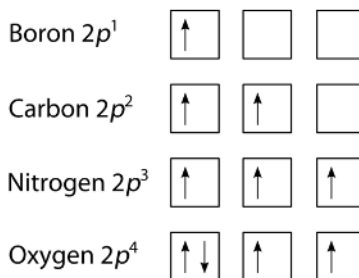


Figure 3.8.1: The $2p$ sublevel, for the elements boron ($Z = 5$), carbon ($Z = 6$), nitrogen ($Z = 7$), and oxygen ($Z = 8$). According to Hund's rule, as electrons are added to a set of orbitals of equal energy, one electron enters each orbital before any orbital receives a second electron.

Orbital Filling Diagrams

An **orbital filling diagram** is the more visual way to represent the arrangement of all the electrons in a particular atom. In an orbital filling diagram, the individual orbitals are shown as circles (or squares) and orbitals within a sublevel are drawn next to each other horizontally. Each sublevel is labeled by its principal energy level and sublevel. Electrons are indicated by arrows inside the circles. An arrow pointing upwards indicates one spin direction, while a downward pointing arrow indicates the other direction. The orbital filling diagrams for hydrogen, helium, and lithium are shown in the figure below.

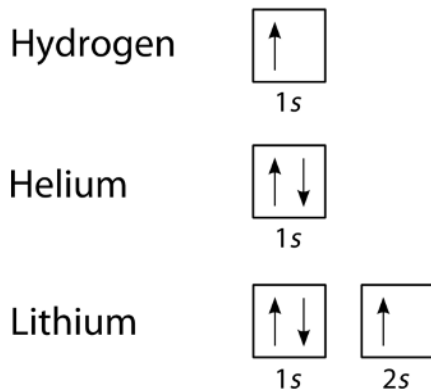


Figure 3.8.2: Orbital filling diagrams for hydrogen, helium, and lithium.

According to the Aufbau process, sublevels and orbitals are filled with electrons in order of increasing energy. Since the s sublevel consists of just one orbital, the second electron simply pairs up with the first electron as in helium. The next element is lithium and necessitates the use of the next available sublevel, the $2s$.

The filling diagram for carbon is shown in the figure below. There are two $2p$ electrons for carbon and each occupies its own $2p$ orbital.

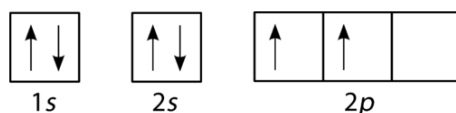


Figure 3.8.3: Orbital filling diagram for carbon.

Oxygen has four **2p** electrons. After each **2p** orbital has one electron in it, the fourth electron can be placed in the first **2p** orbital with a spin opposite that of the other electron in that orbital.

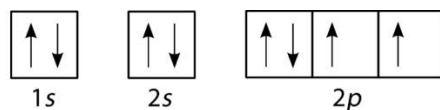


Figure 3.8.4: Orbital filling diagram for oxygen.

Summary

Hund's rule specifies the order of electron filling within a set of orbitals. Orbital filling diagrams are a way of indicating electron locales in orbitals.

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3.9: Electron Configurations

```
C:\Temp> dir
Volume in drive C is C
Volume Serial Number is 74F5-B93C

Directory of C:\Temp

2009-08-25 11:59 <DIR> .
2009-08-25 11:59 <DIR> ..
2007-03-01 11:37 2,321,400 AdobeUpdater12345.exe
2009-04-03 10:01 27,988 dd_depcheckdotnetfx30.txt
2009-04-03 10:01 764 dd_dotnetfx3error.txt
2009-04-03 10:01 32,572 dd_dotnetfx3install.txt
2009-06-09 13:46 35,145 GenProfile.log
2009-08-05 12:11 155 KB969856.log
2009-04-20 08:37 402 MS129e0b.LOG
2009-04-09 16:34 38,895 offcnll.log
2009-04-03 16:02 <DIR> OfficePatches
2009-07-14 14:30 <DIR> Officefix
2009-08-25 10:52 16,384 PerfLib_Perfdata_c30.dat
2009-04-03 10:01 1,744 uxeventlog.txt
2009-08-25 11:42 30,245,632 WVF2F.tmp
2009-04-20 10:07 1,397 {AC76BA86-7AD7-1033-7B44-AA1300000003}.ini
2009-04-20 10:13 617 {AC76BA86-7AD7-1033-7B44-AA1300000003}.ini
13 File(s) 52,723,295 bytes
4 Dir(s) 83,570,208,768 bytes free
```

Figure 3.9.1 (Public Domain; David R. Tribble (User:Loadmaster/Wikimedia Commons) via [Commons wikimedia](#), [Dir Command in Windows Command Prompt](#) [commons.wikimedia.org])

How big is a file?

If you keep your papers in manila folders, you can pick up a folder and see how much it weighs. If you want to know how many different papers (articles, bank records, or whatever else you keep in a folder), you have to take everything out and count. A computer directory, on the other hand, tells you exactly how much you have in each file. We can get the same information on atoms. If we use an orbital filling diagram, we have to count arrows. When we look at electron configuration data, we simply add up the numbers.

Electron Configurations

Electron configuration notation eliminates the boxes and arrows of orbital filling diagrams. Each occupied sublevel designation is written followed by a superscript that is the number of electrons in that sublevel. For example, the hydrogen configuration is $1s^1$, while the helium configuration is $1s^2$. Multiple occupied sublevels are written one after another. The electron configuration of lithium is $1s^2 2s^1$. The sum of the superscripts in an electron configuration is equal to the number of electrons in that atom, which is in turn equal to its atomic number.

Example 3.9.1

Draw the orbital filling diagram for carbon and write its electron configuration.

Solution

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

Known

- atomic number of carbon, $Z=6$

Use the order of fill diagram to draw an orbital filling diagram with a total of six electrons. Follow Hund's rule. Write the electron configuration.

Step 2: Construct diagram.

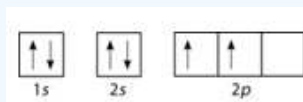


Figure 3.9.2: Orbital filling diagram for carbon. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Electron configuration $1s^2 2s^2 2p^2$

Step 3: Think about your result.

Following the $2s$ sublevel is the $2p$, and p sublevels always consist of three orbitals. All three orbitals need to be drawn even if one or more is unoccupied. According to Hund's rule, the sixth electron enters the second of those p orbitals, with the same spin as the fifth electron.

Second Period Elements

Periods refer to the horizontal rows of the periodic table. The first period of the periodic table contains only the elements hydrogen and helium. This is because the first principal energy level consists of only the s sublevel, and so only two electrons are required in order to fill the entire principal energy level. Each time a new principal energy level begins, as with the third element lithium, a new period is started on the periodic table. As one moves across the second period, electrons are successively added. With beryllium ($Z = 4$), the $2s$ sublevel is complete and the $2p$ sublevel begins with boron ($Z = 5$). Since there are three $2p$ orbitals and each orbital holds two electrons, the $2p$ sublevel is filled after six elements. Table 3.9.1 shows the electron configurations of the elements in the second period.

Table 3.9.1: Electron Configurations of Second-Period Elements

Element Name	Symbol	Atomic Number	Electron Configuration
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	$1s^2 2s^2 2p^6$





Summary

- 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.

Review

1. What does electron configuration notation eliminate?
2. How do we know how many electrons are in each sublevel?
3. An atom has the electron configuration of $1s^2 2s^2 2p^5$. How many electrons are in that atom?
4. Which element has the electron configuration of $1s^2 2s^2 2p^6 3s^2$?

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3.10: Valence Electrons

A chemical reaction involves either electron removal, electron addition, or electron sharing. The path that a specific element will take in a reaction depends on where the electrons are in the atom and how many there are.

Valence Electrons

In the study of chemical reactivity, electrons in the outermost principal energy level are very important and so are given a special name. **Valence electrons** are the electrons in the highest occupied principal energy level of an atom. In the second period elements, the two electrons in the **1s** sublevel are called **inner-shell electrons** and are not involved directly in the element's reactivity, or in the formation of compounds.

Valence electrons for representative elements

Lithium has a single electron in the second principal energy level, and so we say that lithium has one valence electron.

Li: $1s^2 2s^1$ (the electron in the 2s energy level is the valence electron)

Beryllium has two valence electrons:

Be: $1s^2 2s^2$ (the two electrons in the 2s energy level are the valence electrons)

How many valence electrons does boron have? Recognize that the second principal energy level consists of both the **2s** and the **2p** sublevels, and so the answer is three.

B: $1s^2 2s^2 2p^1$ (there are three electrons on the highest occupied energy level $n=2$)

In fact, the number of valence electrons goes up by one for each step across a period, until the last element is reached. Neon, with its configuration ending in **$2s^2 2p^6$** , has eight valence electrons.

Valence electrons for transition elements.

Transition elements are a bit trickier. In this case, we also need to consider the electrons in the highest occupied energy level (n) plus the electrons in the $(n-1)$ d orbital. For example, the electron configuration of iron is Fe is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. The total number of valence electrons for iron is 8: 2 electrons in the highest occupied energy level ($n=4$) plus 6 electrons in the $(n-1)$ d orbital, that is, 3d.

Using abbreviated electron configuration (or noble gas configuration) to identify valence electrons

The best approach to identify the number of valence electrons is to use the abbreviated notation for electron configuration. In this notation, the inner-shell electrons are abbreviated by indicating the noble gas that corresponds to that same number of electrons

For example, the abbreviated electron configuration for Lithium is [He] $2s^1$, so the number of valence electrons is 1 ($2s^1$)

In the case of Boron, the abbreviated electron configuration is [He] $2s^2 2p^1$, so the number of valence electrons is 3 ($2s^2 2p^1$)

In the case of Iron, the abbreviated electron configuration is [Ar] $4s^2 3d^6$, so the number of valence electrons is 8 ($4s^2 3d^6$)

Summary

- Valence electrons are the outer-shell electrons of an atom.
- Valence electrons determine the reactivity of an atom.

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3.11: Noble Gas Configuration



Figure 3.11.1 (Public Domain; User:Danbold/Wikipedia via [Wikipedia](#))

How does it feel to be full after a meal?

Envision that you have nearly finished a great meal, but cannot put another bite in your mouth because there is no place for it to go. The noble gases have the same problem—there is no room for any more electrons in their outer shells. They are completely full and cannot handle any more.

Noble Gas Configuration

Sodium, element number 11, is the first element in the third period of the periodic table. Its electron configuration is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons of the sodium atom are the inner-shell electrons and the configuration of just those ten electrons is exactly the same as the configuration of the element neon ($Z = 10$). This provides the basis for a shorthand notation for electron configurations called the noble gas configuration. The elements that are found in the last column of the periodic table are an important group of elements called the noble gases. They are helium, neon, argon, krypton, xenon, and radon. A **noble gas configuration** of an atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons. So for sodium, we make the substitution of [Ne] for the $1s^2 2s^2 2p^6$ part of the configuration. Sodium's noble gas configuration becomes [Ne] $3s^1$. Table 3.11.1 shows the noble gas configurations of the third period elements.

Table 3.11.1: Electron Configurations of Third-Period Elements

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Sodium	Na	11	[Ne] $3s^1$
Magnesium	Mg	12	[Ne] $3s^2$
Aluminum	Al	13	[Ne] $3s^2 3p^1$
Silicon	Si	14	[Ne] $3s^2 3p^2$
Phosphorus	P	15	[Ne] $3s^2 3p^3$
Sulfur	S	16	[Ne] $3s^2 3p^4$
Chlorine	Cl	17	[Ne] $3s^2 3p^5$
Argon	Ar	18	[Ne] $3s^2 3p^6$

Again, the number of valence electrons increases from one to eight across the third period.

The fourth and subsequent periods follow the same pattern, except for the use of a different noble gas. Potassium has nineteen electrons, one more than the noble gas argon, so its configuration could be written as [Ar] $4s^1$. In a similar fashion, strontium has two more electrons than the noble gas krypton, which would allow us to write its electron configuration as [Kr] $5s^2$. All elements can be represented in this fashion.



Summary

- The noble gas configuration system allows some shortening of the total electron configuration by using the symbol for the noble gas of the previous period as part of the pattern of electrons.

Review

1. What is the element represented by $[\text{Ne}] 3s^2 3p^2$?
2. What element has this electron configuration $[\text{Ar}] 3d^7 4s^2$?
3. What noble gas would be part of the electron configuration notation for Mn?
4. How would you write the electron configuration for Ba?

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3.12: Electronic Structure and the Periodic Table

Learning Objectives

- Relate the electron configurations of the elements to the shape of the periodic table.
- Determine the expected electron configuration of an element by its place on the periodic table.
- Predict group and period for an element given its electron configuration

In Chapter 2, we introduced the periodic table as a tool for organizing the known chemical elements. A periodic table is shown in Figure 3.12.1. The elements are listed by atomic number (the number of protons in the nucleus), and elements with similar chemical properties are grouped together in columns.

1 H 1.00794																		2 He 4.002602			
3 Li 6.941	4 Be 9.012182															5 B 10.811	6 C 12.0107	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797
11 Na 22.989770	12 Mg 24.3050															13 Al 26.581538	14 Si 28.0855	15 P 30.973761	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6534	29 Cu 63.545	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.504	36 Kr 83.80				
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 196.56655	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29				
55 Cs 132.90545	56 Ba 137.327	57 La 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.56655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.58038	84 Po (209)	85 At (210)	86 Rn (222)				
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (277)	114 (289) (287)		116 (289)		118 (293)					

58 Ce 140.116	59 Pr 140.50765	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.964	64 Gd 157.25	65 Tb 158.92534	66 Dy 162.50	67 Ho 164.93032	68 Er 167.26	69 Tm 168.93421	70 Yb 173.04	71 Lu 174.967
90 Th 232.0381	91 Pa 231.035888	92 U 238.0289	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Figure 3.12.1: The Periodic Table

Why does the periodic table have the structure it does? The answer is rather simple, if you understand electron configurations, the shape of the periodic table mimics the filling of the subshells with electrons.

The shape of the periodic table mimics the filling of the subshells with electrons.

Let us start with H and He. Their electron configurations are $1s^1$ and $1s^2$, respectively; with He, the $n = 1$ shell is filled. These two elements make up the first row of the periodic table (Figure 3.12.2)

1																	2						
3																	4						
5																	6						
7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
31	32	33	34	35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78
79	80	81	82	83	84	85	86	87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102

Figure 3.12.2: The 1s Subshell. H and He represent the filling of the 1s subshell.

The next two electrons, for Li and Be, would go into the 2s subshell. Figure 3.12.3 shows that these two elements are adjacent on the periodic table.

3	4																	5					
6	7																	8					
9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30		
31	32	33	34	35	36	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
55	56	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78
79	80	81	82	83	84	85	86	87	88	89	90	91	92	93	94	95	96	97	98	99	100	101	102

Figure 3.12.3: The 2s Subshell. In Li and Be, the 2s subshell is being filled.

For the next six elements, the 2p subshell is being occupied with electrons. On the right side of the periodic table, these six elements (B through Ne) are grouped together (Figure 3.12.4).

Diagram of the periodic table showing the 2p subshell being filled. The 2p subshell is highlighted in green and labeled '2p' with an arrow pointing to it.

Diagram of the 2p subshell, showing a 2x6 grid of cells.

Figure 3.12.4: The 2p Subshell. For B through Ne, the 2p subshell is being occupied.

The next subshell to be filled is the 3s subshell. The elements coinciding this subshell being filled, Na and Mg, are back on the left side of the periodic table (Figure 3.12.5).

Diagram of the periodic table showing the 3s subshell being filled. The 3s subshell is highlighted in green and labeled '3s' with an arrow pointing to it.

Diagram of the 3s subshell, showing a 2x2 grid of cells.

Figure 3.12.5: The 3s Subshell. Now the 3s subshell is being occupied.

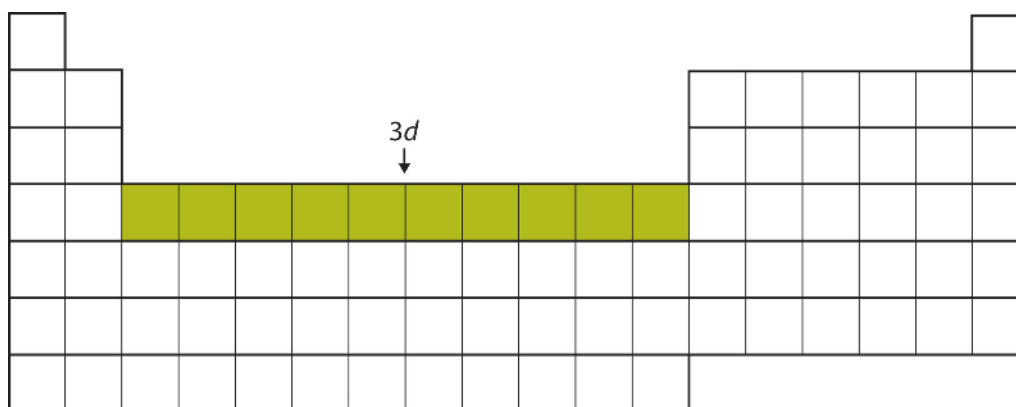
Next, the 3p subshell is filled with the next six elements (Figure 3.12.6).

Figure 8.5.6: The 3p Subshell. Next, the 3p subshell is filled with electrons.

Instead of filling the 3d subshell next, electrons go into the 4s subshell (Figure 3.12.7).

Figure 3.12.7: The 4s Subshell. The 4s subshell is filled before the 3d subshell. This is reflected in the structure of the periodic table.

After the 4s subshell is filled, the 3d subshell is filled with up to 10 electrons. This explains the section of 10 elements in the middle of the periodic table (Figure 3.12.8).



$3d$

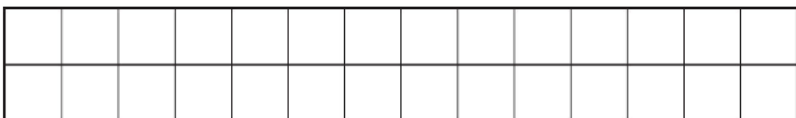
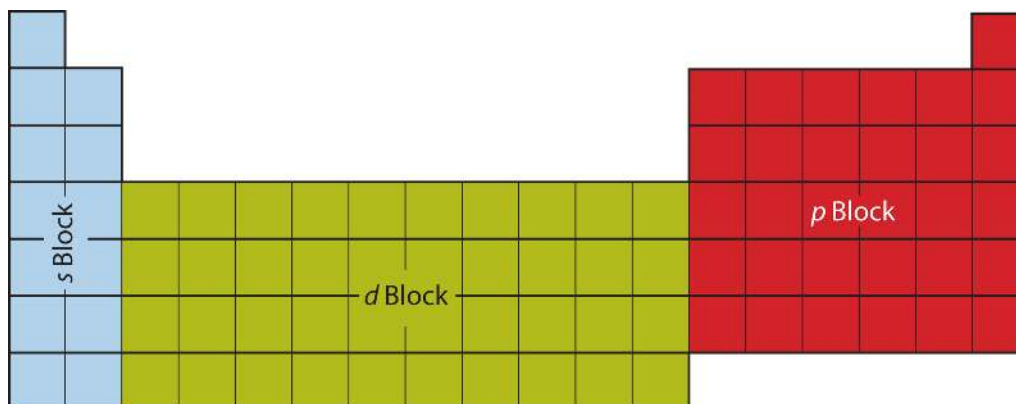


Figure 3.12.8: The 3d Subshell. The 3d subshell is filled in the middle section of the periodic table.

And so forth. As we go across the columns of the periodic table, the overall shape of the table outlines how the electrons are occupying the shells and subshells.

The first two columns on the left side of the periodic table are where the **s** subshells are being occupied. Because of this, the first two rows of the periodic table are labeled the **s block**. Similarly, the **p block** is the right-most six columns of the periodic table, the **d block** is the middle 10 columns of the periodic table, while the **f block** is the 14-column section that is normally depicted as detached from the main body of the periodic table. It could be part of the main body, but then the periodic table would be rather long and cumbersome. Figure 3.12.9 shows the blocks of the periodic table.



s Block d Block p Block

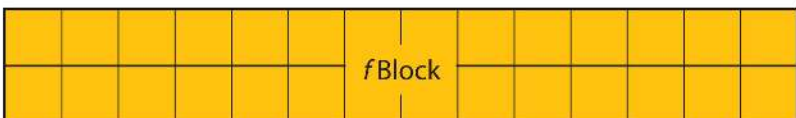


Figure 3.12.9: Blocks on the Periodic Table. The periodic table is separated into blocks depending on which subshell is being filled for the atoms that belong in that section.

The electrons in the highest-numbered shell, plus any electrons in the last unfilled subshell, are called **valence electrons**; the highest-numbered shell is called the **valence shell**. (The inner electrons are called *core electrons*.) The valence electrons largely control the chemistry of an atom. If we look at just the valence shell's electron configuration, we find that in each column, the valence shell's electron configuration is the same. For example, take the elements in the first column of the periodic table: H, Li,

Na, K, Rb, and Cs. Their electron configurations (abbreviated for the larger atoms) are as follows, with the valence shell electron configuration highlighted:

H:	$1s^1$
Li:	$1s^2 2s^1$
Na:	$[\text{Ne}] 3s^1$
K:	$[\text{Ar}] 4s^1$
Rb:	$[\text{Kr}] 5s^1$
Cs:	$[\text{Xe}] 6s^1$

They all have a similar electron configuration in their valence shells: a single s electron. Because much of the chemistry of an element is influenced by valence electrons, we would expect that these elements would have similar chemistry—and *they do*. The organization of electrons in atoms explains not only the shape of the periodic table, but also the fact that elements in the same column of the periodic table have similar chemistry.

The same concept applies to the other columns of the periodic table. Elements in each column have the same valence shell electron configurations, and the elements have some similar chemical properties. This is strictly true for all elements in the s and p blocks. In the d and f blocks, because there are exceptions to the order of filling of subshells with electrons, similar valence shells are not absolute in these blocks. However, many similarities do exist in these blocks, so a similarity in chemical properties is expected.

Similarity of valence shell electron configuration implies that we can determine the electron configuration of an atom solely by its position on the periodic table. Consider Se, as shown in Figure 3.12.10. It is in the fourth column of the p block. This means that its electron configuration should end in a p^4 electron configuration. Indeed, the electron configuration of Se is $[\text{Ar}] 4s^2 3d^{10} 4p^4$, as expected.

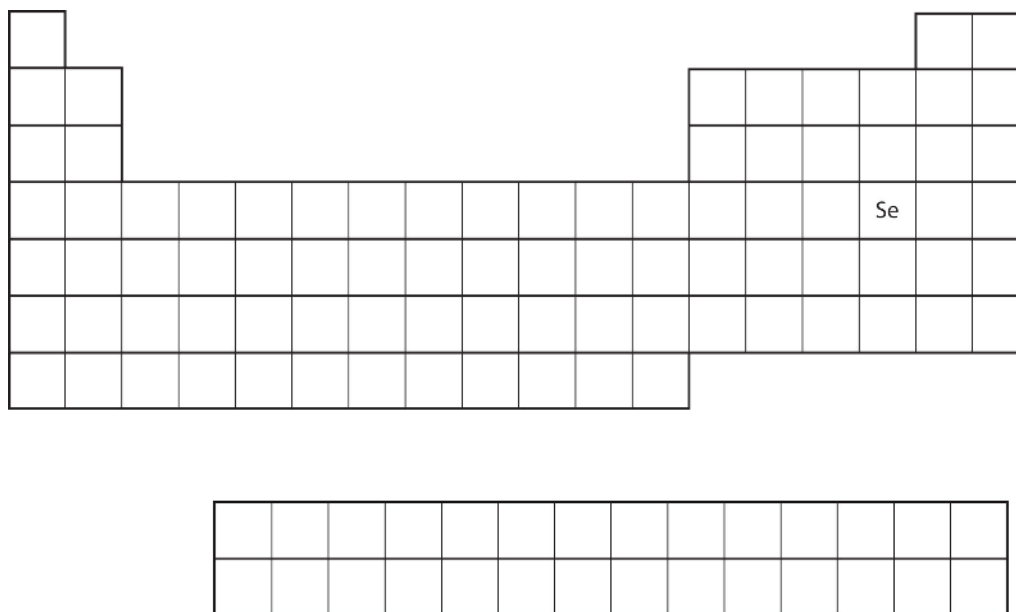


Figure 3.12.10: Selenium on the Periodic Table

Example 3.12.1:

From the element's position on the periodic table, predict the valence shell electron configuration for each atom (Figure 3.12.11).

betanin from beets for purple; and even carmine, a deep red dye that is extracted from the cochineal, a small insect that is a parasite on cacti in Central and South America. (That's right—you may be eating bug juice!)

Some colorings are artificial. In the United States, the Food and Drug Administration currently approves only seven compounds as artificial colorings in food, beverages, and cosmetics:

- FD&C Blue #1: Brilliant Blue FCF
- FD&C Blue #2: Indigotine
- FD&C Green #3: Fast Green FCF
- RD&C Red #3: Erythrosine
- FD&C Red #40: Allura Red AC
- FD&C Yellow #5: Tartrazine
- FD&C Yellow #6: Sunset Yellow FCF

Lower-numbered colors are no longer on the market or have been removed for various reasons. Typically, these artificial colorings are large molecules that absorb certain colors of light very strongly, making them useful even at very low concentrations in foods and cosmetics. Even at such low amounts, some critics claim that a small portion of the population (especially children) is sensitive to artificial colorings and urge that their use be curtailed or halted. However, formal studies of artificial colorings and their effects on behavior have been inconclusive or contradictory. Many people continue to enjoy processed foods with artificial coloring (like those shown in the accompanying figure).



Artificial food colorings. Artificial food colorings are found in a variety of food products, such as processed foods, candies, and egg dyes. Even pet foods have artificial food coloring in them. (CC BY 2.0, [Matthew Bland](#) via [Flickr](#).)

Predicting group and period from the electron configuration

As you can imagine from our discussion above, there is a clear connection between electron configuration and location on the periodic table. We might ask ourselves if there is a simple way to predict the group and period for an element given its electron configuration. The answer is yes, we just need to consider the valence electrons for that element to be able to predict its location on the periodic table. The general rule is that the **period** for an element is the same as the **highest occupied electron shell (n)** for that element, while the **group corresponds to the total number of valence electrons**. In the case of valence electrons with p-type orbital, we need to add 10 to the number of valence electrons in order to determine the group in the 1-18 group system. Examples:

- Lithium has an electron configuration $1s^2 2s^1$. The highest occupied electron shell is 2, so the period number is 2. The number of valence electrons is 1 (no p-valence electrons), so the group number is 1.
- Boron has an electron configuration $1s^2 2s^2 2p^1$. The highest occupied electron shell is 2, so the period number is 2. The number of valence electrons is 3, but they include electrons in a p-type orbital, so to calculate the group number we need to add 10 to the number of valence electrons. So group number = $10 + 3 = 13$.
- Iron has an electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. The highest occupied electron shell is 4, so the period number is 4. The number of valence electrons is 8 (no p-valence electrons), so the group number is 8.

Summary

The arrangement of electrons in atoms is responsible for the shape of the periodic table. Electron configurations can be predicted by the position of an atom on the periodic table.

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3.13: Periodic Trends- Atomic Size, Ionization Energy, Electron Affinity, and Metallic Character

Learning Objectives

- Be able to state how certain properties of atoms vary based on their relative position on the periodic table.

One of the reasons the periodic table is so useful is because its structure allows us to qualitatively determine how some properties of the elements vary versus their position on the periodic table. The variations of properties versus positions on the periodic table are called **periodic trends**. There is no other tool in science that allows us to judge relative properties of a class of objects like this, which makes the periodic table a very useful tool. Many periodic trends are general. There may be a few points where an opposite trend is seen, but there is an overall trend when considered across a whole row or down a whole column of the periodic table.

The first periodic trend we will consider is atomic radius. The **atomic radius** is an indication of the size of an atom. Although the concept of a definite radius of an atom is a bit fuzzy, atoms behave as if they have a certain radius. Such radii can be estimated from various experimental techniques, such as the x-ray crystallography of crystals.

As you go down a column of the periodic table, the atomic radii increase. This is because the valence electron shell is getting larger and there is a larger principal quantum number, so the valence shell lies physically farther away from the nucleus. Going across a row on the periodic table, left to right, the trend is different. Even though the valence shell maintains the same principal quantum number, the number of protons—and hence the nuclear charge—is increasing as you go across the row. The increasing positive charge casts a tighter grip on the valence electrons, so as you go across the periodic table, the atomic radii decrease.

Figure 3.13.1 shows spheres representing the atoms of the s and p blocks from the periodic table to scale, showing the two trends for the atomic radius.

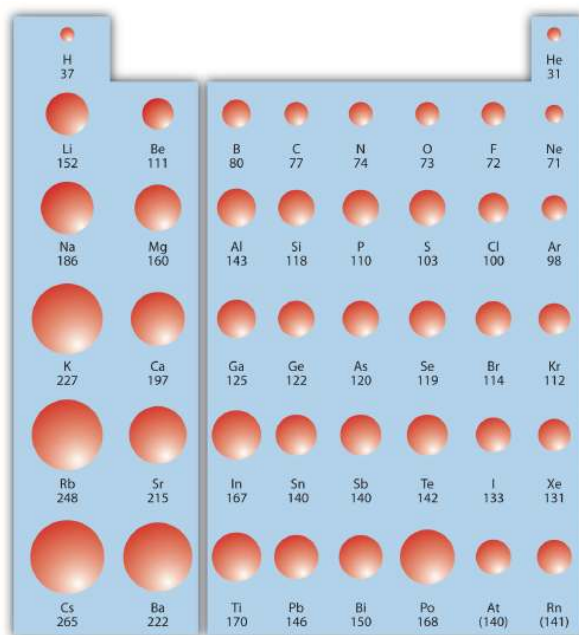


Figure 3.13.1: Atomic Radii Trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

✓ Example 3.13.1: Atomic Radii

Referring only to a periodic table and not to Figure 3.13.1, which atom is larger in each pair?

- Si or S
- S or Te

Solution

- Si is to the left of S on the periodic table; it is larger because as you go across the row, the atoms get smaller.
- S is above Te on the periodic table; Te is larger because as you go down the column, the atoms get larger.

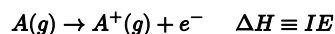
? Exercise 3.13.1: Atomic Radii

Referring only to a periodic table and not to Figure 3.13.1, which atom is smaller, Ca or Br?

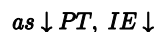
Answer

Br

Ionization energy (IE) is the amount of energy required to remove an electron from an atom in the gas phase:



IE is usually expressed in kJ/mol of atoms. It is always positive because the removal of an electron always requires that energy be put in (i.e., it is endothermic). IE also shows periodic trends. As you go down the periodic table, it becomes easier to remove an electron from an atom (i.e., IE decreases) because the valence electron is farther away from the nucleus. Thus,



However, as you go across the periodic table and the electrons get drawn closer in, it takes more energy to remove an electron; as a result, IE increases:

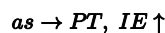


Figure 3.13.2 shows values of IE versus position on the periodic table. Again, the trend is not absolute, but the general trends going across and down the periodic table should be obvious.

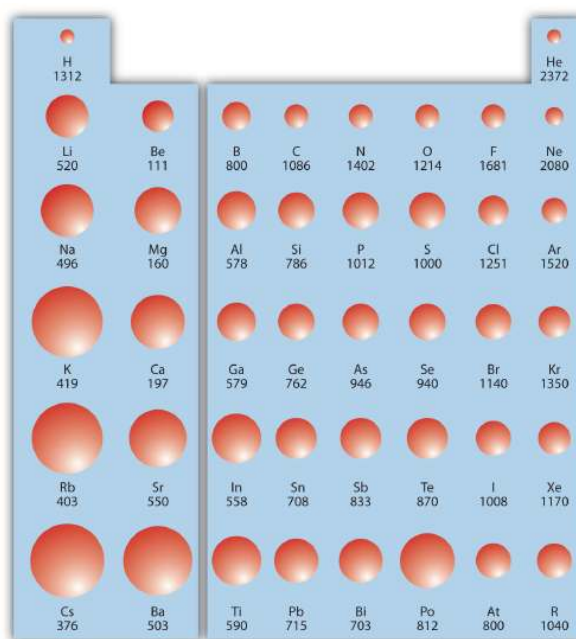
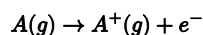


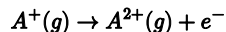
Figure 3.13.2: Ionization Energy on the Periodic Table. Values are in kJ/mol.

IE also shows an interesting trend within a given atom. This is because more than one IE can be defined by removing successive electrons (if the atom has them to begin with):

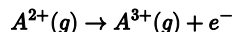
- First Ionization Energy (IE_1):



- Second Ionization Energy (IE_2):



- Third Ionization Energy (IE_3):



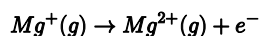
and so forth.

Each successive IE is larger than the previous because an electron is being removed from an atom with a progressively larger positive charge. However, IE takes a large jump when a successive ionization goes down into a new shell. For example, the following are the first three IEs for Mg, whose electron configuration is $1s^2 2s^2 2p^6 3s^2$:

- First Ionization Energy (IE_1) = 738 kJ/mol:



- Second Ionization Energy (IE_2) = 1,450 kJ/mol:



- Third Ionization Energy (IE_3) = 7,734 kJ/mol:



The second IE is twice the first, which is not a surprise: the first IE involves removing an electron from a neutral atom, while the second one involves removing an electron from a positive ion. The third IE, however, is over *five times* the previous one. Why is it so much larger? Because the first two electrons are removed from the 3s subshell, but the third electron has to be removed from the $n = 2$ shell (specifically, the 2p subshell, which is lower in energy than the $n = 3$ shell). Thus, it takes much more energy than just overcoming a larger ionic charge would suggest. It is trends like this that demonstrate that electrons within atoms are organized in groups.

✓ Example 3.13.2: Ionization Energies

Which atom in each pair has the larger first ionization energy?

- Ca or Sr
- K or K^+

Solution

- Because Sr is below Ca on the periodic table, it is easier to remove an electron from it; thus, Ca has the higher IE.
- Because K^+ has a positive charge, it will be harder to remove another electron from it, so its IE is larger than that of K. Indeed, it will be significantly larger because the next electron in K^+ to be removed comes from another shell.

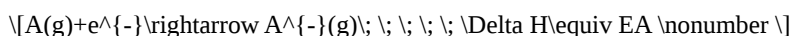
? Exercise 3.13.2: Ionization Energies

Which atom has the lower ionization energy, C or F?

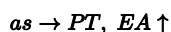
Answer

C

The opposite of IE is described by **electron affinity (EA)**, which is the energy change when a gas-phase atom accepts an electron:



EA is also usually expressed in kJ/mol. EA also demonstrates some periodic trends, although they are less obvious than the other periodic trends discussed previously. Generally, as you go across the periodic table, EA increases its magnitude:



There is not a definitive trend as you go down the periodic table; sometimes EA increases, sometimes it decreases. Figure 3.13.3 shows EA values versus position on the periodic table for the *s*- and *p*-block elements. The trend is not absolute, especially considering the large positive EA values for the second column. However, the general trend going across the periodic table should be obvious.

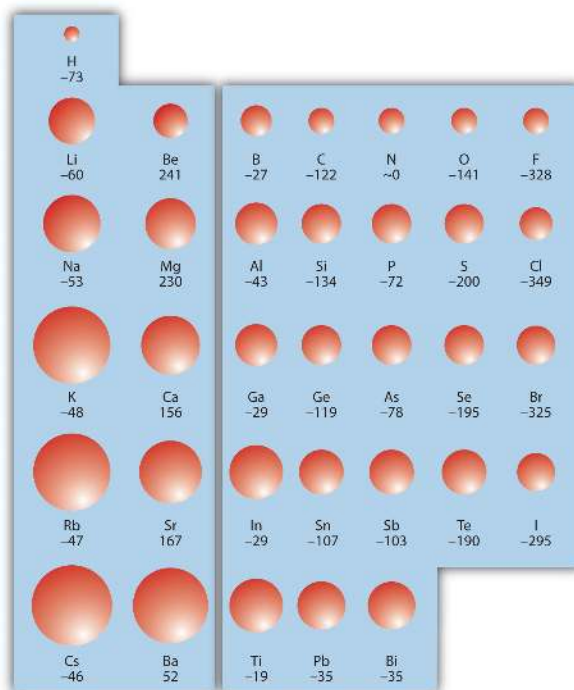


Figure 3.13.3: Electron Affinity on the Periodic Table. Values are in kJ/mol.

✓ Example 3.13.3: Electron Affinities

Predict which atom in each pair will have the highest magnitude of Electron Affinity.

- C or F
- Na or S

Solution

- C and F are in the same row on the periodic table, but F is farther to the right. Therefore, F should have the larger magnitude of EA.
- Na and S are in the same row on the periodic table, but S is farther to the right. Therefore, S should have the larger magnitude of EA.

? Exercise 3.13.3: Electron Affinities

Predict which atom will have the highest magnitude of Electron Affinity: As or Br.

Answer

Br

Metallic Character

The metallic character is used to define the chemical properties that metallic elements present. Generally, metals tend to lose electrons to form cations. Nonmetals tend to gain electrons to form anions. They also have a high oxidation potential—therefore they are easily oxidized and are strong reducing agents. Metals also form basic oxides; the more basic the oxide, the higher the metallic character.

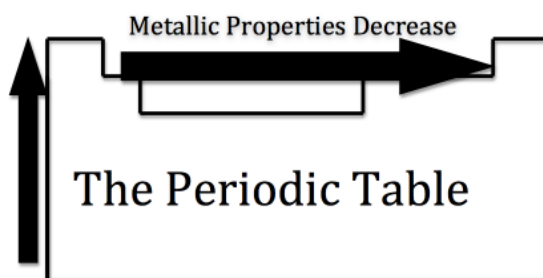


Figure 3.13.4: Courtesy of Jessica Thornton (UCD)

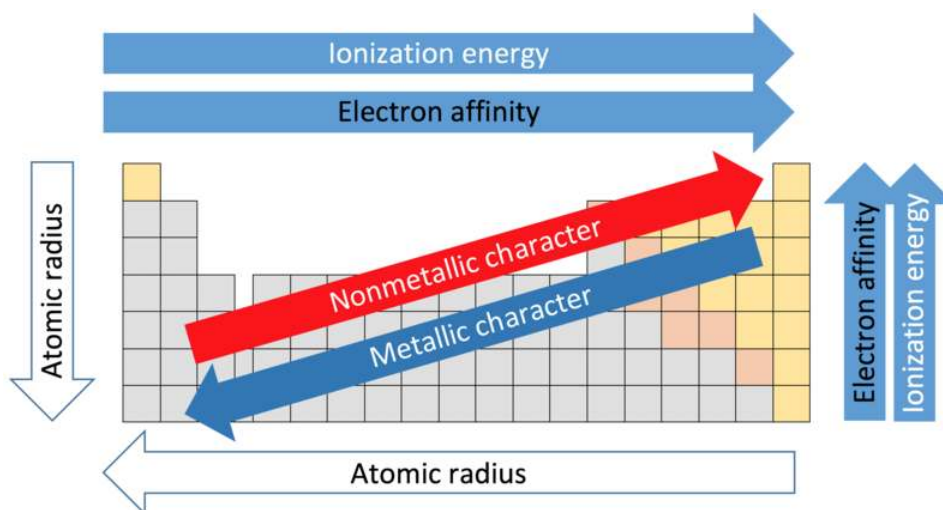
As you move across the table from left to right, the metallic character decreases, because the elements easily accept electrons to fill their valance shells. Therefore, these elements take on the nonmetallic character of forming anions. As you move up the table, the metallic character decreases, due to the greater pull that the nucleus has on the outer electrons. This greater pull makes it harder for the atoms to lose electrons and form cations.

📌 Uses of the Periodic Properties of Elements

1. Predict greater or smaller atomic size and radial distribution in neutral atoms and ions.
2. Measure and compare ionization energies.
3. Compare electron affinities and electronegativities.
 - Predict redox potential.
 - Compare metallic character with other elements; ability to form cations.
 - Predict reactions that may or may not occur due to the trends.
 - Determine greater cell potential (sum of oxidation and reduction potential) between reactions.
 - Complete chemical reactions according to trends.

Summary

- Certain properties—notably atomic radius, ionization energies, and electron affinities—can be qualitatively understood by the positions of the elements on the periodic table. The major trends are summarized in the figure below.
- There are three factors that help in the prediction of the trends in the Periodic Table: number of protons in the nucleus, number of shells, and shielding effect.



Various periodic trends (CC BY-SA 4.0; Sandbh via [Wikipedia](#))

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CHAPTER OVERVIEW

4: Chemical Bond I

[4.1: 4.1-Types of Chemical Bonding](#)

[4.2: Ions](#)

[4.3: Formulas for Ionic Compounds](#)

[4.4: Ionic Nomenclature](#)

[4.5: Transition Metal Ion Formation](#)

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[4.7: Characteristics of Ionic Compounds](#)

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4.1: 4.1-Types of Chemical Bonding

Learning Objectives

- Define the octet rule.
- Describe how ionic bonds are formed.

Atoms can join together by forming a chemical bond, which is a very strong attraction between two atoms. Chemical bonds are formed when electrons in different atoms interact with each other to make an arrangement that is more stable than when the atoms are apart.

What causes atoms to make a chemical bond with other atoms, rather than remaining as individual atoms? A clue comes by considering the noble gas elements, the rightmost column of the periodic table. These elements—helium, neon, argon, krypton, xenon, and radon—do not form compounds very easily, which suggests that they are especially stable as lone atoms. What else do the noble gas elements have in common? Except for helium, they all have eight valence electrons. Chemists have concluded that atoms are especially stable if they have eight electrons in their outermost shell. This useful rule of thumb is called the **octet rule**, and it is a key to understanding why compounds form.

Of the noble gases, only krypton, xenon, and radon have been found to make compounds.

There are three ways for an atom that does not have an octet of valence electrons to obtain an octet in its outer shell. One way is the transfer of electrons to another atom (**losing electrons**). This will result in an atom with a net positive charge known as **cation**. Another option is to capture electrons from another atom (**gaining electrons**). This will result in an atom with a net negative charge known as **anion**. A third option for an atom to obtain an octet of electrons is by **sharing electrons** with another atom. These shared electrons simultaneously occupy the outermost shell of more than one atom. Whether an atom decides to capture, lose, or share electrons with another atom depends on the chemical properties of both atoms, and that will result in different types of chemical bonds

IMPORTANT NOTE: Despite our focus on the octet rule, we must remember that for small atoms, such as hydrogen, helium, and lithium, the first shell is, or becomes, the outermost shell and hold only two electrons. Therefore, these atoms satisfy a “**duet rule**” rather than the octet rule.

Example 4.1.1

A sodium atom has one valence electron. Do you think it is more likely for a sodium atom to lose one electron or gain seven electrons to obtain an octet?

Solution

Although either event is possible, a sodium atom is more likely to lose its single valence electron. When that happens, it becomes an ion with a net positive charge. This can be illustrated as follows:

Sodium atom		Sodium ion	
11 protons	11+	11 protons	11+
11 electrons	11–	10 electrons	10–
	0 overall charge		+1 overall charge

Exercise 4.1.1

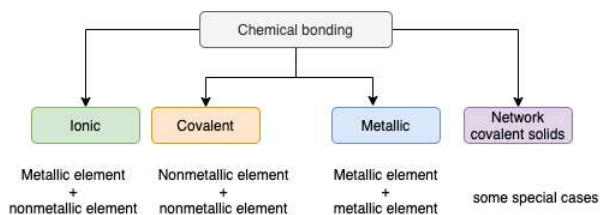
A fluorine atom has seven valence electrons. Do you think it is more likely for a fluorine atom to lose seven electrons or gain one electron to obtain an octet? Write the formula of the resulting ion.

Answer

The process that involves less number of electrons is more favorable. Fluorine would gain one electron. The formula of the resulting ion is F^- .

Types of chemical bonds

There four types of chemical bonds, depending on the nature of the atoms forming the bond (metallic or nonmetallic):



Key Takeaways

- Atoms have a tendency to have eight electrons in their valence shell.
- The attraction of oppositely charged ions is what makes ionic bonds.

Exercises

1. What is the octet rule?
2. How are ionic bonds formed?
3. Why is an ionic compound unlikely to consist of two positively charged ions?
4. Why is an ionic compound unlikely to consist of two negatively charged ions?
5. A calcium atom has two valence electrons. Do you think it will lose two electrons or gain six electrons to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.
6. An aluminum atom has three valence electrons. Do you think it will lose three electrons or gain five electrons to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.
7. A selenium atom has six valence electrons. Do you think it will lose six electrons or gain two electrons to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.
8. An iodine atom has seven valence electrons. Do you think it will lose seven electrons or gain one electron to obtain an octet in its outermost electron shell? Write the formula of the resulting ion.

Answers

1. The octet rule is the concept that atoms tend to have eight electrons in their valence electron shell.
2. Ionic bonds are formed by the attraction between oppositely charged ions.
3. Positive charges repel each other, so an ionic compound is not likely between two positively charged ions.
4. Negative charges repel each other also.
5. Ca atom is more likely to lose two electrons. It will become Ca^{2+} ion.
6. An Al atom is more likely to lose three electrons. It will become Al^{3+} ion.

7. Selenium is more likely to gain two electrons. It will become Se^{2-} ion.
8. Iodine is more likely to gain one electron. It will become I^- ion.

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4.2: Ions

Learning Objectives

- Define the two types of ions.
- Use Lewis diagrams to illustrate ion formation.

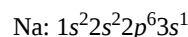
Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite 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.

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.

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

Electron Transfer

We can use electron configurations to illustrate the electron transfer process between sodium atoms and chlorine atoms.



As demonstrated here, a sodium atom (Na) has one valence electron in the third principal energy level. It 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 sodium ion, Na^+ , has the electron configuration with an octet of electrons from the second principal energy level. It is now the same as that of the noble gas neon. The term **isoelectronic** refers to an atom and an ion of a different atom (or two different ions) that have the same electron configuration. The sodium ion is isoelectronic with the neon atom. The equation below illustrates this process.

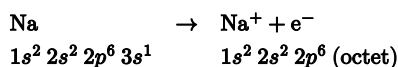


Figure 4.2.1 is a graphical depiction of this process.

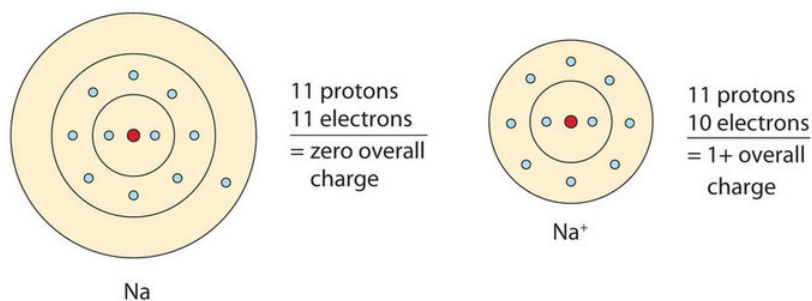


Figure 4.2.1: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge.

Now, let's consider chlorine atom, $\text{Cl: } 1s^2 2s^2 2p^6 3s^2 3p^5$

Only one more electron is needed to achieve an octet in chlorine's valence shell. When a chlorine atom gains an electron, its outermost principal energy level achieves an octet. 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. This process is illustrated below. (In table salt, this electron comes from the sodium atom.)

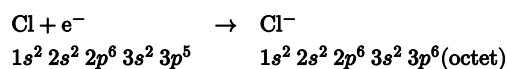


Figure 4.2.2 is a graphical depiction of this process.

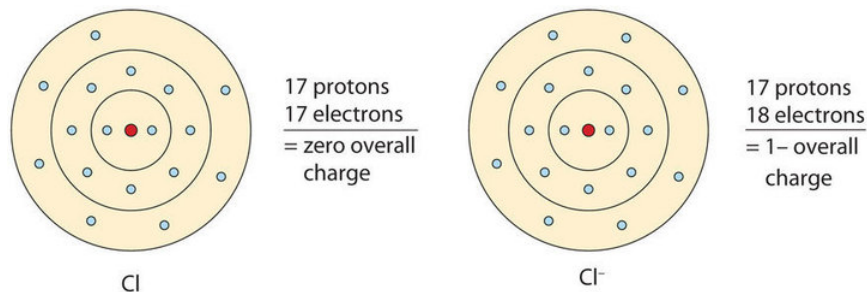


Figure 4.2.2: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1- charge.

With two oppositely charged ions, there is an electrostatic attraction between them because opposite charges attract. The resulting combination is the compound sodium chloride. Notice that there are no leftover electrons. The number of electrons lost by the sodium atom (one) equals the number of electrons gained by the chlorine atom (one), so the compound is electrically neutral. In macroscopic samples of sodium chloride, there are billions and billions of sodium and chloride ions, although there is always the same number of cations and anions.

✓ Example 4.2.1

Write the electron configuration of aluminum atom ($Z=13$). How many electrons must Al lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

Solution

The electron configuration of Al atom is $1s^2 2s^2 2p^6 3s^2 3p^1$. The second shell has octet ($2s^2 2p^6$) while the valence shell has 3 electrons ($3s^2 3p^1$). Al can achieve octet by losing the 3 valence electrons. The resulting cation is Al^{3+} with electron configuration, $1s^2 2s^2 2p^6$.

4.2.1 Exercise

Write the electron configuration of oxygen atom ($Z=8$). How many electrons must O lose/gain to achieve octet? Write the formula of the resulting ion and its electron configuration.

Answer

The electron configuration of O atom is $1s^2 2s^2 2p^4$. The second shell has six electrons ($2s^2 2p^4$) and needs two electrons to achieve octet. Oxygen will gain 2 electrons. The resulting anion is O^{2-} with electron configuration, $1s^2 2s^2 2p^6$.

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1+ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2+ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1- charge. Figure 4.2.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a multiply charged ion. The barium cation is written Ba^{2+} , not Ba^{+2} .

1A						8A
H ⁺	2A					
Li ⁺						
Na ⁺	Mg ²⁺					
K ⁺	Ca ²⁺					
Rb ⁺	Sr ²⁺					

	3A	4A	5A	6A	7A	
			N ³⁻	O ²⁻	F ⁻	
	Al ³⁺		P ³⁻	S ²⁻	Cl ⁻	
				Se ²⁻	Br ⁻	
					I ⁻	

Figure 4.2.3: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

✓ Example 4.2.2

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

- Mg⁺
- K⁺

Solution

(a) Mg is in Group 2A and has two valence electrons. It achieves octet by losing two electrons to form Mg²⁺ cation. Losing only one electron to form Mg⁺ does not make an octet, hence, Mg⁺ is not likely to form.

4.2.2Exercise

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

- S³⁻
- N³⁻

Answer

(a) S is in Group 6A and has six valence electrons. It achieves octet by gaining two electrons to form S²⁻ anion. Gaining three electrons to form S³⁻ does not make it octet, hence, S³⁻ is not likely to form.

Lewis Diagrams

Chemists use simple diagrams to show an atom's valence electrons and how they transfer. These diagrams have two advantages over the electron shell diagrams. First, they show only valence electrons. Second, instead of having a circle around the chemical symbol to represent the electron shell, they have up to eight dots around the symbol; each dot represents a valence electron. These dots are arranged to the right and left and above and below the symbol, **with no more than two dots on a side**. For example, the representation for sodium is as follows:



and the representation for chlorine is as follows:



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

These diagrams are called **Lewis electron dot diagrams**, or simply Lewis diagrams, after Gilbert N. Lewis, the American chemist who introduced them. To write an element's Lewis dot symbol, place the dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more than four valence electrons, are again distributed one at a time, each paired with one of the first four. In other words, **place the dots singly on each side before pairing them**. The Lewis electron dot diagram of fluorine, for example, with seven valence electrons, is constructed as follows:

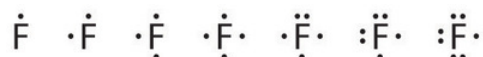
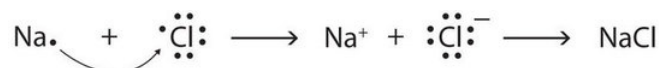


Figure 4.2.4 shows the electron configurations and Lewis diagrams of the elements lithium through neon, which is the entire second period of the periodic table. For the **main group** elements, the **number of valence electrons** is the **same as the group number** listed at the top of the periodic table.



Figure 4.2.4: Lewis Electron Dot Diagrams of the Elements Lithium through Neon

The transfer of electrons can be illustrated easily with Lewis diagrams:



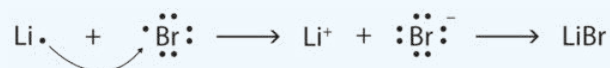
In representing the final formula, the dots are omitted.

✓ Example 4.2.3

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

Solution

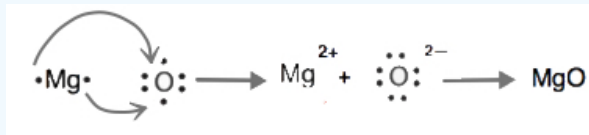
From the periodic table, we see that lithium is in the same column as sodium, so it will have the same valence shell electron configuration. That means that the neutral lithium atom will have the same Lewis diagram that the sodium atom has. Similarly, bromine is in the same column as chlorine, so it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



? Exercise 4.2.3

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

Answer



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

Most of the elements that make ionic compounds form an ion that has a characteristic charge. For example, sodium makes ionic compounds in which the sodium ion always has a $1+$ charge. Chlorine makes ionic compounds in which the chloride ion always has a $1-$ charge. Some elements, especially transition metals, can form ions of multiple charges. Figure 4.2.5 shows the characteristic charges for some of these ions. As we saw in Figure 4.2.1, there is a pattern to the charges on many of the main group ions, but there is no simple pattern for transition metal ions (or for the larger main group elements).

[illegible]

Key Takeaways

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

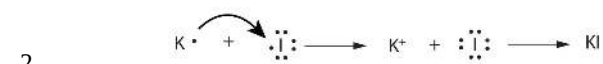
Exercises

1. What are the two types of ions?
2. Use Lewis diagrams to illustrate the formation of an ionic compound from a potassium atom and an iodine atom.
3. When the following atoms become ions, what charges do they acquire?
 - a. Li
 - b. S
 - c. Ca
 - d. F
4. Identify each as a cation, an anion, or neither.
 - a. H^+
 - b. Cl^-
 - c. O_2
 - d. Ba^{2+}
 - e. CH_4
 - f. CS_2
5. Identify each as a cation, an anion, or neither.
 - a. NH_3
 - b. Br^-
 - c. H^-
 - d. Hg^{2+}

- e. CCl_4
 - f. SO_3
6. Write the electron configuration for each ion.
- a. Li^+
 - b. Mg^{2+}
 - c. F^-
 - d. S^{2-}
7. Write the electron configuration for each ion.
- a. Na^+
 - b. Be^{2+}
 - c. Cl^-
 - d. O^{2-}
8. Draw Lewis diagrams for the ions listed in Exercise 6. Also include Lewis diagrams for the respective neutral atoms as a comparison.
9. Draw Lewis diagrams for the ions listed in Exercise 7. Also include Lewis diagrams for the respective neutral atoms as a comparison.
10. Using Lewis diagrams, show the electron transfer for the formation of LiF .
11. Using Lewis diagrams, show the electron transfer for the formation of MgO .
12. Using Lewis diagrams, show the electron transfer for the formation of Li_2O .
13. Using Lewis diagrams, show the electron transfer for the formation of CaF_2 .
14. What characteristic charge do atoms in the first column of the periodic table have when they become ions?
15. What characteristic charge do atoms in the second column of the periodic table have when they become ions?
16. What characteristic charge do atoms in the third-to-last column of the periodic table have when they become ions?
17. What characteristic charge do atoms in the next-to-last column of the periodic table have when they become ions?

Answers

1. Cations (positive charged) and anions (negative charged).



3.
 - a. 1^+
 - b. 2^-
 - c. 2^+
 - d. 1^-
4.
 - a. cation
 - b. anion
 - c. neither
 - d. cation
 - e. neither
 - f. neither
5.
 - a. neither
 - b. anion

- c. anion
- d. cation
- e. neither
- f. neither

6.

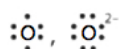
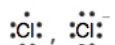
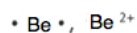
- a. $1s^2$
- b. $1s^2 2s^2 2p^6$
- c. $1s^2 2s^2 2p^6$
- d. $1s^2 2s^2 2p^6 3s^2 3p^6$

7.

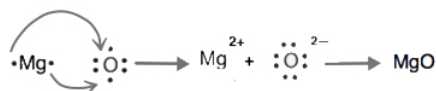
- a. $1s^2 2s^2 2p^6$
- b. $1s^2$
- c. $1s^2 2s^2 2p^6 3s^2 3p^6$
- d. $1s^2 2s^2 2p^6$

8.

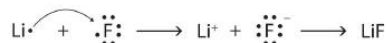
- a. $\text{Li}^\bullet, \text{Li}^+$
- b. $\bullet\text{Mg}^\bullet, \text{Mg}^{2+}$
- c. $\text{:}\ddot{\text{F}}\text{:}, \text{:}\ddot{\text{F}}\text{:}^-$
- d. $\text{:}\ddot{\text{S}}\text{:}, \text{:}\ddot{\text{S}}\text{:}^{2-}$



9.



10.



11.



13.

12.



14. $1+$

15. $2+$

16. $2-$

17. $1-$

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4.3: Formulas for Ionic Compounds

Learning Objectives

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

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

Ionic compounds exist as alternating positive and negative ions in regular, three-dimensional arrays called crystals (Figure 4.3.1). As you can see, there are no individual **NaCl** “particles” in the array; instead, there is a continuous lattice of alternating sodium and chloride ions. However, we can use the ratio of sodium ions to chloride ions, expressed in the lowest possible whole numbers, as a way of describing the compound. In the case of sodium chloride, the ratio of sodium ions to chloride ions, expressed in lowest whole numbers, is 1:1, so we use **NaCl** (one **Na** symbol and one **Cl** symbol) to represent the compound. Thus, **NaCl** is the chemical formula for sodium chloride, which is a concise way of describing the relative number of different ions in the compound. A macroscopic sample is composed of myriads of NaCl pairs; each individual pair called a **formula unit or empirical formula**. The formula unit or empirical formula represents the minimum proportion between cations and anions in the crystal lattice. In sodium chloride, there is one Na^+ cation per each Cl^- , so the formula unit for sodium chloride is **NaCl**. In calcium chloride, there are two Cl^- anions per Ca^{2+} cation, so the empirical formula for calcium chloride is **CaCl₂**.

Although it is convenient to think that **NaCl** crystals are composed of individual **NaCl** units, Figure 4.3.1 shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by ions of opposite charge.

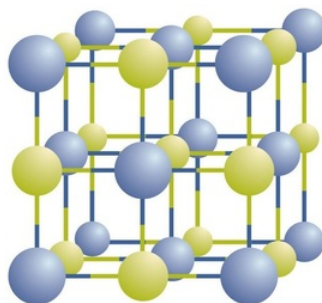


Figure 4.3.1: A Sodium Chloride Crystal. A crystal contains a three-dimensional array of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions.

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

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

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

If we look at the ionic compound consisting of lithium ions and bromide ions, we see that the lithium ion has a 1+ charge and the bromide ion has a 1− charge. Only one ion of each is needed to balance these charges. The formula for lithium bromide is **LiBr**.

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

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



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



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

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

For compounds in which the ratio of ions is not as obvious, the subscripts in the formula can be obtained by **crossing charges**: use the absolute value of the charge on one ion as the subscript for the other ion. This method is shown schematically in Figure 3.3.2.

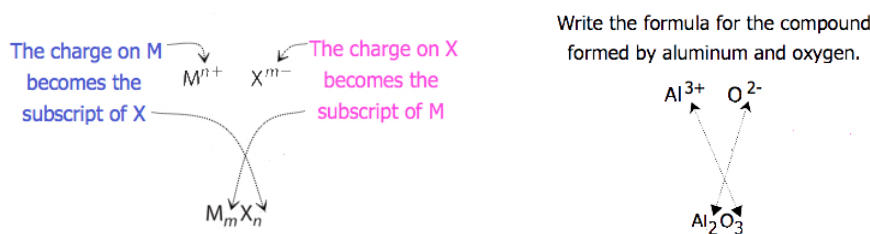


Figure 4.3.2: Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges.

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

Example 4.3.1

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

- the sodium ion and the sulfur ion
- the aluminum ion and the fluoride ion
- the 3+ iron ion and the oxygen ion

Solution

- To obtain a valence shell octet, sodium forms an ion with a 1+ charge, while the sulfur ion has a 2− charge. Two sodium 1+ ions are needed to balance the 2− charge on the sulfur ion. Rather than writing the formula as NaNaS , we shorten it by convention to Na_2S .
- The aluminum ion has a 3+ charge, while the fluoride ion formed by fluorine has a 1− charge. Three fluorine 1− ions are needed to balance the 3+ charge on the aluminum ion. This combination is written as AlF_3 .
- Iron can form two possible ions, but the ion with a 3+ charge is specified here. The oxygen atom has a 2− charge as an ion. To balance the positive and negative charges, we look to the least common multiple—6: two iron 3+ ions will give 6+, while three 2− oxygen ions will give 6−, thereby balancing the overall positive and negative charges. Thus, the formula for this ionic compound is Fe_2O_3 . Alternatively, use the crossing charges method shown in Figure 3.3.2.

Exercise 4.3.1

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:



Polyatomic Ions

Some ions consist of groups of atoms **covalently** bonded together and have an **overall electric charge**. Because these ions contain more than one atom, they are called polyatomic ions. The Lewis structures, names and formulas of some polyatomic ions are found in Table 3.3.1.

POLYATOMIC IONS: NAMES, FORMULAE & CHARGES						
A polyatomic ion is a charged species consisting of two or more atoms covalently bonded together. Here's a guide to some of the most common examples!						
ACETATE	CARBONATE	HYDROGEN CARBONATE	CYANIDE	NITRATE	NITRITE	AMMONIUM
Formula: C ₂ H ₃ O ₂ ⁻	Formula: CO ₃ ²⁻	Formula: HCO ₃ ⁻	Formula: CN ⁻	Formula: NO ₃ ⁻	Formula: NO ₂ ⁻	Formula: NH ₄ ⁺
HYDROXIDE	PEROXIDE	SULFITE	SULFATE	HYDROGEN SULFATE	THIOSULFATE	PHOSPHATE
Formula: OH ⁻	Formula: O ₂ ²⁻	Formula: SO ₃ ²⁻	Formula: SO ₄ ²⁻	Formula: HSO ₄ ⁻	Formula: S ₂ O ₃ ²⁻	Formula: PO ₄ ³⁻
HYPOCHLORITE	CHLORITE	CHLORATE	PERCHLORATE	CHROMATE	DICHROMATE	PERMANGANATE
Formula: ClO ⁻	Formula: ClO ₂ ⁻	Formula: ClO ₃ ⁻	Formula: ClO ₄ ⁻	Formula: CrO ₄ ²⁻	Formula: Cr ₂ O ₇ ²⁻	Formula: MnO ₄ ⁻
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Table 4.3.1: Some Polyatomic Ions

Polyatomic ions have defined formulas, names, and charges that cannot be modified in any way. Table 4.3.2 lists the ion names and ion formulas of the most common polyatomic ions. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1– charge. Figure 4.3.2 lists the most common polyatomic ions.

Table 4.3.2: Ion Names and Ion Formulas of Common Polyatomic Ions

I o n N a m e	Ion Formula
a m m o n i u m i o n	NH_4^+
h y d r o x i d e i o n	OH^-
c y a n i d e i o n	CN^-
c a r b o n a t e i	CO_3^{2-}

Ion Name	Ion Formula
o n	
b i c a r b o n a t e o r h y d r o g e n c a r b o n a t e	HCO_3^-
a c e t a t e i o n	$\text{C}_2\text{H}_3\text{O}_2^-$ or CH_3CO_2^-
n	NO_3^-

Ion Name	Ion Formula
nitrate ion	
nitrite ion	NO_2^-
sulfate ion	SO_4^{2-}
sulfite ion	SO_3^{2-}
phosphates	PO_4^{3-}

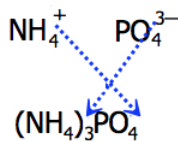
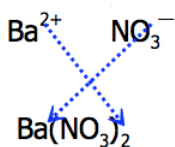
I o n N a m e	Ion Formula
p h a t e i o n	
p h o s p h i t e i o n	PO_3^{3-}

Note that only one polyatomic ion in this Table, the ammonium ion (NH_4^{+1}), is a cation. This polyatomic ion contains one nitrogen and four hydrogens that collectively bear a +1 charge. The remaining polyatomic ions are all negatively-charged and, therefore, are classified as anions. However, only two of these, the hydroxide ion and the cyanide ion, are named using the "-ide" suffix that is typically indicative of negatively-charged particles. The remaining polyatomic anions, which all contain oxygen, in combination with another non-metal, exist as part of a series in which the number of oxygens within the polyatomic unit can vary. As has been repeatedly emphasized in several sections of this text, no two chemical formulas should share a common chemical name. A single suffix, "-ide," is insufficient for distinguishing the names of the anions in a related polyatomic series. Therefore, "-ate" and "-ite" suffixes are employed, in order to denote that the corresponding polyatomic ions are part of a series. Additionally, these suffixes also indicate the relative number of oxygens that are contained within the polyatomic ions. Note that all of the polyatomic ions whose names end in "-ate" contain one more oxygen than those polyatomic anions whose names end in "-ite." Unfortunately, much like the common system for naming transition metals, these suffixes only indicate the *relative* number of oxygens that are contained within the polyatomic ions. For example, the **nitrate** ion, which is symbolized as NO_3^{-1} , has one more oxygen than the **nitrite** ion, which is symbolized as NO_2^{-1} . However, the **sulfate** ion is symbolized as SO_4^{-2} . While both the **nitrate** ion and the **sulfate** ion share an "-ate" suffix, the former contains three oxygens, but the latter contains four. Additionally, both the **nitrate** ion and the **sulfite** ion contain three oxygens, but these polyatomic ions do not share a common suffix. Unfortunately, the relative nature of these suffixes mandates that the ion formula/ion name combinations of the polyatomic ions must simply be memorized.

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

Write the formula for the compound formed by

- a. barium and nitrate b. ammonium and phosphate



Example 4.3.2

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

- the potassium ion and the sulfate ion
- the calcium ion and the nitrate ion

Solution

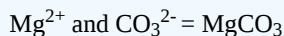
- Potassium ions have a charge of 1+, while sulfate ions have a charge of 2-. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4 .
- Calcium ions have a charge of 2+, while nitrate ions have a charge of 1-. We will need two nitrate ions to balance the charge on each calcium ion. The formula for nitrate must be enclosed in parentheses. Thus, we write $\text{Ca}(\text{NO}_3)_2$ as the formula for this ionic compound.

Exercise 4.3.2

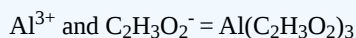
Write the chemical formula for an ionic compound composed of each pair of ions.

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

Answer a:



Answer b:



Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between **metal and nonmetal** elements are usually **ionic**. For example, CaBr_2 contains a metallic element (calcium, a group 2A metal) and a nonmetallic element (bromine, a group 7A nonmetal). Therefore, it is 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 5A, and oxygen, from group 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 4.3.3

Identify each compound as ionic or not ionic.

- Na_2O
- PCl_3

- c. NH_4Cl
- d. OF_2

Solution

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

Exercise 4.3.3

Identify each compound as ionic or not ionic.

- a. N_2O
- b. FeCl_3
- c. $(\text{NH}_4)_3\text{PO}_4$
- d. SOCl_2

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Looking Closer: Blood and Seawater

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

Ion	Percent in Seawater	Percent in Blood
Na^+	2.36	0.322
Cl^-	1.94	0.366
Mg^{2+}	0.13	0.002
SO_4^{2-}	0.09	—
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
HCO_3^-	0.002	0.165
$\text{HPO}_4^{2-}, \text{H}_2\text{PO}_4^-$	—	0.01

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

Key Takeaways

- Proper chemical formulas for ionic compounds balance the total positive charge with the total negative charge.
- Groups of atoms with an overall charge, called polyatomic ions, also exist.

EXERCISES

1. What information is contained in the formula of an ionic compound?
2. Why do the chemical formulas for some ionic compounds contain subscripts, while others do not?
3. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. Mg^{2+} and I^-
 - b. Na^+ and O^{2-}
4. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. Na^+ and Br^-
 - b. Mg^{2+} and Br^-
 - c. Mg^{2+} and S^{2-}
5. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. K^+ and Cl^-
 - b. Mg^{2+} and Cl^-
 - c. Mg^{2+} and Se^{2-}
6. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. Na^+ and N^{3-}
 - b. Mg^{2+} and N^{3-}
 - c. Al^{3+} and S^{2-}
7. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. Li^+ and N^{3-}
 - b. Mg^{2+} and P^{3-}
 - c. Li^+ and P^{3-}
8. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. Fe^{3+} and Br^-
 - b. Fe^{2+} and Br^-
 - c. Au^{3+} and S^{2-}
 - d. Au^+ and S^{2-}
9. Write the chemical formula for the ionic compound formed by each pair of ions.
 - a. Cr^{3+} and O^{2-}
 - b. Cr^{2+} and O^{2-}
 - c. Pb^{2+} and Cl^-
 - d. Pb^{4+} and Cl^-
10. Write the chemical formula for the ionic compound formed by each pair of ions.

- a. Cr^{3+} and NO_3^-
- b. Fe^{2+} and PO_4^{3-}
- c. Ca^{2+} and CrO_4^{2-}
- d. Al^{3+} and OH^-

11. Write the chemical formula for the ionic compound formed by each pair of ions.

- a. NH_4^+ and NO_3^-
- b. H^+ and $\text{Cr}_2\text{O}_7^{2-}$
- c. Cu^+ and CO_3^{2-}
- d. Na^+ and HCO_3^-

12. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

- a. Ba and S
- b. Cs and I

13. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

- a. K and S
- b. Sc and Br

14. Which compounds would you predict to be ionic?

- a. Li_2O
- b. $(\text{NH}_4)_2\text{O}$
- c. CO_2
- d. FeSO_3
- e. C_6H_6
- f. $\text{C}_2\text{H}_6\text{O}$

15. Which compounds would you predict to be ionic?

- a. $\text{Ba}(\text{OH})_2$
- b. CH_2O
- c. NH_2CONH_2
- d. $(\text{NH}_4)_2\text{CrO}_4$
- e. C_8H_{18}
- f. NH_3

Answers

1. the ratio of each kind of ion in the compound

2. Sometimes more than one ion is needed to balance the charge on the other ion in an ionic compound.

3.

- a. MgI_2
- b. Na_2O

4.

- a. NaBr

- b. MgBr_2
- c. MgS

5.

- a. KCl
- b. MgCl_2
- c. MgSe

6.

- a. Na_3N
- b. Mg_3N_2
- c. Al_2S_3

7.

- a. Li_3N
- b. Mg_3P_2
- c. Li_3P

8.

- a. FeBr_3
- b. FeBr_2
- c. Au_2S_3
- d. Au_2S

9.

- a. Cr_2O_3
- b. CrO
- c. PbCl_2
- d. PbCl_4

10.

- a. $\text{Cr}(\text{NO}_3)_3$
- b. $\text{Fe}_3(\text{PO}_4)_2$
- c. CaCrO_4
- d. $\text{Al}(\text{OH})_3$

11.

- a. NH_4NO_3
- b. $\text{H}_2\text{Cr}_2\text{O}_7$

- c. Cu_2CO_3
- d. NaHCO_3

12.

- a. Ba^{2+} , S^{2-} , BaS
- b. Cs^+ , I^- , CsI

13.

- a. K^+ , S^{2-} , K_2S
- b. Sc^{3+} , Br^- , ScBr_3

14.

- a. ionic
- b. ionic
- c. not ionic
- d. ionic
- e. not ionic
- f. not ionic

15.

- a. ionic
- b. not ionic
- c. not ionic
- d. ionic
- e. not ionic
- f. not ionic

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4.4: Ionic Nomenclature

Learning Objectives

- To use the rules for naming 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. 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**, 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 4.4.1: The Common System of Cation Names

Element	Stem	Charge	Name
iron	ferr-	2+	ferrous ion
		3+	ferric ion
copper	cupr-	1+	cuprous ion
		2+	cupric ion
tin	stann-	2+	stannous ion
		4+	stannic ion
lead	plumb-	2+	plumbous ion
		4+	plumbic ion
chromium	chrom-	2+	chromous ion
		3+	chromic ion
gold	aur-	1+	aurous ion
		3+	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 4.4.1 lists the elements that use the common system, along with their respective cation names.

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 4.4.2 lists the names of some common monatomic ions.

Table 4.4.2: Some Monatomic Anions

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

The polyatomic ions have their own characteristic names, as discussed earlier.

✓ Example 4.4.1

Name each ion.

- Ca^{2+}
- S^{2-}
- SO_3^{2-}
- NH_4^+
- Cu^+

Answer a

the calcium ion

Answer b

the sulfide ion (from Table 4.4.2)

Answer c

the sulfite ion

Answer d

the ammonium ion

Answer e

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

? Exercise 4.4.1

Name each ion.

- Fe^{2+}
- Fe^{3+}
- SO_4^{2-}

- d. Ba^{2+}
- e. HCO_3^-

Answer a

the iron (II) or ferrous ion

Answer b

the iron (III) or ferric ion

Answer c

the sulfate ion

Answer d

the barium ion

Answer e

the bicarbonate ion or hydrogen carbonate ion

✓ Example 4.4.2

Write the formula for each ion.

- a. the bromide ion
- b. the phosphate ion
- c. the cupric ion
- d. the magnesium ion

Answer a

Br^-

Answer b

PO_4^{3-}

Answer c

Cu^{2+}

Answer d

Mg^{2+}

? Exercise 4.4.2

Write the formula for each ion.

- a. the fluoride ion
- b. the carbonate ion
- c. the ferrous ion
- d. the potassium ion

Answer a

F^-

Answer b

CO_3^{2-}

Answer c

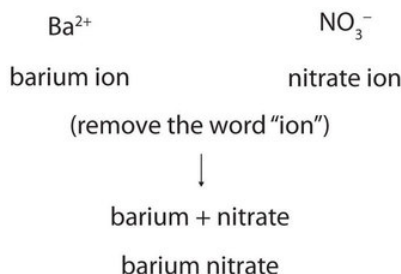


Answer d



Naming Compounds

Now that we know how to name ions, we are ready to name ionic compounds. We do so by placing the name of the cation first, followed by the name of the anion, and dropping the word *ion* from both parts. For example, what is the name of the compound whose formula is **Ba(NO₃)₂**?



The compound's name does not indicate that there are two nitrate ions for every barium ion. You must determine the relative numbers of ions by balancing the positive and negative charges.

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider **FeCl₂** and **FeCl₃**. In the first compound, the iron ion has a 2+ charge because there are two **Cl⁻** ions in the formula (1- charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three **Cl⁻** ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride. If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively.

✓ Example 4.4.3

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

- $\text{Ca}_3(\text{PO}_4)_2$
- $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$
- KCl
- CuCl
- SnF_2

Answer a

calcium phosphate

Answer b

ammonium dichromate (the prefix *di-* is part of the name of the anion)

Answer c

potassium chloride

Answer d

copper(I) chloride or cuprous chloride

Answer e

tin(II) fluoride or stannous fluoride

? Exercise 4.4.3

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

- ZnBr_2
- $\text{Fe}(\text{NO}_3)_3$
- Al_2O_3
- CuF_2
- AgF

Answer a

zinc bromide

Answer b

iron (III) nitrate or ferric nitrate

Answer c

aluminum oxide

Answer d

copper (II) fluoride or cupric fluoride

Answer e

silver fluoride

Figure 4.4.1 is a synopsis of how to name simple ionic compounds.

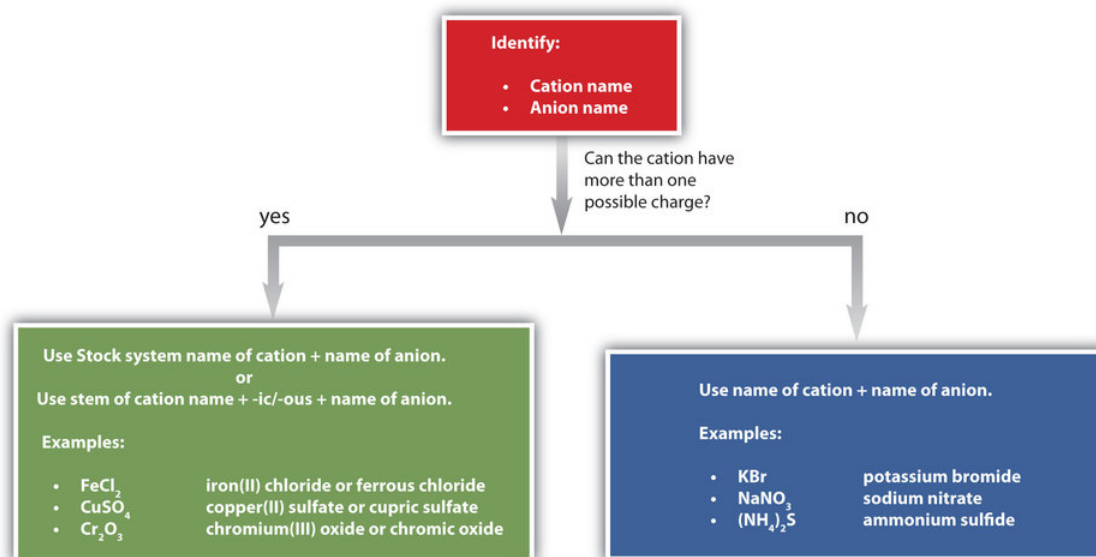


Figure 4.4.1: A Guide to Naming Simple Ionic Compounds. Follow these steps to name a simple ionic compound.

Identify the cation name and the anion name. If the cation can have more than one possible charge, either use the Stock system name of the cation and name of the anion, or use the stem of the cation name and -ic/-ous and the name of the anion. Examples of this would be FeCl_2 , which is iron(II) chloride or ferrous chloride, CuSO_4 , which is copper(II) sulfate or cupric sulfate, and Cr_2O_3 , which is chromium(III) oxide or chromic oxide. If the cation can not have more than one possible charge, use the name of the cation and the name of the anion. Examples of this would be KBr , which is potassium bromide, NaNO_3 , which is sodium nitrate, and $(\text{NH}_4)_2\text{S}$, which is ammonium sulfide.

KEY TAKEAWAY

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

EXERCISES

- Briefly describe the process for naming an ionic compound.
 - In what order do the names of ions appear in the names of ionic compounds?
- Which ionic compounds can be named using two different systems? Give an example.
- Name each ion.
 - Ra^{2+}
 - P^{3-}
 - H_2PO_4^-
 - Sn^{4+}
- Name each ion.
 - Cs^+
 - As^{3-}
 - HSO_4^-
 - Sn^{2+}
- Name the ionic compound formed by each pair of ions.
 - Na^+ and Br^-
 - Mg^{2+} and Br^-
 - Mg^{2+} and S^{2-}
- Name the ionic compound formed by each pair of ions.
 - K^+ and Cl^-
 - Mg^{2+} and Cl^-
 - Mg^{2+} and Se^{2-}
- Name the ionic compound formed by each pair of ions.
 - Na^+ and N^{3-}
 - Mg^{2+} and N^{3-}
 - Al^{3+} and S^{2-}
- Name the ionic compound formed by each pair of ions.
 - Li^+ and N^{3-}
 - Mg^{2+} and P^{3-}
 - Li^+ and P^{3-}
- Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.
 - Fe^{3+} and Br^-
 - Fe^{2+} and Br^-
 - Au^{3+} and S^{2-}
 - Au^+ and S^{2-}
- Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.
 - Cr^{3+} and O^{2-}
 - Cr^{2+} and O^{2-}
 - Pb^{2+} and Cl^-
 - Pb^{4+} and Cl^-
- Name the ionic compound formed by each pair of ions. Use both the Stock and common systems, where appropriate.
 - Cr^{3+} and NO_3^-

- b. Fe^{2+} and PO_4^{3-}
- c. Ca^{2+} and CrO_4^{2-}
- d. Al^{3+} and OH^-

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

- a. NH_4^+ and NO_3^-
- b. K^+ and $\text{Cr}_2\text{O}_7^{2-}$
- c. Cu^+ and CO_3^{2-}
- d. Na^+ and HCO_3^-

14. Give two names for each compound.

- a. $\text{Al}(\text{HSO}_4)_3$
- b. $\text{Mg}(\text{HSO}_4)_2$

15. Give two names for each compound.

- a. $\text{Co}(\text{HCO}_3)_2$
- b. LiHCO_3

Answers

1. Name the cation and then the anion but don't use numerical prefixes.
2. the cation name followed by the anion name
3. Ionic compounds in which the cation can have more than one possible charge have two naming systems. FeCl_3 is either iron(III) chloride or ferric chloride (answers will vary).
4.
 - a. the radium ion
 - b. the phosphide ion
 - c. the dihydrogen phosphate ion
 - d. the tin(IV) ion or the stannic ion
5.
 - a. the cesium ion
 - b. the arsenide ion
 - c. the hydrogen sulfate ion
 - d. the tin(II) ion or the stannous ion
6.
 - a. sodium bromide
 - b. magnesium bromide
 - c. magnesium sulfide
7.
 - a. potassium chloride
 - b. magnesium chloride
 - c. magnesium selenide
8.
 - a. sodium nitride
 - b. magnesium nitride
 - c. aluminum sulfide
- 9.

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

10.

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

11.

- a. chromium(III) oxide or chromic oxide
- b. chromium(II) oxide or chromous oxide
- c. lead(II) chloride or plumbous chloride
- d. lead(IV) chloride or plumbic chloride

12.

- a. chromium(III) nitrate or chromic nitrate
- b. iron(II) phosphate or ferrous phosphate
- c. calcium chromate
- d. aluminum hydroxide

13.

- a. ammonium nitrate
- b. potassium dichromate
- c. copper(I) carbonate or cuprous carbonate
- d. sodium hydrogen carbonate or sodium bicarbonate

14.

- a. aluminum hydrogen sulfate or aluminum bisulfate
- b. magnesium hydrogen sulfate or magnesium bisulfate

15.

- a. cobalt hydrogen carbonate or cobalt bicarbonate
- b. lithium hydrogen carbonate or lithium bicarbonate

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4.5: Transition Metal Ion Formation

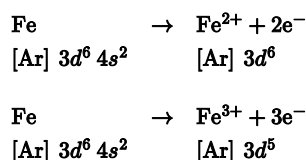
Figure 4.5.1 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via [Wikipedia](#))

How do transition metals form ions?

The transition metals are an interesting and challenging group of elements. They have perplexing patterns of electron distribution that do not always follow the electron filling rules. Predicting how they will form ions is also not always obvious.

Transition Metal Ions

Transition metals belong to the *d* block, meaning that the *d* sublevel of electrons is in the process of being filled with up to ten electrons. Many transition metals cannot lose enough electrons to attain a noble-gas electron configuration. In addition, the majority of transition metals are capable of adopting ions with different charges. Iron, which forms either the Fe^{2+} or Fe^{3+} ions, loses electrons as shown below.

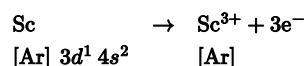


According to the Aufbau process, the electrons fill the *4s* sublevel before beginning to fill the *3d* sublevel. However, the outermost *s* electrons are always the first to be removed in the process of forming transition metal cations. Because most transition metals have two valence electrons, the charge of $2+$ is a very common one for their ions. This is the case for iron above. A half-filled *d* sublevel (d^5) is particularly stable, which is the result of an iron atom losing a third electron.

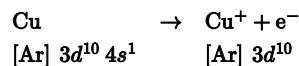
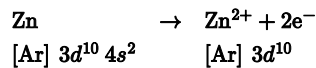


Figure 4.5.2: A. Rust is a complex combination of oxides of iron, among them iron (III) oxide, Fe_2O_3 . B. Iron (II) sulfate, FeSO_4 , has been known since ancient times as green vitriol and was used for centuries in the manufacture of inks. (CC by 2.0; Public Domain; (A) Paulnasca; (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Rust_screw.jpg; (B) <http://commons.wikimedia.org/wiki/File:Iron%2528II%2529-sulfate-heptahydrate-sample.jpg>)

Some transition metals that have relatively few *d* electrons may attain a noble-gas electron configuration. Scandium is an example.



Others may attain configurations with a full d sublevel, such as zinc and copper.



The resulting configuration above, with 18 electrons in the outermost principal energy level, is referred to as a **pseudo noble-gas electron configuration**. It gives particular stability to the Zn^{2+} and Cu^{+} ions.



Summary

- Transition metal ion formation is more complex than simple cation formation.
- Transition metal ions often involve rearrangements of both d and s electrons.

Review

1. What block do the transition metals fall in?
2. When writing the electron configuration for a transition metal, which sublevel (s, p, d, f) is filled first?
3. When forming transition metal ions, which sublevel (s, p, d, f) loses electrons first? What is the pseudo noble-gas electron configuration?

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4.6: Formula Mass

Learning Objectives

- To determine the formula mass of an ionic compound.

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

Let us start by calculating the formula 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:

Na:	22.99 amu
Cl:	+ 35.45 amu
Total:	58.44 amu

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

When an ionic compound has more than one anion or cation, you must remember to use the proper multiple of the atomic mass for the element in question. For the formula mass of calcium fluoride (CaF₂), we must multiply the mass of the fluorine atom by 2 to account for the two fluorine atoms in the chemical formula:

Ca:	1 × 40.08	40.08 amu
F:	2 × 19.00 =	+ 38.00 amu
Total:		78.08 amu

The formula mass of CaF₂ is 78.08 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. For example, potassium nitrate (KNO₃) has one potassium atom, one nitrogen atom, and three oxygen atoms:

K:	1 × 39.10	39.10 amu
N:	1 × 14.01	+ 14.01 amu
O:	3 × 16.00 =	+ 48.00 amu
Total:		101.11 amu

The formula mass of KNO₃ is 101.11 amu.

Potassium nitrate is a key ingredient in gunpowder and has been used clinically as a diuretic.

When a formula contains more than one polyatomic unit in the chemical formula, as in Ca(NO₃)₂, do not forget to multiply the atomic mass of every atom inside the parentheses by the subscript outside the parentheses. This is necessary because the subscript

refers to the *entire polyatomic ion*. Thus, for $\text{Ca}(\text{NO}_3)_2$, the subscript 2 implies two complete nitrate ions, so we must sum the masses of two (1×2) nitrogen atoms and six (3×2) oxygen atoms, along with the mass of a single calcium atom:

Ca:	1×40.08	40.08 amu
N:	$2 \times 14.01 =$	+ 28.02 amu
O:	$6 \times 16.00 =$	+ 96.00 amu
Total:		164.10 amu

The key to calculating the formula mass of an ionic compound is to correctly count each atom in the formula and multiply the atomic masses of its atoms accordingly.

Example 4.6.1

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

- FeCl_3
- $(\text{NH}_4)_3\text{PO}_4$

Solution

a.

Fe:		55.85 amu
Cl:	$3 \times 35.45 =$	+ 106.35 amu
Total:		162.20 amu

The formula mass of FeCl_3 is 162.20 amu.

- When we distribute the subscript 3 through the parentheses containing the formula for the ammonium ion, we see that we have 3 nitrogen atoms and 12 hydrogen atoms. Thus, we set up the sum as follows:

N:	$3 \times 14.01 =$	42.03 amu
H:	$12 \times 1.01 =$	+ 12.12 amu
P:		+ 30.97 amu
O:	$4 \times 16.00 =$	+ 64.00 amu
Total:		149.12 amu

The formula mass for $(\text{NH}_4)_3\text{PO}_4$ is 149.12 amu.

Exercise 4.6.1

Use the atomic masses (**rounded to two decimal places**) from the **Periodic Table** 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 a:

$$47.87 + 2(16.00) = 79.87 \text{ amu}$$

Answer b:

$$107.9 + 79.90 = 187.8 \text{ amu}$$

Answer c:

$$197.0 + 3(14.01) + 9(16.00) = 383.03 \text{ amu}$$

Answer d:

$$3(55.85) + 2(30.97) + 8(16.00) = 357.49 \text{ amu}$$

To Your Health: Hydrates

Some ionic compounds have water (H_2O) incorporated within their formula unit. These compounds, called *hydrates*, have a characteristic number of water units associated with each formula unit of the compound. Hydrates are solids, not liquids or solutions, despite the water they contain.

To write the chemical formula of a hydrate, write the number of water units per formula unit of compound after its chemical formula. The two chemical formulas are separated by a vertically centered dot. The hydrate of copper(II) sulfate has five water units associated with each formula unit, so it is written as $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. The name of this compound is copper(II) sulfate pentahydrate, with the *penta-* prefix indicating the presence of five water units per formula unit of copper(II) sulfate.



Magnesium sulfate heptahydrate. (Public Domain; [Chemicalinterest](#)).

Hydrates have various uses in the health industry. Calcium sulfate hemihydrate ($\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$), known as plaster of Paris, is used to make casts for broken bones. Epsom salt ($\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$) is used as a bathing salt and a laxative. Aluminum chloride hexahydrate is an active ingredient in antiperspirants. The accompanying table lists some useful hydrates.

Table 4.6.1: Names and Formulas of Some Widely Used Hydrates

Formula	Name	Uses
$\text{AlCl}_3 \cdot 6\text{H}_2\text{O}$	aluminum chloride hexahydrate	antiperspirant
$\text{CaSO}_4 \cdot \frac{1}{2}\text{H}_2\text{O}$	calcium sulfate hemihydrate (plaster of Paris)	casts (for broken bones and castings)
$\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$	calcium sulfate dihydrate (gypsum)	drywall component
$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$	cobalt(II) chloride hexahydrate	drying agent, humidity indicator
$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$	copper(II) sulfate pentahydrate	fungicide, algicide, herbicide
$\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$	magnesium sulfate heptahydrate (Epsom salts)	laxative, bathing salt
$\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$	sodium carbonate decahydrate (washing soda)	laundry additive/cleaner

KEY TAKEAWAY

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

EXERCISES

1. What is the relationship between atomic mass and formula mass?
2. How are subscripts used to determine a formula mass when more than one polyatomic ion is present in a chemical formula?
3. What is the formula mass for the ionic compound formed by each pair of ions?
 - a. Na^+ and Br^-
 - b. Mg^{2+} and Br^-
 - c. Mg^{2+} and S^{2-}
4. What is the formula mass for the ionic compound formed by each pair of ions?
 - a. K^+ and Cl^-
 - b. Mg^{2+} and Cl^-
 - c. Mg^{2+} and Se^{2-}
5. What is the formula mass for the ionic compound formed by each pair of ions?
 - a. Na^+ and N^{3-}
 - b. Mg^{2+} and N^{3-}
 - c. Al^{3+} and S^{2-}
6. What is the formula mass for the ionic compound formed by each pair of ions?
 - a. Li^+ and N^{3-}
 - b. Mg^{2+} and P^{3-}
 - c. Li^+ and P^{3-}
7. What is the formula mass for each compound?
 - a. FeBr_3
 - b. FeBr_2
 - c. Au_2S_3
 - d. Au_2S
8. What is the formula mass for each compound?
 - a. Cr_2O_3
 - b. CrO
 - c. PbCl_2
 - d. PbCl_4
9. What is the formula mass for each compound?
 - a. $\text{Cr}(\text{NO}_3)_3$
 - b. $\text{Fe}_3(\text{PO}_4)_2$
 - c. CaCrO_4
 - d. $\text{Al}(\text{OH})_3$
10. What is the formula mass for each compound?
 - a. NH_4NO_3
 - b. $\text{K}_2\text{Cr}_2\text{O}_7$
 - c. Cu_2CO_3
 - d. NaHCO_3
11. What is the formula mass for each compound?

- a. $\text{Al}(\text{HSO}_4)_3$
- b. $\text{Mg}(\text{HSO}_4)_2$

12. What is the formula mass for each compound?

- a. $\text{Co}(\text{HCO}_3)_2$
- b. LiHCO_3

Answers

1. The formula mass is the sum of the atomic masses of the atoms in the formula.
2. The subscript is distributed throughout the parentheses to determine the total number of atoms in the formula.
3.
 - a. 102.89 amu
 - b. 184.11 amu
 - c. 56.37 amu
4.
 - a. 74.55 amu
 - b. 95.21 amu
 - c. 103.28 amu
5.
 - a. 82.98 amu
 - b. 100.95 amu
 - c. 150.14 amu
6.
 - a. 34.83 amu
 - b. 134.87 amu
 - c. 51.79 amu
7.
 - a. 295.55 amu
 - b. 215.65 amu
 - c. 490.18 amu
 - d. 426.06 amu
8.
 - a. 152.00 amu
 - b. 68.00 amu

- c. 278.10 amu
- d. 349.00 amu

9.

- a. 238.03 amu
- b. 357.49 amu
- c. 156.08 amu
- d. 78.01 amu

10.

- a. 80.06 amu
- b. 294.20 amu
- c. 187.11 amu
- d. 84.01 amu

11.

- a. 318.19 amu
- b. 218.45 amu

12.

- a. 180.97 amu
- b. 67.96 amu

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4.7: Characteristics of Ionic Compounds

Learning Objectives

- To describe the physical properties of ionic compounds.

The figure below shows just a few examples of the color and brilliance of naturally occurring ionic crystals. The regular and orderly arrangement of ions in the crystal lattice is responsible for the various shapes of these crystals, while transition metal ions give rise to the colors.



Figure 4.7.1: In nature, the ordered arrangement of ionic solids gives rise to beautiful crystals. (A) Amethyst - a form of quartz, SiO_2 , whose purple color comes from iron ions. (B) Cinnabar - the primary ore of mercury is mercury (II) sulfide, HgS (C) Azurite - a copper mineral, $\text{Cu}_3(\text{CO}_3)_2(\text{OH})_2$. (D) Vanadinite - the primary ore of vanadium, $\text{Pb}_3(\text{VO}_4)_3\text{Cl}$.

Physical Properties of Ionic Compounds

Melting Points

Because of the many simultaneous attractions between cations and anions that occur, ionic crystal lattices are very strong. The process of melting an ionic compound requires the addition of large amounts of energy in order to break all of the ionic bonds in the crystal. For example, sodium chloride has a melting temperature of about 800°C . As a comparison, the molecular compound water melts at 0°C .

Shattering

Ionic compounds are generally hard, but **brittle**. Why? It takes a large amount of mechanical force, such as striking a crystal with a hammer, to force one layer of ions to shift relative to its neighbor. However, when that happens, it brings ions of the same charge next to each other (see below). The repulsive forces between like-

charged ions cause the crystal to shatter. When an ionic crystal breaks, it tends to do so along smooth planes because of the regular arrangement of the ions.

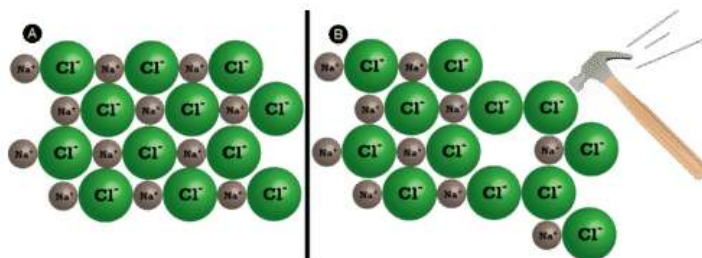


Figure 4.7.2: (A) The sodium chloride crystal is shown in two dimensions. (B) When struck by a hammer, the negatively-charged chloride ions are forced near each other and the repulsive force causes the crystal to shatter.

Solubility in water

A considerable high number of ionic compounds are soluble in water. Table salt, or sodium chloride (NaCl), the most common ionic compound, is soluble in water (360 g/L). Recall that NaCl is a salt crystal composed not of discrete NaCl molecules, but rather of an extended array of Na^+ and Cl^- ions bound together in three dimensions through electrostatic interactions. When a crystal of NaCl comes into contact with water, the water molecules interact with the Na^+ and Cl^- ions on the crystal's surface, as shown in the figure 4.7.3. As a consequence of the interaction between the Na^+ and Cl^- ions and the water molecules, the electrostatic interactions within the crystal are broken.



Figure 4.7.3: Dissolution of sodium chloride in water. Image by [Ahazard.sciencewriter](#), CC BY-SA 4.0, via Wikimedia Commons

Conductivity

Another characteristic property of ionic compounds is their **electrical conductivity**. The figure below shows three experiments in which two electrodes that are connected to a light bulb are placed in beakers containing three different substances.

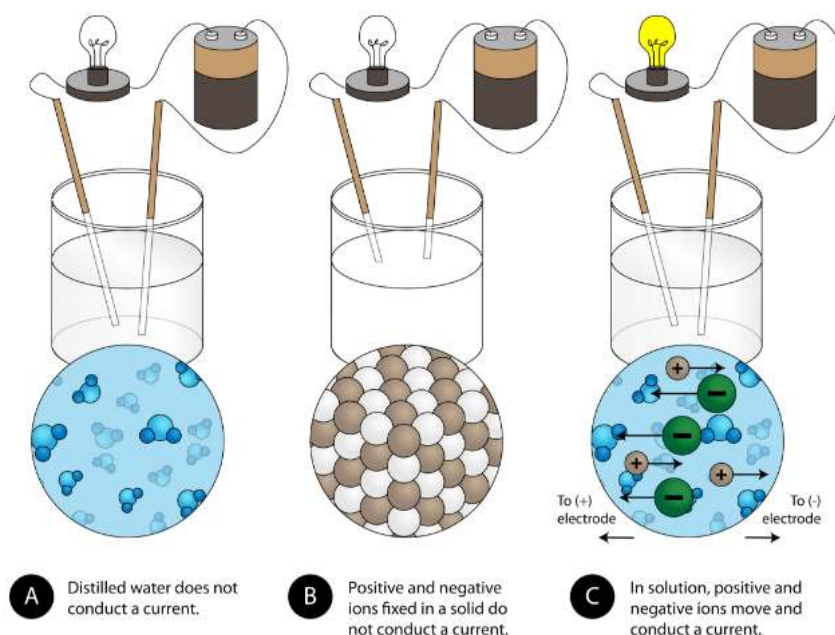


Figure 4.7.4. (A) Distilled water does not conduct electricity. (B) A solid ionic compound also does not conduct. (C) A water solution of an ionic compound conducts electricity well.

In the first beaker, distilled water does not conduct a current because water is a molecular compound. In the second beaker, solid sodium chloride also does not conduct a current. Despite being ionic and thus composed of charged particles, the solid crystal lattice does not allow the ions to move between the electrodes. Mobile charged particles are required for the circuit to be complete and the light bulb to light up. In the third beaker, the NaCl has been dissolved into the distilled water. Now the crystal lattice has been broken apart and the individual positive and negative ions can move. Cations move to one electrode, while anions move to the other, allowing electricity to flow (see figure below). Melting an ionic compound also frees the ions to conduct a current. Ionic compounds conduct an electric current when melted or dissolved in water.

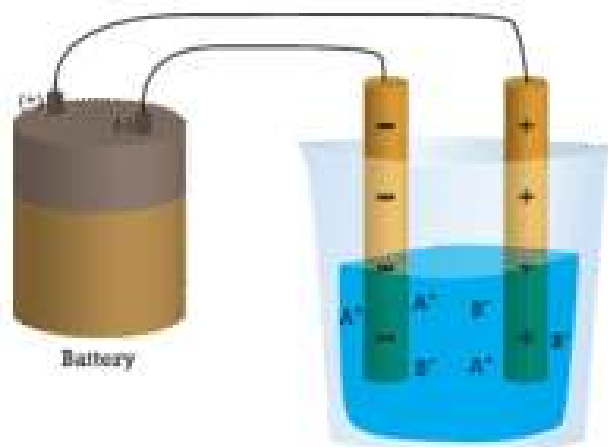
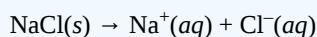


Figure 4.7.5: In an ionic solution, the A^+ ions migrate toward the negative electrode, while the B^- ions migrate toward the positive electrode.

Example 4.7.1

Write the dissociation equation of solid NaCl in water.

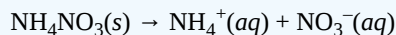
Solution



Exercise 4.7.1

Write the dissociation equation of solid NH_4NO_3 in water.

Answer



Key Takeaways

- Ionic compounds have high melting points.
- Ionic compounds are hard and brittle.
- Ionic compounds dissociate into ions when dissolved in water.
- Solutions of ionic compounds and melted ionic compounds conduct electricity, but solid materials do not.
- An ionic compound can be identified by its chemical formula: metal + nonmetal or polyatomic ions.

Contributors and Attributions

- CK-12 Foundation by Sharon Bewick, Richard Parsons, Therese Forsythe, Shonna Robinson, and Jean Dupon.

Exercises

1. Explain how the bonding in an ionic solid explains some of the properties of these solids.
2. Which type(s) of solid conduct(s) electricity in their liquid state but not in their solid state?
3. Based on chemical formula, identify which of the following is an ionic solid?

- a. Hg
- b. PH_3
- c. $\text{Ba}(\text{NO}_3)_2$

4. Based on chemical formula, identify which of the following will conduct electric current in the melted state but not in the solid state?

- a. H_2S
- b. Si
- c. KBr

5. Based on chemical formula, identify which of the following will dissociate into ions when dissolved in water?

- a. Co
- b. CO
- c. CaCl_2

6. Write dissociation equations for the following ionic compounds.

- a. $\text{LiF}(\text{s})$
- b. $\text{CuSO}_4(\text{s})$
- c. $\text{Na}_2\text{S}(\text{s})$
- d. $\text{MgBr}_2(\text{s})$

Answers

1. The oppositely charged ions are very strongly held together, so ionic crystals have **high melting points**. Ionic crystals are also **brittle** because any distortion of the crystal moves same-charged ions closer to each other, so they repel. When melted and dissolved in solution, the ions become free to **conduct electric current**.

2. Ionic compounds

3. $\text{Ba}(\text{NO}_3)_2$

4. KBr

5. CaCl_2

6.

- a. $\text{LiF}(\text{s}) \rightarrow \text{Li}^+(\text{aq}) + \text{F}^-(\text{aq})$
- b. $\text{CuSO}_4(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
- c. $\text{Na}_2\text{S}(\text{s}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{S}^{2-}(\text{aq})$
- d. $\text{MgBr}_2(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{Br}^-(\text{aq})$

Citations

Solubility of Ionic Compounds: Salts. (2019, June 5). Retrieved May 18, 2021, from <https://chem.libretexts.org/@go/page/52399>

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CHAPTER OVERVIEW

5: Chemical Bond II

- 5.1: Lewis Electron Dot Diagrams
- 5.2: Covalent Bonds
- 5.3: The Covalent structure of Polyatomic Ions
- 5.4: Exceptions to the Octet Rule
- 5.5: Resonance - Equivalent Lewis Structures for the Same Molecule
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- 5.13: Network Covalent Atomic Solids- Carbon and Silicates

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5.1: Lewis Electron Dot Diagrams

Learning Objective

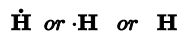
- Draw a Lewis electron dot diagram for an atom or a monatomic ion.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram, or a Lewis diagram, or a Lewis structure) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (The order in which the positions are used does not matter.) For example, the Lewis electron dot diagram for hydrogen is simply



Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:



The electron dot diagram for helium, with two valence electrons, is as follows:



By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1s subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:



Beryllium has two valence electrons in its 2s shell, so its electron dot diagram is like that of helium:



The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:



Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2s subshell and two in the 2p subshell. As usual, we will draw two dots together on one side, to represent the 2s electrons. However, conventionally, we draw the dots for the two p electrons on different sides. As such, the electron dot diagram for carbon is as follows:



With N, which has three p electrons, we put a single dot on each of the three remaining sides:



For oxygen, which has four p electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that a side has no more than two electrons.



Fluorine and neon have seven and eight dots, respectively:



With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the $n = 3$ shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

✓ Example 5.1.1

What is the Lewis electron dot diagram for each element?

- aluminum
- selenium

Solution

- The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3s electrons:



- The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the $n = 4$ shell, there are six electrons. Its electron dot diagram is as follows:



? Exercise 5.1.1

What is the Lewis electron dot diagram for each element?

- phosphorus
- argon

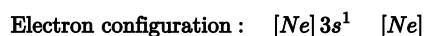
Answer



Elements in the same column of the periodic table have similar Lewis electron dot diagrams because they have the same valence shell electron configuration. The electron dot diagrams for the first column of elements are as follows:



Monatomic ions are atoms that have either lost (for cations) or gained (for anions) electrons. Electron dot diagrams for ions are the same as for atoms, except that some electrons have been removed for cations, while some electrons have been added for anions. Thus, in comparing the electron configurations and electron dot diagrams for the Na atom and the Na^+ ion, we note that the Na atom has a single valence electron in its Lewis diagram, while the Na^+ ion has lost that one valence electron:



Technically, the valence shell of the Na^+ ion is now the $n = 2$ shell, which has eight electrons in it. So why do we not put eight dots around Na^+ ? Conventionally, when we show electron dot diagrams for ions, we show the original valence shell of the atom, which in this case is the $n = 3$ shell and empty in the Na^+ ion.

In making cations, electrons are first lost from the *highest numbered shell*, not necessarily the last subshell filled. For example, in going from the neutral Fe atom to the Fe^{2+} ion, the Fe atom loses its two 4s electrons first, not its 3d electrons, despite the fact that the 3d subshell is the last subshell being filled. Thus, we have

Lewis dot diagram : $\text{Fe} : \text{Fe}^{2+}$

Electron configuration : $[\text{Ar}] 4s^2 3d^6$ $[\text{Ar}] 3d^6$

Anions have extra electrons when compared to the original atom. Here is a comparison of the Cl atom with the Cl^- ion:

Lewis dot diagram : $:\ddot{\text{Cl}}\cdot$ $:\ddot{\text{Cl}}:^-$

Electron configuration : $[\text{Ne}] 3s^2 3p^5$ $[\text{Ne}] 3s^2 3p^6$

✓ Example 5.1.2

What is the Lewis electron dot diagram for each ion?

- Ca^{2+}
- O^{2-}

Solution

- Having lost its two original valence electrons, the Lewis electron dot diagram is simply: Ca^{2+}
- The O^{2-} ion has gained two electrons in its valence shell, so its Lewis electron dot diagram is as follows:



? Exercise 5.1.2

The electronic configuration of the thallium ion is $6s^2 4f^{14} 5d^{10}$. What is the Lewis electron dot diagram for the Tl^+ ion?

Answer



Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.
- Lewis electron dot diagrams for ions have less (for cations) or more (for anions) dots than the corresponding atom.

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5.2: Covalent Bonds

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**.

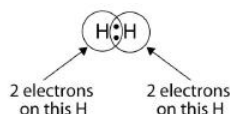
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1s subshell. Each H atom starts with a single electron in its valence shell:



The two H atoms can share their electrons:



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:



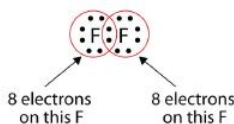
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:



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:

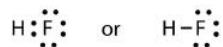


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three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



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

Example 5.2.1:

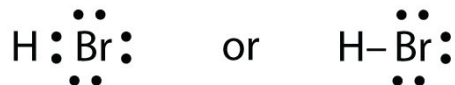
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

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 5.2.1

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

Answer



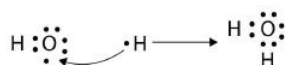
More than two atoms can participate in covalent bonding, although any given covalent bond will be between two atoms only. Consider H and O atoms:



The H and O atoms can share an electron to form a covalent bond:



The H atom has a complete valence shell. However, the O atom has only seven electrons around it, which is not a complete octet. We fix this by including a second H atom, whose single electron will make a second covalent bond with the O atom:



(It does not matter on what side the second H atom is positioned.) Now the O atom has a complete octet around it, and each H atom has two electrons, filling its valence shell. This is how a water molecule, H_2O , is made.

Example 5.2.2:

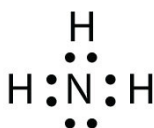
Use a Lewis electron dot diagram to show the covalent bonding in NH_3 .

Solution

The N atom has the following Lewis electron dot diagram:



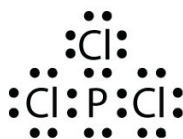
It has three unpaired electrons, each of which can make a covalent bond by sharing electrons with an H atom. The electron dot diagram of NH_3 is as follows:



Exercise 5.2.2

Use a Lewis electron dot diagram to show the covalent bonding in PCl_3 .

Answer:



There is a simple set of steps for determining the Lewis electron dot diagram of a simple molecule. First, you must identify the central atom and the surrounding atoms. The **central atom** is the atom in the center of the molecule, while the **surrounding atoms** are the atoms making bonds to the central atom. Usually, **the atom with the lowest electron affinity is the central atom** (except for H, which can only act as a surrounding atom). After the central and surrounding atoms have been identified, follow these steps:

1. Count the total number of valence electrons. Add extra if the species has negative charges and remove some for every positive charge on the species.
2. Write the central atom and surround it with the surrounding atoms.
3. Put a pair of electrons between the central atom and each surrounding atom.
4. Complete the octets around the surrounding atoms (except for H).
5. Put remaining electrons, if any, around the central atom.
6. Check that every atom has a full valence shell.

Let us try these steps to determine the electron dot diagram for BF_4^- . The B atom is the central atom, and the F atoms are the surrounding atoms. There is a negative sign on the species, so we have an extra electron to consider.

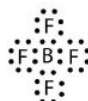
1. **Count the total number of electrons.** B has 3, each F has 7, and there is one extra electron: $3 + 7 + 7 + 7 + 7 + 1 = 32$.
2. **Write the central atom surrounded by surrounding atoms.**



3. **Put a pair of electrons between the central atom and each surrounding atom.** This uses up eight electrons, so we have $32 - 8 = 24$ electrons left.



4. **Complete the octets around the surrounding atoms (except for H).** This uses up 24 more electrons, leaving $24 - 24 = 0$ electrons left.



5. **Put remaining electrons, if any, around the central atom.** There are no additional electrons to add to the central atom.

6. **Check.** The B atom has eight electrons around it, as does each F atom. Each atom has a complete octet. This is a good Lewis electron dot diagram for BF_4^- .

Sometimes, however, these steps do not work. If we were to follow these steps for the compound formaldehyde (CH_2O), we would get the following:



The H and O atoms have the proper number of electrons, but the C atom has only six electrons around it, not the eight electrons for an octet. How do we fix this?

We fix this by recognizing that two atoms can share more than one pair of electrons. In the case of CH_2O , the O and C atoms share two pairs of electrons, with the following Lewis electron dot diagram as a result:



By circling the electrons around each atom, we can now see that the O and C atoms have octets, while each H atom has two electrons:



Each valence shell is full, so this is an acceptable Lewis electron dot diagram. If we were to use lines to represent the bonds, we would use two lines between the C and O atoms:



The bond between the C and O atoms is a **double bond** and represents two bonding pairs of electrons between the atoms. If using the rules for drawing Lewis electron dot diagrams do not work as written, a double bond may be required.

Example 5.2.3: Carbon Dioxide

What is the proper Lewis electron dot diagram for CO_2 ?

Solution

The central atom is a C atom, with O atoms as surrounding atoms. We have a total of $4 + 6 + 6 = 16$ valence electrons. Following the rules for Lewis electron dot diagrams for compounds gives us



The O atoms have complete octets around them, but the C atom has only four electrons around it. The way to solve this dilemma is to make a double bond between carbon and *each* O atom:



Each O atom still has eight electrons around it, but now the C atom also has a complete octet. This is an acceptable Lewis electron dot diagram for CO₂.

Exercise 5.2.3

What is the proper Lewis electron dot diagram for carbonyl sulfide (COS)?

Answer



It is also possible to have a **triple bond**, in which there are three pairs of electrons between two atoms. Good examples of this are elemental nitrogen (N₂) and acetylene (C₂H₂):



Acetylene is an interesting example of a molecule with two central atoms, which are both C atoms. Polyatomic ions are bonded together with covalent bonds. Because they are ions, however, they participate in ionic bonding with other ions. So both major types of bonding can occur at the same time.

Food and Drink App: Vitamins and Minerals

Vitamins are nutrients that our bodies need in small amounts but cannot synthesize; therefore, they must be obtained from the diet. The word *vitamin* comes from "vital amine" because it was once thought that all these compounds had an amine group (NH₂) in it. This is not actually true, but the name stuck anyway.

All vitamins are covalently bonded molecules. Most of them are commonly named with a letter, although all of them also have formal chemical names. Thus vitamin A is also called retinol, vitamin C is called ascorbic acid, and vitamin E is called tocopherol. There is no single vitamin B; there is a group of substances called the *B complex vitamins* that are all water soluble and participate in cell metabolism. If a diet is lacking in a vitamin, diseases such as scurvy or rickets develop. Luckily, all vitamins are available as supplements, so any dietary deficiency in a vitamin can be easily corrected.

A mineral is any chemical element other than carbon, hydrogen, oxygen, or nitrogen that is needed by the body. Minerals that the body needs in quantity include sodium, potassium, magnesium, calcium, phosphorus, sulfur, and chlorine. Essential minerals that the body needs in tiny quantities (so-called *trace elements*) include manganese, iron, cobalt, nickel, copper, zinc, molybdenum, selenium, and iodine. Minerals are also obtained from the diet. Interestingly, most minerals are consumed in ionic form, rather than as elements or from covalent molecules. Like vitamins, most minerals are available in pill form, so any deficiency can be compensated for by taking supplements.

Supplement Facts	
Serving Size 1 Tablet	
Each Tablet Contains	% Daily Value
Vitamin A 5000 I.U. (50% as Beta Carotene)	100%
Vitamin C 60 mg	100%
Vitamin D 400 I.U.	100%
Vitamin E 30 I.U.	100%
Vitamin K 25 mcg	31%
Thiamin (Vitamin B1) 1.5 mg	100%
Riboflavin (Vitamin B2) 1.7 mg	100%
Niacin 20 mg	100%
Vitamin B6 2 mg	100%
Folic Acid 400 mcg	100%
Vitamin B12 6 mcg	100%
Biotin 30 mcg	10%
Pantothenic Acid 10 mg	100%
Calcium 160 mg	16%
Iron 18 mg	100%
Phosphorus 110 mg	11%
Iodine 150 mcg	100%
Magnesium 100 mg	25%
Zinc 15 mg	100%
Selenium 20 mcg	29%
Copper 2 mg	100%
Manganese 2 mg	100%
Chromium 120 mcg	100%
Molybdenum 75 mcg	100%
Chloride 72 mg	2%
Potassium 80 mg	2%
Boron 150 mcg	*
Nickel 5 mcg	*
Silicon 2 mg	*
Tin 10 mcg	*
Vanadium 10 mcg	*
Lutein 250 mcg	*
* Daily Value (DV) not established	

Figure 5.2.1 :Vitamins and Mineral supplements. Every entry down through pantothenic acid is a vitamin, and everything from calcium and below is a mineral.

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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5.3: The Covalent structure of Polyatomic Ions

Our discussion of **ionic compounds** was confined to monatomic ions. However, more complex ions, containing several atoms *covalently bonded* to one another, but having a positive or negative charge, occur quite frequently in chemistry. The charge arises because the total number of valence electrons from the atoms cannot produce a stable structure. With one or more electrons added or removed, a stable structure results. Well-known examples of such **polyatomic ions** are the sulfate ion (SO_4^{2-}),

--	--

the hydroxide ion (OH^-),

--	--

the hydronium ion (H_3O^+),

--	--

and the ammonium ion (NH_4^+).

--	--

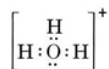
The atoms in these ions are joined together by covalent electron-pair bonds, and we can draw Lewis structures for the ions just as we can for molecules. The only difference is that the number of electrons in the ion does not exactly balance the sum of the nuclear charges. Either there are too many electrons, in which case we have an anion, or too few, in which case we have a cation.

Consider, for example, the hydroxide ion (OH^-) for which the Lewis structure is



O shares a pair of electrons with H. Three pairs of unpaired electrons is also shown. The entire O H molecule is enclosed in a square bracket with a superscript of negative sign.

A neutral molecule containing one O and one H atom would contain only seven electrons, six from O and one from H. The hydroxide ion, though, contains an *octet* of electrons, *one more* than the neutral molecule. The hydroxide ion must thus carry a single negative charge. In order to draw the Lewis structure for a given ion, we must first determine how many valence electrons are involved. Suppose the structure of H_3O^+ is required. The total number of electrons is obtained by adding the valence electrons for each atom, $6 + 1 + 1 + 1 = 9$ electrons. We must now subtract 1 electron since the species under consideration is not H_3O but H_3O^+ . The total number of electrons is thus $9 - 1 = 8$. Since this is an octet of electrons, we can place them all around the O atom. The final structure then follows very easily:

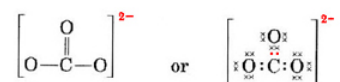


A central O shares three pairs of electrons with three H on the left top and right side. The bottom pair of electron from O is unpaired. The entire O H 3 molecule is enclosed in a square bracket which has a superscript positive sign.

In more complicated cases it is often useful to calculate the number of *shared electron pairs* before drawing a Lewis structure. This is particularly true when the ion in question is an *oxyanion* (i.e., a central atom is surrounded by several O atoms). A well-known oxyanion is the carbonate ion, which has the formula CO_3^{2-} . (Note that the central atom C is written first, as was done earlier for molecules.) The total number of valence electrons available in CO_3^{2-} is

$$4(\text{for C}) + 3 \times 6(\text{for O}) + 2(\text{for the } -2 \text{ charge}) = 24$$

We must distribute these electrons over 4 atoms, giving each an octet, a requirement of $4 \times 8 = 32$ electrons. This means that $32 - 24 = 8$ electrons need to be counted twice for octet purposes; i.e., 8 electrons are shared. The ion thus contains *four* electron-pair bonds. Presumably the C atom is double-bonded to one of the O's and singly bonded to the other two:



A central C is double bonded to one O and single bonded to two other O. The entire structure is enclosed in a square bracket with superscript of negative 2.

In this diagram the 4C electrons have been represented by dots, the 18 O electrons by $\times$'s, and the 2 extra electrons by colored dots, for purposes of easy reference. Real electrons do not carry labels like this; they are all the same.

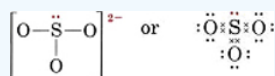
There is a serious objection to the Lewis structure just drawn. How do the electrons know which oxygen atom to single out and form a double bond with, since there is otherwise nothing to differentiate the oxygens? The answer is that they do not. To explain the bonding in the CO_3^{2-} ion and some other molecules requires an extension of the Lewis theory. We pursue this matter further when we discuss [resonance](#). Now we end with an example.

✓ Example 5.3.1 : Lewis Structure

Draw a Lewis structure for the sulfite ion, SO_3^{2-} .

Solution The safest method here is to count electrons. The total number of valence electrons available is

$6(\text{for S}) + 3 \times 6(\text{for O}) + 2(\text{for the charge}) = 26$ To make four octets for the four atoms would require 32 electrons, and so the difference, $32 - 26 = 6$, gives the number of shared electrons. There are thus only *three* electron-pair bonds in the ion. The central S atom must be linked by a single bond to each O atom.



A central S is single bonded to three O. There is a lone pair from S that is not bonded to any other atom. As a result, the overall charge of the molecule is negative 2.

Note that each of the S—O bonds is coordinate covalent.

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5.4: Exceptions to the Octet Rule



Figure 5.4.1 (Public Domain; the U.S. Internal Revenue Service via Wikipedia)

Are rules always followed?

Every spring, millions of Americans file their income tax forms. The different rules determine how much tax a person pays. There are also exceptions to the rules. You pay less tax if you are married and/or have children. There are certain limits on how much money you can make before paying taxes. The rule is that you pay taxes, but there are also exceptions based on your personal situation. The bonding rules for molecules are generally applicable, but there are some exceptions allowed.

Exceptions to the Octet Rule

As the saying goes, all rules are made to be broken. When it comes to the octet rule, this is true. Exceptions to the octet rule fall into one of three categories: (1) an **incomplete octet**, (2) **odd-electron molecules**, and (3) an **expanded octet**.

Incomplete Octet

In some compounds, the number of electrons surrounding the central atom in a stable molecule is fewer than eight. Beryllium is an alkaline earth metal and so may be expected to form ionic bonds. However, its very small size and somewhat higher ionization energy compared to other metals actually leads beryllium to form primarily molecular compounds. Since beryllium only has two valence electrons, it does not typically attain an octet through sharing of electrons. The Lewis structure of gaseous beryllium hydride (BeH_2) consists of two single covalent bonds between Be and H (see figure below).

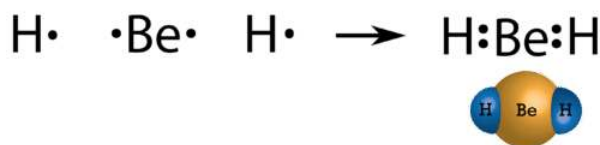


Figure 5.4.2: Beryllium hydride. (CC BY-NC 3.0; Christopher Auyeung, Joy Sheng via CK-12 Foundation)

Boron and aluminum, with three valence electrons, also tend to form covalent compounds with an incomplete octet. The central boron atom in boron trichloride (BCl_3) has six valence electrons, as shown in the figure below.

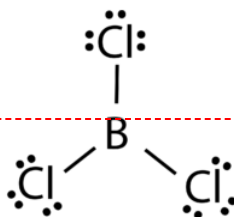


Figure 5.4.3: Boron trichloride. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Odd-Electron Molecules

There are a number of molecules whose total number of valence electrons is an odd number. It is not possible for all of the atoms in such a molecule to satisfy the octet rule. An example is nitrogen dioxide (NO_2). Each oxygen atom contributes six valence electrons and the nitrogen atom contributes five for a total of seventeen. The Lewis structure for NO_2 appears in the figure below.

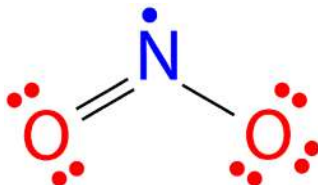


Figure 5.4.4: The Lewis structure for nitrogen dioxide, an odd electron molecule. (CC BY-NC-SA; Todd Helmenstine via about.com)

Expanded Octets

Atoms of the second period cannot have more than eight valence electrons around the central atom. However, atoms of the third period and beyond are capable of exceeding the octet rule by having more than eight electrons around the central atom. Starting with the third period, the d sublevel becomes available, so it is possible to use these orbitals in bonding, resulting in an expanded octet.

Phosphorus and sulfur are two elements that react with halogen elements and make stable compounds with expanded octets. In phosphorus pentachloride, the central phosphorus atom makes five single bonds to chlorine atoms and, as a result, has ten electrons surrounding it (see figure below). In sulfur hexafluoride, the central sulfur atom has twelve electrons from its six bonds to fluorine atoms (see figure below).

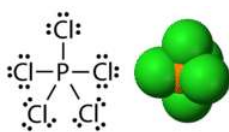


Figure 5.4.5: Phosphorus pentachloride. Left image: Lewis structure. Right image: molecular model. (CC BY-NC 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: <http://commons.wikimedia.org/wiki/File:Phosphorus-pentachloride-3D-vdW.png>)

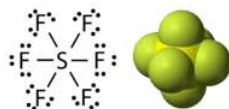


Figure 5.4.6: Sulfur hexafluoride. Left image: Lewis structure. Right image: molecular model. (CC BY-NC 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: <http://commons.wikimedia.org/wiki/File:Sulfur-hexafluoride-3D-vdW.png>)



Summary

- Exceptions exist to the rules for covalent bonding
- These exceptions apply to atoms whose electrons will not accommodate the normal octet rule.

Review

1. What is an incomplete octet?
2. What is an odd-electron molecule?
3. Why are there extra electrons in the expanded octet?

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5.5: Resonance - Equivalent Lewis Structures for the Same Molecule

Learning Objectives

- Explain the concept of resonance and how it works with within molecules.

Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone (O_3) molecule in Figure 5.5.1. There are a total of 18 electrons in the structure and so the following two structures are possible.



Figure 5.5.1: Resonance forms of ozone. Note the use of the double-headed arrow.

The structure on the left (5.5.1) can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as O_3 consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the O_3 and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an $\text{O}-\text{O}$ single bond and a double bond.

Resonance is the use of two or more Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms are best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (Figure 5.5.2).



Figure 5.5.2: "Half-bond" model of ozone molecule. This is a better description of the electronic structure of ozone than either of the resonance structures in Figure 5.5.1.

Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, NO_3^- in Figure 5.5.3.

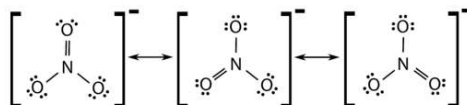


Figure 5.5.3: Resonance structure of nitrate anion.

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

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5.6: Covalent Compounds - Formulas and Names

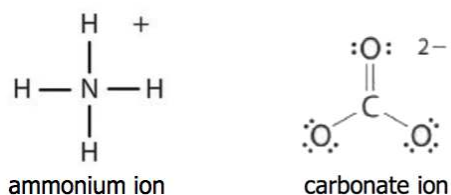
Learning Objectives

- Identify covalent and ionic compounds.
- Determine the chemical formula of a simple covalent compound from its name.
- Determine the name of a simple covalent compound from its chemical formula.

COVALENT AND IONIC COMPOUNDS

What elements make covalent bonds? Covalent bonds form when two or more nonmetals combine. For example, both hydrogen and oxygen are nonmetals, and when they combine to make water, they do so by forming covalent bonds. Compounds that involve a metal binding with either a non-metal will display ionic bonding. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal).

A polyatomic ion is an ion composed of two or more atoms that have a charge as a group (poly = many). The ammonium ion (see figure below) consists of one nitrogen atom and four hydrogen atoms. Together, they comprise a single ion with a 1+ charge and a formula of NH_4^+ . The carbonate ion (see figure below) consists of one carbon atom and three oxygen atoms and carries an overall charge of 2-. The formula of the carbonate ion is CO_3^{2-} .



The atoms of a polyatomic ion are tightly bonded together and so the entire ion behaves as a single unit. Several examples are found in Table 3.3.2. Nonmetal atoms in polyatomic ions are joined by covalent bonds, but the ion as a whole participates in ionic bonding. For example, ammonium chloride (NH_4Cl) has ionic bonds between a polyatomic ion, NH_4^+ , and Cl^- ions, but within the ammonium ion, the nitrogen and hydrogen atoms are connected by covalent bonds.

Example 5.6.1

Is each compound formed from ionic bonds, covalent bonds, or both?

- Na_2O
- Na_3PO_4
- N_2O_4

Answer a

The elements in Na_2O are a metal and a nonmetal, which form ionic bonds.

Answer b

Because sodium is a metal and we recognize the formula for the phosphate ion, we know that this compound is ionic. However, within the polyatomic phosphate ion, the atoms are held together by covalent bonds.

Answer c

The elements in N_2O_4 are both nonmetals, rather than a metal and a nonmetal. Therefore, the atoms form covalent bonds.

Exercise 5.6.1

Is each compound are formed from ionic bonds, covalent bonds, or both?

- $\text{Ba}(\text{OH})_2$

- b. F_2
c. PCl_3

Answer a:

both

Answer b:

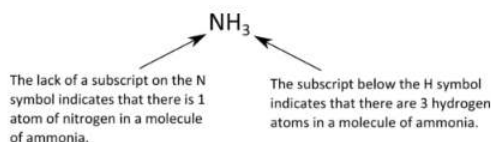
covalent

Answer c:

covalent

MOLECULAR FORMULAS

The chemical formulas for covalent compounds are referred to as **molecular formulas** because these compounds exist as separate, discrete molecules. Typically, a molecular formula begins with the nonmetal that is closest to the lower left corner of the periodic table, except that hydrogen is almost never written first (H_2O is the prominent exception). Then the other nonmetal symbols are listed. Numerical subscripts are used if there is more than one of a particular atom. For example, we have already seen CH_4 , the molecular formula for methane. Below is the molecular formula of ammonia, NH_3 .



NAMING COVALENT COMPOUNDS

Naming *binary* (two-element) covalent compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 5.6.1 lists these numerical prefixes. Normally, no prefix is added to the first element's name if there is only one atom of the first element in a molecule. If the second element is oxygen, the trailing vowel is usually omitted from the end of a polysyllabic prefix but not a monosyllabic one (that is, we would say "monoxide" rather than "monooxide" and "trioxide" rather than "troxide").

Table 5.6.1: Numerical Prefixes for Naming Binary Covalent Compounds

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

*This prefix is not used for the first element's name.

Let us practice by naming the compound whose molecular formula is CCl_4 . The name begins with the name of the first element—carbon. The second element, *chlorine*, becomes *chloride*, and we attach the correct numerical prefix (“tetra-”) to indicate that the molecule contains four chlorine atoms. Putting these pieces together gives the name *carbon tetrachloride* for this compound.

Example 5.6.2

Write the molecular formula for each compound.

- a. chlorine trifluoride
- b. phosphorus pentachloride
- c. sulfur dioxide
- d. dinitrogen pentoxide

Solution

If there is no numerical prefix on the first element’s name, we can assume that there is only one atom of that element in a molecule.

- a. ClF_3
- b. PCl_5
- c. SO_2
- d. N_2O_5 (The *di-* prefix on nitrogen indicates that two nitrogen atoms are present.)

Exercise 5.6.2

Write the molecular formula for each compound.

- a. nitrogen dioxide
- b. dioxygen difluoride
- c. sulfur hexafluoride
- d. selenium monoxide

Answer a:

- a. NO_2

Answer b:

- O_2F_2

Answer c:

- SF_6

Answer d:

- SeO

Because it is so unreactive, sulfur hexafluoride is used as a spark suppressant in electrical devices such as transformers.

Example 5.6.3

Write the name for each compound.

- a. BrF_5
- b. S_2F_2
- c. CO

Solution

- a. bromine pentafluoride
- b. disulfur difluoride

c. carbon monoxide

Exercise 5.6.3

Write the name for each compound.

- a. CF_4
- b. SeCl_2
- c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O : water
- NH_3 : ammonia
- CH_4 : methane

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system that we will introduce in Section 4.6 "Introduction to Organic Chemistry".

Concept Review Exercises

1. How do you recognize a covalent compound?
2. What are the rules for writing the molecular formula of a simple covalent compound?
3. What are the rules for naming a simple covalent compound?

Answers

1. A covalent compound is usually composed of two or more nonmetal elements.
2. It is just like an ionic compound except that the element further down and to the left on the periodic table is listed first and is named with the element name.
3. Name the first element first and then the second element by using the stem of the element name plus the suffix -ide. Use numerical prefixes if there is more than one atom of the first element; always use numerical prefixes for the number of atoms of the second element.

Key Takeaways

- The chemical formula of a simple covalent compound can be determined from its name.
- The name of a simple covalent compound can be determined from its chemical formula.

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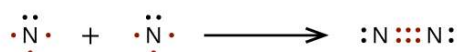
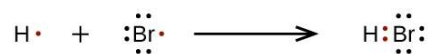
5.7: Multiple Covalent Bonds

Learning Objectives

- To draw Lewis structures.
- To recognize molecules that are likely to have multiple covalent bonds.

DRAWING LEWIS STRUCTURES

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

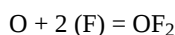


For more complicated molecules and molecular ions, it is helpful to follow the step-by-step procedure outlined here:

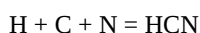
- Determine the total number of valence (outer shell) electrons among all the atoms. For cations, subtract one electron for each positive charge. For anions, add one electron for each negative charge.
- Draw a skeleton structure of the molecule or ion, arranging the atoms around a central atom. (Generally,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

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

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



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



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

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



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

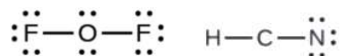
- In OF_2 , six electrons are placed on each F.

- In HCN, six electrons placed on N



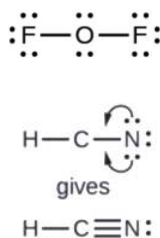
4. Place all remaining electrons on the central atom.

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



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

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

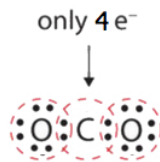


H dash C dash N with 6 dots around the N. 4 of the dots are pointing to the dash connecting the C and the N. This turns the formula into H dash C and three lines connecting the C and N together, with two dots on the outside of the N.

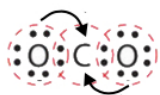
Finally, check to see if the total number of valence electrons are present in the Lewis structure. And then, inspect if the H atom has 2 electrons surrounding it and if each of the main group atoms is surrounded by 8 electrons.

MULTIPLE BONDS

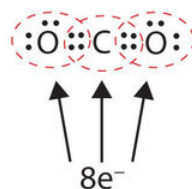
In many molecules, the octet rule would not be satisfied if each pair of bonded atoms shares only two electrons. Review HCN in Step 5 above. Another example is carbon dioxide (CO_2). CO_2 has a total valence of $4e^- + (2 \times 6e^-) = 16e^-$. Following steps 1 to 4, we draw the following:



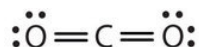
This does not give the carbon atom a complete octet; only four electrons are in its valence shell. This arrangement of shared electrons is far from satisfactory.



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



In this arrangement, the carbon atom shares four electrons (two pairs) with the oxygen atom on the left and four electrons with the oxygen atom on the right. There are now eight electrons around each atom. Two pairs of electrons shared between two atoms make a double bond between the atoms, which is represented by a double dash:



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



✓ Example 5.7.1

Draw the Lewis diagram for each molecule.

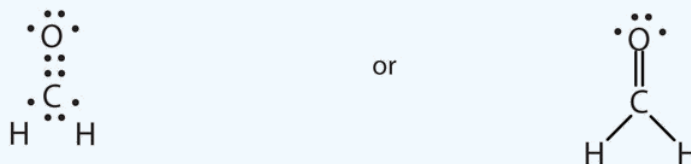
1. N_2
2. CH_2O (The carbon atom is the central atom.) One application of CH_2O , also called formaldehyde, is the preservation of biological specimens. Aqueous solutions of CH_2O are called formalin and have a sharp, characteristic (pungent) odor.

Solution

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



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



? Exercise 5.7.1

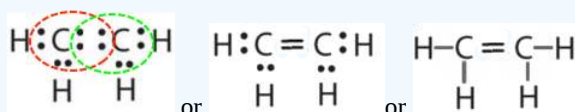
Draw the Lewis diagram for each molecule.

- a. O_2
- b. C_2H_4

Answer a:

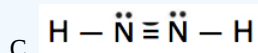
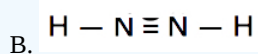
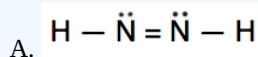


Answer b:



✓ Example 5.7.2

Which is the correct Lewis structure for N_2H_2 ?



Solution

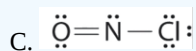
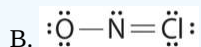
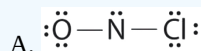
Lewis structure A is the correct answer. It has a total of $(2 \times 5e^-) + (2 \times 1e^-) = 12e^-$. Each of the N atoms satisfy the octet requirement and the H atoms follow the duet rule.

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

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

5.7.2 Exercise

Which is the correct Lewis structure for NOCl ?



Answer

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

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

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

Key Takeaways

- A Lewis structure shows the bonding and nonbonding electrons around individual atoms in a molecule.
- Some molecules must have multiple covalent bonds between atoms to satisfy the octet rule.
- A double bond contains four electrons and a triple bond contains six electrons.

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5.8: Characteristics of Covalent compounds

Learning Objectives

- Recognize bond characteristics of covalent compounds: bond length and bond energy

Bond Length and Bond Energy

Bond length is the separation distance between two nuclei forming a covalent bond. The bond length is usually measured in meters or picometers ($1 \text{ pm} = 1 \times 10^{-12} \text{ m}$). For example, the covalent bond in the hydrogen molecule (H_2) has a certain length (about $7.4 \times 10^{-11} \text{ m}$). Other covalent bonds also have known bond lengths, which are dependent on both the identities of the atoms in the bond and whether the bonds are single, double, or triple bonds. Table 5.8.1 lists the approximate bond lengths for some single covalent bonds. The exact bond length may vary depending on the identity of the molecule but will be close to the value given in the table.

Bond energy is defined as the energy required to break a particular bond in a molecule in the gas phase, expressed in kilojoules per mole of bonds (kJ/mol). This value depends on the identity of the bonded atoms and the nature of the molecule (for example, the energy required to break the O-H bond in H_2O is different than in H_2O_2). Except for diatomic molecules, the bond energies listed in Table 5.8.1 are *average* values for all bonds of a given type in a range of molecules.

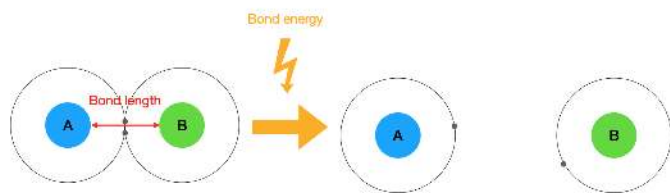


Figure 5.8.1: representation of bond length and bond energy between two atoms A and B

Table 5.8.1: Approximate Bond Lengths and Bond Energies of Some Single Bonds

Bond	Length ($\times 10^{-12} \text{ m}$)	Bond Energy (kJ/mol)
H-H	74	436
H-C	110	414
H-N	100	386
H-O	97	459
H-I	161	295
C-C	154	350
C-N	147	305
C-O	143	358
N-N	145	167
O-O	145	142

Table 5.8.2 compares the lengths and bond energies of single covalent bonds with those of double and triple bonds between the same atoms. Without exception, as the number of covalent bonds between two atoms increases, the bond length decreases and the

bond energy increases.

By comparing both tables, we can conclude that the more energy needed to break a bond, the higher its stability, and the shorter the bond length between the atoms

Table 5.8.2: Comparison of Bond Lengths and Bond Energies for Single and Multiple Bonds

Bond	Length ($\times 10^{-12} m$)	Bond Energy (kJ/mol)
C–C	154	350
C=C	134	614
C≡C	120	839
C–N	147	305
C=N	128	615
C≡N	116	891
C–O	143	358
C=O	120	745
C≡O	113	1072
N–N	145	167
N=N	123	918
N≡N	110	941
O–O	145	142
O=O	121	495

Physical properties of covalent compounds

Covalent compounds form discrete molecules that can exist independently from each other, in contrast to the crystal structure of ionic compounds in which formula units cannot exist individually but only as a part of a lattice of ions. Therefore, **the physical properties of covalent molecules depend heavily on the nature of their interaction with other molecules (intermolecular forces)**. Depending on the nature of these intermolecular interactions, covalent compounds may exist as a solid, a liquid, or a gas at room temperature and normal atmospheric pressure. As a general rule, we can say that covalent compounds have lower melting and boiling points compared to ionic compounds with similar molecular mass.

Since there are no ions in their structure, covalent compounds do not exhibit any electrical conductivity, either in pure form or when dissolved in water. There are however some exceptions that we will discuss in CHE 104. These exceptions involve ionizable covalent compounds such as acids (these molecules dissociate in water forming ions, and therefore, solutions of ionizable covalent compounds can conduct the electricity).

Key Takeaways

- Covalent bonds between different atoms have different bond lengths and bond energies.
- The stronger the covalent bond, the higher its bond energy and the shorter its bond distance
- As a general rule, covalent compounds have lower melting and boiling points compared to ionic compounds with similar molecular mass.
- The physical properties of covalent molecules depend heavily on the nature of their interaction with other molecules

Citations and Attributions

- Characteristics of Covalent Bonds. (2021, March 21). Retrieved May 19, 2021, from <https://chem.libretexts.org/@go/page/16131>
- Bond Energies. (2020, August 21). Retrieved May 19, 2021, from <https://chem.libretexts.org/@go/page/1981>

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5.9: Molecular Geometry

Learning Objectives

- Predict the structures of small molecules using valence shell electron pair repulsion (VSEPR) theory

Thus far, we have used two-dimensional Lewis structures to represent molecules. However, molecular structure is actually three-dimensional, and it is important to be able to describe molecular bonds in terms of their distances, angles, and relative arrangements in space (Figure 5.9.1). A bond angle is the angle between any two bonds that include a common atom, usually measured in degrees. A bond distance (or bond length) is the distance between the nuclei of two bonded atoms along the straight line joining the nuclei. Bond distances are measured in Ångstroms ($1 \text{ Å} = 10^{-10} \text{ m}$) or picometers ($1 \text{ pm} = 10^{-12} \text{ m}$, $100 \text{ pm} = 1 \text{ Å}$).

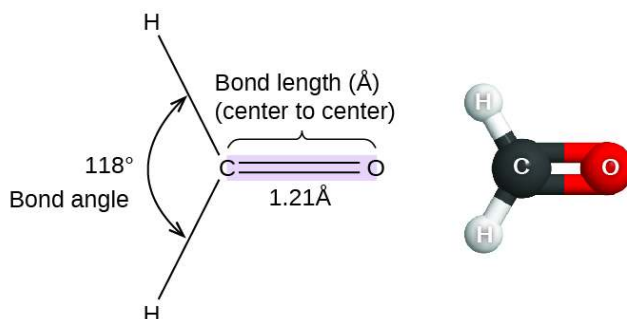


Figure 5.9.1: Bond distances (lengths) and angles are shown for the formaldehyde molecule, H_2CO .

VSEPR Theory

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

VSEPR theory predicts the arrangement of electron pairs around each central atom and, usually, the correct arrangement of atoms in a molecule.

As a simple example of VSEPR theory, let us predict the structure of a gaseous BeF_2 molecule. The Lewis structure of BeF_2 (Figure 5.9.2) shows only two electron pairs around the central beryllium atom. With two bonds and no lone pairs of electrons on the central atom, *the bonds are as far apart as possible*, and the electrostatic repulsion between these regions of high electron density is reduced to a minimum when they are on opposite sides of the central atom. The bond angle is 180° (Figure 5.9.2).

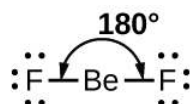


Figure 5.9.2: The BeF_2 molecule adopts a linear structure in which the two bonds are as far apart as possible, on opposite sides of the Be atom.

Figure 5.9.3 illustrates this and other electron-pair geometries that minimize the repulsions among regions of high electron density (bonds and/or lone pairs). Two "electron clouds" or regions of electron density around a central atom in a molecule form a linear geometry; three electron clouds form a trigonal planar geometry; four electron clouds form a tetrahedral geometry; five electron clouds form a trigonal bipyramidal geometry; and six electron clouds form an octahedral geometry (Notice that five and six electron clouds correspond to situations of **octet rule expansion** around the central atom).

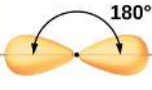
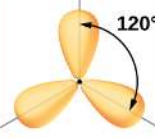
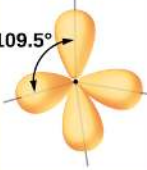
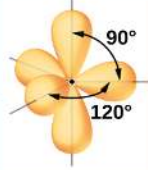
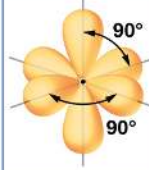
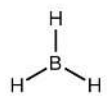
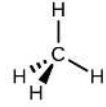
Number of regions	Two regions of high electron density (bonds and/or unshared pairs)	Three regions of high electron density (bonds and/or unshared pairs)	Four regions of high electron density (bonds and/or unshared pairs)	Five regions of high electron density (bonds and/or unshared pairs)	Six regions of high electron density (bonds and/or unshared pairs)
Spatial arrangement					
Line-dash-wedge notation	$\text{H}-\text{Be}-\text{H}$				
Electron pair geometry	Linear; 180° angle	Trigonal planar; all angles 120°	Tetrahedral; all angles 109.5°	Trigonal bipyramidal; angles of 90° or 120° An attached atom may be equatorial (in the plane of the triangle) or axial (above or below the plane of the triangle).	Octahedral; all angles 90° or 180°

Figure 5.9.3: The basic electron-pair geometries predicted by VSEPR theory maximize the space around any region of electron density (bonds or lone pairs).

Electron-pair Geometry versus Molecular Structure

It is important to note that electron-pair geometry around a central atom is *not* the same thing as its molecular structure. The electron-pair geometries shown in Figure 5.9.3 describe all regions where electrons are located, bonds as well as lone pairs. Molecular structure describes the location of the *atoms*, not the electrons.

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

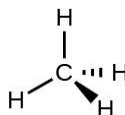


Figure 5.9.4: The molecular structure of the methane molecule, CH_4 , is shown with a tetrahedral arrangement of the hydrogen atoms. VSEPR structures like this one are often drawn using the wedge and dash notation, in which solid lines represent bonds in the plane of the page, solid wedges represent bonds coming up out of the plane, and dashed lines represent bonds going down into the plane.

For example, the methane molecule, CH_4 , which is the major component of natural gas, has four bonding pairs of electrons around the central carbon atom; the electron-pair geometry is tetrahedral, as is the molecular structure (Figure 5.9.4). On the other hand, the ammonia molecule, NH_3 , also has four electron pairs associated with the nitrogen atom, and thus has a tetrahedral electron-pair

geometry. One of these regions, however, is a lone pair, which is not included in the molecular structure, and this lone pair influences the shape of the molecule (Figure 5.9.5).

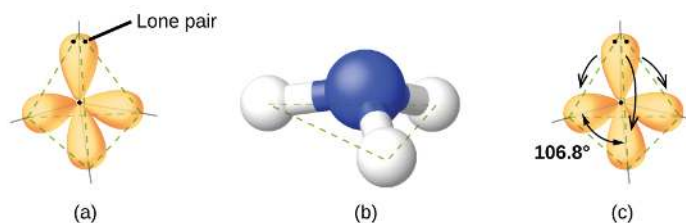


Figure 5.9.5: (a) The electron-pair geometry for the ammonia molecule is tetrahedral with one lone pair and three single bonds. (b) The trigonal pyramidal molecular structure is determined from the electron-pair geometry. (c) The actual bond angles deviate slightly from the idealized angles because the lone pair takes up a larger region of space than do the single bonds, causing the HNH angle to be slightly smaller than 109.5° .

Small distortions from the ideal angles in Figure 5.9.5 can result from differences in repulsion between various regions of electron density. VSEPR theory predicts these distortions by establishing that a lone pair of electrons occupies a larger region of space than the electrons in bonding pair. This is because the movement of a bonding electron pair is limited to the region between the atoms forming the bond, while the lone pair can move more freely.

In the ammonia molecule, the three hydrogen atoms attached to the central nitrogen are not arranged in a flat, trigonal planar molecular structure, but rather in a three-dimensional trigonal pyramid (Figure 5.9.6) with the nitrogen atom at the apex and the three hydrogen atoms forming the base. The ideal bond angles in a trigonal pyramid are based on the tetrahedral electron pair geometry. Again, **there are slight deviations from the ideal because lone pairs occupy larger regions of space than do bonding electrons**. The H–N–H bond angles in NH_3 are slightly smaller than the 109.5° angle in a regular tetrahedron (Figure 5.9.6) because the lone pair-bonding pair repulsion is greater than the bonding pair-bonding pair repulsion. The ideal molecular structures are predicted based on the electron-pair geometries for various combinations of lone pairs and bonding pairs.

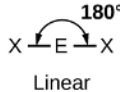
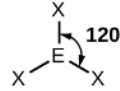
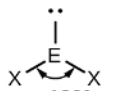
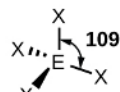
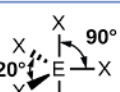
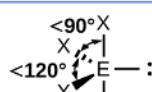
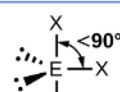
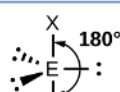
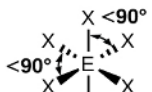
Number of electron pairs	Electron pair geometries: 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal planar	 Bent or angular			
4	 Tetrahedral	 Trigonal pyramid	 Bent or angular		
5	 Trigonal bipyramid	 Sawhorse or seesaw	 T-shape	 Linear	
6	 Octahedral	 Square pyramid	 Square planar	 T-shape	 Linear

Figure 5.9.6: The molecular structures are identical to the electron-pair geometries when there are no lone pairs present (first column). For a particular number of electron pairs (row), the molecular structures for one or more lone pairs are determined based on modifications of the corresponding electron-pair geometry.

Predicting Electron Pair Geometry and Molecular Structure

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

1. Write the Lewis structure of the molecule or polyatomic ion.
2. Count the number of regions of electron density or electron clouds around the central atom. A single, double, or triple bond counts as one region of electron density.
3. Identify the electron-pair geometry based on the number of regions of electron clouds: linear, trigonal planar, tetrahedral, trigonal bipyramidal, or octahedral (Figure 5.9.7, first column).
4. Use the number of lone pairs to determine the molecular structure (Figure 5.9.7). If more than one arrangement of lone pairs and chemical bonds is possible, choose the one that will minimize repulsions, remembering that lone pairs occupy more space than multiple bonds, which occupy more space than single bonds. In trigonal bipyramidal arrangements, repulsion is minimized when every lone pair is in an equatorial position. In an octahedral arrangement with two lone pairs, repulsion is minimized when the lone pairs are on opposite sides of the central atom.

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

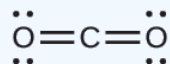
Example 5.9.1: Predicting Electron-pair Geometry and Molecular Structure

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

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

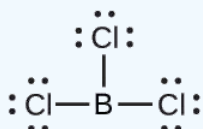
Solution

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

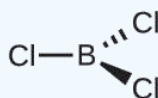


This shows us two regions of high electron density around the carbon atom—each double bond counts as one region, and there are no lone pairs on the carbon atom. Using VSEPR theory, we predict that the two regions of electron density arrange themselves on opposite sides of the central atom with a bond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

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



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



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

Exercise 5.9.1

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

Answer

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

Example 5.9.2: Predicting Electron-pair Geometry and Molecular Structure

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

Solution

We write the Lewis structure of NH_4^+ as:

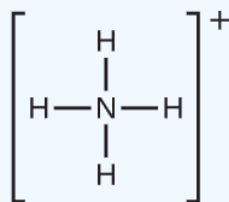


Figure 5.9.7).

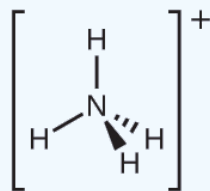


Figure 5.9.8: The ammonium ion displays a tetrahedral electron-pair geometry as well as a tetrahedral molecular structure.

Exercise 5.9.2

Identify a molecule with trigonal bipyramidal molecular structure.

Answer

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

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

Example 5.9.3: Lone Pairs on the Central Atom

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

Solution

The Lewis structure of H_2O indicates that there are four regions of high electron density around the oxygen atom: two lone pairs and two chemical bonds:

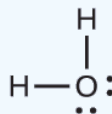


Figure 5.9.9. Thus, the electron-pair geometry is tetrahedral and the molecular structure is bent with an angle slightly less than 109.5° . In fact, the bond angle is 104.5° .

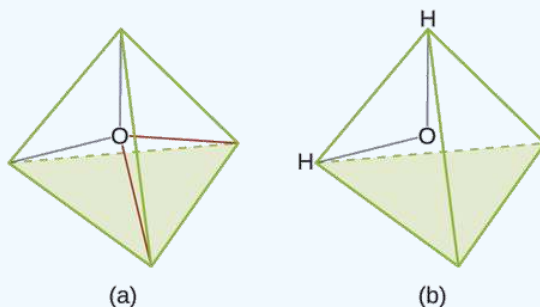


Figure 5.9.9: (a) H_2O has four regions of electron density around the central atom, so it has a tetrahedral electron-pair geometry. (b) Two of the electron regions are lone pairs, so the molecular structure is bent.

Exercise 5.9.3

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

Answer

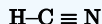
electron pair geometry: tetrahedral; molecular structure: trigonal pyramidal

USING MOLECULAR SHAPE SIMULATIONS

Example 5.9.6: Molecular Simulation

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

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



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

Solution

The molecular structure is linear.

Summary

VSEPR theory predicts the three-dimensional arrangement of atoms in a molecule. It states that valence electrons will assume an electron-pair geometry that minimizes repulsions between areas of high electron density (bonds and/or lone pairs). Molecular structure, which refers only to the placement of atoms in a molecule and not the electrons, is equivalent to electron-pair geometry only when there are no lone electron pairs around the central atom. A dipole moment measures a separation of charge. For one bond, the bond dipole moment is determined by the difference in electronegativity between the two atoms. For a molecule, the overall dipole moment is determined by both the individual bond moments and how these dipoles are arranged in the molecular structure. Polar molecules (those with an appreciable dipole moment) interact with electric fields, whereas nonpolar molecules do not.

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5.10: Electronegativity and Bond Polarity

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

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

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure 5.10.1.

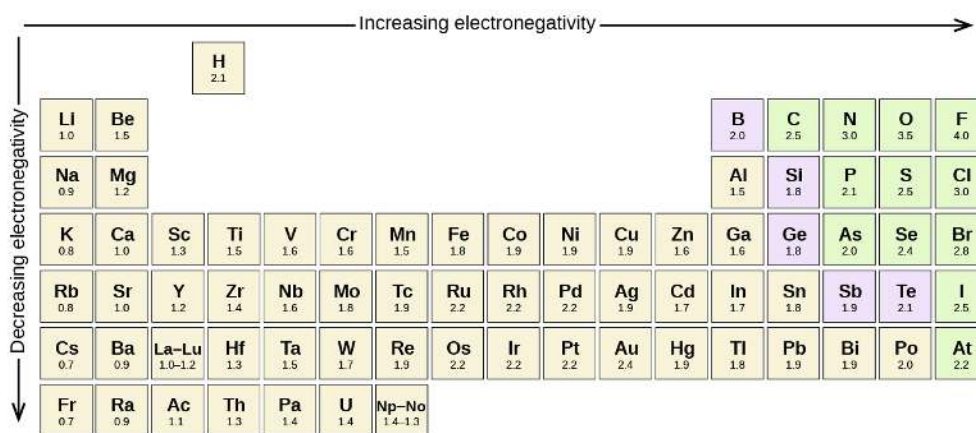


Figure 5.10.1: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

Electronegativity Difference	Bond Type
0	nonpolar covalent
0–0.4	slightly polar covalent
0.5–2.1	definitely polar covalent
>2.1	likely ionic

Nonpolar Covalent Bonds

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

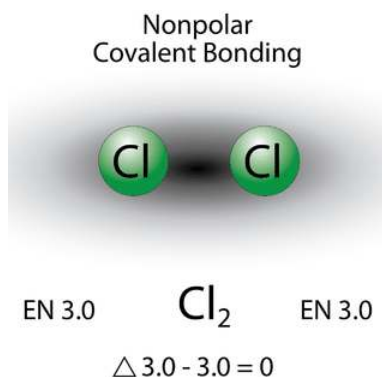


Figure 5.10.2: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

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

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.1 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

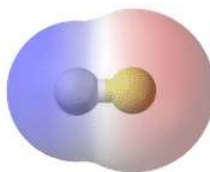


Figure 5.10.3: In the polar covalent bond of HF , the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

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



Figure 5.10.4: Use of δ to indicate partial charge.

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



Figure 5.10.5: Use of crossed arrow to indicate polarity

Example 5.10.1: Bond Polarity

What is the polarity of each bond?

- C–H
- O–H

Solution

Using Figure 5.10.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

- For the C–H bond, the difference in the electronegativities is $2.5 - 2.1 = 0.4$. Thus we predict that this bond will be non polar covalent.

b. For the O–H bond, the difference in electronegativities is $3.5 - 2.1 = 1.4$, so we predict that this bond will be definitely polar covalent.

Exercise 5.10.1

What is the polarity of each bond?

a. Rb–F

b. P–Cl

Answer a

likely ionic

Answer b

polar covalent

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. **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. 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 5.10.6: 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 (**CO₂**) 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 **CO₂** is zero.

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

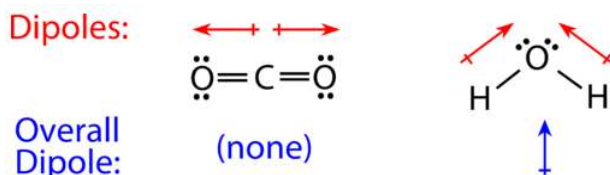


Figure 5.10.7: The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO₂ molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

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

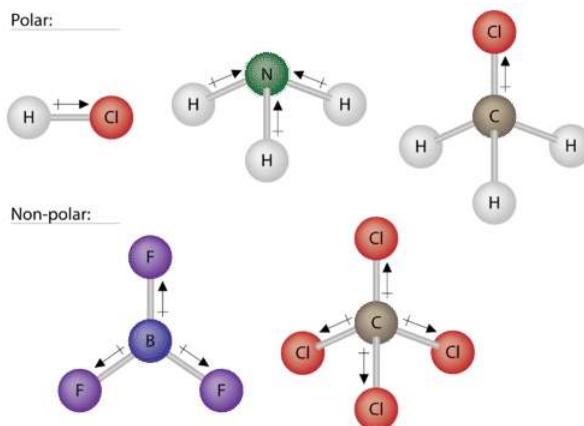


Figure 5.10.8: Some examples of polar and nonpolar molecules based on molecular geometry.

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.

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure 5.10.14). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

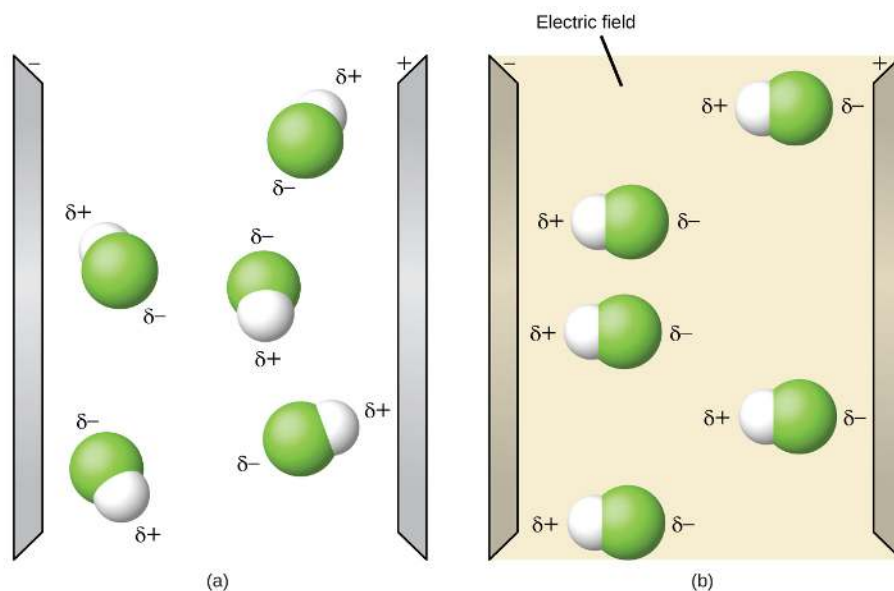


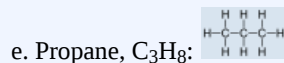
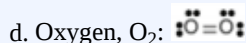
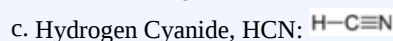
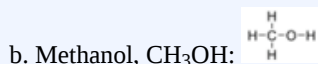
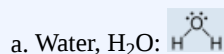
Figure 5.10.9: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

Example 5.10.2:

Label each of the following as polar or nonpolar.



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 5.10.2

Label each of the following as polar or nonpolar.

a. SO_3

b. NH_3

Answer a

non polar

Answer b

polar

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5.11: Ionic Compounds Containing Polyatomic Ions

When polyatomic ions are included, the number of ionic compounds increases significantly. Indeed, most ionic compounds contain polyatomic ions. Well-known examples are sodium hydroxide (NaOH) with OH^- as the polyatomic anion, calcium carbonate (CaCO_3), and ammonium nitrate (NH_4NO_3), which contains two polyatomic ions: NH_4^+ and NO_3^- .

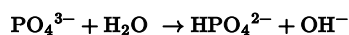
A list of the more important polyatomic ions is given in the following table, which can be used for reference while learning the charges of polyatomic ions. A great many of them are oxyanions (polyatomic ions that contain oxygen).

Table 5.11.1 Ions

Charge	Name	Formula
-3	Phosphate	PO_4^{3-}
	Arsenate	AsO_4^{3-}
-2	Carbonate	CO_3^{2-}
	Peroxide	O_2^{2-}
	Sulfate	SO_4^{2-}
	Sulfite	SO_3^{2-}
	Chromate	CrO_4^{2-}
	Dichromate	$\text{Cr}_2\text{O}_7^{2-}$
	Hydrogen phosphate	HPO_4^{2-}
	Hydrogen carbonate (bicarbonate)	HCO_3^-
-1	Superoxide	O_2^-
	Hydrogen sulfate	HSO_4^-
	Dihydrogen phosphate	H_2PO_4^-
	Hydroxide	OH^-
	Nitrate	NO_3^-
	Nitrite	NO_2^-
	Acetate	$\text{C}_2\text{H}_3\text{O}_2^-$ or CH_3COO^-
	Cyanide	CN^-
	Permanganate	MnO_4^-
	Perchlorate	ClO_4^-
	Chlorate	ClO_3^-
	Chlorite	ClO_2^-
	Hypochlorite	ClO^-
+1	Ammonium	NH_4^+
	Hydronium	H_3O^+

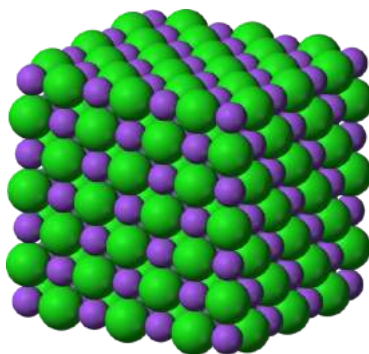
The properties of compounds containing polyatomic ions are very similar to those of binary ionic compounds. The ions are arranged in a regular lattice and held together by coulombic forces of attraction. The resulting crystalline solids usually have high melting points (1500 °F for CaCO_3) and all conduct electricity when molten.

Most are soluble in water and form conducting solutions in which the ions can move around as independent entities. In general polyatomic ions are colorless, unless, like CrO_4^{2-} or MnO_4^- , they contain a transition-metal atom. The more negatively charged polyatomic ions, like their monatomic counterparts, show a distinct tendency to react with water, producing hydroxide ions; for example,

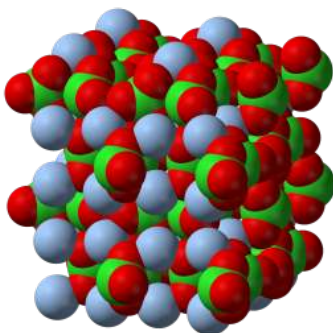


It is important to realize that compounds containing polyatomic ions must be *electrically neutral*. In a crystal of calcium sulfate, for instance, there must be equal numbers of Ca^{2+} and SO_4^{2-} ions in order for the charges to balance. The formula is thus CaSO_4 . In the case of sodium sulfate, by contrast, the Na^+ ion has only a single charge. In this case we need *two* Na^+ ions for each SO_4^{2-} ion in order to achieve electroneutrality. The formula is thus Na_2SO_4 .

Structurally, polyatomic ions are similar to the ionic solids we saw earlier. An example of a simple ionic compound, NaCl , is seen below, alongside a more complex ionic solid, AgClO_3 . Notice how both are tightly packed and form a repeating pattern, which lends both compounds strength and brittleness. In the Silver Chlorate (AgClO_3), however, polyatomic ions are present where the Cl^- ions are present in the Sodium Chloride (NaCl).



A cubic sodium chloride lattice made up of stacked green and purple spheres. The spheres are stacked in an alternating pattern. The green sphere is slightly larger than the purple sphere.



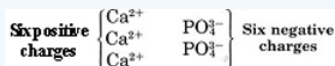
The three dimensional lattice of silver chlorate is complex because each silver sphere arranged in a regular cubic lattice is connected to a one green sphere which is connected to three red spheres.

On the left is the lattice structure of the ionic solid NaCl , with Na represented by the green spheres and Cl represented by the purple spheres. On the right is the lattice structure of AgClO_3 , with the silver spheres representing Ag and the red and green molecules representing the polyatomic ion ClO_3

✓ Example 5.11.1 : Ionic Formula

What is the formula of the ionic compound calcium phosphate?

Solution It is necessary to have the correct ratio of calcium ions, Ca^{2+} , and phosphate ions, PO_4^{3-} , in order to achieve electroneutrality. The required ratio is the *inverse* of the ratio of the charges on ions. Since the charges are in the ratio of 2:3, the numbers must be in the ratio of 3:2. In other words the solid salt must contain three calcium ions for every two phosphate ions:



Composition of calcium phosphate is three moles of Ca ions which make up a total of six positive charges as well as two phosphate ions which has a total of six negative charges.

The formula for calcium phosphate is thus $\text{Ca}_3(\text{PO}_4)_2$.

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5.12: Metallic Bonding



Figure 5.12.1: A nineteenth century copper plate. (Public Domain; John Pearson via [Wikipedia](#))

Why do metals behave the way they do?

The image below is of a copper plate that was made in 1893. The utensil has a great deal of elaborate decoration, and the item is very useful. What would have happened if this plate was made of copper (I) chloride instead? Copper (I) chloride does contain copper, after all. However, the CuCl would end up as a powder when a metalworker pounded on it to shape it. Metals behave in unique ways. The bonding that occurs in a metal is responsible for its distinctive properties: luster, malleability, ductility, and excellent conductivity.

The Metallic Bond

Pure metals are crystalline solids, but unlike ionic compounds, every point in the crystal lattice is occupied by an identical atom. The electrons in the outer energy levels of a metal are mobile and capable of drifting from one metal atom to another. This means that the metal is more properly viewed as an array of positive ions surrounded by a sea of mobile valence electrons. Electrons which are capable of moving freely throughout the empty orbitals of the metallic crystal are called **delocalized electrons** (see below). A **metallic bond** is the attraction of the stationary metal cations to the surrounding mobile electrons.

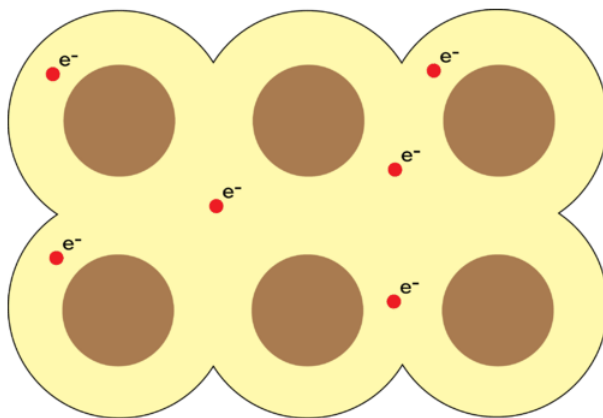


Figure 5.12.2: In a metal, the stationary metal cations are surrounded by a sea of mobile valence electrons that are not associated with any one cation. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)



Properties of Metals

The metallic bonding model explains the physical properties of metals. Metals conduct electricity and heat very well because of their free-flowing electrons. As electrons enter one end of a piece of metal, an equal number of electrons flow outward from the other end. When light is shone onto the surface of a metal, its electrons absorb small amounts of energy and become excited into one of its many empty orbitals. The electrons immediately fall back down to lower energy levels and emit light. This process is responsible for the high **luster** of metals.



Figure 5.12.3: The American Platinum Eagle is the official platinum bullion coin of the United States and was first minted in 1997. The luster of a metal is due to its metallic bonds. (Public Domain; the US Mint via Front: http://commons.wikimedia.org/wiki/File:2005_AEPlat_Proof_Obv.png; Back: http://commons.wikimedia.org/wiki/File:American_Platinum_Eagle_2008_Proof_Rev.jpg)

Recall that ionic compounds are very brittle. Application of a force results in like-charged ions in the crystal coming too close to one another, causing the crystal to shatter. When a force is applied to a metal, the free-flowing electrons can slip in between the stationary cations and prevent them from coming in contact. Imagine ball bearings that have been coated with oil sliding past one another. As a result, metals are very **malleable** and **ductile**. They can be hammered into shapes, rolled into thin sheets, or pulled into thin wires.

Summary

- The metallic bond is responsible for the properties of metals.
- Metals conduct electricity and heat well.
- Metals are ductile and malleable.
- Metals have luster.

Review

1. What is a delocalized electron?
2. Why do metals conduct electricity and heat well?
3. Why do metals have luster?

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5.13: Network Covalent Atomic Solids- Carbon and Silicates

Covalent Network Solids

Covalent solids are formed by networks or chains of atoms or molecules held together by covalent bonds. A perfect single crystal of a covalent solid is therefore a single giant molecule. For example, the structure of diamond, shown in part (a) in Figure 5.13.2, consists of carbon atoms, each bonded to four other carbon atoms in a tetrahedral array to create a giant network. The carbon atoms form six-membered rings. In, graphite, the other common allotrope of carbon, has the structure shown in part (b) in Figure 5.13.1. It contains planar networks of six-membered rings carbon atoms in which each carbon is bonded to three others. This leaves a single electron in a 2p orbital that can be used to form C=C double bonds, resulting in a ring with alternating double and single bonds.

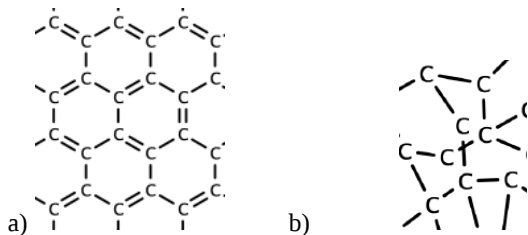


Figure 5.13.1 a) Carbon bonding in graphite; b) Carbon bonding in diamond

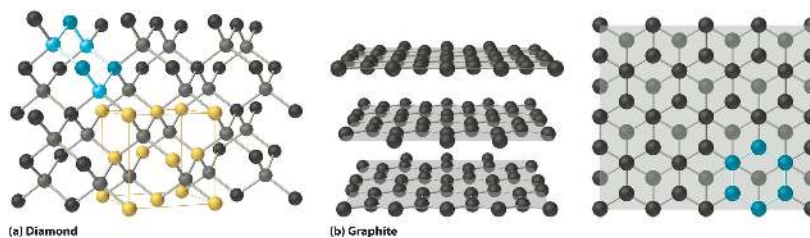


Figure 5.13.2 The Structures of Diamond and Graphite. (a) Diamond consists of sp^3 hybridized carbon atoms, each bonded to four other carbon atoms. The tetrahedral array forms a giant network in which carbon atoms form six-membered rings. (b) These side (left) and top (right) views of the graphite structure show the layers of fused six-membered rings and the arrangement of atoms in alternate layers of graphite. The rings in alternate layers are staggered, such that every other carbon atom in one layer lies directly under (and above) the center of a six-membered ring in an adjacent layer.

Elemental silicon has the same structure, as does silicon carbide (SiC), which has alternating C and Si atoms. The structure of crystalline quartz (silicon dioxide, SiO_2), shown in 5.13.3 can be viewed as being derived from the structure of silicon by inserting an oxygen atom between each pair of silicon atoms.



5.13.3 . Network covalent solid structure of SiO_2 . Image by [Smokefoot](#), [CC BY-SA 4.0](#), via Wikimedia Commons

All compounds with the diamond and related structures are extremely hard, high-melting-point solids that are not easily deformed. Instead, they tend to shatter when subjected to large stresses, and they usually do not conduct electricity very well. In fact, diamond (melting point = $3500^\circ C$ at 63.5 atm) is one of the hardest substances known, and silicon carbide (melting point = $2986^\circ C$) is used commercially as an abrasive in sandpaper and grinding wheels. It is difficult to deform or melt these and related compounds because strong covalent (C–C or Si–Si) or polar covalent (Si–C or Si–O) bonds must be broken, which requires a large input of energy.

Fullerenes

Until the mid 1980's, pure carbon was thought to exist in two forms: graphite and diamond. The discovery of C_{60} molecules in interstellar dust in 1985 added a third form to this list. The existence of C_{60} , which resembles a soccer ball, had been hypothesized

by theoreticians for many years. In the late 1980's synthetic methods were developed for the synthesis of C_{60} , and the ready availability of this form of carbon led to extensive research into its properties.

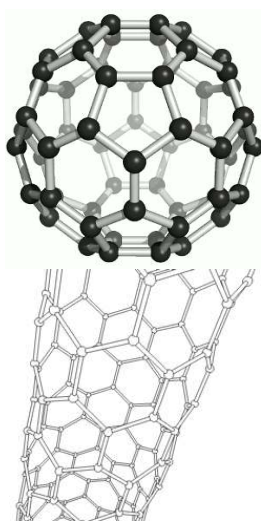


Figure 5.13.4: Example of fullerenes: a buckyball (C_{60} on left and an extended buckytube. Images used with permission from Wikipedia.

The C_{60} molecule (Figure 5.13.4; left), is called buckminsterfullerene, though the shorter name fullerene is often used. The name is a tribute to the American architect R. Buckminster Fuller, who is famous for designing and constructing geodesic domes which bear a close similarity to the structure of C_{60} . As is evident from the display, C_{60} is a sphere composed of six-member and five-member carbon rings. These balls are sometimes fondly referred to as "Bucky balls". It should be noted that fullerenes are an entire class of pure carbon compounds rather than a single compound. A distorted sphere containing more than 60 carbon atoms have also been found, and it is also possible to create long tubes (Figure 5.13.4; right). All of these substances are pure carbon.

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CHAPTER OVERVIEW

6: Intermolecular Forces

[6.1: Intermolecular Forces](#)

[6.2: Intermolecular Forces in Action](#)

[6.3: Intermolecular Forces in Solutions](#)

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6.1: Intermolecular Forces

Learning Objectives

- Describe the types of intermolecular forces possible between atoms or molecules (dispersion forces, dipole-dipole attractions, and hydrogen bonding)
- Identify the types of intermolecular forces experienced by specific molecules based on their structures
- Explain the relation between the intermolecular forces present within a substance and its physical properties (melting point, boiling point, solubility in polar and nonpolar solvents, etc)

In the following description, the term *particle* will be used to refer to an atom, molecule, or ion. Note that we will use the popular phrase “intermolecular attraction” to refer to attractive forces between the particles of a substance, regardless of whether these particles are molecules, atoms, or ions.

Consider these two aspects of the molecular-level environments in solid, liquid, and gaseous matter:

- Particles in a solid are tightly packed together and often arranged in a regular pattern; in a liquid, they are close together with no regular arrangement; in a gas, they are far apart with no regular arrangement.

The differences in the properties of a solid, liquid, or gas reflect the strengths of the attractive forces between the atoms, molecules, or ions that make up each phase. The phase in which a substance exists depends on the relative extents of its intermolecular forces (IMFs) and the kinetic energies (KE) of its molecules. IMFs are the various forces of attraction that may exist between the atoms and molecules of a substance due to electrostatic phenomena, as will be detailed in this module. These forces serve to hold particles close together, whereas the particles’ KE provides the energy required to overcome the attractive forces and thus increase the distance between particles. Figure 6.1.1 illustrates how changes in physical state may be induced by changing the temperature, hence, the average KE, of a given substance.

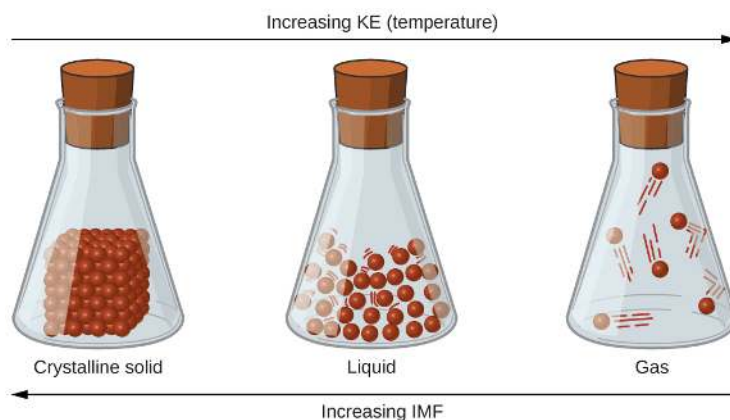


Figure 6.1.1: Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

As an example of the processes depicted in this figure, consider a sample of water. When gaseous water is cooled sufficiently, the attractions between H_2O molecules will be capable of holding them together when they come into contact with each other; the gas condenses, forming liquid H_2O . For example, liquid water forms on the outside of a cold glass as the water vapor in the air is cooled by the cold glass, as seen in Figure 6.1.2.



Figure 6.1.2: Condensation forms when water vapor in the air is cooled enough to form liquid water, such as (a) on the outside of a cold beverage glass or (b) in the form of fog. (credit a: modification of work by Jenny Downing; credit b: modification of work by Cory Zanker)

We can also liquefy many gases by compressing them, if the temperature is not too high. The increased pressure brings the molecules of a gas closer together, such that the attractions between the molecules become strong relative to their KE. Consequently, they form liquids. Butane, C_4H_{10} , is the fuel used in disposable lighters and is a gas at standard temperature and pressure. Inside the lighter's fuel compartment, the butane is compressed to a pressure that results in its condensation to the liquid state, as shown in Figure 6.1.3.



Figure 6.1.3: Gaseous butane is compressed within the storage compartment of a disposable lighter, resulting in its condensation to the liquid state. (credit: modification of work by "Sam-Cat"/Flickr)

Finally, if the temperature of a liquid becomes sufficiently low, or the pressure on the liquid becomes sufficiently high, the molecules of the liquid no longer have enough KE to overcome the IMF between them, and a solid forms. A more thorough discussion of these and other changes of state, or phase transitions, is provided in a later module of this chapter.

Forces between Molecules

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not *intramolecular* forces. *Intramolecular* forces are those *within* the molecule that keep the molecule together, that is, the chemical bonds between the atoms. *Intermolecular* forces are the attractions *between* molecules, which determine many of the physical properties of a substance. Figure 6.1.4 illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the **IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule**. For example, to overcome the IMFs in one mole of liquid HCl and convert it

into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

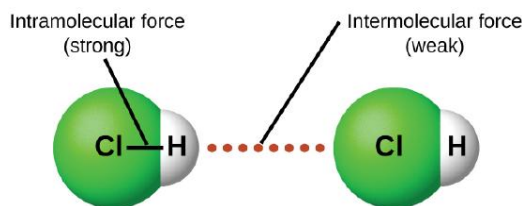
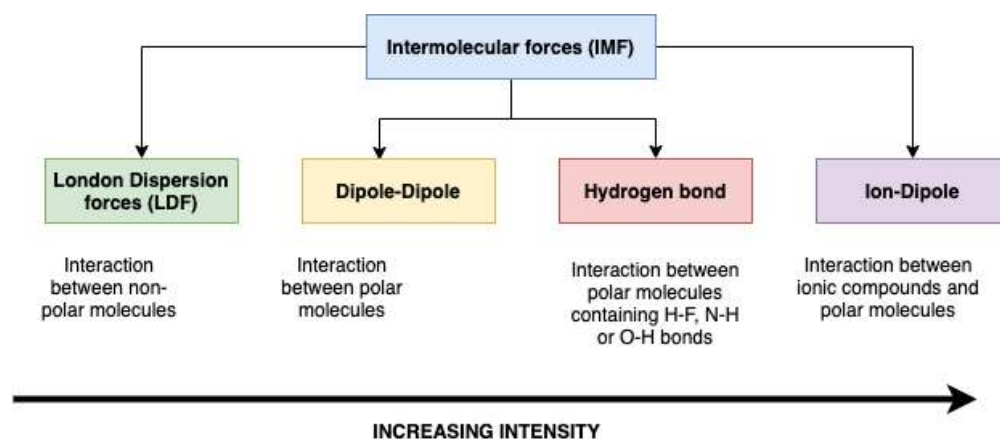


Figure 6.1.4: Intramolecular forces keep a molecule intact. Intermolecular forces hold multiple molecules together and determine many of a substance's properties.

All of the attractive forces between neutral atoms and molecules are known as **van der Waals forces**, although they are usually referred to more informally as **intermolecular forces**. We will consider the various types of IMFs in the next sections of this module.



London Dispersion Forces

One of the three van der Waals forces is present in all condensed phases, regardless of the nature of the atoms or molecules composing the substance. This attractive force is called **London dispersion forces (LDF)** in honor of German-born American physicist Fritz London who, in 1928, first explained it. This force is often referred to as simply the dispersion force. To better understand this type of intermolecular attractions, it is better to think of molecules in terms of Erwin Schrödinger's quantum mechanical model of the atom, in which an atom is considered as a positive nucleus immersed in an electron cloud. In similar terms, molecules can be considered as several nuclei sharing an electron cloud. This electron cloud is not static, but rather mobile as electrons are moving at a rapid pace. Because the electrons of an atom or molecule are in constant motion, at any moment in time, an atom or molecule can develop a **temporary, instantaneous dipole** if its electrons are distributed asymmetrically. The presence of this dipole can, in turn, distort the electrons of a neighboring atom or molecule, producing an induced dipole. These two rapidly fluctuating, temporary dipoles thus result in a relatively weak electrostatic attraction between the species—a so-called dispersion force like that illustrated in Figure 6.1.5.

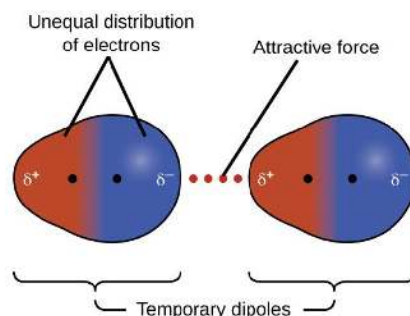


Figure 6.1.5: London dispersion forces result from the formation of temporary dipoles, as illustrated here for two nonpolar diatomic molecules.

London 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 6.1.1.

Table 6.1.1: 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

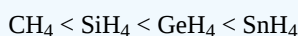
The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther 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.

Example 6.1.1: London Forces and Their Effects

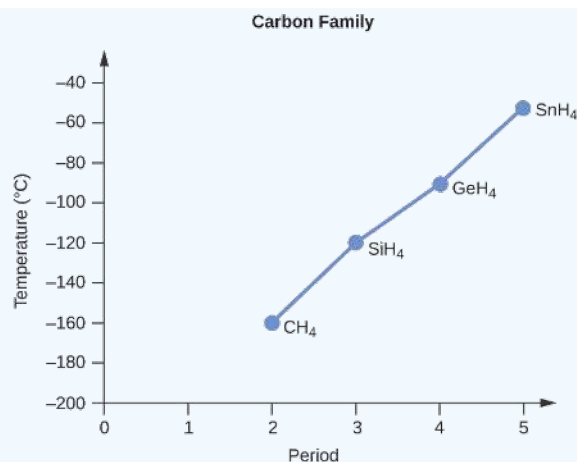
Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4 , SiH_4 , GeH_4 , and SnH_4 . Explain your reasoning.

Solution

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



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



Dipole-Dipole Attractions

Recall from the chapter on chemical bonding and molecular geometry that *polar* molecules have a partial positive charge on one side and a partial negative charge on the other side of the molecule—a separation of charge called a **dipole**. Consider a polar molecule such as hydrogen chloride, HCl. In the HCl molecule, the more electronegative Cl atom bears the partial negative charge, whereas the less electronegative H atom bears the partial positive charge. An attractive force between HCl molecules results from the attraction between the positive end of one HCl molecule and the negative end of another. This attractive force is called a dipole-dipole attraction—the electrostatic force between the partially positive end of one polar molecule and the partially negative end of another, as illustrated in Figure 6.1.8.

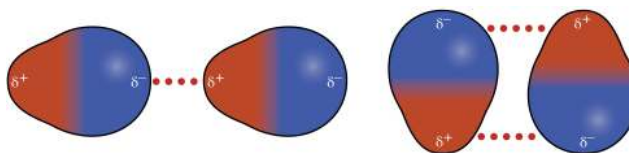


Figure 6.1.8: This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another. These dipoles are permanent, and not temporary as in the case for London Dispersion Forces.

The effect of a dipole-dipole attraction is apparent when we compare the properties of HCl molecules to nonpolar F₂ molecules. Both HCl and F₂ consist of the same number of atoms and have approximately the same molecular mass. At a temperature of 150 K, molecules of both substances would have the same average KE. However, the dipole-dipole attractions between HCl molecules are sufficient to cause them to “stick together” to form a liquid, whereas the relatively weaker dispersion forces between nonpolar F₂ molecules are not, and so this substance is gaseous at this temperature. The higher normal boiling point of HCl (188 K) compared to F₂ (85 K) is a reflection of the greater strength of dipole-dipole attractions between HCl molecules, compared to the attractions between nonpolar F₂ molecules. We will often use values such as boiling or freezing points, or enthalpies of vaporization or fusion, as indicators of the relative strengths of IMFs of attraction present within different substances.

Example 6.1.2: 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 6.1.2

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

Answer

ICl and Br_2 have similar masses (~ 160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br_2 is nonpolar and does not. The relatively stronger dipole-dipole attractions require more energy to overcome, so ICl will have the higher boiling point.

Hydrogen Bonding

Nitrosyl fluoride (ONF , molecular mass 49 amu) is a gas at room temperature. Water (H_2O , molecular mass 18 amu) is a liquid, even though it has a lower molecular mass. We clearly cannot attribute this difference between the two compounds to dispersion forces. Both molecules have about the same shape and ONF is the heavier and larger molecule. It is, therefore, expected to experience more significant dispersion forces. Additionally, we cannot attribute this difference in boiling points to differences in the dipole moments of the molecules. Both molecules are polar and exhibit comparable dipole moments. The large difference between the boiling points is due to a particularly strong dipole-dipole attraction that may occur when a molecule contains a hydrogen atom bonded to a fluorine, oxygen, or nitrogen atom (the three most electronegative elements). The very large difference in electronegativity between the H atom (2.1) and the atom to which it is bonded (4.0 for an F atom, 3.5 for an O atom, or 3.0 for a N atom), combined with the very small size of a H atom and the relatively small sizes of F, O, or N atoms, leads to *highly concentrated partial charges* with these atoms. **Molecules with F-H, O-H, or N-H moieties are very strongly attracted to similar moieties in nearby molecules, a particularly strong type of dipole-dipole attraction called hydrogen bonding.** Examples of hydrogen bonds include $\text{HF}\cdots\text{HF}$, $\text{H}_2\text{O}\cdots\text{HOH}$, and $\text{H}_3\text{N}\cdots\text{HNH}_2$, in which the hydrogen bonds are denoted by dots. Figure 6.1.9 illustrates hydrogen bonding between water molecules.

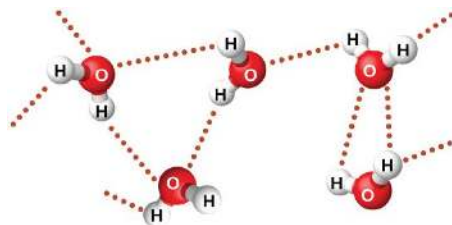


Figure 6.1.9: Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

Despite use of the word “bond,” keep in mind that hydrogen bonds are *intermolecular* attractive forces, not *intramolecular* attractive forces (covalent bonds). Hydrogen bonds are much weaker than covalent bonds, only about 5 to 10% as strong, but are generally much stronger than other dipole-dipole attractions and dispersion forces.

Hydrogen bonds have a pronounced effect on the properties of condensed phases (liquids and solids). For example, consider the trends in boiling points for the binary hydrides of group 15 (NH_3 , PH_3 , AsH_3 , and SbH_3), group 16 hydrides (H_2O , H_2S , H_2Se , and H_2Te), and group 17 hydrides (HF , HCl , HBr , and HI). The boiling points of the heaviest three hydrides for each group are plotted in Figure 6.1.10. As we progress down any of these groups, the polarities of the molecules decrease slightly, whereas the sizes of the molecules increase substantially. The effect of increasingly stronger dispersion forces dominates that of increasingly weaker dipole-dipole attractions, and the boiling points are observed to increase steadily. For the group 15, 16, and 17 hydrides, the boiling points for each class of compounds increase with increasing molecular mass for elements in periods 3, 4, and 5.

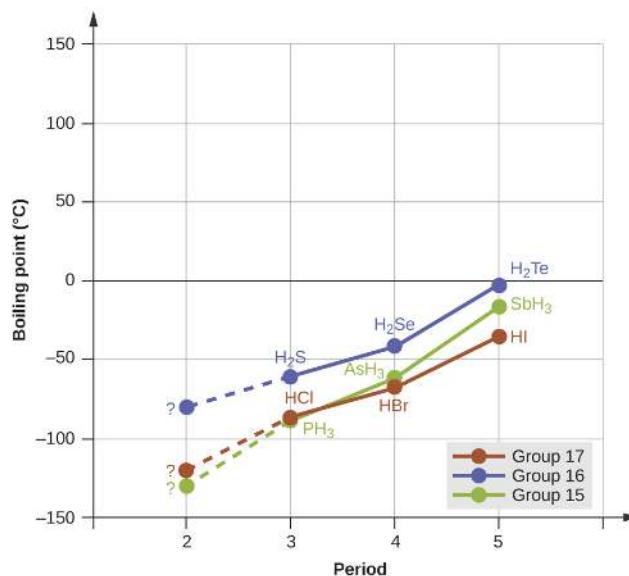


Figure 6.1.10: In comparison to periods 3–5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding.

If we use this trend to predict the boiling points for the lightest hydride for each group, we would expect NH₃ to boil at about -120 °C, H₂O to boil at about -80 °C, and HF to boil at about -110 °C. However, when we measure the boiling points for these compounds, we find that they are dramatically higher than the trends would predict, as shown in Figure 6.1.10. The stark contrast between our naïve predictions and reality provides compelling evidence for the strength of hydrogen bonding.

Hydrogen Bonding and DNA

Deoxyribonucleic acid (DNA) is found in every living organism and contains the genetic information that determines the organism's characteristics, provides the blueprint for making the proteins necessary for life, and serves as a template to pass this information on to the organism's offspring. A DNA molecule consists of two (anti-)parallel chains of repeating nucleotides, which form its well-known double helical structure, as shown in Figure 6.1.10.

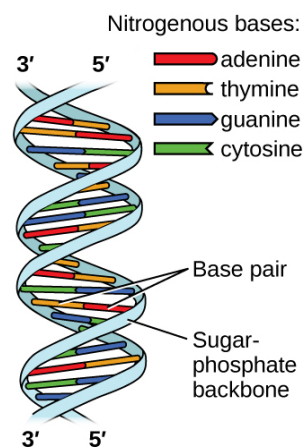


Figure 6.1.11: Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (credit: modification of work by Jerome Walker, Dennis Myts)

Each nucleotide contains a (deoxyribose) sugar bound to a phosphate group on one side, and one of four nitrogenous bases on the other. Two of the bases, cytosine (C) and thymine (T), are single-ringed structures known as pyrimidines. The other two, adenine (A) and guanine (G), are double-ringed structures called purines. These bases form complementary base pairs consisting of one purine and one pyrimidine, with adenine pairing with thymine, and cytosine with guanine. Each base pair is held together by

hydrogen bonding. A and T share two hydrogen bonds, C and G share three, and both pairings have a similar shape and structure
Figure 6.1.12

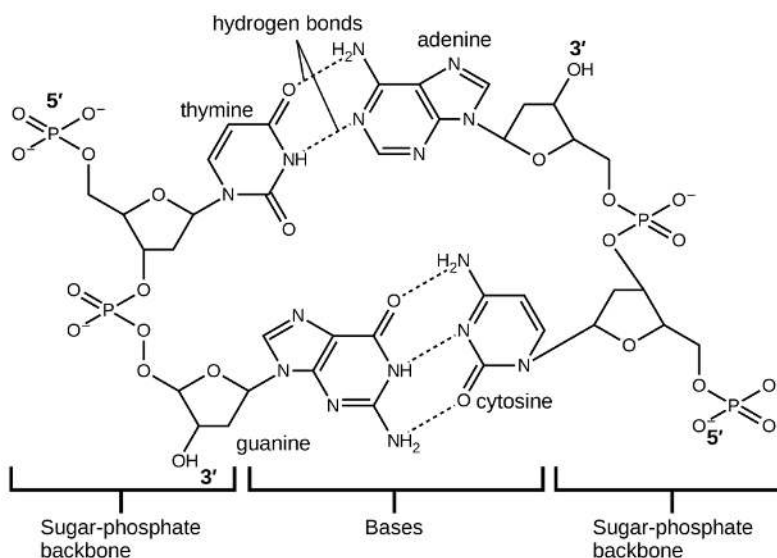


Figure 6.1.12: The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called “complementary base pairs.”

The cumulative effect of millions of hydrogen bonds effectively holds the two strands of DNA together. Importantly, the two strands of DNA can relatively easily “unzip” down the middle since hydrogen bonds are relatively weak compared to the covalent bonds that hold the atoms of the individual DNA molecules together. This allows both strands to function as a template for replication.

Ion-Dipole

The last type of intermolecular force that we need to discuss is the interaction between ionic compounds and polar molecules. Note that this intermolecular involves two different species: an ion and a polar molecule. Ion-dipole interaction results electrostatic interactions between an ion and the charged ends of a dipole, as indicated in Figure 6.1.13. The cation will be attracted to the partial negative portion of the polar molecule, while the anion will be attracted to the partial positive portion of the dipole. Because this IMF involves a fully charged species (an ion), it is the strongest of all IMFs.

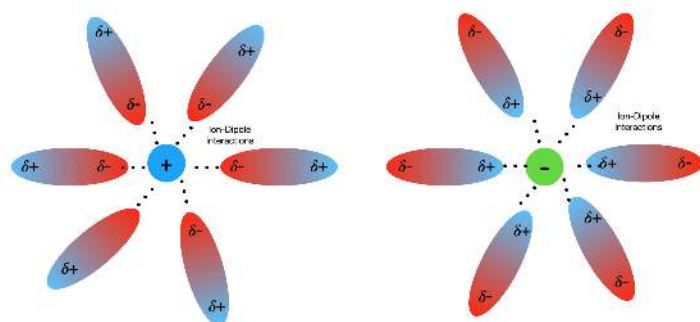


Figure 6.1.13. Formation of ion-dipole interactions

Ion-dipole is the interaction involved in solutions where an ionic compound is dissolved into a polar solvent, like that of a solution of table salt (NaCl) in water. When the ions in the NaCl crystal are added into water, the ions can interact with the dipoles in the H₂O molecules by ion-dipole forces, so the crystal disintegrates, the ions get surrounded by water molecules, and are now free to move in the solution.



Figure 6.1.14. Dissolution of NaCl in water. Image by [Ahazard.sciencewriter](#), CC BY-SA 4.0, via Wikimedia Commons

Summary

The physical properties of condensed matter (liquids and solids) can be explained in terms of the kinetic molecular theory. In a liquid, intermolecular attractive forces hold the molecules in contact, although they still have sufficient kinetic energy to move past each other. Intermolecular attractive forces, collectively referred to as van der Waals forces, are responsible for the behavior of liquids and solids and are electrostatic in nature. Dipole-dipole attractions result from the electrostatic attraction of the partial negative end of one dipolar molecule for the partial positive end of another. The temporary dipole that results from the motion of the electrons in an atom can induce a dipole in an adjacent atom and give rise to the London dispersion force. London forces increase with increasing molecular size. Hydrogen bonds are a special type of dipole-dipole attraction that results when hydrogen is bonded to one of the three most electronegative elements: F, O, or N.

Contributors and Attributions

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- Ion-Dipole Forces. (2020, July 8). Retrieved May 19, 2021, from <https://chem.libretexts.org/@go/page/221367>

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6.2: Intermolecular Forces in Action

Learning Objectives

- Understand how intermolecular forces affect the physical properties of matter, such as viscosity and surface tension

Boiling points and melting point in covalent compounds

As described previously, whether a covalent substance is a solid, a liquid, or a gas at room temperature depends on the relative strength of its intermolecular forces (IMFs) compared to the kinetic energies (KE) of its molecules at room temperature. When the IMFs are stronger than the KE, the molecules tend to be close together, with very restricted ability to move, then the substance is likely to be a solid. When the IMF and the KE are similar in magnitude, the molecules can move, but their movement is limited to short distances, so the covalent compound is likely to be liquid. When the IMFs are very weak compared to the KE, the attraction between molecules is minimal, so the molecules can move freely. Therefore the substance is most likely to be a gas.

 States of matter En

Figure 6.2.1: In a gas, the kinetic energy of the molecules is higher than the intermolecular forces. In a liquid, intermolecular forces and kinetic energy have comparable values. In a solid, intermolecular forces are stronger than the kinetic energy of the molecules. Image by Yelod - Wikimedia Commons * Yelod - Wikipedia (En) * ויקיפדיה העברית - ילוד, CC BY-SA 3.0, via Wikimedia Commons

Since the kinetic energy of the molecules increases with the temperature, if a covalent compound is in solid state and its temperature is increased gradually, there will be a temperature value at which the molecular kinetic energy will overcome the intermolecular forces, and the solid will melt into a liquid. This temperature is known as the melting point. Similarly, if the liquid is further heated, there will be a temperature value at which the molecular kinetic energy will be much higher than the intermolecular forces and the liquid will vaporize into a gas. This temperature is known as the boiling point. We can predict that the melting point and boiling point values for a given substance depend on the intensity of its intermolecular forces: the stronger the intermolecular forces, the higher the melting point and boiling point values.

In the case of ionic compounds, metallic substances, and network covalent solids, the melting and boiling processes involve the separation of particles within crystal structures, and therefore, chemical bonds must be broken. As a consequence, ionic compounds, metallic substances, and network covalent solid substances always exhibit very high melting and boiling points values compared to covalent compounds. This is because the melting and boiling processes for covalent compounds do not involve breaking the covalent bond, but rather separating the molecules by overcoming the acting intermolecular forces.

Table 6.2.1 shows the boiling point and melting for some substances and the forces that must be overcome in each case.

Table 6.2.1. Boiling points and melting points compared to intermolecular forces

Substance	Melting point (°C)	Boiling point (°C)	Forces that must be overcome
molecular oxygen O ₂	-218	-183	London dispersion forces
hydrogen sulfide, H ₂ S	-82	-60	Dipole-Dipole
water, H ₂ O	0	100	Hydrogen bonding
metallic aluminum, Al	660	2743	metallic bond
Sodium Chloride, NaCl	800	1465	ionic bond
Carbon in diamond, C	4027	4830	covalent bonds in a network covalent solid

We can see that breaking chemical bonds requires much more energy than overcoming intermolecular forces. There are a few places in the universe in which the temperature is high enough to have ionic or metallic compounds in gas or liquid phase. Some examples, the magma produced by volcanoes have a high content of Fe₂O₃ in a liquid state, while the sun contains some metals in gas state (iron, magnesium, etc)



Figure 6.2.2. Volcano in Hawaii producing lava. Lava contains a high proportion of melted ionic compounds such as Fe₂O₃, Al₂O₃, and FeO. Image by [Hawaii Volcano Observatory \(DAS\)](#), Public domain, via Wikimedia Commons

Solubility

The ability of a given substance to dissolve in a given solvent depends on the strength of the intermolecular forces that can act between solute and solvent molecules. In general, polar substances will dissolve in polar solvents, while non-polar substances will dissolve in non-polar solvents. This is because non-polar substances can interact very well through London Dispersion forces, while polar substances can interact well through dipole-dipole or hydrogen bonding interactions. Therefore, the general rule is "like dissolves like".

Table 6.2.2 shows the solubility values for a series of substances in water, a polar solvent

Table 6.2.2 shows the solubility values for a series of substances in water at 25°C and 1 atm

Substance	Solubility in H ₂ O (g/100 mL, at 25°C)	Acting intermolecular forces
Oxygen, O ₂	0.0041	LDF
Sulfur dioxide, SO ₂	9.4	Dipole-Dipole
Ammonia, NH ₃	31	Hydrogen bonding
Sodium chloride, NaCl	36	Ion-Dipole

As we can see, the solubility values increase when the intermolecular interactions between the solvent and solute molecule become stronger. Switching the solvent into a non-polar one, such as carbon tetrachloride (CCl₄), carbon disulfide (CS₂) or hexane (C₆H₆) will certainly show the opposite trend in solubility for ten above listed substances.

Viscosity

The viscosity of a liquid is a measure of its resistance to flow. It is similar to the concept of "thickness". For example, maple syrup has a higher viscosity than sweet tea. Viscosity can be measured in units of millipascal-second (mPa·s). As the intensity of intermolecular forces increases, so does the viscosity. Table 6.2.3 shows the viscosity values for a series of liquid substances.

Table 6.2.3: viscosity values for a series of liquid substances at 25°C

Substance	Viscosity (mPa·s) at 25°C	Intermolecular forces
Benzene, C ₆ H ₆	0.604	LDF
Water, H ₂ O	0.89	Hydrogen bonding
Mercury, Hg	1.526	Metallic bond

Surface Tension

The next time you are by a still body of water, take a close look at what is scooting along on the surface. You may see insects seemingly floating on top of the water. These creatures are known by a variety of names including water skaters, water striders, pond skaters, and other equally descriptive names. They take advantage of a property called surface tension to stay above the water and not sink. The force they exert downward is less than the forces exerted among the water molecules on the surface of the pond, so the insect does not penetrate beneath the surface of the water.

Molecules within a liquid are pulled equally in all directions by intermolecular forces. However, molecules at the surface are pulled downwards and sideways by other liquid molecules, but not upwards away from the surface. The overall effect is that the surface molecules are pulled into the liquid, creating a surface that is tightened like a film (Figure 6.2.3A). The **surface tension** of a liquid is a measure of the elastic force in the liquid's surface. Liquids that have strong intermolecular forces, like the hydrogen bonding in water, exhibit the greatest surface tension. Surface tension allows objects that are denser than water, such as the paper clip shown in B in the figure below, to nonetheless float on its surface. It is also responsible for the beading up of water droplets on a freshly waxed car because there are no attractions between the polar water molecules and the nonpolar wax.

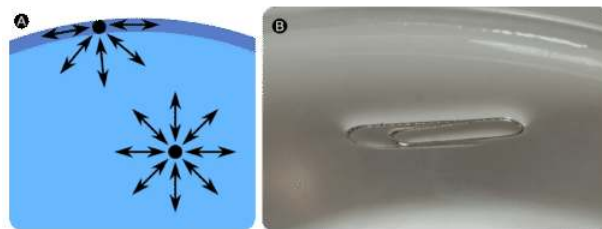


Figure 6.2.3: (A) Molecules at the surface of a liquid are pulled downwards into the liquid, creating a tightened surface. (B) Surface tension allows a paper clip to float on water's surface.

Other liquids, such as diethyl ether, do not demonstrate strong surface tension interactions. The intermolecular forces for the ether are the relatively weak dipole-dipole interactions that do not draw the molecules together as tightly as hydrogen bonds would.

Summary

- The surface tension of a liquid is a measure of the elastic force in the liquid's surface.
- Liquids with strong intermolecular forces have higher surface tensions than liquids with weaker forces.

Contributions & Attributions

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6.3: Intermolecular Forces in Solutions

Learning Objectives

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

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table 6.3.1 lists some common types of solutions, with examples of each.

Table 6.3.1 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 6.3.1: Sugar and Water

A solution is made by dissolving 1.00 g of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) 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 6.3.1

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

Answer

solute: $\text{HCl}(\text{g})$; 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 6.3.2 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 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 6.3.2 Water (clear liquid) and oil (yellow) do not form liquid solutions. Image used with permission (CC BY-SA 1.0 Generic; [Victor Blacus](#))

Example 6.3.2: 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 6.3.2

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 6.3.3

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

- methanol (CH_3OH)
- sodium sulfate (Na_2SO_4)
- octane (C_8H_{18}), a non-polar organic compound

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

Exercise 6.3.3

Toluene ($C_6H_5CH_3$) 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.

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.

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CHAPTER OVERVIEW

7: Overview of Inorganic Compounds

[7.1: Comparing Ionic and Molecular Substances](#)

[7.2: Nomenclature of Inorganic Compounds](#)

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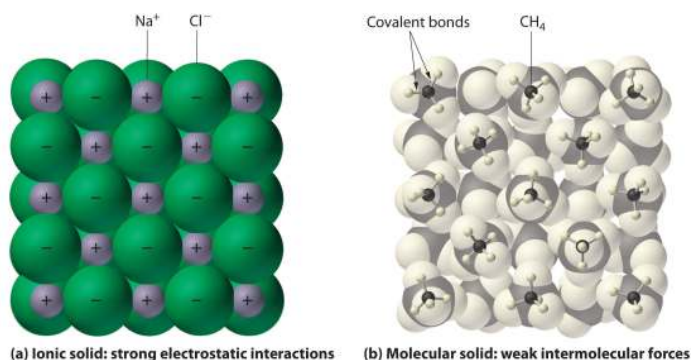
7.1: Comparing Ionic and Molecular Substances

Learning Objectives

- Know the physical properties of ionic and molecular substances.

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays.

Molecular compounds, sometimes called



(a) The positively and negatively charged ions in an ionic solid such as sodium chloride (NaCl) are held together by strong electrostatic interactions. (b) In this representation of the packing of methane (CH_4) molecules in solid methane, a prototypical molecular solid, the methane molecules are held together in the solid only by relatively weak intermolecular forces, even though the atoms within each methane molecule are held together by strong covalent bonds.

covalent compounds, display a wide range of physical properties due to the different types of intermolecular attractions such as different kinds of polar interactions. The melting and boiling points of molecular compounds are generally quite low compared to those of **ionic compounds**. This is because the energy required to disrupt the **intermolecular forces** between molecules is far less than the energy required to break the ionic bonds in a crystalline ionic compound (Figure 7.1.1). Ionic solids typically melt at high temperatures and boil at even higher temperatures. For example, sodium chloride melts at 801°C and boils at 1413°C . (As a comparison, the molecular compound water melts at 0°C and boils at 100°C .) The water solubility of molecular compounds is variable and depends primarily on the type of intermolecular forces involved.

Figure 7.1.1 Interactions in Ionic and Covalent Solids.

Since molecular compounds are composed of neutral molecules, their electrical conductivity is generally quite poor, whether in the solid or liquid state. In solid form, an ionic compound is not electrically conductive because its ions are unable to flow ("electricity" is the flow of charged particles). When molten, however, it can conduct electricity because its ions are able to move freely through the liquid (Figure 7.1.2; Video 7.1.1).

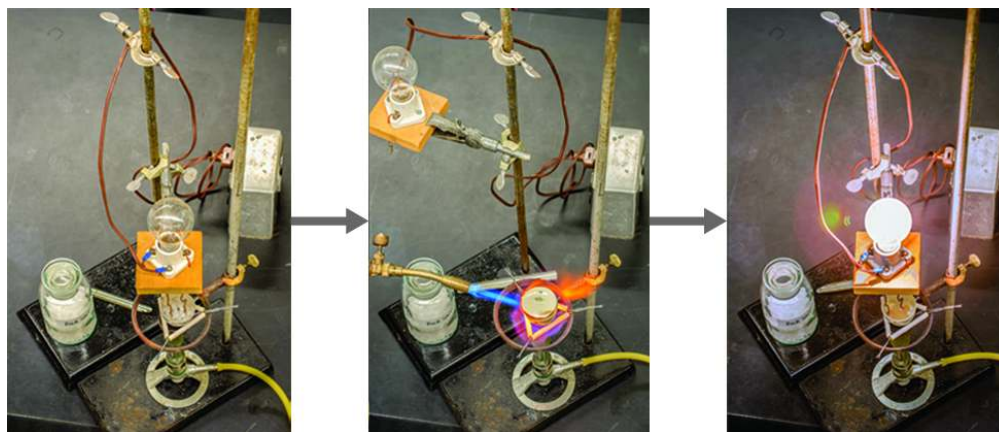


Figure 7.1.2 Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)

This figure shows three photos connected by right-facing arrows. The first shows a light bulb as part of a complex lab equipment setup. The light bulb is not lit. The second photo shows a substance being heated or set on fire. The third shows the light bulb again which is lit.

Conductivity of Molten Salt



Video 7.1.1 Watch this video to see a mixture of salts melt and conduct electricity.

The table below summarizes some of the differences between ionic and molecular compounds.

Table 7.1.1 Comparison of Ionic and Molecular Compounds

Property	Ionic Compounds	Molecular Compounds
Type of elements	Metal and nonmetal	Nonmetals only
Bonding	Ionic - transfer of electron(s) between atoms	Covalent - sharing of pair(s) of electrons between atoms
Representative unit	Formula unit	Molecule
Physical state at room temperature	Solid	Gas, liquid, or solid
Water solubility	Usually high	Variable
Melting and boiling temperatures	Generally high	Generally low
Electrical conductivity	Good when molten or in solution	Poor

One type of molecular compound behaves quite differently than that described so far. A covalent network solid is a compound in which all of the atoms are connected to one another by covalent bonds. Diamond is composed entirely of carbon atoms, each bonded to four other carbon atoms in a tetrahedral geometry. Melting a covalent network solid is not accomplished by overcoming the relatively weak intermolecular forces. Rather, all of the covalent bonds must be broken, a process that requires extremely high temperatures. Diamond, in fact, does not melt at all. Instead, it vaporizes to a gas at temperatures above **3500°C**.

Summary

- The physical properties of a material are affected by the intermolecular forces holding the molecules together.
- Ionic compounds usually form hard crystalline solids with high melting points. Covalent molecular compounds, in contrast, consist of discrete molecules held together by weak intermolecular forces and can be gases, liquids, or solids at room temperature and pressure.
- Ionic compounds in molten form or in solution can conduct electricity while molecular compounds do not..

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7.2: Nomenclature of Inorganic Compounds

Generally, there are two types of inorganic compounds that can be formed: ionic compounds and molecular compounds. Nomenclature is the process of naming chemical compounds with different names so that they can be easily identified as separate chemicals. Inorganic compounds are compounds that do not deal with the formation of carbohydrates, or simply all other compounds that do not fit into the description of an organic compound. For example, organic compounds include molecules with carbon rings and/or chains with hydrogen atoms (see picture below). Inorganic compounds, the topic of this section, are every other molecule that does not include these distinctive carbon and hydrogen structures.

Compounds between Metals and Nonmetals (Cation and Anion)

Compounds made of a metal and nonmetal are commonly known as **Ionic Compounds**, where the compound name has an ending of **-ide**. Cations have positive charges while anions have negative charges. The net charge of any ionic compound must be zero which also means it must be electrically neutral. For example, one Na^+ is paired with one Cl^- ; one Ca^{2+} is paired with two Br^- . There are two rules that must be followed through:

- The **cation** (metal) is always named first with its name unchanged
- The **anion** (nonmetal) is written after the cation, modified to end in **-ide**

Table 1: Cations and Anions:

+1 Charge	+2 Charge	-1 Charge	-2 Charge	-3 Charge	-4 Charge
Group 1A elements	Group 2A elements	Group 7A elements	Group 6A elements	Group 5A elements	Group 4A elements
Hydrogen: H^+	Beryllium: Be^{2+}	Hydride: H^-	Oxide: O^{2-}	Nitride: N^{3-}	Carbide: C^{4-}
Lithium: Li^+	Magnesium: Mg^{2+}	Fluoride: F^-	Sulfide: S^{2-}	Phosphide: P^{3-}	
Sodium: Na^+	Calcium: Ca^{2+}	Chloride: Cl^-			
Potassium: K^+	Strontium: Sr^{2+}	Bromide: Br^-			
Rubidium: Rb^+	Barium: Ba^{2+}	Iodide: I^-			
Cesium: Cs^+					

✓ Example 1



Sodium + Chlorine = Sodium Chloride; Calcium + Bromine = Calcium Bromide

The **transition metals** may form more than one ion, thus it is needed to be specified which particular ion we are talking about. This is indicated by assigning a Roman numeral after the metal. The Roman numeral denotes the charge and the oxidation state of the transition metal ion. For example, iron can form two common ions, Fe^{2+} and Fe^{3+} . To distinguish the difference, Fe^{2+} would be named iron (II) and Fe^{3+} would be named iron (III).

Table of Transition Metal and Metal Cations:

+1 Charge	+2 Charge	+3 Charge	+4 Charge
Copper(I): Cu^+	Copper(II): Cu^{2+}	Aluminum: Al^{3+}	Lead(IV): Pb^{4+}
Silver: Ag^+	Iron(II): Fe^{2+}	Iron(III): Fe^{3+}	Tin(IV): Sn^{4+}
	Cobalt(II): Co^{2+}	Cobalt(III): Co^{3+}	
	Tin(II): Sn^{2+}		
	Lead(II): Pb^{2+}		
	Nickel: Ni^{2+}		
	Zinc: Zn^{2+}		

✓ Example 2

Ions:	$\text{Fe}^{2+} + 2\text{Cl}^-$	$\text{Fe}^{3+} + 3\text{Cl}^-$
Compound:	FeCl_2	FeCl_3
Nomenclature	Iron (II) Chloride	Iron (III) Chloride

However, some of the transition metals' charges have specific Latin names. Just like the other nomenclature rules, the ion of the transition metal that has the lower charge has the Latin name ending with **-ous** and the one with the higher charge has a Latin name ending with **-ic**. The most common ones are shown in the table below:

Transition Metal Ion with Roman Numeral	Latin name
Copper (I): Cu^+	Cuprous
Copper (II): Cu^{2+}	Cupric
Iron (II): Fe^{2+}	Ferrous
Iron (III): Fe^{3+}	Ferric
Lead (II): Pb^{2+}	Plumbous
Lead (IV): Pb^{4+}	Plumbic
Mercury (I): Hg_2^{2+}	Mercurous
Mercury (II): Hg^{2+}	Mercuric
Tin (II): Sn^{2+}	Stannous
Tin (IV): Sn^{4+}	Stannic

Several exceptions apply to the Roman numeral assignment: Aluminum, Zinc, and Silver. Although they belong to the transition metal category, these metals do not have Roman numerals written after their names because these metals only exist in one ion. Instead of using Roman numerals, the different ions can also be presented in plain words. The metal is changed to end in **-ous** or **-ic**.

- **-ous** ending is used for the **lower** oxidation state
- **-ic** ending is used for the **higher** oxidation state

✓ Example 3

Compound	Cu_2O	CuO	FeCl_2	FeCl_3
Charge	Charge of copper is +1	Charge of copper is +2	Charge of iron is +2	Charge of iron is +3
Nomenclature	Cuprous Oxide	Cupric Oxide	Ferrous Chloride	Ferric Chloride

However, this **-ous/-ic** system is inadequate in some cases, so the Roman numeral system is preferred. This system is used commonly in naming acids, where H_2SO_4 is commonly known as Sulfuric Acid, and H_2SO_3 is known as Sulfurous Acid.

Compounds between Nonmetals and Nonmetals

Compounds that consist of a nonmetal bonded to a nonmetal are commonly known as **Molecular Compounds**, where the element with the positive oxidation state is written first. In many cases, nonmetals form more than one binary compound, so **prefixes** are used to distinguish them.

# of Atoms	1	2	3	4	5	6	7	8	9	10
Prefixes	Mono-	Di-	Tri-	Tetra-	Penta-	Hexa-	Hepta-	Octa-	Nona-	Deca-

✓ Example 4

CO = carbon **monoxide** BCl_3 = boron**trichloride**
 CO_2 = carbon **dioxide** N_2O_5 = **dinitrogen pentoxide**

The prefix **mono-** is not used for the first element. If there is not a prefix before the first element, it is assumed that there is only one atom of that element.

Binary Acids

Although HF can be named hydrogen fluoride, it is given a different name for emphasis that it is an acid. An **acid** is a substance that dissociates into hydrogen ions (H^+) and anions in water. A quick way to identify acids is to see if there is an H (denoting hydrogen) in front of the molecular formula of the compound. To name acids, the prefix **hydro-** is placed in front of the nonmetal modified to end with **-ic**. The state of acids is aqueous (aq) because acids are found in water.

Some common binary acids include:

HF (g) = hydrogen fluoride → HF (aq) = **hydrofluoric acid**
HBr (g) = hydrogen bromide → HBr (aq) = **hydrobromic acid**
HCl (g) = hydrogen chloride → HCl (aq) = **hydrochloric acid**
 H_2S (g) = hydrogen sulfide → H_2S (aq) = **hydrosulfuric acid**

It is important to include (aq) after the acids because the same compounds can be written in gas phase with hydrogen named first followed by the anion ending with **-ide**.

✓ Example 5

hypo____ite ____ite ____ate per____ate
 ClO^- ClO_2^- ClO_3^- ClO_4^-
hypochlorite chlorite chlorate perchlorate
----->

As indicated by the arrow, moving to the right, the following trends occur:

Increasing number of oxygen atoms

Increasing oxidation state of the nonmetal

(Usage of this example can be seen from the set of compounds containing Cl and O)

This occurs because the number of oxygen atoms are increasing from hypochlorite to perchlorate, yet the overall charge of the polyatomic ion is still -1. To correctly specify how many oxygen atoms are in the ion, prefixes and suffixes are again used.

Polyatomic Ions

In **polyatomic ions**, polyatomic (meaning two or more atoms) are joined together by **covalent bonds**. Although there may be a element with positive charge like H^+ , it is not joined with another element with an ionic bond. This occurs because if the atoms formed an ionic bond, then it would have already become a compound, thus not needing to gain or loose any electrons. Polyatomic anions are more common than polyatomic cations as shown in the chart below. Polyatomic anions have negative charges while polyatomic cations have positive charges. To indicate different polyatomic ions made up of the same elements, the name of the ion is modified according to the example below:

Table: Common Polyatomic ions

Name: Cation/Anion	Formula
Ammonium ion	NH_4^+
Hydronium ion	H_3O^+
Acetate ion	$C_2H_3O_2^-$
Arsenate ion	AsO_4^{3-}
Carbonate ion	CO_3^{2-}
Hypochlorite ion	ClO^-
Chlorite ion	ClO_2^-
Chlorate ion	ClO_3^-
Perchlorate ion	ClO_4^-
Chromate ion	CrO_4^{2-}
Dichromate ion	$Cr_2O_7^{2-}$
Cyanide ion	CN^-
Hydroxide ion	OH^-
Nitrite ion	NO_2^-
Nitrate ion	NO_3^-
Oxalate ion	$C_2O_4^{2-}$
Permanganate ion	MnO_4^-
Phosphate ion	PO_4^{3-}
Sulfite ion	SO_3^{2-}
Sulfate ion	SO_4^{2-}
Thiocyanate ion	SCN^-
Thiosulfate ion	$S_2O_3^{2-}$

To combine the topic of acids and polyatomic ions, there is nomenclature of aqueous acids. Such acids include sulfuric acid (H_2SO_4) or carbonic acid (H_2CO_3). To name them, follow these quick, simple rules:

1. If the ion ends in **-ate** and is added with an acid, the acid name will have an **-ic** ending. Examples: nitrate ion (NO_3^-) + H^+ (denoting formation of acid) = **nitric** acid (HNO_3)
2. If the ion ends in **-ite** and is added with an acid, then the acid name will have an **-ous** ending. Example: nitrite ion (NO_2^-) + H^+ (denoting formation of acid) = **nitrous** acid (HNO_2)

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Problems

1. What is the correct formula for Calcium Carbonate?
 - a. $\text{Ca}^+ + \text{CO}_2^-$
 - b. CaCO_2^-
 - c. CaCO_3
 - d. 2CaCO_3
2. What is the correct name for FeO ?
 - a. Iron oxide
 - b. Iron dioxide
 - c. Iron(III) oxide
 - d. Iron(II) oxide
3. What is the correct name for $\text{Al}(\text{NO}_3)_3$?
 - a. Aluminum nitrate
 - b. Aluminum(III) nitrate
 - c. Aluminum nitrite
 - d. Aluminum nitrogen trioxide
4. What is the correct formula of phosphorus trichloride?
 - a. P_2Cl_2
 - b. PCl_3
 - c. PCl_4
 - d. P_4Cl_2
5. What is the correct formula of lithium perchlorate?
 - a. Li_2ClO_4
 - b. LiClO_2
 - c. LiClO
 - d. None of these

6. Write the correct name for these compounds.

a. BeC_2O_4 :

b. NH_4MnO_4 :

c. CoS_2O_3 :

7. What is $\text{W}(\text{HSO}_4)_5$?

8. How do you write diphosphorus trioxide?

9. What is H_3P ?

10. By adding oxygens to the molecule in number 9, we now have H_3PO_4 ? What is the name of this molecule?

Answer

1.C; Calcium + Carbonate $\rightarrow \text{Ca}^{2+} + \text{CO}_3^{2-} \rightarrow \text{CaCO}_3$

2.D; $\text{FeO} \rightarrow \text{Fe} + \text{O}^{2-} \rightarrow$ Iron must have a charge of +2 to make a neutral compound $\rightarrow \text{Fe}^{2+} + \text{O}^{2-} \rightarrow$ Iron(II) Oxide

3.A; $\text{Al}(\text{NO}_3)_3 \rightarrow \text{Al}^{3+} + (\text{NO}_3^-)_3 \rightarrow$ Aluminum nitrate

4.B; Phosphorus trichloride $\rightarrow \text{P} + 3\text{Cl} \rightarrow \text{PCl}_3$

5.D, LiClO_4 ; Lithium perchlorate $\rightarrow \text{Li}^+ + \text{ClO}_4^- \rightarrow \text{LiClO}_4$

6. a. Beryllium Oxalate; $\text{BeC}_2\text{O}_4 \rightarrow \text{Be}^{2+} + \text{C}_2\text{O}_4^{2-} \rightarrow$ Beryllium Oxalate

b. Ammonium Permanganate; $\text{NH}_4\text{MnO}_4 \rightarrow \text{NH}_4^+ + \text{MnO}_4^- \rightarrow$ Ammonium Permanganate

c. Cobalt (II) Thiosulfate; $\text{CoS}_2\text{O}_3 \rightarrow \text{Co} + \text{S}_2\text{O}_3^{2-} \rightarrow$ Cobalt must have +2 charge to make a neutral compound $\rightarrow \text{Co}^{2+} + \text{S}_2\text{O}_3^{2-} \rightarrow$ Cobalt(II) Thiosulfate

7. Tungsten (V) hydrogen sulfate

8. P_2O_3

9. Hydrophosphoric Acid

10. Phosphoric Acid

Contributors and Attributions

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CHAPTER OVERVIEW

8: Chemical Reactions

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8.1: Chemical Changes and Chemical Reactions

In [Chapter 2](#), we classified changes in our environment utilizing the concepts of *physical* and *chemical changes*. We said that a **physical change** alters the *appearance* of a substance without changing its molecular structure. Ice melts, water evaporates and mountains are slowly weathered into dust. All of these change the *characteristics* of substances, but they do not alter its basic structure. A **chemical change**, however, results in the transformation of one molecular substance into another. Gasoline burns, reacting with oxygen in the atmosphere, generating light, heat, and converting the carbon-based molecules into carbon dioxide gas and water vapor. When substances combine like this and undergo chemical changes, we say that a **chemical reaction** has occurred. Some chemical reactions are quite evident, like the burning of gasoline, and involve the production of **heat** or **light**. In other types of chemical reactions, **gases** are evolved, **color changes** occur and clear solutions become cloudy, with the ultimate formation of an insoluble substance (a **precipitate**). Chemical changes can also be quite obscure and their occurrence can only be detected by sophisticated chemical analysis.

Sometimes chemical changes occur spontaneously, others require the input of energy (heat) in order to occur. Chemical reactions can occur rapidly, like the explosive reaction of sodium metal in the presence of water, and others occur very slowly, like the rusting of iron or the tarnish that slowly develops on some metal surfaces exposed to air. In this chapter we will learn to represent chemical reactions using **chemical equations**. We will learn to **balance** these equations, explore **types** of reactions and learn to **predict products** from simple reactions. Central to all of this is the concept of the *chemical equation*.

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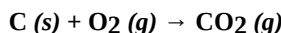
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8.2: Chemical Equations

The processes that occur during a chemical change can be represented using a **chemical equation**. In a chemical equation, the chemical formulas for the substance or substances that *undergo* the chemical reaction (the **reactants**) and the formulas for the new substance or substances that are formed (the **products**) are both shown, and are linked by an **arrow**. The arrow in a chemical equation has the properties of an “equals sign” in mathematics, and because of this, ***in a chemical equation, there must be the same number and types of atoms on each side of the arrow.***

Reactants → Products

As an example of a chemical reaction, consider the reaction between solid carbon and oxygen gas to form carbon dioxide gas. This chemical equation for this reaction can be written as shown below.



In this equation, we have used **(s)** and **(g)** to represent the physical state of the reactants at the time of the reaction (solid and gas). Other abbreviations that are often used include **(l)** for liquid and **(aq)** to indicate that the reactant or product is dissolved in aqueous solution.

As we inspect this equation we see that there is one carbon atom on each side of the arrow and that there are two oxygen atoms on each side. An equation in which there are the same number and types of atoms on both sides of the arrow is referred to as **balanced**. As you write chemical equations, it is important to remember those elements that naturally occur as diatomic molecules (Table 1.1). Remember that when you include these as reactants or products, remember to indicate that they are diatomic by using the subscript “2”.

Table 1.1 Common Diatomic Elements

Element	Chemical Formula
Hydrogen	H ₂
Oxygen	O ₂
Nitrogen	N ₂
Fluorine	F ₂
Chlorine	Cl ₂
Bromine	Br ₂
Iodine	I ₂

? Exercise 8.2.1

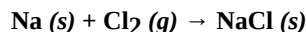
- Write a chemical equation for the reaction of solid iron with solid sulfur to form solid iron(II) sulfide.
- Write a chemical equation for the reaction of solid carbon with solid magnesium oxide to form carbon monoxide gas and magnesium metal.

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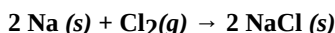
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8.3: Balancing Chemical Equations

In another example of a chemical reaction, sodium metal reacts with chlorine gas to form solid sodium chloride. An equation describing this process is shown below.



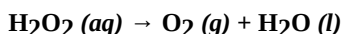
Inspection of this equation, however, shows that, while there is one sodium atom on each side of the arrow, there are two chlorine atoms in the reactants and only one in the products. This equation is **not balanced**, and is therefore *not* a valid chemical equation. In order to balance this equation, we must insert **coefficients** (not *subscripts*) in front of the appropriate reactants or products so that *the same number and types of atoms appear on both sides of the equation*. Because chlorine is diatomic, there are two chlorines in the reactants and there must also be two chlorines in the products. In order to accomplish this, we place the coefficient “2” in front of the product, NaCl. Now we are balanced for chlorine, but there are two atoms of sodium in the products and only one shown in the reactants. To resolve this, we need to place the coefficient “2” in front of the sodium in the reactant, to give the equation shown below.



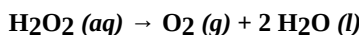
In this equation, there are two sodiums in the reactants, two sodiums in the products, two chlorines in the reactants and two chlorines in the products; **the equation is now balanced**.

There are many different strategies that people use in order to balance chemical equations. The simplest methods, where you examine and modify coefficients in some systematic order, is generally called “balancing by inspection”. These methods are generally useful for most simple chemical equations, although mathematical algorithms are often necessary for highly complex reactions. One version of the “inspection” method that we will use in this section can be called the “**odd-even**” approach. Looking at the first equation that we wrote for the sodium-chlorine reaction, we note that there are an **odd** number of chlorines in the products and an **even** number of chlorines in the reactants. The first thing we did in balancing this equation was to insert the multiplier “2” in front of the product (NaCl) so that there were now an *even* number of chlorines on both sides of the equation. Once we did that, we simply had to balance the other element (Na) which was “odd” on both sides, and the equation was easily balanced. When you are using this approach with more complicated equations, it is often useful to begin by balancing the **most complex molecule** in the equation first (the one with the most atoms), and focus on the element in this compound that is present in the **greatest amount**.

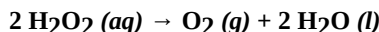
Another example where the “odd-even” approach works well is the decomposition of hydrogen peroxide to yield water and oxygen gas, as shown below.



As we inspect this equation, we note that there are an even number of oxygen atoms in the reactants and an odd number of oxygens in the products. Specifically, water has only one oxygen (in the products) and the number of oxygen atoms in the products can be made “even” by inserting the coefficient “2” in front of H₂O. Doing this (shown below) we note that we now have four hydrogens in the products and only two in the reactants.



Balancing the hydrogens by inserting “2” in front of H₂O₂ in the reactants gives us an equation with four hydrogens on both sides on four oxygens on both sides; **the equation is now balanced**.



Contributors

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8.4: Classifying Chemical Reactions

The reactions we have examined in the previous sections can be classified into a few simple types. Organizing reactions in this way is useful because it will assist us in predicting the products of unknown reactions. There are many different classifications of chemical reactions, but here we will focus on the following types: **synthesis**, **decomposition**, **single replacement** and **double replacement**. In addition, we will see that some of these reactions involve changes in the oxidation numbers of the reactants and products; these will be referred to as **oxidation-reduction**, or “**redox**” reactions (see figure below)

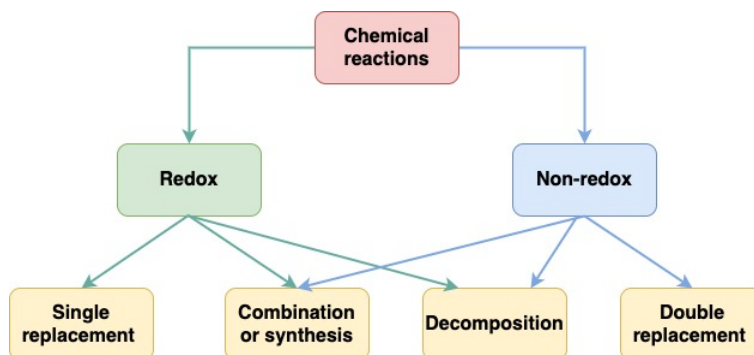
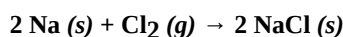


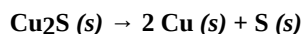
Figure 8.1. Classification of Chemical reactions

The first type of reaction we will consider is a **synthesis** reaction (also called a **combination reaction**). In a synthesis reaction, elements or compounds undergo reaction and combine to form a single new substance. The reaction of sodium metal with chlorine gas to give sodium chloride is an example of a synthesis reaction where both reactants are elements.

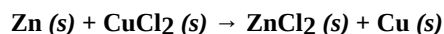


In this reaction, sodium metal and chlorine gas have *combined* to yield (synthesize) the more complex molecule, sodium chloride.

In **decomposition** reactions, a single compound will break down to form two or more new substances. The substances formed can be elements, compounds, or a mixture of both elements and compounds. Two simple examples of decomposition reactions are shown below.

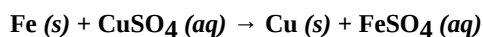


In a **single-replacement reaction** (also called a single-displacement reaction) an element and a compound will react so that their elements are switched. In other words, an element will typically displace another element from within a compound. As a general rule, **metals** will replace **metals** in compounds and **non-metals** will typically replace **non-metals**. An example of a single replacement reaction is shown below.



In this example, elemental zinc has *displaced* the metal, copper, from copper(II) chloride to form zinc chloride and elemental copper. In the reactants, zinc was elemental and in the products, it is present within the compound, zinc chloride. Likewise, copper was present in a compound in the reactants and is elemental in the products.

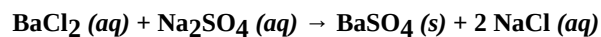
In another example, iron metal will react with an aqueous solution of copper sulfate to give copper metal and iron(II) sulfate.



In this reaction, elemental iron replaces copper in a compound with sulfate anion and elemental copper metal is formed; metal replaces metal. The tendency of metals to replace other metals in single-replacement reactions is often referred to as an activity series, as we will learn next semester.

A **double-replacement** reaction (or double-displacement) two ionic compounds in aqueous solution switch anions and form two new compounds. In order for a chemical reaction to occur in a double-replacement reaction, one of the new compounds that is

formed must be *insoluble* in water, forming a solid **precipitate** . If both of the new compounds which are formed are water-soluble, then no reaction has occurred (as we will learn later). An example of a double-replacement reaction is shown below.



In this reaction, solid barium sulfate is formed as a precipitate. This is a chemical change and this is a valid chemical reaction.

- [Paul R. Young](#), Professor of Chemistry, University of Illinois at Chicago, Wiki: AskTheNerd; PRY askthenerd.com - pyoung @uic.edu; [ChemistryOnline.com](#)

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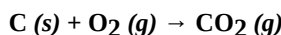
8.5: Redox Reactions

The **oxidation number** (O.N.) of an element represents the total number of electrons which have been *removed* (a positive oxidation state) or *added* (a negative oxidation state) to get the element into its present state. The term **oxidation** describes the *loss* of electrons by an element and an *increase* in oxidation state; the term **reduction** describes the *gain* of electrons and a *decrease* in oxidation state. Oxidation numbers for elements in compounds can be calculated using a simple set of rules, which are reproduced below in Table 5.3.

Table 5.3 Rules for Assigning Oxidation Numbers

1. The oxidation number of an element in the free state is **zero**.
2. A monoatomic ion will have an oxidation number that is equal to its charge.
3. In most compounds, the oxidation number of hydrogen is **+1** (exception: when hydrogen is bonded to a metal, the oxidation number **-1**.)
4. Oxygen, within a compound, will generally have an oxidation number of **-2** (exception: in peroxides, the oxidation number of oxygen is **-1**)
5. The oxidation numbers of combined elements from groups I and II are **+1** and **+2**, respectively.
6. The algebraic sum of oxidation numbers of all elements in a neutral compound formula equals zero.
7. The algebraic sum of the oxidation numbers of all elements in a polyatomic ion is equal to the charge of the ion

In many chemical reactions, the oxidation number of elements change. Consider the synthesis reaction shown below. In the reactants, carbon and oxygen are both elements and their oxidation numbers are zero (Rule 1). In the product, oxygen will have an oxidation number of **-2** (Rule 4), therefore, carbon in CO_2 must have an oxidation number of **+4** in order to balance the four negative charges on the oxygens. During this reaction, the oxidation number of carbon has changed from zero in the reactants to **+4** in the products and the oxidation number of oxygen has changed from zero to **-2**. This is an example of a **redox reaction**; a **chemical reaction in which the oxidation numbers of elements change on going from reactants to products**.



In a redox reaction, the element that “**loses electrons**” is said to be **oxidized** and will have an *increase in its oxidation number*. In the example above, the oxidation number of carbon *increases* from zero to **+4**; it has “lost electrons” and has been **oxidized**. The element that “**gains electrons**” in a redox reaction is said to be **reduced** and will have a *decrease in its oxidation number*. In the reaction above, the oxidation number of oxygen has *decreased* from zero to **-2**; it has “gained electrons” and has been **reduced**.

An examination of the rules for assigning oxidation numbers reveals that there are many elements for which there are no specific rules, such as nitrogen, sulfur, and chlorine. These elements, as well as some others, can have variable oxidation numbers depending on the other atoms to which they are covalently bonded in a molecular compound. It is useful to analyze a few molecules in order to see the strategy to follow in assigning oxidation numbers to other atoms.

Oxidation numbers for the atoms in a binary ionic compound are easy to assign because they are equal to the charge of the ion (rule 2). In CuCl_2 , the oxidation number of copper is **+2**, while the oxidation number of chlorine is **-1**. This is because CuCl_2 is an ionic compound that is actually composed of these ions.

Assigning oxidation numbers for molecular compounds is trickier. The key is to remember rule 6: that the sum of all the oxidation numbers for any neutral species must be zero. Make sure to account for any subscripts which appear in the formula. As an example, consider the compound nitric acid, HNO_3 . According to rule 3, the oxidation number of hydrogen is **+1**. According to rule 4, the oxidation number of oxygen is **-2**. There is no rule regarding nitrogen, but its oxidation number can be calculated as follows.

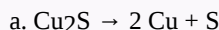
$1(+1) + x + 3(-2) = 0$, where x is the oxidation number of nitrogen.

Solving for x , we obtain $x = +5$

The oxidation number of the nitrogen atom in HNO_3 is **+5**

Examples 8.5.1

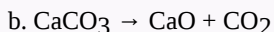
For each of the reactions given below, calculate the oxidation number of each of the elements in the reactants and the products and determine if the reaction involves oxidation-reduction. If it is a redox reaction, identify the elements that have been oxidized and reduced.



Reactants: Cu O.N.= +1 S O.N. = -2

Products: Cu O.N.= 0 S O.N.= 0

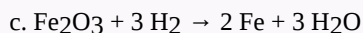
Element oxidized: sulfur Element Reduced: copper. This is a redox reaction.



Reactants: Ca O.N.= +2 C O.N.= +4 O O.N.= -2

Products: Ca O.N.= +2 C O.N.= +4 O O.N.= -2

Element oxidized: none. Element Reduced: none. This is non-redox reaction.



Reactants: Fe O.N.=+3 O O.N.= -2 H O.N.=0

Products: Fe O.N.=0 O O.N.= -2 H O.N.=+1

Element oxidized: hydrogen. Element Reduced: iron. This is a redox reaction



Reactants: Ag O.N.=+1 (cation) N O.N.=+5 (rule 6) O O.N.= -2 Na O.N.=+1 Cl O.N.= -1

Ag O.N.=+1 (cation) N O.N.=+5 (rule 6) O O.N.= -2 Na O.N.=+1 Cl O.N.= -1

Element oxidized: none. Element Reduced: none. This is a non-redox reaction

Contributors

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- [Oxidation numbers](#). Contributed by [Allison Soult](#) , Senior Lecturer (Chemistry) at [University of Kentucky](#)

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8.6: The Law of Conservation of Matter

Learning Objectives

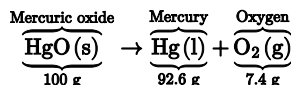
- Correctly define a law as it pertains to science.
- State the law of conservation of matter.

In science, a law is a general statement that explains a large number of observations. Before being accepted, a law must be verified many times under many conditions. Laws are therefore considered the highest form of scientific knowledge and are generally thought to be inviolable. Scientific laws form the core of scientific knowledge. One scientific law that provides the foundation for understanding in chemistry is the law of conservation of matter. It states that in any given system that is closed to the transfer of matter (in and out), the amount of matter in the system stays constant. A concise way of expressing this law is to say that the amount of matter in a system is *conserved*.

With the development of more precise ideas on elements, compounds and mixtures, scientists began to investigate how and why substances react. French chemist A. Lavoisier laid the foundation to the scientific investigation of matter by describing that substances react by following certain laws. These laws are called the laws of chemical combination. These eventually formed the basis of Dalton's Atomic Theory of Matter.

Law of Conservation of Mass

According to this law, during any physical or chemical change, the total mass of the products remains equal to the total mass of the reactants.



Another way of stating this is, "In a chemical reaction, matter is neither created nor destroyed." The law of conservation of mass is also known as the "law of indestructibility of matter."

✓ Example 8.6.1

If heating 10 grams of CaCO_3 produces 4.4 g of CO_2 and 5.6 g of CaO , show that these observations are in agreement with the law of conservation of mass.



A sample of calcium carbonate (CaCO_3). (Public Domain; [Walkerma](#)).

Solution

- Mass of the reactants, CaCO_3 : 10 g
- Mass of the products, CO_2 and CaO : $4.4 \text{ g} + 5.6 \text{ g} = 10 \text{ g}$.

Because the mass of the reactants = the mass of the products, the observations are in agreement with the law of conservation of mass.

What does this mean for chemistry? In any chemical change, one or more initial substances change into a different substance or substances. Both the initial and final substances are composed of atoms because all matter is composed of atoms. According to the

law of conservation of matter, matter is neither created nor destroyed, so we must have the **same number and kind of atoms** after the chemical change as were present before the chemical change.

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Figure 5.1.1 shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose we had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose we had been able to measure the oxygen used by the fire and the gases produced by the fire. What would we find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

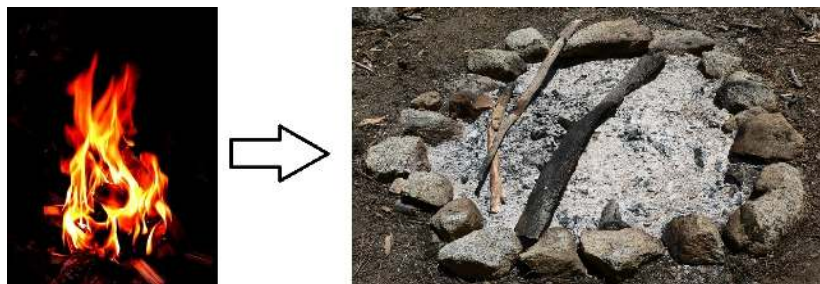


Figure 5.1.1: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). Images used with permission (CC BY-SA 2.5; Einar Helland Berger for fire and for ash).

? Exercise 8.6.1

- What is the law of conservation of matter?
- How does the law of conservation of matter apply to chemistry?

Answer a:

The law of conservation of matter states that in any given system that is closed to the transfer of matter, the amount of matter in the system stays constant

Answer b:

The law of conservation of matter says that in chemical reactions, the total mass of the products must equal the total mass of the reactants.

Key Takeaway

The amount of matter in a closed system is conserved.

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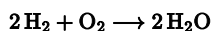
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8.7: Mole Calculations in Chemical Reactions

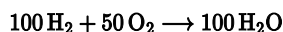
Learning Objectives

- Balance a chemical equation in terms of moles.
- Use the balanced equation to construct conversion factors in terms of moles.
- Calculate moles of one substance from moles of another substance using a balanced chemical equation.

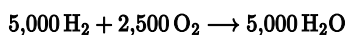
Consider this balanced chemical equation:



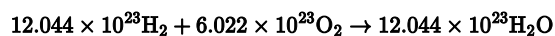
We interpret this as "two molecules of hydrogen react with one molecule of oxygen to make two molecules of water." The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:



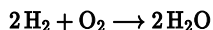
This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:



Again, this is not conventional, but it is still balanced. Suppose we use a much larger number:



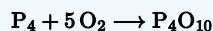
These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro's number, while the second number is Avogadro's number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?



is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. We can just as easily read this chemical equation as "two moles of hydrogen react with one mole of oxygen to make two moles of water." All balanced chemical reactions are balanced in terms of moles.

Example 8.7.1

Interpret this balanced chemical equation in terms of moles.

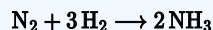


Solution

The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as "one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide."

Exercise 8.7.1

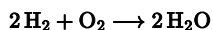
Interpret this balanced chemical equation in terms of moles.



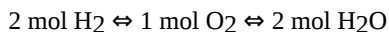
Answer

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

In Section 4.1, we stated that a chemical equation is simply a recipe for a chemical reaction. As such, chemical equations also give us equivalents—equivalents between the reactants and the products. However, now we understand that *these equivalents are expressed in terms of moles*. Consider the chemical equation



This chemical reaction gives us the following equivalents:



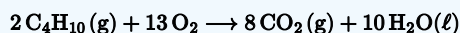
Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance. If, for example, we want to know how many moles of oxygen will react with 17.6 mol of hydrogen, we construct a conversion factor between 2 mol of H_2 and 1 mol of O_2 and use it to convert from moles of one substance to moles of another:

$$17.6 \cancel{\text{mol H}_2} \times \frac{1 \text{ mol O}_2}{2 \cancel{\text{mol H}_2}} = 8.80 \text{ mol O}_2$$

Note how the mol H_2 unit cancels, and mol O_2 is the new unit introduced. This is an example of a **mole-mole calculation**, when you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the numbers are small, but numbers won't always be so simple!

✓ Example 8.7.2

For the balanced chemical equation



if 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution

We are relating an amount of oxygen to an amount of carbon dioxide, so we need the equivalence between these two substances. According to the balanced chemical equation, the equivalence is



We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

$$154 \cancel{\text{mol O}_2} \times \frac{8 \text{ mol CO}_2}{13 \cancel{\text{mol O}_2}} = 94.8 \text{ mol CO}_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they do not contribute to the number of significant figures in the final answer.

? Exercise 8.7.2

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer

118 mol

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or liters—but moles. Any stoichiometry problem will likely need to work through the mole unit at some point, especially if you are working with a balanced chemical reaction.

Summary

Balanced chemical reactions are balanced in terms of moles. A balanced chemical reaction gives equivalents in moles that allow stoichiometry calculations to be performed.

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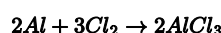
8.8: Mole-Mass and Mass-Mass Calculations

Learning Objectives

- From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation



Suppose we know we have 123.2 g of Cl_2 . How can we determine how many moles of AlCl_3 we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of AlCl_3 , we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2 \cancel{\text{g Cl}_2} \times \frac{1 \text{ mol Cl}_2}{70.90 \cancel{\text{g Cl}_2}} = 1.738 \text{ mol Cl}_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of AlCl_3 . The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2 \text{ mol AlCl}_3}{3 \text{ mol Cl}_2}$$

Using this conversion factor with the molar quantity we calculated above, we get

$$1.738 \cancel{\text{mol Cl}_2} \times \frac{2 \text{ mol AlCl}_3}{3 \cancel{\text{mol Cl}_2}} = 1.159 \text{ mol AlCl}_3$$

So, we will get 1.159 mol of AlCl_3 if we react 123.2 g of Cl_2 .

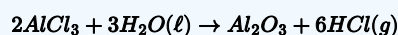
In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

$$123.2 \cancel{\text{g Cl}_2} \times \frac{1 \text{ mol Cl}_2}{70.90 \cancel{\text{g Cl}_2}} \times \frac{2 \text{ mol AlCl}_3}{3 \cancel{\text{mol Cl}_2}} = 1.159 \text{ mol AlCl}_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

✓ Example 8.8.1

How many moles of HCl will be produced when 249 g of AlCl_3 are reacted according to this chemical equation?



Solution

We will do this in two steps: convert the mass of AlCl_3 to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of AlCl_3 is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$249 \cancel{\text{g AlCl}_3} \times \frac{1 \text{ mol AlCl}_3}{133.33 \cancel{\text{g AlCl}_3}} = 1.87 \text{ mol AlCl}_3$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of AlCl_3 and the number of moles of HCl :

$$\frac{6 \text{ mol HCl}}{2 \text{ mol AlCl}_3}$$

Applying this conversion factor to the quantity of AlCl_3 , we get

$$1.87 \cancel{\text{mol AlCl}_3} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{mol AlCl}_3}} = 5.61 \text{ mol HCl}$$

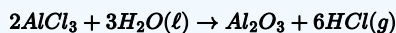
Alternatively, we could have done this in one line:

$$249 \cancel{\text{g AlCl}_3} \times \frac{1 \cancel{\text{mol AlCl}_3}}{133.33 \cancel{\text{g AlCl}_3}} \times \frac{6 \text{ mol HCl}}{2 \cancel{\text{mol AlCl}_3}} = 5.60 \text{ mol HCl}$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

? Exercise 8.8.1

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?



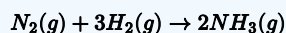
Answer

0.442 mol

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

✓ Example 8.8.2

How many grams of NH_3 will be produced when 33.9 mol of H_2 are reacted according to this chemical equation?



Solution

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have

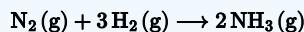
$$33.9 \cancel{\text{mol H}_2} \times \frac{2 \text{ mol NH}_3}{3 \cancel{\text{mol H}_2}} = 22.6 \text{ mol NH}_3$$

Now, using the molar mass of NH_3 , which is 17.03 g/mol, we get

$$22.6 \cancel{\text{mol NH}_3} \times \frac{17.03 \text{ g NH}_3}{1 \cancel{\text{mol NH}_3}} = 385 \text{ g NH}_3$$

? Exercise 8.8.2

How many grams of N₂ are needed to produce 2.17 mol of NH₃ when reacted according to this chemical equation?

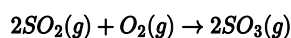


Answer

30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO₃ that can be produced by the reaction of 45.3 g of SO₂ and O₂:



First, we convert the given amount, 45.3 g of SO₂, to moles of SO₂ using its molar mass (64.06 g/mol):

$$45.3 \cancel{\text{g SO}_2} \times \frac{1 \text{ mol SO}_2}{64.06 \cancel{\text{g SO}_2}} = 0.707 \text{ mol SO}_2$$

Second, we use the balanced chemical reaction to convert from moles of SO₂ to moles of SO₃:

$$0.707 \cancel{\text{mol SO}_2} \times \frac{2 \text{ mol SO}_3}{2 \cancel{\text{mol SO}_2}} = 0.707 \text{ mol SO}_3$$

Finally, we use the molar mass of SO₃ (80.06 g/mol) to convert to the mass of SO₃:

$$0.707 \cancel{\text{mol SO}_3} \times \frac{80.06 \text{ g SO}_3}{1 \cancel{\text{mol SO}_3}} = 56.6 \text{ g SO}_3$$

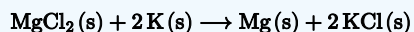
We can also perform all three steps sequentially, writing them on one line as

$$45.3 \cancel{\text{g SO}_2} \times \frac{1 \text{ mol SO}_2}{64.06 \cancel{\text{g SO}_2}} \times \frac{2 \text{ mol SO}_3}{2 \cancel{\text{mol SO}_2}} \times \frac{80.06 \text{ g SO}_3}{1 \cancel{\text{mol SO}_3}} = 56.6 \text{ g SO}_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO₃, which is what we are looking for, as our final answer.

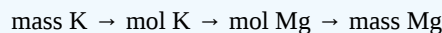
✓ Example 8.8.3

What mass of Mg will be produced when 86.4 g of K are reacted?



Solution

We will simply follow the steps

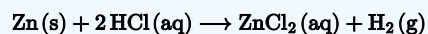


In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4 \cancel{\text{g K}} \times \frac{1 \text{ mol K}}{39.09 \cancel{\text{g K}}} \times \frac{1 \cancel{\text{mol Mg}}}{2 \cancel{\text{mol K}}} \times \frac{24.31 \text{ g Mg}}{1 \cancel{\text{mol Mg}}} = 26.87 \text{ g Mg}$$

? Exercise 8.8.3

What mass of H₂ will be produced when 122 g of Zn are reacted?



Answer

3.77 g

Summary

- Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.
- Mass quantities of one substance can be related to mass quantities using a balanced chemical equation.
- In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

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8.9: Limiting Reagents

Learning Objectives

- Identify a limiting reagent from a set of reactants.
- Calculate how much product will be produced from the limiting reagent.
- Calculate how much reactant(s) remains when the reaction is complete.

In addition to the assumption that reactions proceed all the way to completion, one additional assumption we have made about chemical reactions is that all the reactants are present in the proper quantities to react to products; this is not always the case. In Figure 8.9.2 we are taking hydrogen atoms and oxygen atoms (left) to make water molecules (right). However, there are not enough oxygen atoms to use up all the hydrogen atoms. We run out of oxygen atoms and cannot make any more water molecules, so the process stops when we run out of oxygen atoms.

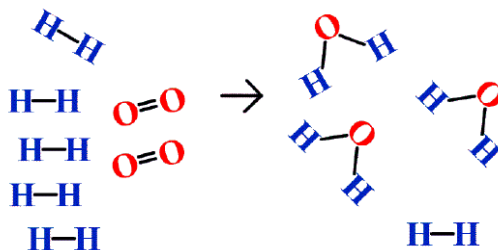
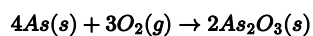


Figure 8.9.2: Making Water. In this scenario for making water molecules, we run out of O atoms before we use up all the H atoms. Similar situations exist for many chemical reactions when one reactant runs out before the other.

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reagent; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reagent and which is in excess.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reagent. What we need to do is determine an amount of one product (either moles or mass), assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reagent. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

For example, consider this reaction:



Suppose we start a reaction with 50.0 g of As and 50.0 g of O₂. Which one is the limiting reagent? We need to perform two mole-mass calculations, each assuming that each reactant reacts completely. Then we compare the amount of the product produced by each and determine which is less.

The calculations are as follows:

$$50.0 \cancel{g As} \times \frac{1 \cancel{mol As}}{74.92 \cancel{g As}} \times \frac{2 \cancel{mol As_2O_3}}{4 \cancel{mol As}} = 0.334 \cancel{mol As_2O_3}$$

$$50.0 \cancel{g O_2} \times \frac{1 \cancel{mol O_2}}{32.00 \cancel{g O_2}} \times \frac{2 \cancel{mol As_2O_3}}{3 \cancel{mol O_2}} = 1.04 \cancel{mol As_2O_3}$$

Comparing these two answers, it is clear that 0.334 mol of As₂O₃ is less than 1.04 mol of As₂O₃, so arsenic is the limiting reagent. If this reaction is performed under these initial conditions, the arsenic will run out before the oxygen runs out. We say that the oxygen is "in excess."

Identifying the limiting reagent, then, is straightforward. However, there are usually two associated questions: (1) what mass of product (or products) is then actually formed? and (2) what mass of what reactant is left over? The first question is straightforward to answer: simply perform a conversion from the number of moles of product formed to its mass, using its molar mass. For As_2O_3 , the molar mass is 197.84 g/mol; knowing that we will form 0.334 mol of As_2O_3 under the given conditions, we will get

$$0.334 \cancel{\text{mol As}_2\text{O}_3} \times \frac{197.84 \text{ g As}_2\text{O}_3}{1 \cancel{\text{mol As}_2\text{O}_3}} = 66.1 \text{ g As}_2\text{O}_3$$

The second question is somewhat more convoluted to answer. First, we must do a mass-mass calculation relating the limiting reagent (here, As) to the other reagent (O_2). Once we determine the mass of O_2 that reacted, we subtract that from the original amount to determine the amount left over. According to the mass-mass calculation,

$$50.0 \cancel{\text{g As}} \times \frac{1 \cancel{\text{mol As}}}{74.92 \cancel{\text{g As}}} \times \frac{3 \cancel{\text{mol O}_2}}{4 \cancel{\text{mol As}}} \times \frac{32.00 \text{ g O}_2}{1 \cancel{\text{mol O}_2}} = 16.0 \text{ g O}_2 \text{ reacted}$$

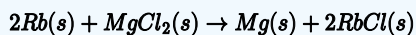
Because we reacted 16.0 g of our original O_2 , we subtract that from the original amount, 50.0 g, to get the mass of O_2 remaining:

$$50.0 \text{ g O}_2 - 16.0 \text{ g O}_2 \text{ reacted} = 34.0 \text{ g O}_2 \text{ left over}$$

You must remember to perform this final subtraction to determine the amount remaining; a common error is to report the 16.0 g as the amount remaining.

✓ Example 8.9.1

A 5.00 g quantity of Rb is combined with 3.44 g of MgCl_2 according to this chemical reaction:



What mass of Mg is formed, and what mass of what reactant is left over?

Solution

Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less.

$$5.00 \cancel{\text{g Rb}} \times \frac{1 \cancel{\text{mol Rb}}}{85.47 \cancel{\text{g Rb}}} \times \frac{1 \cancel{\text{mol Mg}}}{2 \cancel{\text{mol Rb}}} \times \frac{24.31 \text{ g Mg}}{1 \cancel{\text{mol Mg}}} = 0.711 \text{ g Mg}$$

$$3.44 \cancel{\text{g MgCl}_2} \times \frac{1 \cancel{\text{mol MgCl}_2}}{95.21 \cancel{\text{g MgCl}_2}} \times \frac{1 \cancel{\text{mol Mg}}}{1 \cancel{\text{mol MgCl}_2}} \times \frac{24.31 \text{ g Mg}}{1 \cancel{\text{mol Mg}}} = 0.878 \text{ g Mg}$$

The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of MgCl_2 reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount.

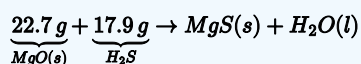
$$5.00 \cancel{\text{g Rb}} \times \frac{1 \cancel{\text{mol Rb}}}{85.47 \cancel{\text{g Rb}}} \times \frac{1 \cancel{\text{mol MgCl}_2}}{2 \cancel{\text{mol Rb}}} \times \frac{95.21 \text{ g MgCl}_2}{1 \cancel{\text{mol MgCl}_2}} = 2.78 \text{ g MgCl}_2 \text{ reacted}$$

Because we started with 3.44 g of MgCl_2 , we have

$$3.44 \text{ g MgCl}_2 - 2.78 \text{ g MgCl}_2 \text{ reacted} = 0.66 \text{ g MgCl}_2 \text{ left}$$

? Exercise 8.9.1

Given the initial amounts listed, what is the limiting reagent, and what is the mass of the leftover reagent?



Answer

H₂S is the limiting reagent; 1.5 g of MgO are left over.

Summary

The limiting reagent is the reactant that produces the least amount of product. Mass-mass calculations can determine how much product is produced and how much of the other reactants remain.

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8.10: Yields

Learning Objective

- Define and determine theoretical yields, actual yields, and percent yields.

In all the previous calculations we have performed involving balanced chemical equations, we made two assumptions:

- The reaction goes exactly as written.
- The reaction proceeds completely.

In reality, such things as side reactions occur that make some chemical reactions rather messy. For example, in the actual combustion of some carbon-containing compounds, such as methane, some CO is produced as well as CO_2 . However, we will continue to ignore side reactions, unless otherwise noted. The second assumption, that the reaction proceeds completely, is more troublesome. Many chemical reactions do not proceed to completion as written, for a variety of reasons (some of which we will consider in Chapter 13). When we calculate an amount of product assuming that all the reactant reacts, we calculate the **theoretical yield**, an amount that is theoretically produced as calculated using the balanced chemical reaction.

In many cases, however, this is not what really happens. In many cases, less—sometimes, much less—of a product is made during the course of a chemical reaction. The amount that is actually produced in a reaction is called the **actual yield**. By definition, the actual yield is less than or equal to the theoretical yield. If it is not, then an error has been made.

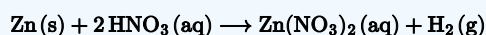
Both theoretical yields and actual yields are expressed in units of moles or grams. It is also common to see something called a percent yield. The **percent yield** is a comparison between the actual yield and the theoretical yield and is defined as

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \quad (8.10.1)$$

It does not matter whether the actual and theoretical yields are expressed in moles or grams, as long as they are expressed in the same units. However, the percent yield always has units of percent. Proper percent yields are between 0% and 100%. Again, if percent yield is greater than 100%, an error has been made.

✓ Example 8.10.1

A worker reacts 30.5 g of Zn with nitric acid and evaporates the remaining water to obtain 65.2 g of $\text{Zn}(\text{NO}_3)_2$. What are the theoretical yield, the actual yield, and the percent yield?



Solution

A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of Zn (65.39 g/mol) and $\text{Zn}(\text{NO}_3)_2$ (189.41 g/mol). In three steps, the mass-mass calculation is:

$$30.5 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{1 \text{ mol Zn}(\text{NO}_3)_2}{1 \text{ mol Zn}} \times \frac{189.41 \text{ g Zn}(\text{NO}_3)_2}{1 \text{ mol Zn}(\text{NO}_3)_2} = 88.3 \text{ g Zn}(\text{NO}_3)_2$$

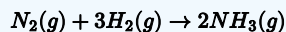
Thus, the theoretical yield is 88.3 g of $\text{Zn}(\text{NO}_3)_2$. The actual yield is the amount that was actually made, which was 65.2 g of $\text{Zn}(\text{NO}_3)_2$. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100 (Equation 8.10.1):

$$\frac{65.2 \text{ g Zn}(\text{NO}_3)_2}{88.3 \text{ g Zn}(\text{NO}_3)_2} \times 100\% = 73.8\%$$

The worker achieved almost three-fourths of the possible yield.

? Exercise 8.10.1

A synthesis produced 2.05 g of NH_3 from 16.5 g of N_2 . What is the theoretical yield and the percent yield?



*Technically, this is a reversible reaction (with double arrows), but for this exercise consider it irreversible (single arrow).

Answer

theoretical yield = 20.1 g; percent yield = 10.2%

📌 Chemistry is Everywhere: Actual Yields in Drug Synthesis and Purification

Many drugs are the product of several steps of chemical synthesis. Each step typically occurs with less than 100% yield, so the overall percent yield might be very small. The general rule is that the overall percent yield is the product of the percent yields of the individual synthesis steps. For a drug synthesis that has many steps, the overall percent yield can be very tiny, which is one factor in the huge cost of some drugs. For example, if a 10-step synthesis has a percent yield of 90% for each step, the overall yield for the entire synthesis is only 35%. Many scientists work every day trying to improve percent yields of the steps in the synthesis to decrease costs, improve profits, and minimize waste.

Even purifications of complex molecules into drug-quality purity are subject to percent yields. Consider the purification of impure albuterol. Albuterol ($\text{C}_{13}\text{H}_{21}\text{NO}_2$; accompanying figure) is an inhaled drug used to treat asthma, bronchitis, and other obstructive pulmonary diseases. It is synthesized from norepinephrine, a naturally occurring hormone and neurotransmitter. Its initial synthesis makes very impure albuterol that is purified in five chemical steps. The details of the steps do not concern us; only the percent yields do:

A child using an albuterol inhaler, a container of albuterol medication, and a molecular model of albuterol are shown in three combined images.

impure albuterol → intermediate A	percent yield = 70%
intermediate A → intermediate B	percent yield = 100%
intermediate B → intermediate C	percent yield = 40%
intermediate C → intermediate D	percent yield = 72%
intermediate D → purified albuterol	percent yield = 35%
overall percent yield = $70\% \times 100\% \times 40\% \times 72\% \times 35\% = 7.5\%$	

That is, only about *one-fourteenth* of the original material was turned into the purified drug. This demonstrates one reason why some drugs are so expensive—a lot of material is lost in making a high-purity pharmaceutical.

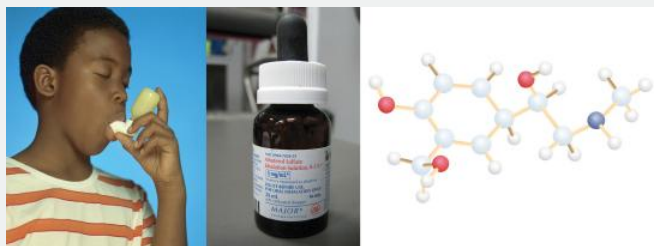


Figure 8.10.1 A child using an albuterol inhaler, the container of albuterol medication, and a molecular model of the albuterol molecule. Source: Photo on far left © Thinkstock. Photo in center courtesy of Intropin, http://commons.wikimedia.org/wiki/File:Albuterol_molecule.png.

Summary

Theoretical yield is the calculated yield using the balanced chemical reaction. Actual yield is what is actually obtained in a chemical reaction. Percent yield is a comparison of the actual yield with the theoretical yield.

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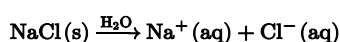
8.11: Ionic Equations

Learning Objectives

- Write ionic equations for chemical reactions between ionic compounds.
- Write net ionic equations for chemical reactions between ionic compounds.

For single-replacement and double-replacement reactions, many of the reactions included ionic compounds—compounds between metals and nonmetals, or compounds that contained recognizable polyatomic ions. Now, we take a closer look at reactions that include ionic compounds.

One important aspect about ionic compounds that differs from molecular compounds has to do with dissolution in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process—for example, with NaCl:



When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 8.11.1). This process is called **dissociation**; we say that the ions *dissociate*.

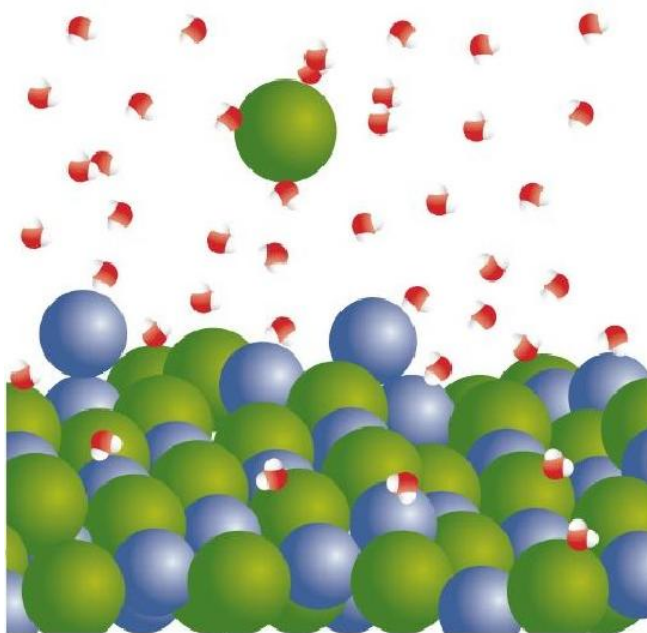
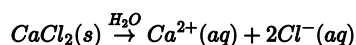
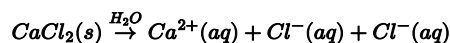


Figure 8.11.1 Ionic Solutions. When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. This behavior was first suggested by the Swedish chemist Svante August Arrhenius [1859–1927] as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry. Keep in mind that when the ions separate, *all* the ions separate. Thus, when CaCl_2 dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from one another:



That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions), and they do not stick together to make Cl_2^- or Cl_2^{2-} . They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

✓ Example 8.11.1

Write the chemical equation that represents the dissociation of each ionic compound.

1. KBr
2. Na_2SO_4

Solution

1. $\text{KBr(s)} \rightarrow \text{K}^+(\text{aq}) + \text{Br}^-(\text{aq})$
2. Not only do the two sodium ions go their own way, but the sulfate ion stays together as the sulfate ion. The dissolving equation is $\text{Na}_2\text{SO}_4(\text{s}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$

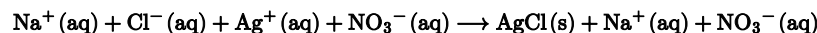
? Exercise 8.11.1

Write the chemical equation that represents the dissociation of $(\text{NH}_4)_2\text{S}$.

Answer



When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when NaCl(aq) reacts with $\text{AgNO}_3(\text{aq})$ in a double-replacement reaction to precipitate AgCl(s) and form $\text{NaNO}_3(\text{aq})$, the complete ionic equation includes NaCl , AgNO_3 , and NaNO_3 written as separate ions:



This is more representative of what is occurring in the solution.

✓ Example 8.11.1

Write the complete ionic equation for each chemical reaction.

1. $\text{KBr(aq)} + \text{AgC}_2\text{H}_3\text{O}_2(\text{aq}) \rightarrow \text{KC}_2\text{H}_3\text{O}_2(\text{aq}) + \text{AgBr(s)}$
2. $\text{MgSO}_4(\text{aq}) + \text{Ba(NO}_3)_2(\text{aq}) \rightarrow \text{Mg(NO}_3)_2(\text{aq}) + \text{BaSO}_4(\text{s})$

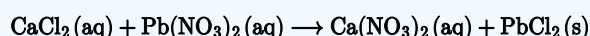
Solution

For any ionic compound that is aqueous, we will write the compound as separated ions.

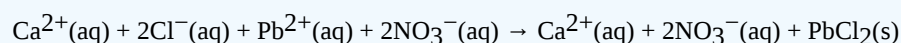
1. The complete ionic equation is $\text{K}^+(\text{aq}) + \text{Br}^-(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{AgBr(s)}$
2. The complete ionic equation is $\text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{BaSO}_4(\text{s})$

? Exercise 8.11.1

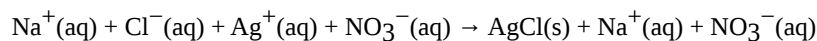
Write the complete ionic equation for



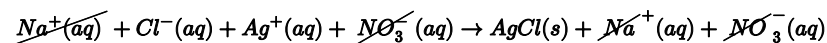
Answer



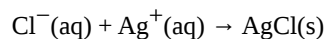
You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in



the $\text{Ag}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions become $\text{AgCl}(\text{s})$, but the $\text{Na}^+(\text{aq})$ ions and the $\text{NO}_3^-(\text{aq})$ ions stay as $\text{Na}^+(\text{aq})$ ions and $\text{NO}_3^-(\text{aq})$ ions. These two ions are examples of spectator ions—ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:



What remains when the spectator ions are removed is called the **net ionic equation**, which represents the actual chemical change occurring between the ionic compounds:



It is important to reiterate that the spectator ions are still present in solution, but they do not experience any net chemical change, so they are not written in a net ionic equation.

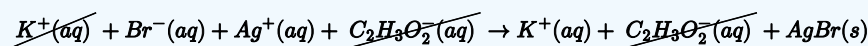
✓ Example 8.11.1

Write the net ionic equation for each chemical reaction.

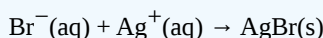
- $\text{K}^+(\text{aq}) + \text{Br}^-(\text{aq}) + \text{Ag}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) \rightarrow \text{K}^+(\text{aq}) + \text{C}_2\text{H}_3\text{O}_2^-(\text{aq}) + \text{AgBr}(\text{s})$
- $\text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + \text{BaSO}_4(\text{s})$

Solution

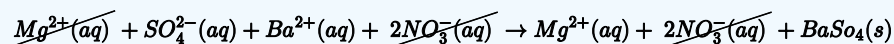
- In the first equation, the $\text{K}^+(\text{aq})$ and $\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$ ions are spectator ions, so they are canceled:



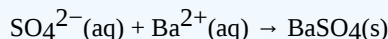
The net ionic equation is



- In the second equation, the $\text{Mg}^{2+}(\text{aq})$ and $\text{NO}_3^-(\text{aq})$ ions are spectator ions, so they are canceled:

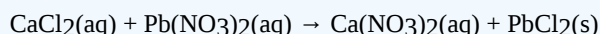


The net ionic equation is

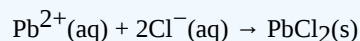


? Exercise 8.11.1

Write the net ionic equation for



Answer



✓ Chemistry is Everywhere: Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of compound that can be dissolved in a sample of water. For example, you can dissolve a

maximum of 36.0 g of NaCl in 100 g of water at room temperature, but you can dissolve only 0.00019 g of AgCl in 100 g of water. We consider **NaCl** soluble but **AgCl** insoluble.

One place where solubility is important is in the tank-type water heater found in many homes in the United States. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate (CaCO_3). However, CaCO_3 has the relatively unusual property of being less soluble in hot water than in cold water. So as the water heater operates by heating water, CaCO_3 can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.



Figure 8.11.2 Tank water heater © Thinkstock. Most homes in the United States have a tank-type water heater like this one.

Another place where solubility versus insolubility is an issue is the Grand Canyon. We usually think of rock as insoluble. But it is actually ever so slightly soluble. This means that over a period of about two billion years, the Colorado River carved rock from the surface by slowly dissolving it, eventually generating a spectacular series of gorges and canyons. And all because of solubility!



Figure 8.11.3: The Grand Canyon was formed by water running through rock for billions of years, very slowly dissolving it. Note the Colorado River is still present in the lower part of the photo. (Sonaal Bangera via [unsplash](#))

Key Takeaways

- Ionic compounds that dissolve separate into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

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