INTRODUCTION TO CHEMISTRY



Lance Lund
Anoka-Ramsey Community College

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Introduction to Chemistry

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Licensing

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Periodic Tables

Versions of Periodic Table below:

- PNG Image Version (Figure 1)
- Downloadable PDF Version (Figure 2)
- Interactive Periodic Table from PubChem (Figure 3)

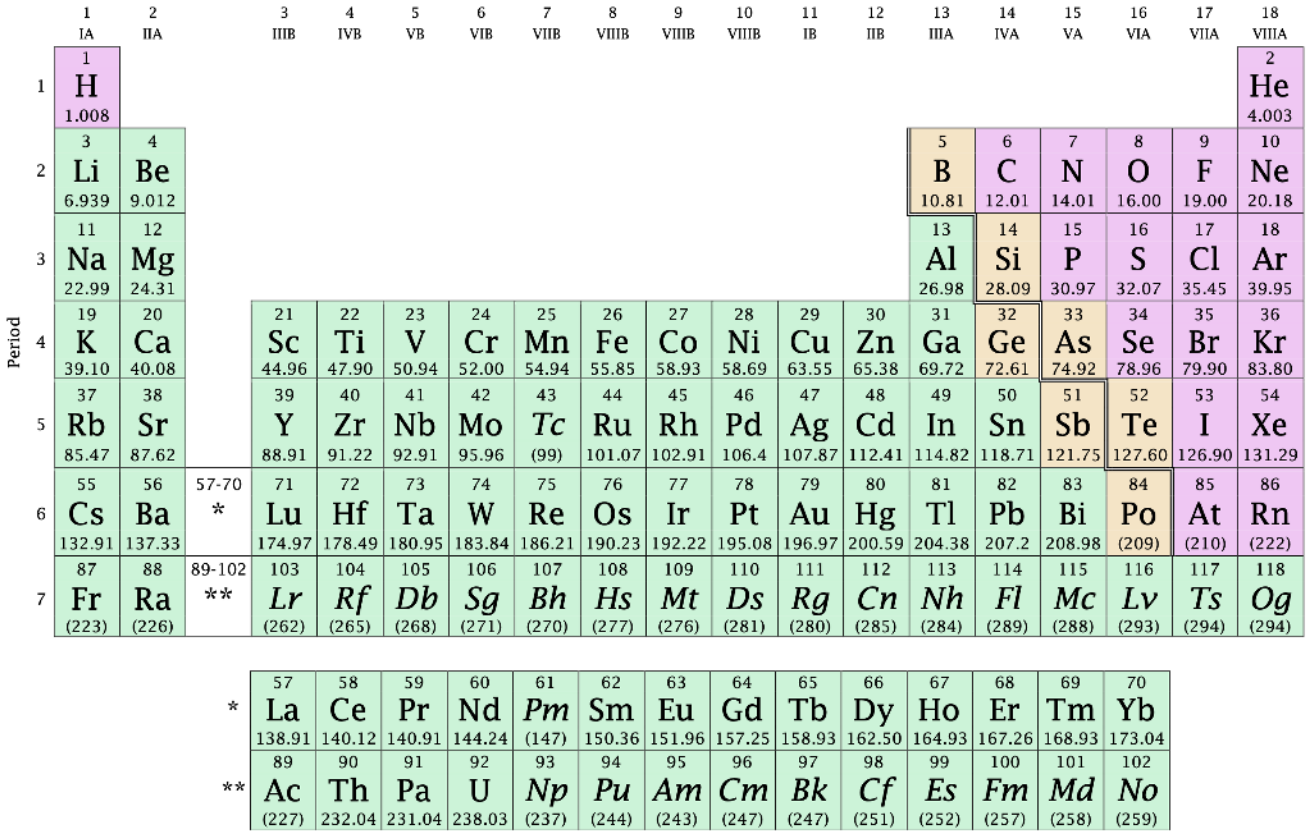


Figure 1: PNG Version of the Periodic Table.

Figure 2: Downloadable PDF Version of the Periodic Table.

Periodic Table of Elements

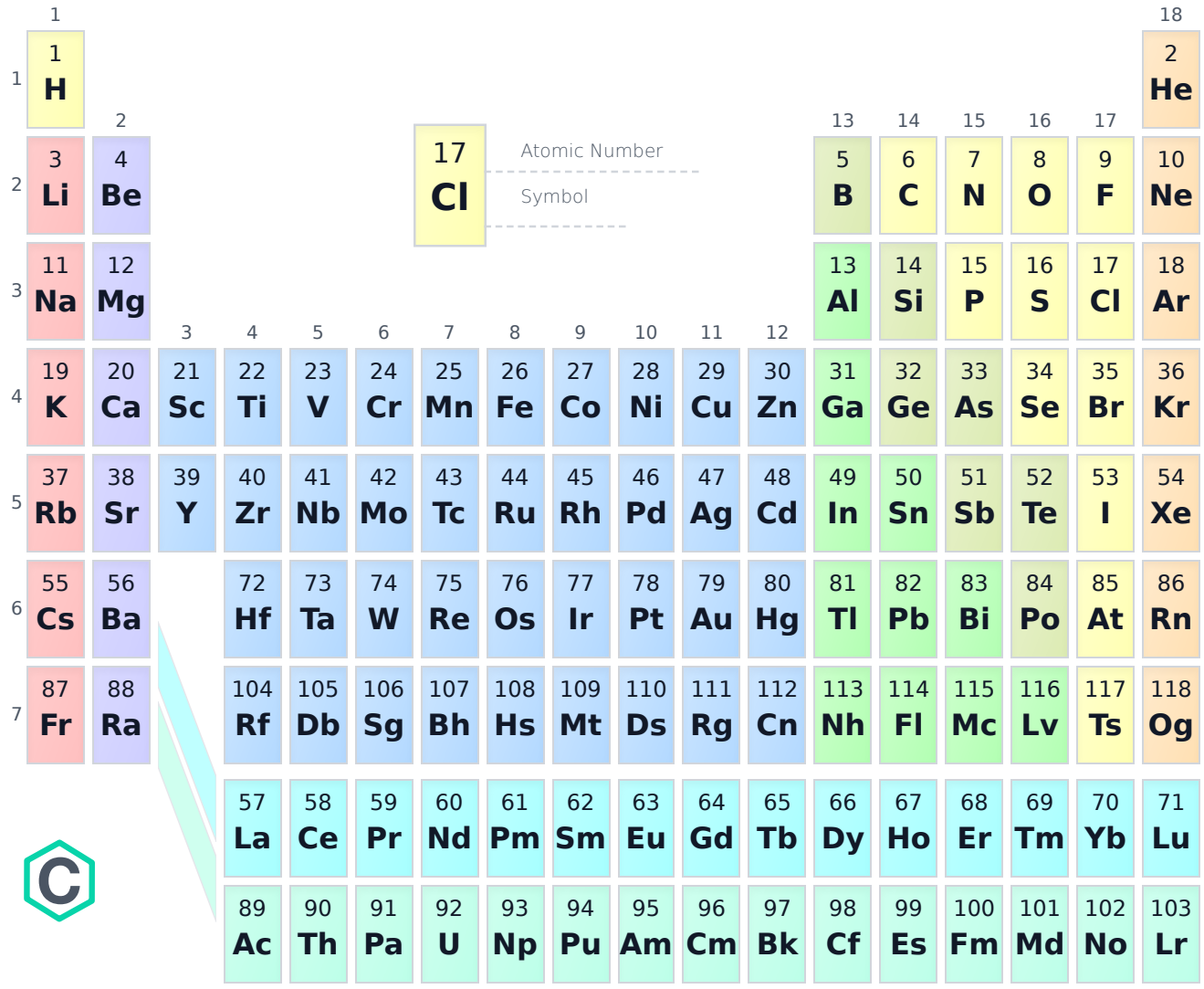


Figure **3**: An interactive modern periodic table. The displayed property may be changed by using the dropdown menu to the left of helium, He. (Public Domain; PubChem via NIH)

Conversions and Constants

Versions of Conversion Tables below:

- PNG Version (Figure 1)
- Downloadable PDF Version (Figure 2)

Common Chemistry Conversions

English to Metric Conversions

The mass, length, volume, and area conversions are given to 4 significant figures. The temperature conversions are exact.

Mass	Length	Volume	Area	Temperature
1 lb = 453.6 g	1 in. = 2.540 cm	1 fl oz = 29.57 mL	1 $in.^2$ = $(2.54\ cm)^2$ = 6.452 cm^2	$T_{°C} = \frac{5}{9}(T_{°F} - 32)$
1 oz = 28.35 g	1 ft = 30.48 cm	1 L = 1.057 qt	1 m^2 = $(3.281\ ft)^2$ = 10.76 ft^2	$T_{°F} = \frac{9}{5}T_{°C} + 32$
1 kg = 2.205 lb	1 m = 3.281 ft	1 gal = 3.785 L		$T_K = T_{°C} + 273.15$
1 metric ton = 1000 kg	1 mi = 1.609 km	1 in^3 = $(2.54\ cm)^3$ = 16.39 cm^3		

English to English Conversions. *All of these conversions are exact.*

Mass	Length	Volume	Area
1 lb = 16 oz	1 ft = 12 in.	1 cup = 8 fl oz	1 ft^2= $(12\ in)^2$ = 144 $in.^2$
1 ton = 2000 lb	1 yd = 3 ft	1 pt = 2 cups	1 mi^2 = 640 acres
	1 mi = 5280 ft	1 qt = 2 pt	
		1 gal = 4 qt	

Other Conversions

Energy	Pressure
1 cal = 4.184 J	1 atm = 760 mm Hg = 760 torr = 29.92 in. Hg
1 J = 1 $\frac{kg \cdot m^2}{s^2}$	1 atm = 14.7 psi = 101,325 Pa = 1.01325 bars
	1 Pa = 1 $\frac{kg}{m \cdot s^2}$

Constants

speed of light (in a vacuum)	$c = 2.998 \times 10^8$ m/s
Planck's constant	$h = 6.626 \times 10^{-34}$ J·s
electron mass proton mass neutron mass	$m_e = 9.109 \times 10^{-31}$ kg $m_p = 1.673 \times 10^{-27}$ kg $m_n = 1.675 \times 10^{-27}$ kg
Avogadro's number	$N_A = 6.0221367 \times 10^{23}$ particles/mol
Gas Constant	$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $= 8.315\ \frac{J}{mol \cdot K}$ $= 8.315\ \frac{kPa \cdot dm^3}{mol \cdot K}$
Faraday Constant	$F = 9.65 \times 10^4$ C/mol
Electronic charge	$e = 1.602 \times 10^{-19}$ C

Figure 1: Common Chemistry Conversions. (Lance S. Lund)

Figure **2**: Common Chemistry Conversions. (Lance S. Lund)

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CHAPTER OVERVIEW

1: What is Chemistry?

1.1: The Scope of Chemistry
1.2: Chemistry in Agriculture
1.3: Chemistry and the Environment
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Template:HideTOC

Thumbnail Chapter 1: Household chemicals (MPCA Photos via Flickr)

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1.1: The Scope of Chemistry

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

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.

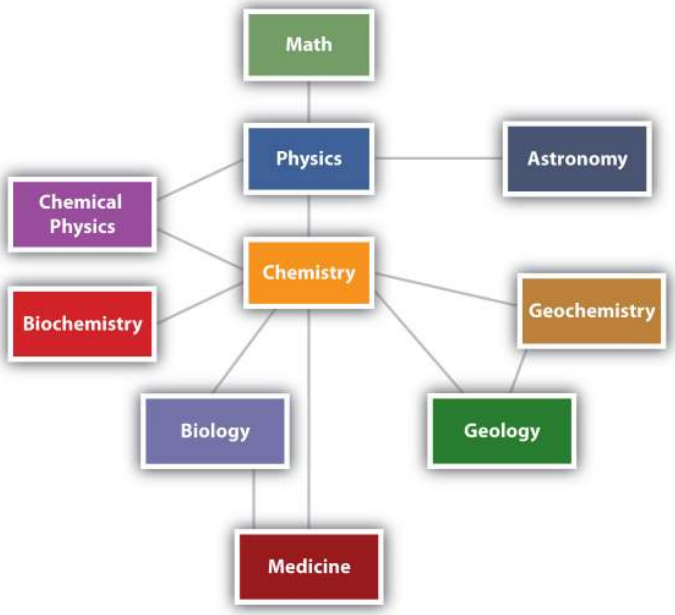


Figure **1.1.1**: The relationships between some of the major branches of science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

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. At some level, all of these fields depend on matter because they all involve "stuff"; because of this, chemistry has been called the "central science," linking them all together.

⚡ Fields of Science

There are many other fields of science, in addition to the ones (biology, medicine, etc.) listed here.

✅ Example 1.1.1

Which fields of study are branches of science? Explain.

A. sculpture
B. astronomy

Solution

A. Sculpture is not considered a science because it is not a study of some aspect of the natural universe.
B. Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

✏️ Exercise 1.1.1

Which fields of study are branches 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

yes

Answer B

yes

Answer C

yes

Answer D

no

Areas of Chemistry

The study of modern chemistry has many branches, but can generally be broken down into five main disciplines, or areas of study:

- **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- **Organic chemistry:** Organic chemistry is the study of chemicals containing carbon. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- **Inorganic chemistry:** Inorganic chemistry is the study of chemicals that, in general, are not primarily based on carbon. Inorganic chemicals are commonly found in rocks and minerals. One current important area of inorganic chemistry deals with the design and properties of materials involved in energy and information technology.
- **Analytical chemistry:** Analytical chemistry is the study of the composition of matter. It focuses on separating, identifying, and quantifying chemicals in samples of matter. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components.
- **Biochemistry:** Biochemistry is the study of chemical processes that occur in living things. Research may cover anything from basic cellular processes up to understanding disease states so that better treatments can be developed.



Figure 1.1.2: (left) Measurement of trace metals using atomic spectroscopy. (right) Measurement of hormone concentrations.

In practice, chemical research is often not limited to just one of the five major disciplines. A chemist may use biochemistry to isolate a particular chemical found in the human body such as hemoglobin, the oxygen carrying component of red blood cells. He or she may then proceed to analyze the hemoglobin using methods that would pertain to the areas of physical or analytical

chemistry. Many chemists specialize in areas that are combinations of the main disciplines, such as bioinorganic chemistry or physical organic chemistry.

History of Chemistry

The history of chemistry is an interesting and challenging one. Very early chemists were often motivated mainly by the achievement of a specific goal or product. Making perfume or soaps did not need a lot of **theory**, just a good recipe and careful attention to detail. There was no standard way of naming materials (and no periodic table that we could all agree on). It was often difficult to figure out exactly what a particular person was using. However, the science developed over the centuries by trial and error.

Major progress was made toward putting chemistry on a solid foundation when Robert Boyle (1637-1691) began his **research** in chemistry (Figure 1.1.3). He developed the basic ideas about the behavior of gases. He could then describe gases mathematically. Boyle also helped form the idea that small particles could combine to form molecules. Many years later, John Dalton used these ideas to develop the atomic theory.



Figure **1.1.3**: Robert Boyle.

The field of chemistry began to develop rapidly in the 1700's. Joseph Priestley (1733-1804) isolated and characterized several gases: oxygen, carbon monoxide, and nitrous oxide. It was later discovered that nitrous oxide ("laughing gas") worked as an anesthetic. This gas was used for that purpose for the first time in 1844 during a tooth extraction. Other gases discovered during that time were chlorine, by C.W. Scheele (1742-1786), and nitrogen, by Antoine Lavoisier (1743-1794). Lavoisier has been considered by many scholars to be the "father of chemistry." Among other accomplishments, he discovered the role of oxygen in combustion and definitively formulated the law of conservation of matter.

Chemists continued to discover new compounds in the 1800's. The science also began to develop a more theoretical foundation. John Dalton (1766-1844) put forth his atomic theory in 1807. This idea allowed scientists to think about chemistry in a much more systematic way. Amadeo Avogadro (1776-1856) laid the groundwork for a more quantitative approach to chemistry by calculating the number of particles in a given amount of a gas. A lot of effort was put forth in studying chemical reactions. These efforts led to new materials being produced. Following the invention of the battery by Alessandro Volta (1745-1827), the field of electrochemistry (both theoretical and applications) developed through major contributions by Humphry Davy (1778-1829) and Michael Faraday (1791-1867). Other areas of the discipline also progressed rapidly.

It would take a large book to cover developments in chemistry during the twentieth century and up to today. One major area of expansion was in the area of the chemistry of living processes. Research in photosynthesis in plants, the discovery and characterization of enzymes as biochemical catalysts, elucidation of the structures of biomolecules such as insulin and DNA—these efforts gave rise to an explosion of information in the field of biochemistry.

The practical aspects of chemistry were not ignored. The work of Volta, Davy, and Faraday eventually led to the development of batteries that provided a source of electricity to power a number of devices (Figure 1.1.4).



Figure **1.1.4**: Battery developed by Volta.

Charles Goodyear (1800-1860) discovered the process of vulcanization, allowing a stable rubber product to be produced for the tires of all the vehicles we have today. Louis Pasteur (1822-1895) pioneered the use of heat sterilization to eliminate unwanted microorganisms in wine and milk. Alfred Nobel (1833-1896) invented dynamite (Figure 1.1.5). After his death, the fortune he made from this product was used to fund the Nobel Prizes in science and the humanities. J.W. Hyatt (1837-1920) developed the first plastic. Leo Baekeland (1863-1944) developed the first synthetic resin, widely used for inexpensive and sturdy dinnerware.



Figure **1.1.5**: Dynamite explosion in Panama, Central America (1908).

Today, chemistry continues to be essential to the development of new materials and technologies, from semiconductors for electronics to powerful new medicines, and beyond. Chemists explore the macroscopic world – the world we can see – to understand the submicroscopic world that we cannot see – the realm of atoms and compounds.

Summary

- Chemistry is the study of matter and the changes it undergoes and considers both *macroscopic* and *submicroscopic* information.
- Matter is anything that has mass and occupies space.
- The five main disciplines of chemistry are physical chemistry, organic chemistry, inorganic chemistry, analytical chemistry and biochemistry.
- Many civilizations contributed to the growth of chemistry. A lot of early chemical research focused on practical uses. Basic chemistry theories were developed during the nineteenth century. New materials and batteries are a few of the products of modern chemistry.

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1.2: Chemistry in Agriculture

In the spring, many people begin to plant their gardens. They see ads in catalogs or shop the gardening section of a local store to get ideas. The right place in the garden is selected, seeds or plants are put in the soil, and then they wait. Whether it is a small home garden or a large thousand acre farm, chemistry contributes greatly to the success of the crop.

Agriculture

Crops need three things for good growth: **water**, **nutrients** from the soil, and **protection** from predators such as insects. Chemistry has made major contributions in all three areas. Water purification uses a number of chemical and physical techniques to remove salts and contaminants that would pollute the soil. Chemical analysis of soil allows the grower to see what nutrients are lacking, so that they can be added. In the spring, grocery stores, hardware stores, and gardening centers have high stacks of bags containing fertilizers and weed killers that enrich the soil and keep down unwanted plants. These same stores also provide a number of sprays or solid treatments for insects that might otherwise eat the plants.



Figure 1.2.1: Idaho wheat field.

Water Purification

Fresh water is essential for good crops. In some areas of the world, there is enough rain to accomplish this task. In other locales, water must be provided so the crops will grow. Worldwide, irrigation covers about 18% of farm land and produces some 40% of crops. Obtaining clean water in many parts of the world is made possible by the process of **desalinization**.



Figure 1.2.2: Desalinization equipment.

Sea water is treated to remove salts and the resulting water can then be used for irrigation without contaminating the soil with materials that harm the growing plants.

Soil Nutrients

In many areas of the world, the soil is deficient in essential nutrients. A number of minerals such as phosphorus, potassium, calcium, and magnesium may not be present in large enough amounts for plants to grow well. Nitrogen is also extremely important for good crops.



Figure 1.2.3: Man spreading chemicals on soil.

Soil analysis is available from a variety of labs. Local university extension services can provide valuable information as to the composition of a soil, and will also make suggestions as to the types and amounts of needed nutrients. Fertilizers can be purchased and added to the soil to enrich it and ensure better yield of crops.

Insect Control

Even if the crop grows well, there is still the possibility of insect or pest damage. The insect or pest can consume the crop or can damage it to the point where it will not grow well. Infestations of army worms can do major damage to corn and grain crops. Aphids and boll weevils are major predators of cotton crops. Failure to control these pests results in widespread crop damage and financial loss for the farmer.

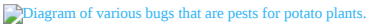


Figure **1.2.4**: Potato pests.

A wide variety of pesticides have been developed by chemists and other scientists to deal with pests. The basic approach is to develop pesticides that interfere with some biochemical process in the pest. Ideally, the pesticide will not affect other living organisms, but this is not always the case. It is very important to read the labels and observe all precautions when using pesticides.

Summary

- Sea water may be purified through the process of desalination.
- Plant nutrients are very important for good plant growth.
- Chemical analysis of soil can tell a farmer or gardener what nutrients the soil needs.
- Chemists have developed many pesticides that will kill plant predators such as the army worm and the boll weevil.

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1.3: Chemistry and the Environment

Where do scientists dump garbage? Where should scientists put unwanted waste materials, after they have made what they want? For many years, scientists simply dumped the waste into the nearest lake, river, or empty spot of ground, with no second thoughts. However, as scientists learn more about the effects of chemicals on living systems, they have become more concerned about the effects of dumping. In many cases, scientists did not know what the long-term effects would be. Today, dumping is generally illegal, and offenders pay heavy penalties of fines and jail sentences.

Environment

Many chemicals that were once commonly used were later found out to be harmful to the **environment**, to human health, or to both. The element lead was once a common additive to gasoline and to paint. Plumbing pipes were once typically made of lead. Only since the 1970s has the danger of lead become apparent. It causes brain damage; small children (who often chewed on objects painted with lead-based paint) are particularly susceptible. The use of lead in gasoline, paint, and plumbing pipes is now banned; new materials are being developed to replace hazardous lead components.

Lead is still widely used in car batteries. In recent years, battery **recycling** has become common – about 98% of car batteries are recycled today. The used batteries go to a processing plant where they are crushed and placed in a tank. The lead sinks to the bottom and can be separated, and melted. The plastic can then be used in a number of applications. The battery acids are treated with chemicals to neutralize them before they are disposed of.



Figure **1.3.1**: Batteries for recycling.

How Chemists Help the Environment

Chemists are involved in all aspects of environmental protection. In the case of lead, chemists measured the amount of lead in soil, paint, plumbing, blood, and other materials. Chemists studied the chemical processes in the body to see how lead did its damage. Methods for removal of lead from the body were developed with the help of chemists. New gasoline products were developed that fostered fuel efficiency and smooth running of engines, without the presence of lead.

Chemists continue to look for threats to our health and the environment – and to search for alternatives – so that harmful chemicals can be replaced with others that will do the job as effectively, but without the harm.

Summary

- Many chemicals that were once commonly used were later found out to be harmful to the **environment**, to human health, or to both.
- Lead contamination has serious harmful effects on the body.
- Chemists contribute in many ways to helping with issues of lead contamination.

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1.4: Chemistry and Society

Examples of the practical applications of chemistry are everywhere (Figure 1.4.1). Engineers need to understand the chemical properties of the substances needed to design biologically compatible implants for joint replacements; or to design roads, bridges, buildings, and nuclear reactors that do not collapse because of weakened structural materials such as steel and cement. Archeology and paleontology rely on chemical techniques to date bones and artifacts and identify their origins. Although law is not normally considered a field related to chemistry, forensic scientists use chemical methods to analyze blood, fibers, and other evidence as they investigate crimes.

In particular, DNA matching – comparing biological samples of genetic material to see whether they could have come from the same person – has been used to solve many high-profile criminal cases as well as clear innocent people who have been wrongly accused or convicted. Forensics is a rapidly growing area of applied chemistry. In addition, the proliferation of chemical and biochemical innovations in industry is producing rapid growth in the area of patent law. Ultimately, the dispersal of information in all of the fields in which chemistry plays a part requires experts who are able to explain complex chemical issues to the public through television, print journalism, the Internet, and popular books.

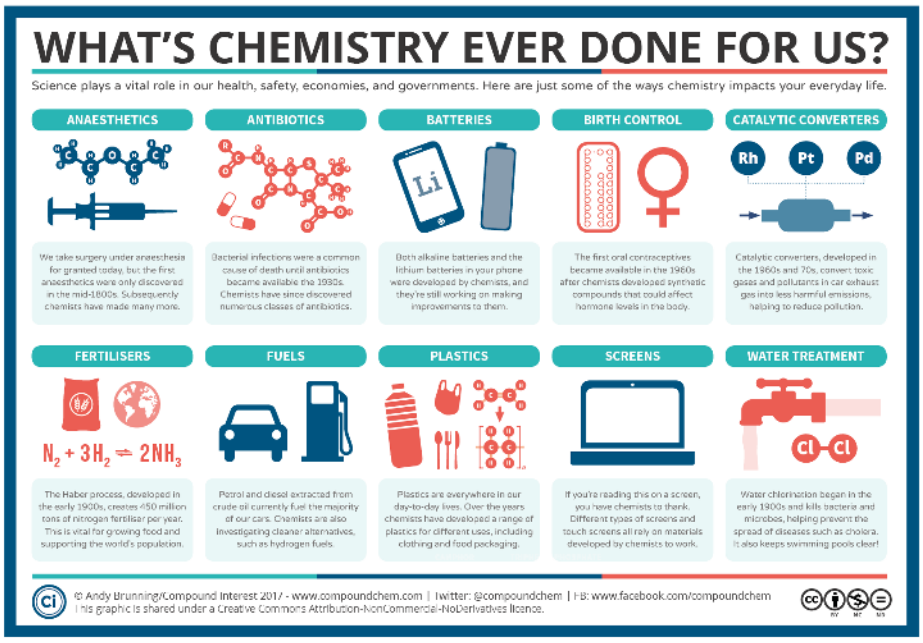


Figure 1.4.1: Chemistry and chemical compounds play an important role in almost everything around us. (Andy Brunning via compoundchem.com)

Hopefully at this point you are fully convinced of how important and useful the study of chemistry can be. You may, however, still be wondering exactly *what* it is that a chemist does. Chemistry is the study of matter and the changes that matter undergoes. In general, chemists are interested in both characteristics that you can test and observe, like a chemical's smell or color, and characteristics that are far too small to see, like what the oxygen you inhale or the carbon dioxide you exhale looks like under a microscope 1,000 times more powerful than any existing in the world today.

How can a chemist *know* what oxygen and carbon dioxide look like under a microscope that doesn't even exist? What happened to the scientific method? What happened to relying on *observations* and *careful measurements*? In fact, because chemists *can't* see the underlying structure of different materials, they have to rely on the scientific method even more! Chemists are a lot like detectives. Suppose a detective is trying to solve a murder case – what do they do? Obviously, the detective starts by visiting the site of the crime and looking for evidence. If the murderer has left enough clues behind, the detective can piece together a theory explaining what happened.

Even though the detective wasn't at the crime scene when the crime was committed and didn't actually see the murderer kill the victim, with the right evidence the detective can be pretty sure of how the crime took place. It is the same with chemistry. When chemists go into the laboratory, they collect evidence by making measurements. Once chemists have collected enough clues from the properties that they can observe, they use that evidence to piece together a theory explaining the properties that they cannot observe – the properties that are too small to see.

⚡ Scientific Method

Chemists use the scientific method, which is a process of establishing facts through testing and experimentation. The basic process involves asking a question, forming a hypothesis, conducting an experiment, and then analyzing the results.

What kinds of properties do chemists actually measure in the laboratory? Well, you can probably guess a few. Imagine that you go to dinner at a friend's house and are served something that you don't recognize, what types of observations might you make to determine exactly *what* you've been given? You might smell the food. You might note the color of the food. You might try to decide whether the food is a liquid or a solid because if it's a liquid, it's probably soup or a drink. The temperature of the food could be useful if you wanted to know whether or not you had been served ice cream! You could also pick up a small amount of food with your fork and try to figure out how much it weighs – a light dessert might be something like an angel cake, while a heavy dessert is probably a pound cake. The quantity of food you have been given might be a clue too. Finally, you might want to know something about the food's texture – is it hard and granular like sugar cubes, or soft and easy to spread, like butter?

Believe it or not, the observations you are likely to make when trying to identify an unknown food are very similar to the observations that a chemist makes when trying to learn about a new material. Chemists rely on smell, color, state (whether it is a solid or liquid or gas), temperature, volume, mass (which is related to weight – as will be discussed in a later section), and texture. There is, however, one property possibly used to learn about a food, but that should *definitely* not be used to learn about a chemical – taste!

In the sections on the Atomic Theory, you will see exactly how measurements of certain properties helped early scientists to develop theories about the chemical structure of matter on a scale much smaller than they could ever hope to see. You will also learn how these theories, in turn, allow us to make predictions about new materials that humankind has not yet created.

The video below gives you some important tips on how to study chemistry in this class. With practice, you too can learn to think like a chemist, and you may even enjoy it!



Video **1.4.1**: How To Study Chemistry.

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1.6: Exercises

1.1: The Scope of Chemistry

1. Identify each as either matter or not matter.
    a. a book
    b. hate
    c. light
    d. a car
    e. a fried egg

Answer

    a. matter
    b. not matter
    c. not matter
    d. matter
    e. matter

2. Name an example of a field that is not considered a science.

3. Which of the following fields of study are branches of science?
    a. biophysics (a mix of biology and physics)
    b. art
    c. business

Answer

    a. yes
    b. no
    c. no

4. Which of the following fields of study are branches of science?
    a. accounting
    b. geochemistry (a mix of geology and chemistry)
    c. astronomy (the study of stars and planets [but not the earth])

Answer

    a. no
    b. yes
    c. yes

Additional Exercises

5. Describe the scientific method.

Answer

Simply stated, the scientific method includes three steps: (1) stating a hypothesis, (2) testing the hypothesis, and (3) refining the hypothesis.

6. "A hypothesis is just a guess." Is this an adequate definition?

Answer

No, it is not. A hypothesis is a proposed explanation for something that can actually be tested. It is based on data that is available. A guess is not necessarily able to be tested.

7. Why do scientists need to perform experiments?

Answer

Scientists perform experiments to test their hypotheses because sometimes the nature of natural universe is not obvious.

8. What is the scientific definition of a theory? How is this word misused in general conversation?

Answer

A theory in science is an explanation of some aspect of the natural world that has been substantiated through repeated experiments or testing. In general conversation, a theory is an idea that has not been tested.

9. What is the scientific definition of a law? How does it differ from the everyday definition of a law?

Answer

A scientific law is a specific statement that is thought to be never violated by the entire natural universe. Everyday laws are arbitrary limits that society puts on its members.

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CHAPTER OVERVIEW

2: Measurements

Chemistry, like all sciences, is quantitative. It deals with *quantities*, things that have amounts and units. Dealing with quantities is very important in chemistry, as is relating quantities to each other. In this chapter, we will discuss how we deal with numbers and units, including how they are combined and manipulated.

2.1: Taking Measurements
2.2: Scientific Notation
2.3: Significant Figures
2.4: Significant Figures in Calculations
2.5: The Metric System
2.6: Dimensional Analysis
2.7: Multi-Step Conversions
2.8: Units Raised to a Power
2.9: Density
2.10: For Future Use
2.11: Exercises

Template:HideTOC

Thumbnail Chapter 2: Balance scale (winnifredxoxo via Flickr)

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2.1: Taking Measurements

Learning Objectives

- Express quantities properly, using a number and a unit.

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


US Navy 050310-N-1159B-114 Lt. Cmdr. Lisa Kromanaker, a nurse assigned to Operational Health Support Unit Great Lakes, Ill., measures the blood pressure of a Nicaraguan woman during a Medical Readiness Training Exercise.

Figure **2.1.1**: Lt. Cmdr. Lisa Kromanaker measures the blood pressure during a Medical Readiness Training Exercise.

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

✅ Example 2.1.1

Identify the number and the unit in each quantity.

A. one dozen eggs
B. 2.54 centimeters
C. a box of pencils
D. 88 meters per second

Solution

A. The number is one dozen, and the unit is eggs.
B. The number is 2.54, and the unit is centimeter.
C. The number 1 is implied because the quantity is only *a* box. The unit is box of pencils.
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.

Summary

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

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2.2: Scientific Notation

Astronomers are accustomed to really big numbers. While the moon is only 406,697 km from earth at its maximum distance, the sun is much further away (150 million km). Proxima Centauri, the star nearest the earth, is 39,900,000,000,000 km away – and this is just the start of long distances. On the other end of the scale, some biologists deal with very small numbers: a typical fungus could be as small as 30 micrometers (0.000030 meters) in length and a virus might only be 0.03 micrometers (0.00000003 meters) long.



Figure **2.2.1**: The sun. (NASA/SDO (AIA), Public domain, via Wikimedia Commons)

A system called **scientific notation** avoids much of the tedium and awkwardness of dealing with numbers that are very large or very small. Consider the numbers 1,000,000, 3,000,000, and 6,700,000:

Number	Also	Meaning	Scientific Notation
1,000,000	$1 \times 1{,}000{,}000$	$\mathbf{1 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10}$	1×10^6
3,000,000	$3 \times 1{,}000{,}000$	$\mathbf{3 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10}$	3×10^6
6,700,000	$6.7 \times 1{,}000{,}000$	$\mathbf{6.7 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10}$	6.7×10^6

Likewise, consider the numbers 0.000001, 0.000003, and 0.0000067:

Number	Also	Meaning	Scientific Notation
0.000001	$\frac{1}{1{,}000{,}000}$	$\frac{1}{10\times10\times10\times10\times10\times10}$	1×10^{-6}
0.000003	$\frac{3}{1{,}000{,}000}$	$\frac{3}{10\times10\times10\times10\times10\times10}$	3×10^{-6}
0.0000067	$\frac{6.7}{1{,}000{,}000}$	$\frac{6.7}{10\times10\times10\times10\times10\times10}$	6.7×10^{-6}

As we can see,

- Really large numbers have positive exponents.
- Really small numbers have negative exponents.

Therefore, you should identify 6.022×10^{23} atoms as a very large number of atoms and 9.11×10^{-31} kg as a very small mass.

In scientific notation, numbers are expressed in the form

$$N \times 10^x \tag{2.2.1}$$

where N is greater than or equal to 1 and less than 10 ($1 \le N < 10$), and x is an integer value.

To convert a quantity to scientific notation, count the number of places required to move the decimal point to yield a number, N, between 1 and 10. The exponent, x, is equal to the number of places that the decimal was moved, and will be positive integer value if the quantity is large (>1) and negative integer value if the quantity is small (<1).

To convert a quantity from scientific notation to decimal notation (without scientific notation), move the decimal the number of places indicated by the exponent in the direction that makes either a really large number (when x is positive) or a really small

number (when x is negative).

✅ Example 2.2.1

If the given quantity is in decimal notation, convert it to scientific notation. If the given quantity is in scientific notation, convert it to decimal notation.

A. 8.62×10^{-6}
B. 92,300,000
C. 4.2×10^{10}
D. 0.000049

Solution

	Explanation	Answer
A	8.62×10^{-6} is a really small number since x is negative. Move the decimal 6 places to make a really small number. Fill in the spaces with zeros. $8.62 \times 10^{-6} = 0.00000862$	0.00000862
B	92,300,000 is a really large number, so the exponent will be positive. Move the decimal 7 places to make a number between 1 and 10. $92,300,000 = 92,300,000$	9.23×10^{7}
C	4.2×10^{10} is a really large number since x is positive. Move the decimal 10 places to make a really large number. Fill in the spaces with zeros. $4.2 \times 10^{10} = 42,000,000,000$	42,000,000,000
D	0.000049 is a really small number, so the exponent will be negative. Move the decimal 5 places to make a number between 1 and 10. $0.000049 = 0.000049$	4.9×10^{-5}

✏️ Exercise 2.2.1

If the given quantity is in decimal notation, convert it to scientific notation. If the given quantity is in scientific notation, convert it to decimal notation.

A. 0.00505
B. 1.7×10^{-8}
C. 3.33×10^{5}
D. 860,000,000,000

Answer A

5.05×10^{-3}

Answer B

0.000000017

Answer C

333,000

Answer D

8.6×10^{11}

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2.3: Significant Figures

Learning Objectives

- Identify the number of significant figures in a reported value.

After a bank robbery has been committed, police will ask witnesses to describe the thieves. They will usually get an answer such as "medium height". Others may say "between 5 foot 8 inches and 5 foot 10 inches". In each case, there is some level of uncertainty about the height of the criminals.

Measurement Uncertainty

Error or **uncertainty** exists in almost all measurements. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest ±0.1 g, other highly sensitive balances are capable of measuring to the nearest ±0.0001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 2.3.1 shows two rulers making the same measurement of an object (indicated by the dotted line).

With either ruler, it is clear that the length of the object is between 2 and 3 cm – closer to 3 cm than 2 cm. With the first ruler, the tenths digit may be estimated and the length reported as 2.9 cm. However, another person may judge that the measurement is 2.8 cm. While the 2 is known for certain, the value of the tenths digit is uncertain.

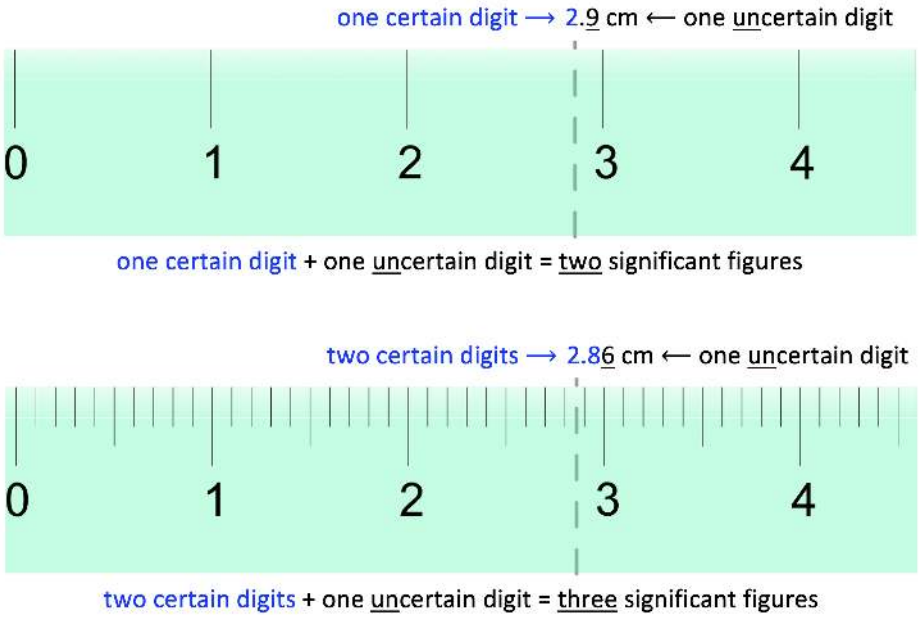


Figure **2.3.1**: The same measurement with two different rulers. The second ruler is more precise, since it has marks that represent tenths of a centimeter.

The second ruler has additional markings that represent tenths of a centimeter. Now the same object may be measured as 2.86 cm. The measurer is capable of estimating the hundredths digit because she can be certain that the tenths digit is an 8. Again, another measurer may report the length to be 2.85 cm or 2.87 cm. In this case, there are two certain digits (the 2 and the 8), with the hundredths digit being uncertain. Clearly, the second ruler is a superior choice for measuring lengths, since it is more precise.

Significant Figures

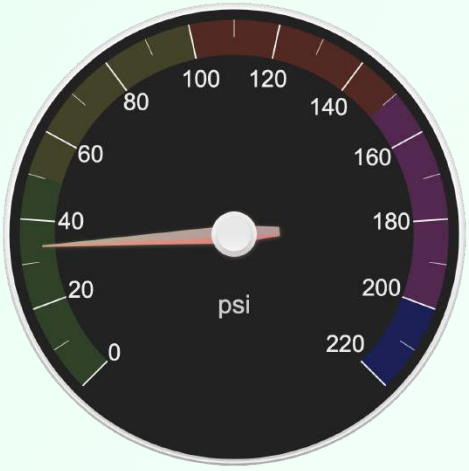
The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler illustration above, the first ruler yielded a length showing 2 significant figures, while the second ruler yielded a length showing 3 significant figures. In a correctly reported measurement, the final digit is significant but *not* certain. Insignificant digits are not reported. With either ruler, it would not be possible to report the length as 2.857 cm, since there is no possible way that the thousandths digit could be estimated. The 7 is *not* significant and would *not* be reported.

⚡ Significant Figures

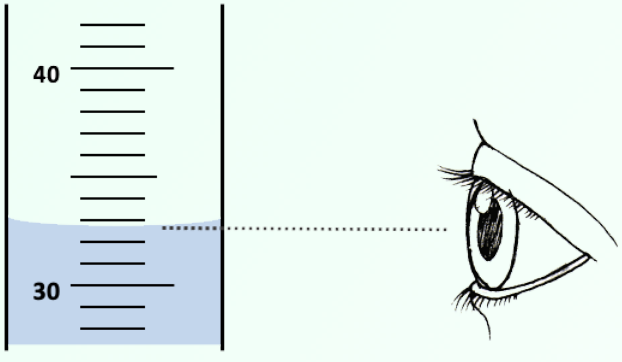
All of the certain digits in a measurement plus one digit that is uncertain.

✅ Example 2.3.1

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



A.



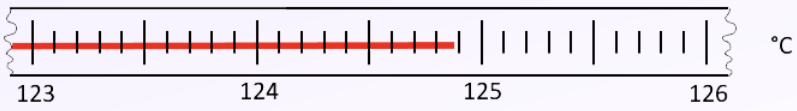
B.

Solutions

	Explanation	Answer
A.	The arrow is between 30 and 40 psi, so the measurement is greater than 30 psi. The arrow appears a little less than one-half of the way across the space between 30 and 40, so the ones place may be estimated as 4. The measurement is reported to two significant figures.	34 psi
B.	The meniscus (lowest level of the liquid) is between 30 and 40 mL, so the measurement is greater than 30 mL. It also appears to be between the second and third small tick marks, so it's at least 32 mL, but less than 33 mL. Since the meniscus appears to be about $\frac{4}{5}$ of the distance above the second tick mark, the tenths place may be estimated as 8. The measurement is reported to three significant figures.	32.8 mL

Exercise 2.3.1

What temperature should be reported?



Answer

124.89 °C

Counting Significant Figures

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

Table 2.3.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	• **237** has three significant figures. • **1.897** has four significant figures.
2. Zeros that appear between other nonzero digits are always significant.	• **39,004** has five significant figures. • **5.02** has three significant figures.
3. **Leading zeros** (zeros that appear in front of all of the nonzero digits) are **not** significant.	• **0.008** has one significant figure. • **0.000416** has three significant figures.
4. **Trailing zeros** (zeros that appear after all nonzero digits) in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	• **620.0** has four significant figures. • **19.300** has five significant figures.
5. A number with trailing zeros that lacks a decimal point may or may not be significant. **Use scientific notation to indicate the appropriate number of significant figures.**	• **3000** has an ambiguous number of significant figures. However: ◦ **3×10^3** has one significant figure. ◦ **3.0×10^3** has two significant figures. ◦ **3.00×10^3** has three significant figures. ◦ **3.000×10^3** has four significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have an infinite number of significant figures. For example, if we count four corners in a rectangle, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000…). Similarly, 1 foot is defined to be the same as 12 inches, so the number 12 in the following equation has an infinite number of significant figures:

$$74.5 \cancel{\text{ inches}} \times \frac{1 \text{ foot}}{12 \cancel{\text{ inches}}} = 6.21 \text{ feet} \tag{2.3.1}$$

✅ Example 2.3.2

Give the number of significant figures in each measurement.

A. 5.87 km
B. 0.031 g
C. 52.90 ms
D. 0.00210 L
E. 800 pounds
F. 5 fingers

Solution

	Explanation	Answer
A	All three digits are significant.	5.87 km has three significant figures.
B	The leading zeros are not significant. The 3 and the 1 are significant.	0.031 g has two significant figures.
C	The 5, 2, and 9 are significant. The trailing zero is also significant.	52.90 ms has four significant figures.
D	The leading zeros are not significant. The 2 and the 1 are significant. The trailing zero is also significant.	0.00210 L has three significant figures.
E	The trailing zeros may or may not be significant. The number could have one, two, or three significant figures.	800 pounds has an ambiguous number of significant figures.
F	5 is a counting number in this instance. A counting number is an exact number.	5 fingers has an infinite number of significant figures in this instance.

Exercise 2.3.2

Give the number of significant figures in each measurement.

A. 60.0500 m
B. 0.00326 s
C. 246.3 kg
D. 3.5×10^{12}
E. 24 students in a lab
F. 3800 g

Answer A

6 significant figures

Answer B

3 significant figures

Answer C

4 significant figures

Answer D

2 significant figures

Answer E

infinite (exact number)

Answer F

Ambiguous, could be two, three or four significant figures.

Accuracy and Precision

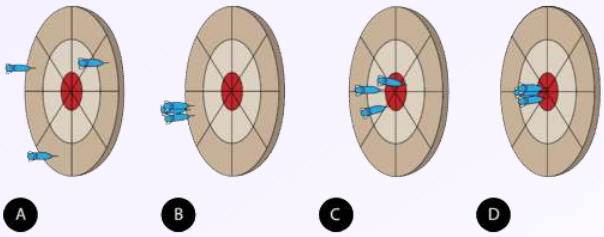
Measurements may be accurate, meaning that the measured value is close to the true or known value. They may be precise, meaning that multiple measurements result in nearly identical values (i.e., they are reproducible). They may be both accurate and precise, or they may be neither accurate nor precise. The goal of scientists is to obtain measured values that are both accurate and precise. The video below demonstrates the concepts of accuracy and precision.



Video **2.3.1**: Difference between precision and accuracy.

Exercise 2.3.3

The following archery targets show marks that represent the results of four sets of measurements.



Which target shows

1. a precise, but inaccurate set of measurements?
2. a set of measurements that is both precise and accurate?
3. a set of measurements that is neither precise nor accurate?

Answer 1

Target **B** is precise, but inaccurate.

Answer 2

Target **D** is both precise and accurate.

Answer 3

Target **A** is neither precise nor accurate.

Summary

- Uncertainty exists in all measurements. The degree of uncertainty is affected in part by the quality of the measuring tool.
- Significant figures give an indication of the certainty of a measurement. Rules allow decisions to be made about how many digits to use in any given situation.

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2.4: Significant Figures in Calculations

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Calculators do just what is asked of them – no more and no less. However, they can sometimes get a little out of hand. If you multiply 2.49 by 6.3, you get an answer of 15.687, a value that ignores the number of significant figures in either number. Division with a calculator is even worse. When you divide 12.2 by 1.7, the answer you obtain is 7.176470588. Neither piece of data is accurate to nine decimal places, but the calculator does not know that. The human being operating the instrument has to make the decision about how the answer should be reported.



Figure **2.4.1**: A TI-84 Plus graphing calculator. (Asimzb via Wikimedia Commons)

Rounding

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 207.518 m. As it stands, the measurement contains six significant figures. How would this number be rounded to a fewer number of significant figures? Follow the process as outlined in Table **2.4.1**.

Table 2.4.1: Rounding Examples

Number of Significant Figures	Rounded Value	Reasoning
6	207.518 m ⇒ 207.518 m	All digits are significant.
5	207.518 m ⇒ 207.52 m	Round up, since the 6th digit is an "8" (≥5).
4	207.518 m ⇒ 207.5 m	The last two digits are simply dropped, since the 5th digit is a "1" (<5).
3	207.518 m ⇒ 208 m	Round up, since the 4th digit is a "5" (≥5).
2	207.518 m ⇒ 210 m ⇒ 2.1×10^2 m	Round up, since the 3rd digit is a "7" (≥5). However, trailing zeros in a number that has no decimal point are ambiguous, so express in scientific notation.
1	207.518 m ⇒ 200 m ⇒ 2×10^2 m	The last five digits are dropped, since the 2nd digit is a "0" (<5). The ones place and tens place are held with zeros. However, trailing zeros in a number that has no decimal point are ambiguous, so express in scientific notation.

Notice that the further a number is rounded, the less reliable that measurement becomes. An approximate value may be sufficient for some purposes, but scientific work requires a much higher level of detail.

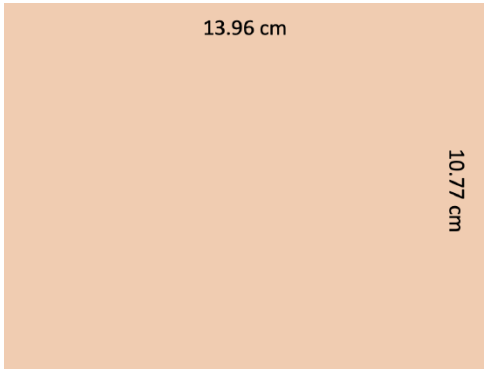
It is important to be aware of significant figures when mathematically manipulating numbers. For example, dividing 125 by 307 on a calculator gives 0.4071661238… to an infinite number of digits. But do the digits in this answer have any practical meaning, especially when you are starting with numbers 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.

⚡ Significant Figures and Rounding

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.

Multiplication and Division

Suppose you wanted to find the area of a rectangle and you measured the dimensions as 13.96 cm by 10.77 cm. Multiplying the two lengths together on a calculator yields an area of 150.349 cm^2. Should the answer be reported as 150.349 cm^2 or should it be rounded to 150.35 cm^2 or 150.3 cm^2 or even more?



Recall that all measurements have uncertainty. Assuming the last digit is the uncertain digit and may be estimated to the nearest ±0.01 cm, a second student may measure the dimensions of the rectangle as 13.95 cm by 10.76 cm. A third student may measure it as 13.97 cm by 10.78 cm. A fourth student may come up with 13.96 cm by 10.78 cm. The measurements are tabulated in the table below, along with the calculated areas.

If the calculated areas are examined more closely, it is easy to see that the hundreds place (the 1), tens place (the 5), and ones place (the 0) are all certain. They have no variation. We begin to see variation or uncertainty in the tenths place. Since significant figures are defined as all of the certain digits in a measurement plus one uncertain digit, the calculated areas should be rounded and reported to the nearest ±0.1 cm^2.

Student	Length	Width	Calculated Area	Reported Area
1	13.96 cm	10.77 cm	150.349 cm^2	150.3 cm^2
2	13.95 cm	10.76 cm	150.102 cm^2	150.1 cm^2
3	13.97 cm	10.78 cm	150.597 cm^2	150.6 cm^2
4	13.96 cm	10.78 cm	150.489 cm^2	150.5 cm^2

This is because the length and width each have four significant figures. Multiplying them together results in the area having four significant figures, i.e. three certain digits plus one uncertain digit. When the multiplied values have a different number of significant figures, the answer should be limited to the factor that has the lowest count of significant figures. The same rule applies to division.

⚡ Rounding Rule for Multiplication and Division

The answer should be rounded so it contains the same number of significant figures as the measurement having the fewest number of significant figures.

✅ Example 2.4.1

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

A. $\frac{346.1\ \text{mi}}{5.3\ \text{h}} =$

B. $14.58\ \text{ft} \times 5.73\ \text{ft}$

C. $36\ \text{in} \times 26\ \text{in} \times 16\ \text{in}$

Solution

A

Explanation	Answer
The calculated answer is 65.301887 mi/h. The reported answer should have two significant figures, since 5.3 h has the fewest significant figures (two) between the two measurements.	65 mi/h

B

Explanation	Answer
The calculated answer is 83.5434 ft^2. The reported answer should have three significant figures, since 5.73 ft has the fewest significant figures (three) between the two measurements.	83.5 ft^2

C

Explanation	Answer
The calculated answer is 14,976 in^3. The reported answer should have two significant figures, since all three measurements have two significant figures. Since an answer of 15,000 in^3 would represent an ambiguous number of significant figures, the answer should be reported in scientific notation with two significant figures.	1.5×10^4 in^3

Addition and Subtraction

Now that we know how to report answers that involved multiplication and division, you may wonder if answers resulting from addition and subtraction are handled in the same manner. What if we wanted to know the combined distance of 385.17 m, 6.2 m, and 45.86 m? Notice that the first and last lengths were measured to the nearest ±0.01 m, while the second length was measured to the nearest ±0.1 m.

Original Measurements	Change 1st Measurement	Change 2nd Measurement	Change 3rd Measurement
385.17 m +6.2 m +45.86 m 437.23 m	385.18 m +6.2 m +45.86 m 437.24 m	385.17 m +6.3 m +45.86 m 437.33 m	385.17 m +6.2 m +45.87 m 437.24 m

The table above shows what happens to the sum when the uncertain digit in each measurement is changed, one by one. Notice when the second measurement is changed from 6.2 m to 6.3 m that the sum also changes in the tenths place, going from 437.23 m to 437.33 m. In other words, the second measurement has the greatest impact on the calculated result since it is the least precise of the three measurements, ±0.1 m vs. ±0.01 m.

±0.01 m
±0.1 m
±0.01 m
±0.1 m

Therefore, the sum of 385.17 m, 6.2 m, and 45.86 m should be reported as 437.2 m.

```
385.17  m
 +6.2   m
+45.86  m
-------
437.2   m
```

This leads us to the rounding rule for addition and subtraction. While different, it is worded similarly to the rule for multiplication and division. Can you spot the difference? It is likely the rule with which you are most familiar.

⚡ Rounding Rule for Addition and Subtraction

The answer should be rounded so it contains the same number of decimal places as the measurement having the fewest number of decimal places.

✅ Example 2.4.2

A.
```
275.71 kg
−  8.4  kg
```

B. 1,027 mL + 611 mL + 363.06 mL

Solution

A

Explanation	Answer
A calculator provides an answer of 267.31 kg, but because 8.4 kg is known to the nearest ±0.1 kg, the final answer should be expressed to the nearest ±0.1 kg.	267.3 kg

B

Explanation	Answer
A calculator provides an answer of 2,001.06 mL, but because 1027 mL and 611 mL are both known to the nearest ±1 mL, the final answer should be expressed to the nearest ±1 mL.	2001 mL

✏️ Exercise 2.4.1

Write the answer for each expression using the correct number of significant figures. As a reminder, calculators do not understand significant figures. *You* are the one who must apply the rules for calculated answers to a result obtained from a calculator.

A. $\frac{165.110\ \text{g}}{8.35\ \text{mL}} =$
B. $8.6\ \text{g} + 32.06\ \text{g} + 88.7\ \text{g}$
C. $255.0\ \text{km} - 99\ \text{km}$
D. $44\ \text{cm} \times 43\ \text{cm}$

Answer A

$19.8\ \text{g/mL}$

Answer B

$129.4\ \text{g}$

Answer C

$156\ \text{km}$

Answer D

$1.9 \times 10^3\ \text{cm}^2$

Calculations Involving Mixed Operations

In practice, chemists generally work with a calculator and carry all digits forward through subsequent calculations. When working on paper, however, we often want to minimize the number of digits we have to write out. Because successive rounding can compound inaccuracies, intermediate rounding needs to be handled correctly. When working on paper, always round an intermediate result so as to retain *at least* one more digit than can be justified and carry this number into the next step in the calculation. The final answer is then rounded to the correct number of significant figures at the very end.

✅ Example 2.4.3

A cylinder filled with water was combined with 5.2 cm^3 of water that had been drawn into a syringe. If the cylinder had an inner height of 16.8 cm and an inner radius of 1.86 cm, what is the combined volume of water? Report the answer with the correct number of significant figures. (Note: For a cylinder, $V = \pi r^2 h$. You may use the π function on your calculator or estimate the value of π as 3.1416.)

Solution

Volume	Explanation	Result
Syringe	The volume of water in the syringe is 5.2 cm^3, known to the nearest ±0.1 cm^3.	5.2 cm^3
Cylinder	The volume of water in the cylinder = $\pi r^2 h$ = $3.1416(1.86\ cm)^2(16.8\ cm)$ = $182.594\ cm^3$. The rules for multiplication and division apply. Since 16.8 cm and 1.86 cm both have three significant figures, the calculated volume has three significant figures, or 183 cm^3. Compounding of rounding errors in the final answer may be avoided by carrying extra digits along in the intermediate results. The uncertain digit is underlined (the ones place).	182.594 cm^3
Combined	The sum of 5.2 cm^3 and 182.594 cm^3 is 187.794 cm^3. The rules for addition and subtraction apply. Since 5.2 cm^3 is uncertain in the tenths place (±0.1 cm^3) and 182.594 cm^3 is uncertain in the ones place (±1 cm^3), the sum is uncertain in the ones place (±1 cm^3), or 187.794 cm^3.	188 cm^3

The volume should be reported as 188 cm^3. Check out the video below for additional examples.



Video **2.4.1**: Significant figures in mixed operations (https://www.youtube.com/watch?v=yBntMndXQWA).

Summary

- Rounding
  - If the number to be dropped is greater than or equal to 5, increase the number to its left by 1.
  - If the number to be dropped is less than 5, there is no change.
- The rule in multiplication and division is that the final answer should have the same number of significant figures as the measurement having the fewest significant figures.
- The rule in addition and subtraction is that the final answer is should have the same number of decimal places as the measurement having the fewest decimal places.

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2.5: The Metric System

Learning Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit.

How long is a yard? It depends on whom you ask and when the question was asked. Today, we have a standard definition of the yard, which one may see marked on every football field. The distance from goal line to goal line is 100 yards. It does not matter whether the field is in Minneapolis, Green Bay, or Detroit. However, at one time that yard was arbitrarily defined as the distance from the tip of the king's nose to the end of his outstretched hand. The problem with this is simple: new king, new distance, and then all those football fields would need to be re-marked.

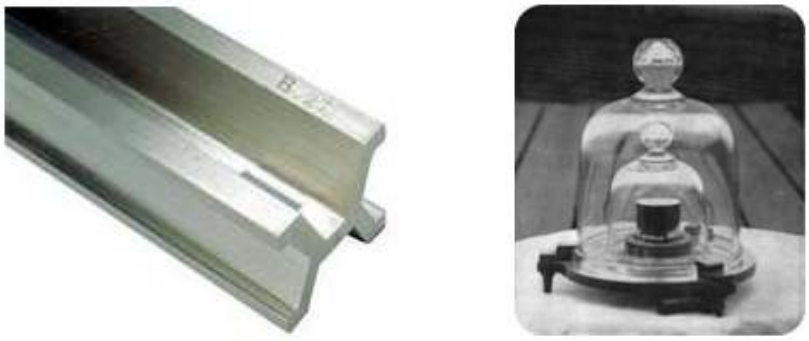


Figure 2.5.1: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions rather simple by comparison. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of the seven base units in Table 2.5.1.

Table 2.5.1: SI Base Units of Measurement

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

The first five units are commonly used in chemistry. All other measured quantities, such as volume, force, and energy, may be derived from these seven base units and are referred to as *derived units*.

⚓ Unfortunately, the Metric System is Not Ubiquitous

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

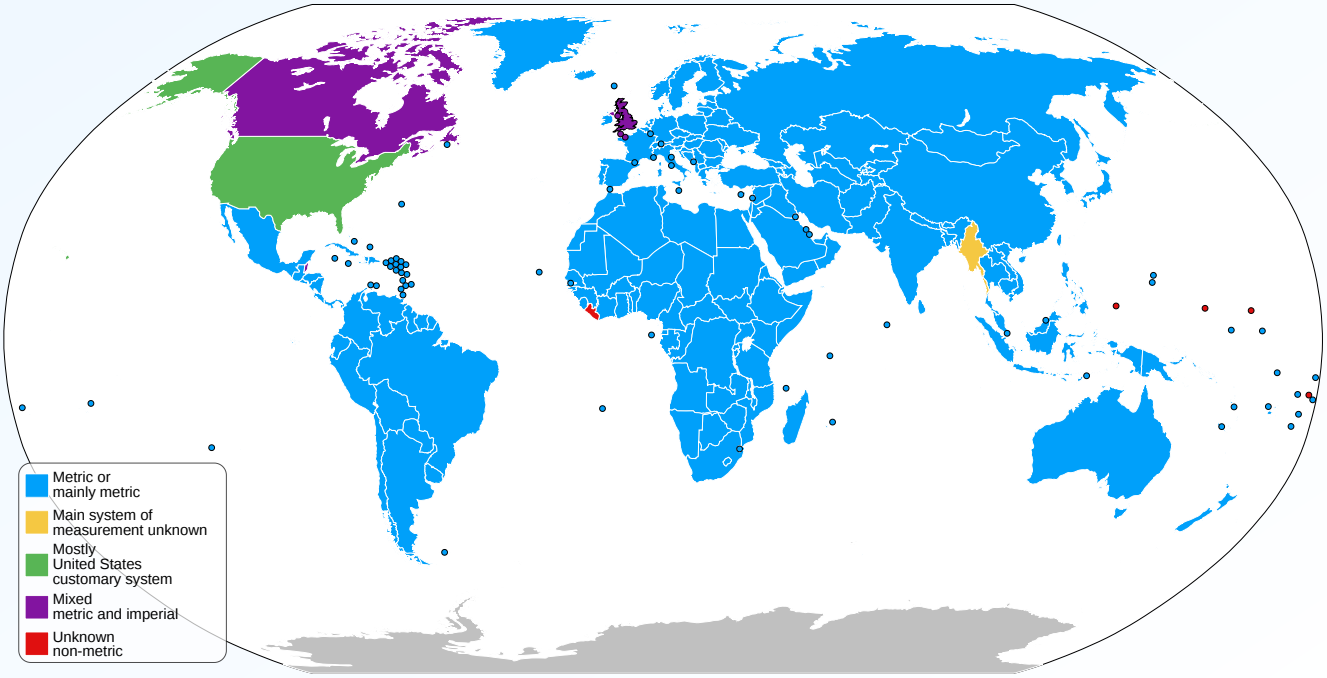


Figure 2.5.1: Areas of world using metric system. Only a few countries are slow or resistant to adoption, including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word μεγας, meaning "great". Table 2.5.2 lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 2.5.2: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Meaning	Example
tera-	$\mathbf{T}$	1,000,000,000,000	10^{12}	1 terameter $(\mathbf{Tm}) = 10^{12}$ $\mathbf{m}$
giga-	$\mathbf{G}$	1,000,000,000	10^{9}	1 gigameter $(\mathbf{Gm}) = 10^{9}$ $\mathbf{m}$
mega-	$\mathbf{M}$	1,000,000	10^{6}	1 megameter $(\mathbf{Mm}) = 10^{6}$ $\mathbf{m}$
kilo-	$\mathbf{k}$	1000	10^{3}	1 kilometer $(\mathbf{km}) = 1000$ $\mathbf{m}$
hecto-	$\mathbf{h}$	100	10^{2}	1 hectometer $(\mathbf{hm}) = 100$ $\mathbf{m}$
deka-	$\mathbf{da}$	10	10^{1}	1 dekameter $(\mathbf{dam}) = 10$ $\mathbf{m}$
		$\mathbf{1}$		1 meter $(\mathbf{m})$
deci-	$\mathbf{d}$	$\frac{1}{10}$	10^{-1}	1 decimeter $(\mathbf{dm}) = 0.1$ $\mathbf{m}$
centi-	$\mathbf{c}$	$\frac{1}{100}$	10^{-2}	1 centimeter $(\mathbf{cm}) = 0.01$ $\mathbf{m}$
milli-	$\mathbf{m}$	$\frac{1}{1000}$	10^{-3}	1 millimeter $(\mathbf{mm})$ $= 0.001$ $\mathbf{m}$
micro-	μ	$\frac{1}{1,000,000}$	10^{-6}	1 micrometer $(\mu\mathbf{m})$ $= 10^{-6}$ $\mathbf{m}$
nano-	$\mathbf{n}$	$\frac{1}{1,000,000,000}$	10^{-9}	1 nanometer $(\mathbf{nm}) = 10^{-9}$ $\mathbf{m}$
pico-	$\mathbf{p}$	$\frac{1}{1,000,000,000,000}$	10^{-12}	1 picometer $(\mathbf{pm}) = 10^{-12}$ $\mathbf{m}$

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

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs 6.5 g, this is easier than saying it weighs 6500 mg or 0.0065 kg. All three are correct, but units of g in this case make for a small and easily managed number. However, if a specific problem needs milligrams instead of grams, go with the milligrams for consistency.

✅ Example 2.5.1

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

A. kiloliter
B. microsecond
C. decimeter
D. nanogram

Solutions

	Explanation	Answer
A	The prefix kilo- means 1000, so 1 kL equals 1000 L.	kL
B	The prefix micro- means $\frac{1}{1,000,000}$, so 1 µs equals 0.000001 s.	µs
C	The prefix deci- means $\frac{1}{10}$, so 1 dm equals 0.1 m.	dm
D	The prefix nano- means $\frac{1}{1,000,000,000}$, so 1 ng equals 0.000000001 g.	ng

Exercise 2.5.1

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

A. kilometer
B. milligram
C. nanosecond
D. centiliter

Answer A

km ⇒ 1 km = 10^3 m = 1000 m

Answer B

mg ⇒ 1 mg = $\frac{1}{1000}$ **g** = 0.001 g

Answer C

ns ⇒ 1 ns = $\frac{1}{1,000,000,000}$ **s** = 10^{-9} s

Answer D

cL ⇒ 1 cL = $\frac{1}{100}$ **L** = 0.01 L

Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.

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2.6: Dimensional Analysis

Learning Objectives

- To convert a value reported in one unit to a corresponding value in a different unit using conversion factors.

During your study of chemistry, you will notice the frequent use of mathematical equations. These equations often use multiple variables and usually require measurements with specific units. Algebra skills are very important! However, conversions may often be accomplished through the use of a technique called dimensional analysis without using equations or algebraic manipulation.

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can be used to convert meters into centimeters, or a more complex one may be used to convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. A conversion always represents a fact and may either be simple or complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/s. Either one of these can be used as a conversion factor depending on the type of calculation (Table 2.6.1).

Table 2.6.1: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid ounce (fl oz)	29.57 milliliters (mL)	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.540 centimeters (cm)	length
1 mile (mi)	1.609 kilometers (km)	length
1.057 quarts (qt)	1 liter (L)	volume

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

Common Chemistry Conversions

A more complete list of conversions may be found in Appendix 2: Conversions.

Of course, there are other ratios which are not listed in either Table 2.6.1 or Appendix 2: Conversions. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %).
- Conversions using metric units, as covered in 2.5: The Metric System.
- Common knowledge ratios (such as 60 seconds in 1 minute or 60 minutes in 1 hour).

If you learned the SI units and prefixes described, then you know that 1 cm is $\frac{1}{100}$ of a meter.

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

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:

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

We know that 100 cm *is* 1 m, so we have the same quantity on the top and the bottom of our fraction, although they are expressed in different units. The fraction can then be used for calculations converting m to cm.

Performing Dimensional Analysis

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

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

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

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

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

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

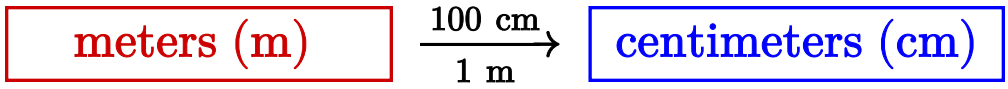
quantity (in old units) × conversion factor = quantity (in new units)

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

In the previous example, we used the fraction $\frac{100\ \text{cm}}{1\ \text{m}}$ as a conversion factor, but the conversion factor $\frac{1\ \text{m}}{100\ \text{cm}}$ is also equal to 1. $\frac{1\ \text{m}}{100\ \text{cm}}$ has the same quantity in the numerator as in the denominator, except that they are expressed in different units. However, had we used the second conversion factor, the original unit would not have canceled, and the result would have been meaningless. Here is what it would look like:

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

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



Notice that the units from which you are converting go on the bottom, while the units to which you are converting go on the top.

General Steps in Performing Dimensional Analysis

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

Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

- Metric-to-metric conversions are *not* considered in the determination of significant figures in a calculation since they are exact.
- Other defined relationships, not measured quantities, are also exact and are considered to have an infinite number of significant figures. In other words, there are exactly 12 inches in 1 foot or there are exactly 60 minutes in 1 hour.
- Counted numbers are also exact. If one counts 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 that are approximations have a limited number of significant figures and should be considered in determining the significant figures of the final answer.

✅ Examples 2.6.1

A. The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?

B. A table is 42 inches long. What is this length in feet?

Solution

Steps for Problem Solving	Part A	Part B
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 42 inches Find: feet
List other known quantities. Metric prefixes are found in 2.5: The Metric System. Other conversions may be found in Appendix 2: Conversions.	1 mL = 10^{-3} L	1 ft = 12 in
Prepare a concept map and use the proper conversion factor(s).	L $\xrightarrow[1\ \text{L}]{1000\ \text{mL}}$ mL	in $\xrightarrow[12\ \text{in}]{1\ \text{ft}}$ ft
Cancel units and calculate.	$4.7\ \cancel{\text{L}} \times \frac{1\ \text{mL}}{10^{-3}\ \cancel{\text{L}}} = \boxed{4700\ \text{mL}}$ or $4.7\ \cancel{\text{L}} \times \frac{1000\ \text{mL}}{1\ \cancel{\text{L}}} = \boxed{4700\ \text{mL}}$	$42\ \cancel{\text{in}} \times \frac{1\ \text{ft}}{12\ \cancel{\text{in}}} = \boxed{3.5\ \text{ft}}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in feet should be less than the given amount in inches.

Exercise 2.6.1

Perform each conversion.

A. 101,000 ns to seconds
B. 3.208 kg to grams
C. 215 lb to kg

Answer A

1.01×10^{-4} s

Answer B

3208 g

Answer C

97.5 kg

Summary

- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce the appropriate units.

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2.7: Multi-Step Conversions

Learning Objectives

- To convert a value reported in one unit to a corresponding value in a different unit using multiple conversion factors.

Multiple Conversions

In some cases, multiple conversions are required to obtain the desired unit. For example, suppose we want to convert 54.7 km to millimeters. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence.

We may first convert the given amount in km to meters, knowing that 1 km = 1000 m, then from meters to millimeters, knowing that 1 mm = 10^{-3} m.

Concept Map

$$\boxed{\text{km}} \xrightarrow{\frac{1000\ \text{m}}{1\ \text{km}}} \boxed{\text{m}} \xrightarrow{\frac{1\ \text{mm}}{10^{-3}\ \text{m}}} \boxed{\text{mm}}$$

Calculation

$$54.7\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1\ \text{mm}}{10^{-3}\ \cancel{\text{m}}} = \boxed{54,700,000\ \text{mm}}$$
$$= \boxed{5.47 \times 10^{7}\ \text{mm}}$$

Multi-Step Conversions

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

Example 2.7.1

Convert 58.2 milligrams to megagrams in one multi-step calculation.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 58.2 mg Find: Mg
List other known quantities. Metric prefixes are found in 2.5: The Metric System. Other conversions may be found in Appendix 2: Conversions.	1 mg = 10^{-3} g 1 Mg = 10^{6} g
Prepare a concept map and use the proper conversion factor(s).	$\boxed{\text{mg}} \xrightarrow{\frac{10^{-3}\ \text{g}}{1\ \text{mg}}} \boxed{\text{g}} \xrightarrow{\frac{1\ \text{Mg}}{10^{6}\ \text{g}}} \boxed{\text{Mg}}$
Calculate the answer.	$58.2\ \cancel{\text{mg}} \times \frac{10^{-3}\ \cancel{\text{g}}}{1\ \cancel{\text{mg}}} \times \frac{1\ \text{Mg}}{1,000,000\ \cancel{\text{g}}} = \boxed{0.0000000582\ \text{Mg}}$ $= \boxed{5.82 \times 10^{-8}\ \text{Mg}}$ Neither conversion factor affects the number of significant figures in the final answer.
Think about your result.	Milligrams are a tiny unit. Megagrams are a huge unit. It makes sense that the larger number is associated with the smaller unit (mg) and that the smaller number is associated with the larger unit (Mg).

✅ Example 2.7.2

Express a speed of 65 mi/h in SI units (m/s).

Solution 1

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 65 mi/h Find: m/s
List other known quantities. Metric prefixes are found in 2.5: The Metric System. Other conversions may be found in Appendix 2: Conversions. Pay close attention to the distinction between m, mi, and min.	1 mi = 1.609 km 1 km = 1000 m 1 h = 60 min 1 min = 60 s
Prepare a concept map and use the proper conversion factor(s).	$\boxed{\frac{\text{mi}}{\text{h}}} \xrightarrow{\frac{1.609\ \text{km}}{1\ \text{mi}}} \boxed{\frac{\text{km}}{\text{h}}} \xrightarrow{\frac{1000\ \text{m}}{1\ \text{km}}} \boxed{\frac{\text{m}}{\text{h}}} \xrightarrow{\frac{1\ \text{h}}{60\ \text{min}}} \boxed{\frac{\text{m}}{\text{min}}} \xrightarrow{\frac{1\ \text{m}}{60\ \text{s}}}$
Calculate the answer.	$\frac{65\ \cancel{\text{mi}}}{\cancel{\text{h}}} \times \frac{1.609\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times \frac{1\ \cancel{\text{min}}}{60\ \text{s}} = \boxed{29\ \frac{\text{m}}{\text{s}}}$
Think about your result.	It matches the answer calculated using a different set of conversion factors. (see Solution 2)

Solution 2

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 65 mi/h Find: m/s
List other known quantities. Metric prefixes are found in 2.5: The Metric System. Other conversions may be found in Appendix 2: Conversions. Pay close attention to the distinction between m, mi, and min.	1 mi = 5280 ft 1 m = 3.281 ft 1 h = 3600 s
Prepare a concept map and use the proper conversion factor(s).	$\boxed{\frac{\text{mi}}{\text{h}}} \xrightarrow{\frac{1\ \text{h}}{3600\ \text{s}}} \boxed{\frac{\text{mi}}{\text{s}}} \xrightarrow{\frac{5280\ \text{ft}}{1\ \text{mi}}} \boxed{\frac{\text{ft}}{\text{s}}} \xrightarrow{\frac{1\ \text{m}}{3.281\ \text{ft}}} \boxed{\frac{\text{m}}{\text{s}}}$
Calculate the answer.	$\frac{65\ \cancel{\text{mi}}}{\cancel{\text{h}}} \times \frac{1\ \cancel{\text{h}}}{3600\ \text{s}} \times \frac{5280\ \cancel{\text{ft}}}{1\ \cancel{\text{mi}}} \times \frac{1\ \text{m}}{3.281\ \cancel{\text{ft}}} = \boxed{29\ \frac{\text{m}}{\text{s}}}$
Think about your result.	It matches the answer calculated using a different set of conversion factors. (see Solution 1)

Example 2.7.2 above shows two different solutions. However, the end result is the same, 65 mi/h is the same as 29 m/s. This demonstrates one of the major

advantages of using dimensional analysis as a problem-solving tool. There are no formulas to learn or memorize. One must simply pay attention to the conversion factors to ensure that the proper units cancel out and leave the desired units for the answer.

Example 2.7.2 also shows the importance of placing the correct units in the numerator and denominator. When using units that show one type of unit "per" another type of unit, the unit that precedes the / symbol belongs in the numerator, while the unit that follows the / symbol belongs in the denominator. Just like we think of 65 mi/h as $65\frac{\text{mi}}{\text{h}}$, we should also think of 2.70 g/mL as $2.70\frac{\text{g}}{\text{mL}}$. By doing so, you will ensure that units cancel properly when setting up a dimensional analysis calculation.

Exercise 2.7.1

Perform each conversion in one multi-step calculation.

A. Convert 43.007 nL to kL.
B. One lap around a modern running track is 400 m. Track events used to be measured in yards. Convert 400 m to yd. Round your answer to three significant figures.
C. How many seconds are in one day?

Answer A

4.3007×10^{-11} kL

Answer B

437 yd

Answer C

8.64×10^{4} s

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 in order to 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.



Figure 2.7.1: A woman consulting with a pharmacist. (Public Domain; Rhoda Baer via National Cancer Institute, an agency that is part of the National Institutes of Health.)

Summary

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

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2.8: Units Raised to a Power

Learning Objectives

- To convert a value reported in one unit raised to a power of 10, to a corresponding value in a different unit raised to the same power of 10, using conversion factors.

Conversion factors for area and volume may also be produced using dimensional analysis. Just remember that if a quantity is raised to a power of 10, both the number and the unit must be raised to the same power of 10. For example, to convert 1500 cm^2 to m^2, we need to start with the relationship between centimeter and meter. We know that 1 cm = 10^{-2} m or 100 cm = 1 m, but since we are given the quantity in 1500 cm^2, then we have to use the relationship:

$$1\ cm^2 = (10^{-2}\ m)^2 = 10^{-4}\ m^2$$

CONCEPT MAP

$$\boxed{cm^2} \quad \left(\frac{10^{-2}\ m}{1\ cm}\right)^2 \quad \boxed{m^2}$$

CALCULATION

$$1500\ \cancel{cm^2} \times \left(\frac{10^{-2}\ m}{1\ \cancel{cm}}\right)^2 = \boxed{0.15\ m^2}$$

or

$$1500\ \cancel{cm^2} \times \left(\frac{1\ m}{100\ \cancel{cm}}\right)^2 = \boxed{0.15\ m^2}$$

or

$$1500\ \cancel{cm^2} \times \frac{1\ m^2}{10,000\ \cancel{cm^2}} = \boxed{0.15\ m^2}$$

✅ Example 2.8.1

What is the volume of a sphere (radius = 4.30 inches) in cubic centimeters (cm^3), given $V = \frac{4}{3}\pi r^3$?

Solution

Steps for Problem Solving	Result
Identify the "given" information and what the problem is asking you to "find."	Given: radius = 4.30 in Find: cm^3 (volume)
List other known quantities. Metric prefixes are found in 2.5: The Metric System. Other conversions may be found in Appendix 2: Conversions.	$V = \frac{4}{3}\pi r^3$ $= \frac{4}{3} \times 3.1416 \times (4.30\ \text{in})^3$ $= 333.04\ \text{in}^3$ 1 in = 2.540 cm
Prepare a concept map and use the proper conversion(s).	$\text{in}^3 \quad \left(\frac{2.540\ \text{cm}}{1\ \text{in}}\right)^3 \quad \text{cm}^3$
Calculate the answer.	$333.04\ \cancel{\text{in}^3} \times \left(\frac{2.540\ \text{cm}}{1\ \cancel{\text{in}}}\right)^3 = \boxed{5.46 \times 10^3\ \text{cm}^3}$
Think about your result.	A centimeter is a smaller unit than an inch, so the answer in cubic centimeters is larger than the given value in cubic inches.

Exercise 2.8.1

Lake Tahoe has a surface area of 191 square miles. What is the area in square kilometers (km^2)?

Answer

494 km^2

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2.9: Density

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



Figure **2.9.1**: Driving logs on the Fraser river near Vancouver, BC (Tony Hisgett via Flickr)

Density

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

$$\text{density} = \frac{\text{mass}}{\text{volume}} \tag{2.9.1}$$

or simply

$$d = \frac{m}{V} \tag{2.9.2}$$

Density is an intensive property, meaning that it does not depend on the amount of material present in the sample. Water has a density of $1.0\,\text{g/mL}$. That density is the same whether you have a small glass of water or a swimming pool full of water. Density is a property that is constant for the particular identity of the matter being studied.

The SI units of density are kilograms per cubic meter $\left(\text{kg/m}^3\right)$, since the kg and the m are the SI units for mass and length respectively. In everyday usage in a laboratory, this unit is awkwardly large. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter $\left(\text{g/cm}^3\right)$. Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L. Densities of some common substances at 20°C (unless otherwise noted) are listed in Table 2.9.1 below.

Table 2.9.1: Densities of Some Common Substances

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084
Ice @ 0°C	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water @ 4°C	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.96		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

The density of a substance is temperature dependent and usually decreases as temperature increases due to the fact that most materials expand as temperature increases. Since we know ice floats in water, it should make sense that it is less dense. Likewise, we would expect corn syrup to sink in water since it is more dense.

✅ Example 2.9.1

A block of wood measures 16.6 cm by 8.7 cm by 3.6 cm and has a mass of 234.1 g. What is the density of the wood block?

Solution

Steps for Problem Solving	
List the known quantities.	Mass = 234.1 g Dimensions = 16.6 cm × 8.7 cm × 3.6 cm
Plan the problem.	$d = \frac{m}{V}$ $V = l \times w \times h$
Calculate the answer.	$d = \frac{m}{V} = \frac{m}{l \times w \times h} = \frac{234.1\ \text{g}}{16.6\ \text{cm} \times 8.7\ \text{cm} \times 3.6\ \text{cm}} = 0.4503\ \text{g/cm}^3 = \boxed{0.45\ \text{g/cm}^3}$
Think about your result.	Since the density of water is 1 g/cm^3 and we know that almost all wood floats on water, a density of 0.45 g/cm^3 makes sense for the block of wood. Also, the answer is rounded to two significant figures, since 3.6 cm and 8.7 cm both have two significant figures and all of the measurements and all of the operations involve multiplication or division.

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

✅ Example 2.9.2

If 55.8 mL of water at 22°C is added to an empty graduated cylinder that has a mass of 100.62 g, what is the combined mass of the cylinder and the water? The density of water at 22°C is 0.998 g/mL.

Solution

Steps for Problem Solving	
List the known quantities.	$d_{water} = 0.998\ \text{g/mL}$ $V_{water} = 55.8\ \text{mL}$ $m_{cylinder} = 100.62\ \text{g}$
Plan the problem.	$m_{total} = m_{cylinder} + m_{water} \Rightarrow m_{total} = 100.62\ \text{g} + m_{water}$ $d = \frac{m}{V}$ Solve for *m*.
Calculate the answer.	Rearrange the equation to solve for the mass of the water: $V \cdot d = \frac{m}{\cancel{V}} \cdot \cancel{V}$ $m = V \cdot d \ \Rightarrow\ m_{water} = V_{water} \cdot d_{water}$ Plug in the volume and density of the water to find the mass of the water. $m_{total} = (55.8\ \cancel{\text{mL}}) \left(0.998\ \frac{\text{g}}{\cancel{\text{mL}}} \right) = 55.\underline{6}9\ \text{g} = 55.7\ \text{g}$ The mass of the water is known to three significant figures, or 55.7 g. The total mass may be found by summing the mass of the cylinder and the mass of the water. To avoid potential rounding issues, we should carry through at least one extra digit in the mass of the water. $m_{total} = 100.62\ \text{g} + 55.\underline{6}9\ \text{g} = 156.\underline{3}1\ \text{g} = \boxed{156.3\ \text{g}}$ The combined mass is 156.3 g. The answer is rounded to the nearest tenth, since the mass of the water was known to the nearest tenth (the last operation involves addition, so we pay attention to decimal places when rounding rather than significant figures).
Think about your result.	Since the density of water at 22°C is about 1 g/mL, its mass (in grams) is numerically close to its volume (in mL). Therefore, a mass of 55.7 g for the water makes sense. Adding the mass of the water to the mass of the cylinder, which was already known, is very straightforward.

Exercise 2.9.1

A. Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.

B. What is the volume of 100.0 g of air at 20°C? Consult Table 2.9.1 for the density.

Answer A

0.274 kg/L

Answer B

83.3 L

Density as a Conversion Factor

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

13.6 g mercury = 1 mL mercury

This relationship can be used to construct two conversion factors:

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

and

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

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 2.0 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:

$$2.0\ \cancel{\text{mL}} \times \frac{13.6\ \text{g}}{1\ \cancel{\text{mL}}} = 27.2\ \text{g} = \boxed{27\ \text{g}}$$

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.

✅ Example 2.9.3

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

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 0.750 g Find: mL
List other known quantities (see Table 2.9.1).	density of mercury = 13.6 g/mL
Prepare a concept map and use the proper conversion(s).	g → (1 mL / 13.6 g) → mL
Calculate the answer.	$0.750\ \cancel{\text{g}} \times \frac{1\ \text{mL}}{13.6\ \cancel{\text{g}}} = 0.055147...\ \text{mL} = \boxed{0.0551\ \text{mL}}$ The final answer is rounded to three significant figures.

Exercise 2.9.2

Several copper coins were placed dropped into a graduated cylinder, displacing 18.1 mL of water. Using Table 2.9.1, what is the mass of the coins?

Answer

162 g

Summary

- Density is the ratio of the mass of an object to its volume.
- Gases are less dense than either solids or liquids.
- Both liquid and solid materials can have a variety of densities.
- For liquids and gases, the temperature will affect the density to some extent.
- Density can be used as a conversion factor between mass and volume.

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2.11: Exercises

2.1: Taking Measurements

1. Identify the unit in each quantity.
    a. 2 boxes of crayons
    b. 3.5 grams of gold

Answer

    a. boxes of crayons
    b. grams of gold

2. Identify the unit in each quantity.
    a. 32 oz of cheddar cheese
    b. 0.045 cm^3 of water

Answer

    a. oz of cheddar cheese
    b. cm^3 of water

3. Identify the unit in each quantity.
    a. 9.58 s (the current world record in the 100 m dash)
    b. 6.14 m (the current world record in the pole vault)

Answer

    a. s
    b. m

4. Identify the unit in each quantity.
    a. 2 dozen eggs
    b. 2.4 km/s (the escape velocity of the moon, which is the velocity you need at the surface to escape the moon's gravity)

Answer

    a. dozen eggs
    b. km/s

2.2: Scientific Notation

5. Express these numbers in scientific notation.
    a. 56.9
    b. 563,100
    c. 0.0804
    d. 0.00000667

Answer

    a. 5.69×10^1

b. 5.631×10^{5}
c. 8.04×10^{-2}
d. 6.67×10^{-6}

6. Express these numbers in scientific notation.
   a. −890,000
   b. 602,000,000,000
   c. 0.0000004099
   d. 0.000000000000011

Answer

a. -8.9×10^{5}
b. 6.02×10^{11}
c. 4.099×10^{-7}
d. 1.1×10^{-14}

7. Express these numbers in scientific notation.
   a. 0.00656
   b. 65,600
   c. 4,567,000
   d. 0.000005507

Answer

a. 6.56×10^{-3}
b. 6.56×10^{4}
c. 4.567×10^{6}
d. 5.507×10^{-6}

8. Express these numbers in scientific notation.
   a. 65
   b. −321.09
   c. 0.000077099
   d. 0.000000000218

Answer

a. 6.5×10^{1}
b. -3.2109×10^{2}
c. 7.7099×10^{-5}
d. 2.18×10^{-10}

9. Express these numbers in decimal notation.
   a. 1.381×10^{5}
   b. 5.22×10^{-7}
   c. 9.998×10^{4}

Answer

a. 138,100
b. 0.000000522
c. 99,980

10. Express these numbers in decimal notation.

a. 7.11×10^{-2}
b. 9.18×10^{2}
c. 3.09×10^{-10}

Answer

a. 0.0711
b. 918
c. 0.000000000309

11. Express these numbers in decimal notation.

a. 8.09×10^{0}
b. 3.088×10^{-5}
c. -4.239×10^{2}

Answer

a. 8.09
b. 0.00003088
c. −423.9

12. Express these numbers in decimal notation.

a. 2.87×10^{-8}
b. 1.78×10^{11}
c. 1.381×10^{-23}

Answer

a. 0.0000000287
b. 178,000,000,000
c. 0.00000000000000000000001381

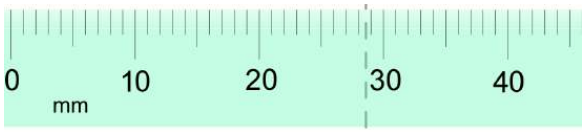
2.3: Significant Figures

13. Express each measurement to the correct number of significant figures.

a.



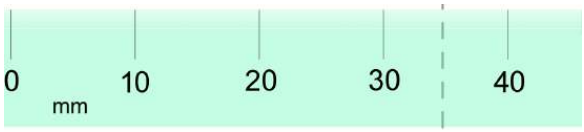
b.



c.



d.



Answer

a. 6.3 psi
b. 28.6 mm
c. 477 psi
d. 35 mm

14. How many significant figures do these numbers have?

a. 23
b. 23.0
c. 0.00023
d. 0.0002302

Answer

a. 2
b. 3
c. 2
d. 4

15. How many significant figures do these numbers have?
    a. 5.44×10^{8}
    b. 1.008×10^{-5}
    c. 43.09
    d. 0.0000001381

Answer

a. 3
b. 4
c. 4
d. 4

16. How many significant figures do these numbers have?
    a. 765,890
    b. 765,890.0
    c. 1.2000×10^{5}
    d. 0.0005060

Answer

a. 5 or 6, ambiguous
b. 7
c. 5
d. 4

17. How many significant figures do these numbers have?
    a. 0.009
    b. 0.0000009
    c. 65,444
    d. 65,040

Answer

a. 1
b. 1
c. 5
d. 4 or 5, ambiguous

18. Write the number 87,449 in scientific notation with four significant figures.

Answer

8.745×10^{4}

19. Write the number 0.000066600 in scientific notation with five significant figures.

Answer

6.6600×10^{-5}

20. Write the number 0.0000558 in scientific notation with two significant figures.

Answer

5.6×10^{-5}

2.4: Significant Figures in Calculations

21. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 56.0 + 3.44 = ?
b. 0.00665 + 1.004 = ?
c. 45.99 − 32.8 = ?
d. 45.99 − 32.8 + 75.02 = ?

Answer

a. 59.4
b. 1.011
c. 13.2
d. 88.2

22. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 1.005 + 17.88 = ?
b. 5,670.0 − 324 = ?
c. 405,007 − 123.3 = ?
d. 55.5 + 66.66 − 77.777 = ?

Answer

a. 18.89
b. 5,346
c. 404,884
d. 44.4

23. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 56.7 × 66.99 = ?
b. 1.00 ÷ 77 = ?
c. 1.00 ÷ 77.0 = ?
d. 6.022 × 1.89 = ?

Answer

a. 3.80×10^{3}
b. 0.013

c. 0.0130
d. 11.4

24. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 0.000440 × 17.22 = ?
b. 203,000. ÷ 0.044 = ?
c. 67 × 85.0 × 0.0028 = ?
d. 999,999 ÷ 33.1 = ?

Answer

a. 0.00758
b. 4.6×10^6
c. 16
d. 3.02×10^4

25. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 67,883 × 0.004321 = ?
b. $(9.67 \times 10^3) \times 0.0055087$ = ?

Answer

a. 293.3
b. 53.3

26. Compute and express each answer with the proper number of significant figures, rounding as necessary.

a. 18,900. × 76.33 ÷ 0.00336 = ?
b. 0.77604 ÷ 76,003 × 8.888 = ?

Answer

a. 4.29×10^8
b. 9.075×10^{-5}

2.5: The Metric System

27. Indicate what multiplier each prefix represents.

a. k
b. m
c. M

Answer

a. 10^3 or 1,000 ×
b. 10^{-3} or $\frac{1}{1000}$ ×
c. 10^6 or 1,000,000 ×

28. Indicate what multiplier each prefix represents.

a. c

b. G
c. μ

Answer

a. 10^{-2} or $\frac{1}{100}$ ×
b. 10^{9} or 1,000,000,000 ×
c. 10^{-6} or $\frac{1}{1{,}000{,}000}$ ×

29. Give the prefix that represents each multiplier.

a. 10^{-3} or $\frac{1}{1000}$ ×
b. 10^{3} or 1,000 ×
c. 10^{9} or 1,000,000,000 ×

Answer

a. m or milli
b. k or kilo
c. G or giga

30. Give the prefix that represents each multiplier.

a. $\frac{1}{1{,}000{,}000{,}000}$ ×
b. $\frac{1}{100}$ ×
c. 1,000,000 ×

Answer

a. n or nano
b. c or centi
c. M or mega

31. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 3.44×10^{-9} s
b. 3,500 L
c. 0.045 m

Answer

a. 3.44 ns
b. 3.5 kL
c. 45 mm

32. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 0.000066 m/s (Hint: you need consider only the unit in the numerator.)
b. 4.66×10^{6} s
c. 7,654 L

Answer

a. 66 μm/s
b. 4.66 Ms

c. 7.654 kL

33. Express each quantity in a more appropriate unit. There may be more than one acceptable answer.

a. 43,600 mL
b. 0.0000044 m
c. 1,438 ms

Answer

a. 43.6 L
b. 4.4 µm
c. 1.438 s

2.6: Problem Solving and Unit Conversions

34. Write the two conversion factors that exist between the two given units.

a. milliliters and liters
b. nanoseconds and seconds
c. kilometers and meters

Answer

a. $\frac{1000\ \text{mL}}{1\ \text{L}}$ and $\frac{1\ \text{L}}{1000\ \text{mL}}$

b. $\frac{1{,}000{,}000{,}000\ \text{ns}}{1\ \text{s}}$ and $\frac{1\ \text{s}}{1{,}000{,}000{,}000\ \text{ns}}$

c. $\frac{1\ \text{km}}{1000\ \text{m}}$ and $\frac{1000\ \text{m}}{1\ \text{km}}$

35. Write the two conversion factors that exist between the two given units.

a. kilograms and grams
b. milliseconds and seconds
c. centimeters and meters

Answer

a. $\frac{1\ \text{kg}}{1000\ \text{g}}$ and $\frac{1000\ \text{g}}{1\ \text{kg}}$

b. $\frac{1000\ \text{ms}}{1\ \text{s}}$ and $\frac{1\ \text{s}}{1000\ \text{ms}}$

c. $\frac{100\ \text{cm}}{1\ \text{m}}$ and $\frac{1\ \text{m}}{100\ \text{cm}}$

36. Perform the following conversions.

a. 5.4 km to meters
b. 0.665 m to millimeters
c. 0.665 m to kilometers

Answer

a. 5,400 m
b. 665 mm
c. 6.65×10^{-4} km

37. Perform the following conversions.

a. 90.6 mL to liters
b. 0.00066 ML to liters
c. 750 L to kiloliters

Answer

a. 0.0906 L
b. 660 L
c. 0.75 kL

38. Perform the following conversions.

a. 17.8 µg to grams
b. 7.22×10^{2} kg to grams
c. 0.00118 g to nanograms

Answer

a. 1.78×10^{-5} g
b. 7.22×10^{5} g
c. 1.18×10^{6} ng

39. Perform the following conversions.

a. 833 ns to seconds
b. 5.809 s to milliseconds
c. 2.77×10^{6} s to megaseconds

Answer

a. 8.33×10^{-7} s
b. 5,809 ms
c. 2.77 Ms

40. Perform the following conversions.

a. 45.0 m/min to meters/second
b. 0.000444 m/s to micrometers/second
c. 60.0 km/h to kilometers/second

Answer

a. 0.750 m/s
b. 444 µm/s
c. 0.0167 km/s

41. Perform the following conversions.

a. 3.4×10^2 cm/s to centimeters/minute
b. 26.6 mm/s to millimeters/hour
c. 13.7 kg/L to kilograms/milliliters

Answer

a. 2.0×10^4 cm/min
b. 9.58×10^4 mm/h
c. 0.0137 kg/mL

2.7: Solving Multi-Step Conversion Problems

42. Perform the following conversions.

a. 0.674 kL to milliliters
b. 2.81×10^{12} mm to kilometers
c. 94.5 kg to milligrams

Answer

a. 6.74×10^5 mL
b. 2.81×10^6 km
c. 9.45×10^7 mg

43. Perform the following conversions.

a. 6.79×10^{-6} kg to micrograms
b. 1.22 mL to kiloliters
c. 9.508×10^{-9} ks to milliseconds

Answer

a. 6.79×10^3 µg
b. 1.22×10^{-6} kL
c. 9.508×10^{-3} ms

44. Perform the following conversions.

a. 6.77×10^{14} ms to kiloseconds
b. 34,550,000 cm to kilometers

Answer

a. 6.77×10^8 ks
b. 345.5 km

45. Perform the following conversions.

a. 4.701×10^{15} mL to kiloliters
b. 8.022×10^{-11} ks to microseconds

Answer

a. 4.701×10^9 kL

b. 0.08022 μs

46. Perform the following conversions. Note that you will have to convert units in both the numerator and the denominator.
    a. 88 ft/s to miles/hour
    b. 0.00667 km/h to meters/second

Answer
    a. 6.0×10^{1} mi/h
    b. 1.85×10^{-3} m/s

47. Perform the following conversions. Note that you will have to convert units in both the numerator and the denominator.
    a. 3.88×10^{2} mm/s to kilometers/hour
    b. 1.004 kg/L to grams/milliliter

Answer
    a. 1.40 km/h
    b. 1.004 g/mL

2.8: Units Raised to a Power

48. Perform the following conversions.
    a. 9.44 m^2 to square centimeters
    b. 3.44×10^{8} mm^3 to cubic meters

Answer
    a. 9.44×10^{4} cm^2
    b. 0.344 m^3

49. Perform the following conversions.
    a. 0.00444 cm^3 to cubic meters
    b. 8.11×10^{2} m^2 to square nanometers

Answer
    a. 4.44×10^{-9} m^3
    b. 8.11×10^{20} nm^2

50. Why would it be inappropriate to convert square centimeters to cubic meters?

Answer

One is a unit of area and the other is a unit of volume.

51. The formula for the area of a triangle is ½ × base × height. What is the area of a triangle in square meters if its base is 166 mm and its height is 930.0 mm? Express the answer to the proper number of significant figures.

Answer

$0.0772\ m^2$

52. What is the area in square millimeters of a rectangle whose sides are 2.44 cm × 6.077 cm? Express the answer to the proper number of significant figures.

Answer

$1.48 \times 10^3\ mm^2$

53. What is the volume in cubic centimeters of a cube with sides of 0.774 m? Express the answer to the proper number of significant figures.

Answer

$4.64 \times 10^5\ cm^3$

54. The formula for the area of a triangle is ½ × base × height. What is the area of a triangle in square centimeters if its base is 1.007 m and its height is 0.665 m? Express the answer to the proper number of significant figures.

Answer

$3.35 \times 10^3\ cm^2$

2.9: Density

55. A block of metal alloy has a mass of 34.96 g. Its dimensions are 3.9 cm by 4.2 cm by 1.6 cm. What is the density of the metal alloy?

Answer

$1.3\ g/cm^3$

56. A plastic cylinder with a mass of 26.7 g is added to a graduated cylinder containing 45.8 mL of water. Once the cylinder was submerged, the volume increased to 61.3 mL. What was the density of the plastic cylinder?

Answer

1.72 g/mL

57. A sample of iron has a volume of 48.2 cm^3. If the density of iron is 7.87 g/cm^3, what is its mass?

Answer

379 g

58. A sample of air has a volume of 1,015 mL. What is its mass? Consult Table 2.9.1 for the density.

Answer

1.22 g

59. The volume of hydrogen used by the *Hindenburg*, the German airship that exploded in New Jersey in 1937, was 2.000×10^8 L. What mass of hydrogen was used by the airship? Consult Table 2.9.1 for the density.

Answer

1.7×10^7 g

60. The volume of an Olympic-sized swimming pool is 2.50×10^9 cm^3. If the pool is filled with alcohol (d = 0.789 g/cm^3), what mass of alcohol is in the pool?

Answer

1.97×10^9 g

61. A typical engagement ring has 0.77 cm^3 of gold. What mass of gold is present? Consult Table 2.9.1 for the density.

Answer

15 g

62. A typical mercury thermometer has 0.039 mL of mercury in it. What mass of mercury is in the thermometer? Consult Table 2.9.1 for the density.

Answer

0.53 g

63. What is the volume of 100.0 g of lead? Consult Table 2.9.1 for the density.

Answer

8.811 cm^3

64. What is the volume of 255.0 g of uranium if uranium has a density of 19.05 g/cm^3?

Answer

13.39 cm^3

65. What is the volume in liters of 222 g of neon if neon has a density of 0.900 g/L?

Answer

247 L

66. What is the volume in liters of 20.5 g of sulfur hexafluoride if sulfur hexafluoride has a density of 6.164 g/L?

Answer

3.33 L

67. Which has the greater volume, 100.0 g of iron (d = 7.87 g/cm^3) or 75.0 g of gold (d = 19.3 g/cm^3)?

Answer

Iron. The volume of the iron is 12.7 cm^3 and the volume of the gold is 3.89 cm^3.

68. Which has the greater volume, 25.0 g of hydrogen gas (d = 0.0840 g/L) or 100.0 g of argon gas (d = 1.78 g/L)?

Answer

Hydrogen. The volume of the hydrogen is 298 L and the volume of the argon is 56.2 L.

Additional Exercises

69. Evaluate 0.00000000552 × 0.0000000006188 and express the answer in scientific notation. You may have to rewrite the original numbers in scientific notation first.

Answer

3.42×10^{-18}

70. Evaluate 333,999,500,000 ÷ 0.00000000003396 and express the answer in scientific notation. You may need to rewrite the original numbers in scientific notation first.

Answer

9.835×10^{21}

71. Fill in the blank: 1 km = _____ µm.

Answer

10^9 µm

72. Fill in the blank: 1 Ms = _____ ns.

Answer

10^{15} ns

73. Fill in the blank: 1 cL = _____ ML.

Answer

10^{-8} ML

74. Fill in the blank: 1 mg = _____ kg.

Answer

10^{-6} kg

75. Convert a speed of 60.0 mi/h into kilometers/hour.

Answer

96.5 km/h

76. Convert a speed of 60.0 km/h into miles/hour.

Answer

37.3 mi/h

77. Convert 52.09 km/h into meters/second.

Answer

14.47 m/s

78. Convert 2.155 m/s into kilometers/hour.

Answer

7.758 km/h

79. What is the mass of 12.67 L of mercury? Consult Table 2.9.1 for the density.

Answer

1.72×10^5 g

80. What is the mass of 0.663 m^3 of air? Consult Table 2.9.1 for the density.

Answer

796 g

81. What is the volume of 2.884 kg of gold? Consult Table 2.9.1 for the density.

Answer

149 cm^3

82. What is the volume of 40.99 kg of cork? Assume a density of 0.22 g/cm^3.

Answer

$1.9 \times 10^5\ \text{cm}^3$

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CHAPTER OVERVIEW

3: Matter and Energy

3.1: What is Matter?
3.2: Physical States of Matter
3.3: Composition of Matter
3.4: Physical and Chemical Properties
3.5: Physical and Chemical Changes
3.6: Conservation of Mass
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Thumbnail Chapter 3: Rocket Launch (ChemLancer adaptation of NASA via Giphy)

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3.1: What is Matter?

Learning Objectives

- Define matter and explain how it is composed of building blocks known as "atoms."

We are all familiar with matter. **Matter** is anything that has mass and volume (takes up space). Matter includes atoms and anything made up of atoms, but not other energy phenomena or waves such as light or sound. While this simple definition is easily applied, the way people view matter is often broken down into two characteristic scales: the macroscopic and the microscopic.



Figure **3.1.1**: Moseley Hall Chemistry Lab. (Mbrickn via Wikimedia Commons)

The **macroscopic scale** is the scale on which objects or phenomena are large enough to be visible practically with the naked eye, without magnifying optical instruments. Everything that one can see, touch, and handle in the chemistry laboratory in Figure **3.1.1** is within the macroscopic scale. To describe each of these objects, only a few macroscopic properties are required. However, each of these items can be decomposed into smaller microscopic scale properties.

The **microscopic scale** is the scale of objects and events smaller than those that can easily be seen by the naked eye, requiring a lens or microscope to see them clearly. All of the everyday objects that we can bump into, touch, or squeeze are ultimately composed of atoms. This ordinary atomic matter is in turn made up of interacting subatomic particles – usually a nucleus of protons and neutrons, and a cloud of orbiting electrons. Because of this, a large number of variables are needed to describe such a system which complicates the characterization.

⚡ Matter vs. Mass

Matter should not be confused with mass as the two are not the same in modern physics. Matter is a physical substance of which systems may be composed; while mass is not a substance, but rather a quantitative property of matter and other substances or systems.

For most common objects that we encounter, it is fairly simple to demonstrate that they have mass and take up space. You might be able to imagine, however, the difficulty for people several hundred years ago to demonstrate that air had mass and volume. Air (and all other gases) are invisible to the eye, have very small masses compared to equal volumes of solids and liquids, and are quite easy to compress (change volume). Without sensitive equipment, it would have been difficult to convince people that gases are matter. Today, we can measure the mass of a small balloon when it is deflated and then blow it up, tie it off, and measure its mass again to detect the additional mass due to the air inside. The mass of air, under room conditions, that occupies a one quart jar is approximately 0.0002 pounds. This small amount of mass would have been difficult to measure in times before balances were designed to accurately measure very small masses. Later, scientists were able to compress gases into such a small volume that the gases turned into liquids, which made it clear that gases are matter.

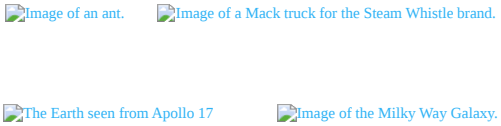


Figure **3.1.2**: Everything from an ant, to a truck, to the earth, and even the entire galaxy is composed of matter.

Even though the universe consists of "things" as wildly different as ants and galaxies, the matter that makes up all of these "things" is composed of a very limited number of building blocks. These building blocks are known as **atoms**, and so far, scientists have discovered or created a grand total of 118 different types of atoms. Scientists have given a name to each different type of atom. A

substance that is composed of only one type of atom is called an element. At this point, what should amaze us is that all forms of matter in our universe are made with only 118 different building blocks. In some ways, it's sort of like cooking a gourmet, five-course meal using only three ingredients! How is it possible? To answer that question, you have to understand the ways in which different elements are put together to form matter.

The most important method that nature uses to organize atoms into matter is the formation of **molecules**. Molecules are groups of two or more atoms that have been bonded together. There are millions of different ways to bond atoms together, which means that there are millions of different possible molecules. Each of these molecules has its own set of chemical properties, and it's these properties with which chemists are most concerned. You will learn a lot more about atoms and molecules, including how they were discovered, in a later part of the textbook.

Summary

- All matter has mass and occupies space. All physical objects are made of matter.
- Matter itself is composed of tiny building blocks known as "atoms". There are only 118 different types of atoms known to man.
- Frequently, atoms are bonded together to form "molecules".

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3.2: Physical States of Matter

Learning Objectives

- To describe the solid, liquid and gas phases.

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



Figure **3.2.1**: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances, such as oxygen and carbon dioxide, exist as gases at room temperature. Others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure 3.2.2 shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

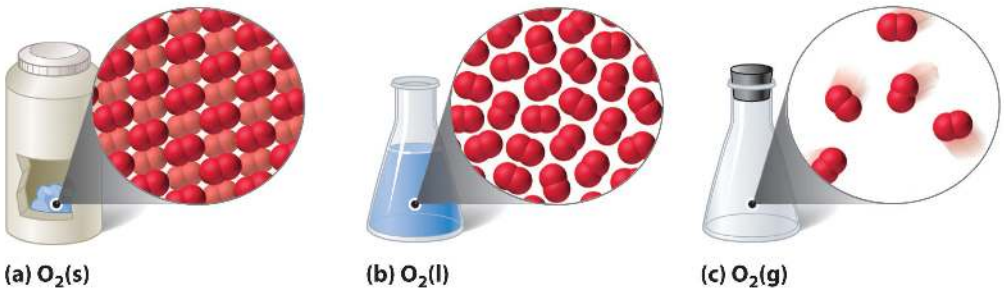


Figure **3.2.2**: A representation of the solid, liquid, and gas states. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 fills its container completely – regardless of the container's size or shape – and consists of widely separated molecules.

Plasma: A Fourth State of Matter

Technically speaking, a fourth state of matter called plasma exists, but it rarely occurs naturally on earth, so we will omit it from our study here.



Figure **3.2.3**: A plasma globe operating in a darkened room. (CC BY-SA 3.0; Chocolateoak).

Solids

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

Solids are defined by the following characteristics:

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

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

Solids usually have their constituent particles arranged in a regular, three-dimensional array called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure 3.2.4. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as **amorphous** (literally, “without form”) solids. Glass is one example of an amorphous solid.

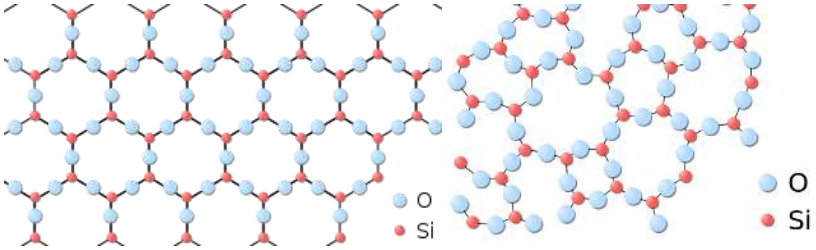


Figure **3.2.4**: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

Liquids

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

Liquids have the following characteristics:

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

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature. Mercury also has an ability to stick to itself (surface tension) – a property that all liquids exhibit. Mercury has a relatively high surface tension, which makes it very unique. Here you see mercury in its common liquid form. When liquid mercury heated to its boiling point of 357°C under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.



Video 3.2.1: Mercury boiling to become a gas.

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. This describes the gas state, which we will consider in more detail elsewhere. Like liquids, gases have no definite shape; but unlike solids and liquids, gases have no definite volume either. The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1000 or more.

Gases have the following characteristics:

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

Table 3.2.1: Characteristics of the Three States of Matter

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

✅ Example 3.2.1

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

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

Solution

A. This statement describes the liquid state.
B. This statement describes the gas state.

C. This statement describes the liquid state.

Exercise 3.2.1

What state or states of matter does each statement describe?

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

Answer A

solid

Answer B

gas

Answer C

solid

Summary

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

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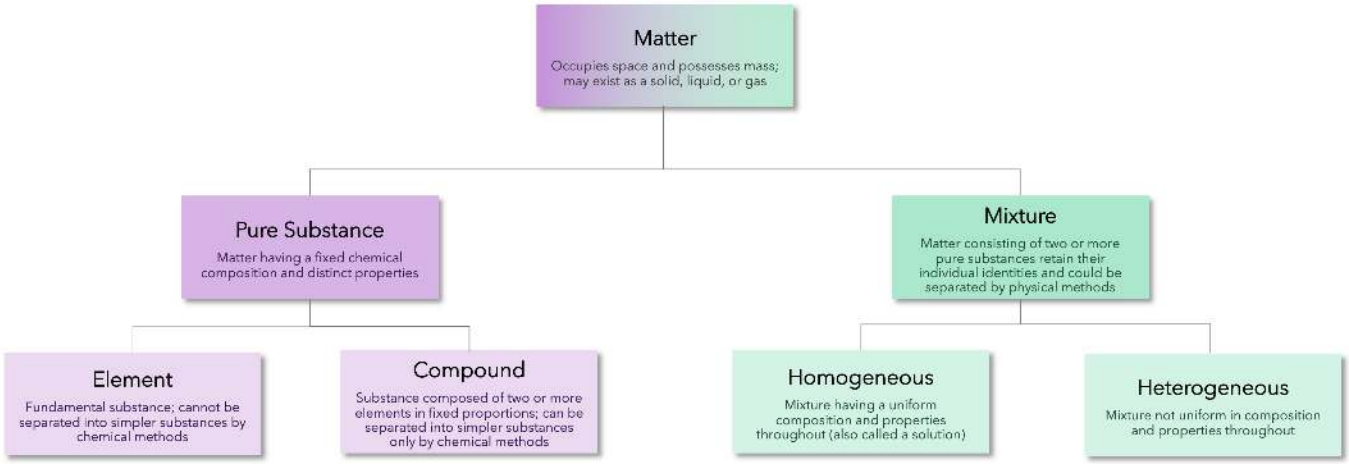
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3.3: Composition of Matter

Learning Objectives

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

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure 3.3.1). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.





gold



sodium chloride



apple cider vinegar



sweet chili sauce

Figure **3.3.1**: Relationships between the types of matter and the methods used to separate mixtures.

Ordinary table salt is called sodium chloride. It is considered a **pure substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a pure substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own

identity and properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

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

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup.

⚡ Phase

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

✅ Example 3.3.1

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

A. filtered tea
B. freshly squeezed orange juice
C. a compact disc
D. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
E. selenium

Given: a chemical substance

Asked for: its classification

Strategy:

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

Solution

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

✏️ Exercise 3.3.1

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

A. white wine
B. mercury
C. ranch-style salad dressing
D. table sugar (sucrose)

Answer A

homogeneous mixture (solution)

Answer B

element

Answer C

heterogeneous mixture

Answer D

compound

✅ Example 3.3.2

How would a chemist categorize each example of matter?

A. saltwater
B. soil
C. water
D. oxygen

Solution

A. Saltwater acts as if it were a single substance even though it contains two substances – salt and water. Saltwater is a homogeneous mixture, or a solution.
B. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
C. Water is a pure substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
D. Oxygen, a pure substance, is an element.

✏️ Exercise 3.3.2

How would a chemist categorize each example of matter?

A. coffee
B. hydrogen
C. an egg

Answer A

a homogeneous mixture (solution), assuming it is filtered coffee

Answer B

element

Answer C

heterogeneous mixture

Summary

- Matter can be classified into two broad categories: pure substances and mixtures.
- A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample.
- Elements and compounds are both examples of pure substances.
- Elements are the simplest substances made up of only one type of atom. Compounds are substances that are made up of more than one type of atom.
- Mixtures are physical combinations of two or more elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous.

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3.4: Physical and Chemical Properties

Learning Objectives

- To distinguish physical from chemical properties.

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

Physical Property

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

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 g/cm^3 at 25°C. The average densities of some common substances are in Table 3.4.1. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Table 3.4.1: Densities of Common Substances

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

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

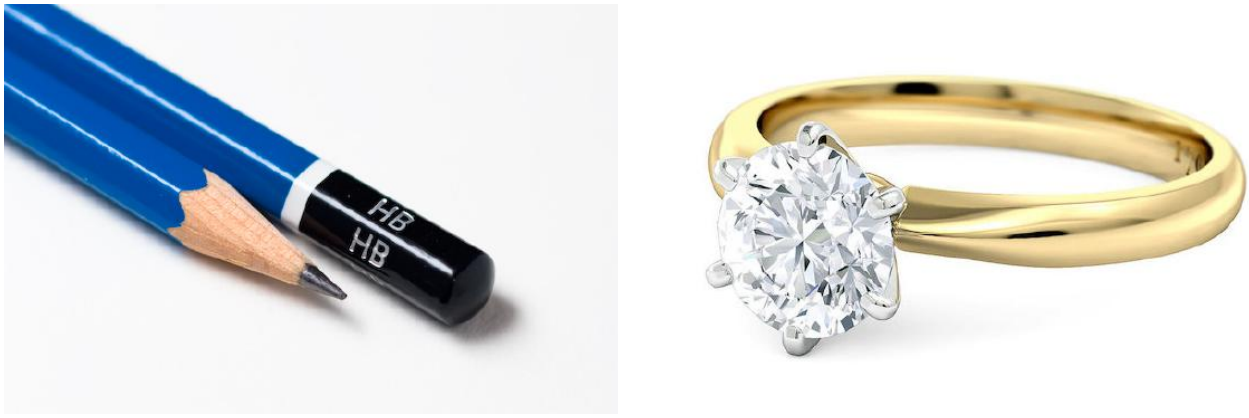


Figure **3.4.1**: Pencil (left) and diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

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

Chemical Properties

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



Figure **3.4.2**: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco. It was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure **3.4.2**). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 3.4.2: A Comparison of Physical and Chemical Properties

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

✅ Example 3.4.1

Which of the following is a chemical property of iron?

A. Iron corrodes in moist air
B. Density = 7.874 g/cm^3
C. Iron is soft when pure.
D. Iron melts at 1808 K.

Solution

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

✏️ Exercise 3.4.1

Which of the following is a physical property of matter?

A. corrosiveness
B. pH (acidity)
C. density
D. flammability

Answer

C. Density

Which of the following is a chemical property of matter?

A. flammability
B. melting point
C. boiling point
D. density

Answer

A. flammability

Summary

- A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance.
- Physical properties include color, density, hardness, and melting and boiling points.
- A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change.
- A chemical change always produces one or more types of matter that differ from the matter present before the change.

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3.5: Physical and Chemical Changes

Learning Objectives

- Label a change as chemical or physical.
- List evidence that can indicate a chemical change occurred.

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

Physical Change

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

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

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



Figure **3.5.1**: 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 composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

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

Chemical Change

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

One good example of a chemical change is burning a candle. The act of burning a candle actually results in the formation of new chemicals (carbon dioxide and water) from the burning of the wax. In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.



Figure **3.5.2**: Burning of wax to generate water and carbon dioxide is a chemical reaction.

We can't actually see molecules breaking and forming bonds, although that's what defines chemical changes. We have to make

other observations to indicate that a chemical change has happened. Some of the evidence for chemical change will involve the energy changes that occur in chemical changes, but some evidence involves the fact that new substances with different properties are formed in a chemical change.

Observations that help to indicate chemical change include:

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

✅ Example 3.5.1

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

A. Boiling water.
B. A nail rusting.
C. Dissolving salt in water. The resulting mixture is a clear solution.
D. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

Solution

A. Physical: Boiling is a physical change. When water boils, no bonds are broken or formed.
B. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change.
C. Physical: Because none of the properties changed, this is a physical change. The salt is easily retrieved by evaporating or boiling off the water. Any water that is evaporated or boiled off may be retrieved by condensing the vapor back to a liquid. No color *change* occurred or other evidence of chemical change.
D. Chemical: The formation of a precipitate and the color change from colorless to yellow indicate a chemical change.

✏️ Exercise 3.5.1

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

A. A mirror is broken.
B. An iron nail corrodes in moist air
C. Copper metal is melted.
D. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer A

physical change

Answer B

chemical change

Answer C

physical change

Answer D

chemical change

Separating Mixtures Through Physical Changes

Homogeneous mixtures (solutions) can be separated into their component substances by physical processes that rely on differences in one or more physical properties, such as differences in their boiling points. Two of these separation methods are distillation and filtration. **Distillation** makes use of differences in volatility, a measure of how easily a substance is converted to a gas at a given temperature. A simple distillation apparatus for separating a mixture of substances, at least one of which is a liquid, is shown in Figure 3.5.3. The most volatile component boils first and is condensed back to a liquid in the water-cooled condenser, from which it flows into the receiving flask. If a solution of salt and water is distilled, for example, the more volatile component, pure water, collects in the receiving flask while the salt remains in the distillation flask.

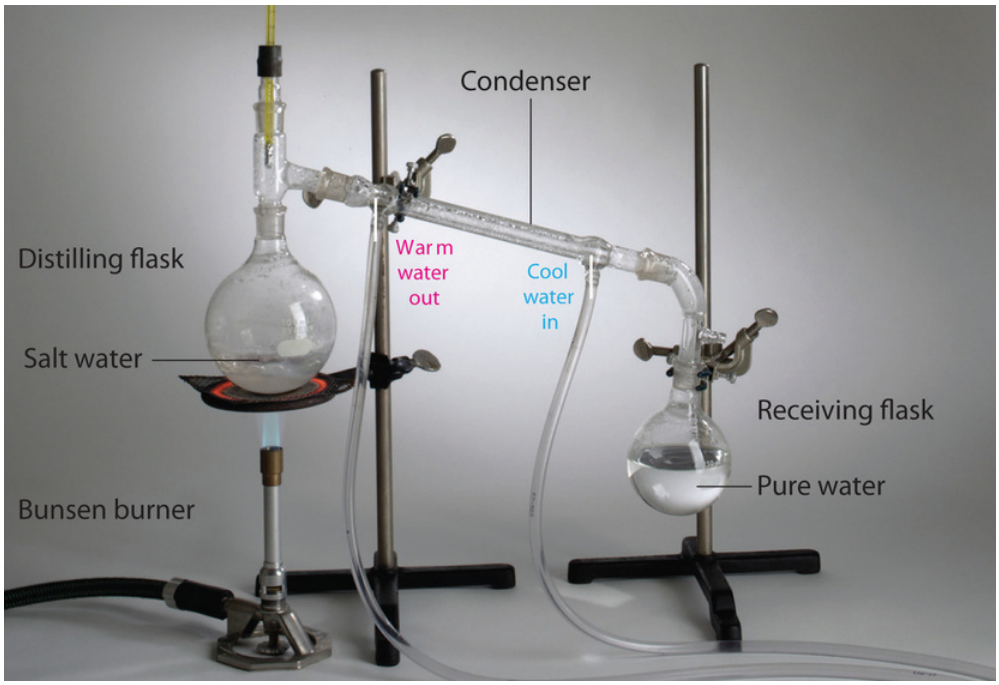


Figure **3.5.3:** The distillation of a solution of table salt in water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.

Mixtures of two or more liquids with different boiling points can be separated with a more complex distillation apparatus. One example is the refining of crude petroleum into a range of useful products: aviation fuel, gasoline, kerosene, diesel fuel, and lubricating oil (in the approximate order of decreasing volatility). Another example is the distillation of alcoholic spirits such as brandy or whiskey. This relatively simple procedure caused more than a few headaches for federal authorities in the 1920s during the era of Prohibition, when illegal stills proliferated in remote regions of the United States.

Another example for using physical properties to separate mixtures is filtration (Figure 3.5.4). **Filtration** is any mechanical, physical or biological operation that separates solids from fluids (liquids or gases) by adding a medium through which only the fluid can pass. The fluid that passes through is called the filtrate. There are many different methods of filtration; all aim to attain the separation of substances. Separation is achieved by some form of interaction between the substance or objects to be removed and the filter. The substance that is to pass through the filter must be a fluid, i.e. a liquid or gas. Methods of filtration vary depending on the location of the targeted material, i.e. whether it is dissolved in the fluid phase or suspended as a solid.

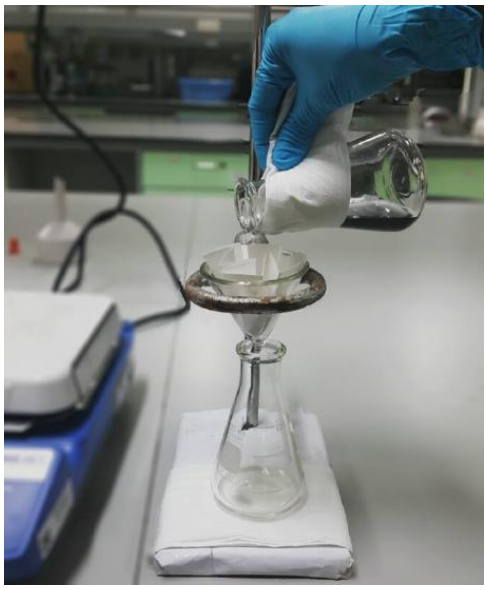


Figure **3.5.4:** Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).

Summary

- Chemists make a distinction between two different types of changes that they study – physical changes and chemical changes.
- Physical changes are changes that do not alter the identity of a substance.
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes. Observations that indicate a chemical change has occurred include color change, temperature change, light given off, formation of bubbles, formation of a precipitate, etc.

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3.6: Conservation of Mass

Combustion is the process of burning something in the presence of oxygen. In Figure 3.6.1, the first image shows a pyrotechnic display at a P!nk concert caused by burning a combustible fuel. The second image shows roasting marshmallows over hot coals and ashes that remain from burning wood. In each case, it seems as if the material (fuel) being burned simply disappears. Sure, heat is obtained and it is obvious that ashes remain after the wood burns, but what really happens? Does the matter simply disappear?

In each case, oxygen gas (obtained from the air) is used to drive the combustion by undergoing a chemical reaction with the fuel as it burns. The combustion leads to the formation of carbon dioxide gas and water vapor. These gases float off into the air. In the case of burning wood, ashes remain as well. Had we been able to measure the combined mass of the fuel and oxygen before they were burned and compared it to the combined mass of the carbon dioxide, water vapor, and any other product (such as the remaining ashes in the fire pit) after they were burned, we would find out that the total mass of the matter before their combustion is equal to the total mass of the matter after combustion.



Figure **3.6.1**: Flames in a pyrotechnic display at a P!nk concert. Roasting marshmallows over a backyard fire pit. (Lance S. Lund)

Law of Conservation of Mass

The law of conservation of mass was described in 1789 by a French chemist, Antoine Lavoisier. The **law of conservation of mass** states that matter can neither be created nor destroyed in a chemical reaction. For example, when wood burns, the mass of the soot, ashes, and gases equals the original mass of the wood and the oxygen when it first reacted. So the mass of the product equals the mass of the reactant. Reactants are the substances present before a change occurs while product(s) are the substances formed as the result of a chemical reaction (Video 3.6.1). Matter and its corresponding mass may not be able to be created or destroyed, but it may change forms to other substances like liquids, gases, and solids.



Video 3.6.1: A demonstration showing the conservation of mass in action.

If you witness a 300 kg tree burn to the ground, there are only ashes left after burning and all of them together weigh 10 kg. It may make you wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke, so the only thing left that you can see is the 10 kg of ash. If you know the law of conservation of mass, then you know that the other 290 kg has to go somewhere because the products have to equal the mass of the tree before it burnt down.

✅ Example 3.6.1

If heating 10.0 grams of calcium carbonate ($CaCO_3$) produces 4.4 g of carbon dioxide (CO_2) and 5.6 g of calcium oxide (CaO), show that these observations are in agreement with the law of conservation of mass.

Solution

$$\text{Mass of the reactants} = \text{Mass of the products}$$

$$10.0\,\text{g of } CaCO_3 = 4.4\,\text{g of } CO_2 + 5.6\,\text{g of CaO}$$

$$10.0\,\text{g of reactant} = 10.0\,\text{g of products}$$

Because the mass of the reactant is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

✏️ Exercise 3.6.1

Potassium hydroxide (**KOH**) readily reacts with carbon dioxide, CO_2, to produce potassium carbonate, K_2CO_3, and water, H_2O. How many grams of potassium carbonate are produced if 224.4 g of KOH reacts with 88.0 g of CO_2? The reaction also produces 36.0 g of water.

Answer

276.4 g of potassium carbonate

The law is also applicable to physical changes. For example, if you have an ice cube that melts into a liquid and you heat that liquid up, it becomes a gas. It will appear to have disappeared but is still there. The mass of the ice cube, the mass of the liquid, and the mass of the gas will all be the same.

Summary

- Burning and other changes in matter do not destroy matter.
- The mass of matter is always the same before and after the changes occur.

- The law of conservation of mass states that matter cannot be created or destroyed.

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3.7: Energy

Learning Objectives

- Define heat and work.
- Distinguish between kinetic energy and potential energy.

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we are awake or asleep.

Ability to Do Work or Produce Heat

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

When scientists speak of **heat**, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature, as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot – the pot and the food inside the pot are cold. As a result, heat moves from the hot stove element to the cold pot. After a while, enough heat is transferred from the stove to the pot, raising the temperature of the pot and all of its contents (Figure 3.7.1).



Figure **3.7.1**: Energy is transferred as heat from the hot stove element to the cooler pot until the pot and its contents become just as hot as the element. The energy that is transferred into the pot as heat is then used to cook the pasta.

Heat is only one way in which energy can be transferred. Energy can also be transferred as **work**. The scientific definition of work is a force (any push or pull) applied over a distance. When you push an object and cause it to move, you do work, and you transfer some of your energy to the object. At this point, it's important to warn you of a common misconception. Sometimes we think that the amount of work done can be measured by the amount of effort put in. This may be true in everyday life, but it is not true in science. By definition, scientific work requires that force be applied over a distance. It does not matter how hard you push or how hard you pull. If you have not moved the object, you haven't done any work.

So far, we've talked about the two ways in which energy can be transferred from one place, or object, to another. Energy can be transferred as heat, and energy can be transferred as work. But the question still remains – what *is* energy?

Kinetic Energy

Energy is simply the capacity to supply heat or do work. Machines use energy, our bodies use energy, energy comes from the sun, energy comes from volcanoes, energy causes forest fires, and energy helps us to grow food. With all of these seemingly different types of energy, it's hard to believe that there are really only two different forms of energy: kinetic energy and potential energy. **Kinetic energy** is energy associated with motion. When an object is moving, it has kinetic energy. When the object stops moving, it has no kinetic energy. While all moving objects have kinetic energy, not all moving objects have the same amount of kinetic

energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has.

Kinetic energy is very common, and it's easy to spot examples of it in the world around you. Sometimes we even try to capture kinetic energy and use it to power things like our home appliances. If you have driven through southwestern Minnesota or northern Iowa, you have probably seen the windmills out in the farm fields or along the highways (Figure 3.7.2). As wind rushes over the hills, the kinetic energy of the moving air particles turns the windmills, trapping the wind's kinetic energy so that people can use it in their houses and offices.



Figure **3.7.2**: A wind farm in northwestern Iowa harnesses the kinetic energy of the wind.

Potential Energy

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

Any stored energy is potential energy. There are a lot of different ways in which energy can be stored, and this can make potential energy very difficult to recognize. In general, an object has potential energy because of its position relative to another object. For example, when a rock is held above the earth, it has potential energy because of its position relative to the ground. This is potential energy because the energy is stored for as long as the rock is held in the air. Once the rock is dropped, though, the stored energy is released as kinetic energy as the rock falls.

Chemical Energy

There are other common examples of potential energy. A ball at the top of a hill stores potential energy until it is allowed to roll to the bottom. When two magnets are held next to one another, they store potential energy too. For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called **chemical potential energy**. Chemical potential energy is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position?

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

⚖ Chemical Potential Energy

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

At this point, you may wonder just how useful chemical potential energy is. If you want to release the potential energy stored in an object held above the ground, you just drop it. But how do you get potential energy out of chemicals? It's actually not difficult. Use the fact that different chemicals have different amounts of potential energy. If you start with chemicals that have a lot of potential energy and allow them to react and form chemicals with less potential energy, all the extra energy that was in the chemicals at the

beginning, but not at the end, is released.

Units of Energy

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

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

$$1\ \text{kilocalorie} = 1\ \text{Calorie} = 1000\ \text{calories} \tag{3.7.1}$$

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

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

$$1\ J = 0.2390\ cal \text{ or } 1\ cal = 4.184\ J \tag{3.7.2}$$

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

$$400\ Cal = 400\ \cancel{kcal} \times \frac{4.184\ kJ}{1\ \cancel{kcal}} = \boxed{1.67 \times 10^3\ kJ}$$

Summary

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

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3.8: Energy and Chemical and Physical Change

Learning Objectives

- State the law of conservation of energy.
- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.
- Determine whether a reaction is endothermic or exothermic through observations, temperature changes, or an energy diagram.

So far, we have talked about how energy exists as either kinetic energy or potential energy and how energy can be transferred as either heat or work. While it's important to understand the difference between kinetic energy and potential energy and the difference between heat and work, the truth is, energy is constantly changing. Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy. Likewise, energy that is transferred as work might later end up transferred as heat, while energy that is transferred as heat might later end up being used to do work.

Even though energy can change form, it must still follow one fundamental law: energy cannot be created or destroyed, it can only be changed from one form to another. This law is known as the **law of conservation of energy**. In a lot of ways, energy is like money. You can exchange quarters for dollar bills and dollar bills for quarters, but no matter how often you convert between the two, you will not end up with any more or any less money than you started with. Similarly, you can transfer (or spend) money using cash, or transfer money using a credit card; but you still spend the same amount of money, and the store still makes the same amount of money.

A campfire is an example of basic thermochemistry. The reaction is initiated by the application of heat from a match. The reaction converting wood to carbon dioxide and water (among other things) continues, releasing heat energy in the process. This heat energy can then be used to cook food, roast marshmallows, or just keep warm when it's cold outside.



Figure **3.8.1**: A looped campfire. [Frank Vincentz (looped gif Jahobr) via Wikimedia Commons]

Exothermic and Endothermic Processes

When physical or chemical changes occur, they are generally accompanied by a transfer of energy. The law of conservation of energy states that in any physical or chemical process, energy is neither created nor destroyed. In other words, the entire energy in the universe is conserved. In order to better understand the energy changes taking place during a reaction, we need to define two parts of the universe: the system and the surroundings. The **system** is the specific portion of matter in a given space that is being studied during an experiment or an observation. The **surroundings** are everything in the universe that is not part of the system. In practical terms for a laboratory chemist, the system is the particular chemicals being reacted while the surroundings are the immediate vicinity within the room. During most processes, energy is exchanged between the system and the surroundings. If the system loses a certain amount of energy, that same amount of energy is gained by the surroundings. If the system gains a certain amount of energy, that energy is supplied by the surroundings.

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings and so the temperature of the surroundings decreases. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases.

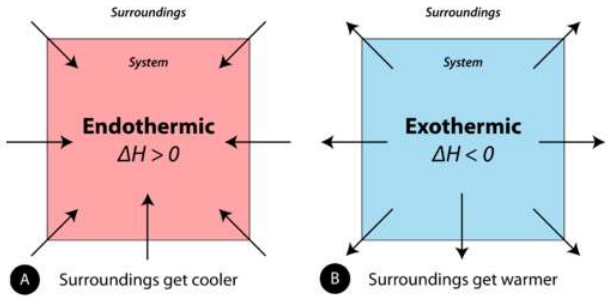


Figure **3.8.2**: (A) Endothermic reaction. (B) Exothermic reaction.

During phase changes, energy changes are usually involved. For example, when solid dry ice vaporizes (physical change), carbon dioxide molecules must absorb energy in order to be able to move about freely as a gas. When liquid water becomes ice, energy is released as the molecules slow their movement and form a solid. Remember that all chemical reactions involve a change in the bonds of the reactants. The bonds in the reactants are broken and the bonds of the products are formed. Chemical bonds have potential energy or "stored energy." Because we are changing the bonding, this means we are also changing how much of this "stored energy" there is in a reaction.

Energy changes are frequently shown by drawing an energy diagram. Energy diagrams show the stored/hidden energy of the reactants and products as well as the activation energy. If, on an energy diagram, the products have more stored energy than the reactants started with, the reaction is endothermic. You had to give the reaction energy. If, on the energy diagram, the products have less stored energy than the reactants started with, the reaction is exothermic.

✅ Example 3.8.1

Label each of the following processes as endothermic or exothermic.

A. water boiling
B. gasoline burning
C. ice forming on a pond

Solution

A. Endothermic – you must put a pan of water on the stove and provide heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
B. Exothermic – when you burn something, it feels hot to you because it is giving off heat into the surroundings.
C. Exothermic – think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.

✏️ Exercise 3.8.1

Label each of the following processes as endothermic or exothermic.

A. The temperature increases when an acid and base are mixed together.
B. An ice cube melts in your hand.

Answer A

exothermic

Answer B

endothermic

Summary

- Phase changes involve changes in energy.
- All chemical reactions involve changes in energy. This may be a change in heat, electricity, light, or other forms of energy. The total amount of energy, though, remains constant.
- Reactions that absorb energy are endothermic. Reactions that release energy are exothermic.

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3.9: Temperature

Learning Objectives

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

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



Figure **3.9.1**: The particles in glowing charcoal have a much greater kinetic energy than the particles in frost and snow.

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



Figure 3.9.2: Daniel Gabriel Fahrenheit (left), Anders Celsius (center), and Lord Kelvin (right).

The Fahrenheit Scale

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

The Celsius Scale

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

The Kelvin Scale

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

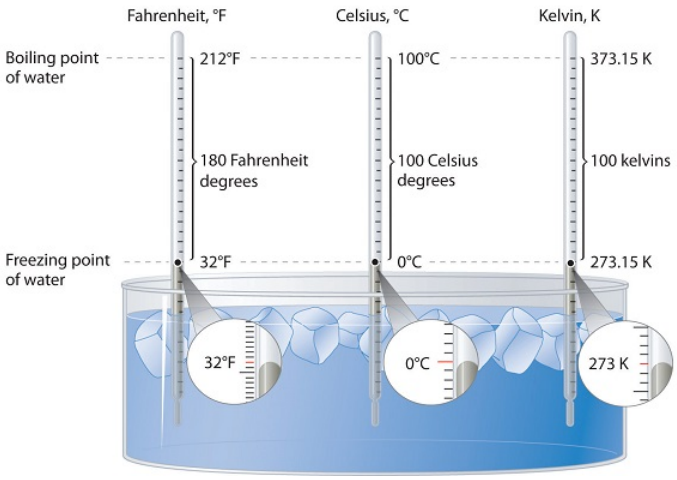


Figure 3.9.3: A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are $\frac{9}{5}$ the size of a degree Fahrenheit (°F). (CC BY-SA-NC 3.0; anonymous)

Converting Between Scales

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

$$T_K = T_{°C} + 273.15 \tag{3.9.1}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: −40°C = −40°F. The relationship between the scales is as follows, where the coefficient for degrees Fahrenheit is exact:

$$T_{°C} = \tfrac{5}{9}(T_{°F} - 32) \tag{3.9.2}$$

$$T_{°F} = \tfrac{9}{5}T_{°C} + 32 \tag{3.9.3}$$

✅ Example 3.9.1

A recipe calls for preheating the oven to 175°C. What is this temperature in K and °F?

Solution

Converting from Celsius to Kelvin requires the use of Equation 3.9.1:

$$T_K = (T_{°C} + 273.15) = 175 + 273.15 = \boxed{448\ \text{K}}$$

Converting from Celsius to Fahrenheit requires the use of Equation 3.9.3:

$$T_{°F} = \tfrac{9}{5}T_{°C} + 32 = \tfrac{9}{5}(175) + 32 = 315 + 32 = \boxed{347°\text{F}}$$

✏️ Exercise 3.9.1

Perform the following temperature conversions.

A. The standard temperature for several chemical and physical relationships and equations is 25°C. What is this temperature in kelvins?
B. The wind chill on a cold winter day is −40.0°F. Express this temperature in °C.
C. While traveling to Africa, a temperature scan is required to determine if a person has a fever. You are told your temperature is 37.0°C. What is this temperature in °F?

Answer A

298 K

Answer B

−40.0°C

Answer C

98.6°F

Summary

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

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3.11: Exercises

3.1: What is Matter?

1. What is matter?

Answer

Matter is anything that has mass and takes up space.

2. Atoms are the building blocks of all matter in the universe. As of 2011, scientists only know of 118 different types of atoms. How do you think it's possible to generate so many different forms of matter using only 118 types of building blocks?

Answer

Atoms can combine in many different ways to form compounds.

3. Which do you think has more matter, a cup of water or a cup of mercury? Explain.

Answer

A cup of mercury. Both have the same volume. However, the cup of mercury will weigh more so it should have more matter in it.

3.2: Physical States of Matter

4. Give an example of matter in each phase: solid, liquid, or gas.

Answer

Answers will vary. Solid = steel; liquid = water; gas = air.

3.3: Composition of Matter

5. Distinguish between an element and a compound. About how many of each are known?

Answer

Elements cannot be broken down to chemically simplier components. Compounds can because they are made up of more than one element. There are 118 known elements. There are millions of known compounds.

6. What is the difference between a homogeneous mixture and a heterogeneous mixture?

Answer

A homogeneous mixture is uniform throughout. A heterogeneous mixture is not uniform throughout.

7. Identify each as a heterogeneous mixture or a homogeneous mixture.
   a. Salt is mixed with pepper.
   b. Sugar is dissolved in water.
   c. Pasta is cooked in boiling water.

Answer

Answer

a. heterogeneous
b. homogeneous
c. heterogeneous

8. Identify each as a heterogeneous mixture or a homogeneous mixture.
   a. air
   b. dirt
   c. a big screen television

Answer

a. homogeneous
b. heterogeneous
c. heterogeneous

9. In Exercise 7, which choices are also solutions?

Answer

b, sugar dissolved in water

10. In Exercise 8, which choices are also solutions?

Answer

a, air

11. Identify each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.
   a. apple juice
   b. copper
   c. carbon dioxide, a 1:2 combination of carbon and oxygen
   d. ham and cheese sandwich

Answer

a. homogeneous mixture
b. element
c. compound
d. heterogeneous mixture

12. Identify each of the following as an element, compound, homogeneous mixture, or heterogeneous mixture.
   a. chocolate chip ice cream
   b. magnesium nitride, a 3:2 combination of magnesium and nitrogen
   c. gasoline
   d. fluorine

Answer

a. heterogeneous mixture
b. compound

c. homogeneous mixture
d. element

3.4: Physical and Chemical Properties

13. Does each statement represent a physical property or a chemical property?

a. Sulfur is yellow.
b. Steel wool burns when ignited by a flame.
c. A gallon of milk weighs over eight pounds.

Answer

a. physical
b. chemical
c. physical

14. Does each statement represent a physical property or a chemical property?

a. A pile of leaves slowly rots in the backyard.
b. In the presence of oxygen, hydrogen can interact to make water.
c. Gold can be stretched into very thin wires.

Answer

a. chemical
b. chemical
c. physical

3.5: Physical and Chemical Changes

15. Does each statement represent a physical change or a chemical change?

a. Water boils and becomes steam.
b. Food is converted into usable form by the digestive system.
c. The alcohol in many thermometers freezes at about −40 degrees Fahrenheit.

Answer

a. physical
b. chemical
c. physical

16. Does each statement represent a physical change or a chemical change?

a. Plants use photosynthesis to convert carbon dioxide into glucose.
b. The house across the street has been painted a new color.
c. The elements sodium and chlorine come together to make a new substance called sodium chloride.

Answer

a. chemical
b. physical
c. chemical

3.6: Conservation of Mass

17. Octane is a major component of gasoline. When octane is burned by your car engine, the octane reacts with oxygen to form carbon dioxide and water. If 100.0 g of octane reacts with 350.2 g of oxygen to make 142.0 g of water, how many grams of carbon dioxide are produced.

Answer

308.2 g

18. Baking soda ($NaHCO_3$) decomposes over time to form sodium carbonate, water, and carbon dioxide. How many grams of baking soda must decompose to form 250.0 g of sodium carbonate? In addition to the sodium carbonate, 42.5 g of water and 103.8 g of carbon dioxide will also form.

Answer

396.3 g

3.7: Energy

19. Classify each of the following as energy primarily transferred as heat, or energy primarily transferred as work:
    a. The energy transferred from your body to a shopping cart as you push the shopping cart down the aisle.
    b. The energy transferred from a wave to your board when you go surfing.
    c. The energy transferred from the flames to your hotdog when you cook your hotdog over a campfire.

Answer

a. work
b. work
c. heat

20. Decide whether each of the following statements is true or false:
    a. When heat is transferred to an object, the object cools down.
    b. Any time you raise the temperature of an object, you have done work.
    c. Any time you move an object by applying force, you have done work.
    d. Any time you apply force to an object, you have done work.

Answer

a. false, it heats up
b. false, you have added heat
c. true
d. false, it must be moved to do the work

21. Classify each of the following energies as kinetic energy or potential energy:
    a. The energy in a chocolate bar.
    b. The energy of rushing water used to turn a turbine or a water wheel.

c. The energy of a skater gliding on the ice.
d. The energy in a stretched rubber band.

Answer

a. potential
b. kinetic
c. kinetic
d. potential

3.8: Energy and Chemical and Physical Change

22. A pot of water is set on a hot burner of a stove. What is the direction of heat flow?

Answer

Heat flows into the pot of water.

23. Some uncooked macaroni is added to a pot of boiling water. What is the direction of heat flow?

Answer

Heat flows from the water to the macaroni.

24. Label each of the following processes as endothermic or exothermic.

a. an ice cube melts in a cup
b. a chemical ice pack is activated by breaking the inner pouch and shaking

Answer

a. endothermic (the ice must be heated to melt so it absorbs energy)
b. endothermic (the ice pack feels cold)

25. Label each of the following processes as endothermic or exothermic.

a. sweat evaporates from your skin
b. a bonfire

Answer

a. endothermic (the liquid must be heated to evaporate so it absorbs energy)
b. exothermic (the fire feels warm so it gives off heat)

3.9: Temperature

26. Perform the following conversions.

a. 255°F to degrees Celsius
b. −255°F to degrees Celsius
c. 50.0°C to degrees Fahrenheit
d. −50.0°C to degrees Fahrenheit

Answer

a. 124°C
b. −159°C
c. 122.0°F
d. −58.0°F

27. Perform the following conversions.

a. 1,065°C to degrees Fahrenheit
b. −222°C to degrees Fahrenheit
c. 400.0°F to degrees Celsius
d. 200.0°F to degrees Celsius

Answer

a. 1949°F
b. −368°F
c. 204.4°C
d. 93.33°C

28. Perform the following conversions.

a. 100.0°C to kelvins
b. −100.0°C to kelvins
c. 100. K to degrees Celsius
d. 300. K to degrees Celsius

Answer

a. 373.2 K
b. 173.2 K
c. -173°C
d. 27°C

29. Perform the following conversions.

a. 1,000.0 K to degrees Celsius
b. 50.0 K to degrees Celsius
c. 37.0°C to kelvins
d. −37.0°C to kelvins

Answer

a. 726.9°C
b. −223.2°C
c. 310.2 K
d. 236.2 K

30. Convert 0 K to degrees Celsius.

Answer

−273°C

31. Convert 0 K to degrees Fahrenheit.

Answer

−460°F

32. The hottest temperature ever recorded on the surface of the earth was 136°F in Libya in 1922. What is the temperature in degrees Celsius and in kelvins?

Answer

57.8°C, 330.9 K

33. The coldest temperature ever recorded on the surface of the earth was −128.6°F in Vostok, Antarctica, in 1983. What is the temperature in degrees Celsius and in kelvins?

Answer

-89.22°C, 183.93 K

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CHAPTER OVERVIEW

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4.1: Cutting Aluminum Until You Get Atoms

Take some aluminum foil. Cut it in half. Now there are two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil. It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can this exercise be taken, at least in theory? Can one continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? Thought experiments like this – and the conclusions based on them – were debated as far back as the fifth century BC by Democritus and other ancient Greek philosophers (Figure 4.1.1).

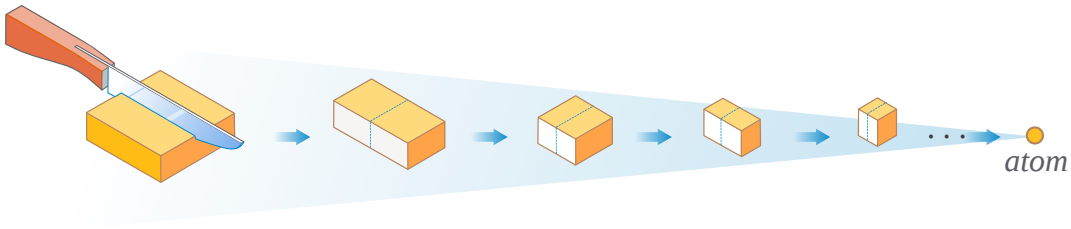


Figure **4.1.1**: Democritus argued that matter, like an aluminum block, cannot be repeatedly cut in half perpetually. At some point, a limit is reached before the substance can no longer be called "aluminum"; this is the atomic limit. (CC BY-NC; Ümit Kaya)

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called diatomic molecules. In representing a diatomic molecule, we use the symbol of the element and include the subscript 2 to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

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4.2: Early 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.

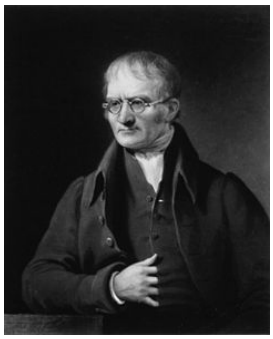



Figure **4.2.1**: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek 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.

John Dalton and Atomic Theory

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 1800s with the work of the British scientist 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.

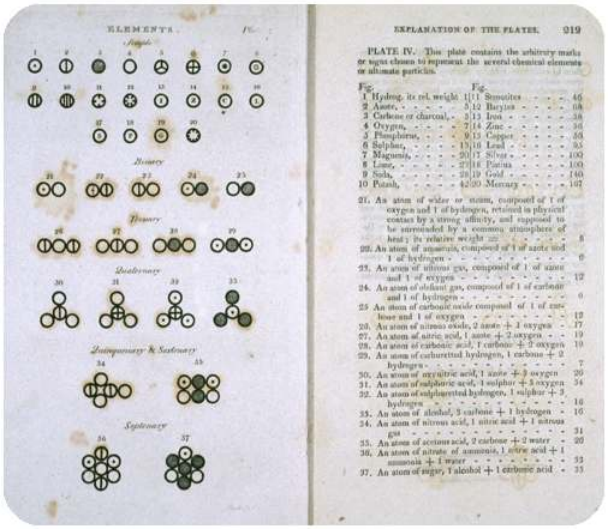
The modern atomic theory, first proposed about 1803 by the English chemist John Dalton (Figure 4.2.1), is a fundamental concept that states that all elements are composed of atoms. An atom is the smallest particle of an element that maintains the identity of that element. Individual atoms are extremely 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 supported the law of definite proportions proposed in 1797 by Joseph Proust. 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.


ELEMENTS.

EXPLANATION OF THE PLATES. 219

PLATE IV. This plate contains the arbitrary marks or signs chosen to represent the several chemical elements or ultimate particles.

Figure **4.2.2**: 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 4.2.1,

control of individual atoms can be used to create animations.



Video **4.2.1**: 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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4.3: Discovery of the Nucleus

Learning Objectives

- Explain the observations that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" model of the atom and the evidence for it.
- Draw a diagram of Thomson's "plum pudding" model of the atom and explain why it has this name.
- Describe Rutherford's gold foil experiment and explain how this experiment altered the "plum pudding" model.
- Draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory was not the entire story. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.



Figure **4.3.1**: J.J. Thomson (center) conducted experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the nuclear model of the atom.

Electrons and Plums

The electron was discovered by J.J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This was a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or similar to a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges, and the electron would be pulled back more toward the center region of the atom.

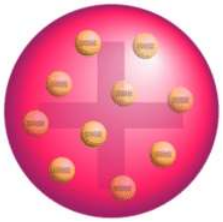


Figure 4.3.2: The "plum pudding" model.

However, this model of the atom soon gave way to a new model developed by New Zealander Ernest Rutherford (1871-1937) about five years later. Thomson did still receive many honors during his lifetime, including being awarded the Nobel Prize in Physics in 1906 and a knighthood in 1908.

Atoms and Gold

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. They bombarded very thin sheets of gold foil with fast moving **alpha particles**. Alpha particles, a type of natural radioactive particle, are positively charged particles with a mass about four times that of a hydrogen atom.

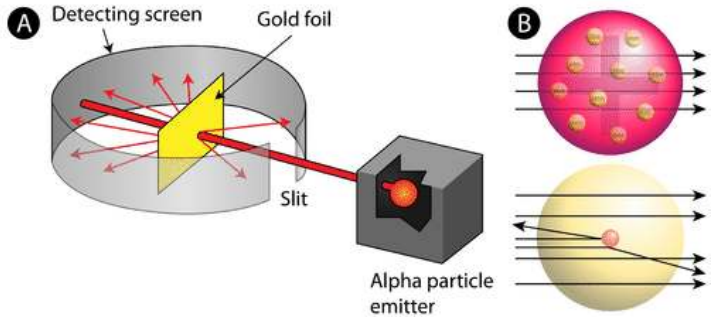


Figure 4.3.3: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top), all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed not deflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.



Video **4.3.1**: Animation of Rutherford's gold foil experiment. (https://youtu.be/5pZj0u_XMbc)

Rutherford's atomic model became known as the **nuclear model**. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupy the vast space around the nucleus. It was not until some years later that a full understanding of the electron was achieved. This proved to be the key to understanding the chemical properties of elements.

Atomic Nucleus

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

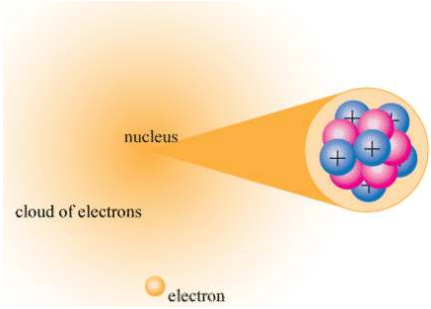


Figure **4.3.4**: The nuclear atom.

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

⛓ Holding it all Together

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

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

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4.4: Protons, Neutrons, and Electrons

Learning Objectives

- Describe the locations, charges, and masses of the three main subatomic particles.
- Determine the number of protons and electrons in an atom.
- Define atomic mass unit (amu).

Dalton's Atomic Theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it was not entirely accurate because, contrary to what Dalton believed, atoms can, in fact, be broken apart into smaller subunits or subatomic particles. We have been talking about the electron in great detail, but there are two other particles of interest to us: protons and neutrons. We already learned that J. J. Thomson discovered a negatively charged particle, called the **electron**. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus, called a **proton**. There is also a third subatomic particle, known as a neutron.

Electrons

Electrons are one of three main types of particles that make up atoms. Electrons are extremely small. The mass of an electron is only about $\frac{1}{2000}$ the mass of a proton or neutron, so electrons contribute virtually nothing to the total mass of an atom. Electrons have an electric charge of 1–, which is equal but opposite to the charge of a proton, which is 1+. All atoms have the same number of electrons as protons, so the positive and negative charges "cancel out", making atoms electrically neutral. While protons and neutrons are located inside the nucleus at the center of the atom, electrons are located outside the nucleus in what is often called the electron cloud.



Figure **4.4.1**: Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

Protons

A proton is another one of three main particles that make up the atom. Protons are found in the nucleus of the atom – the tiny, extremely dense region at the center of the atom. Protons have a positive electrical charge of one (1+) and a mass of about 1 atomic mass unit (1 amu). Together with neutrons, they make up virtually all of the mass of an atom.

Atomic Mass Units

An **atomic mass unit (amu)** is defined as one-twelfth of the mass of a carbon-12 atom, which is about 1.660×10^{-27} kg. Atomic mass units (amu) are useful, since the mass of a proton and the mass of a neutron are almost equal to 1 amu in this unit system.

Neutrons

The last of the three main particles that make up the atom is the neutron. Atoms of all elements (except for most atoms of hydrogen) have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge – they are electrically neutral. That's why the neutrons in the table below are labeled n^0. The zero stands for "zero charge".

A neutron is slightly more massive than a proton (see below) and has about the same diameter. Neutrons are found in all atoms (except for most atoms of hydrogen) and may be found bound together with other neutrons and protons in the atomic nucleus. Since protons and neutrons are so much more massive than electrons, almost all of the mass of an atom comes from the nucleus, which contains all of the neutrons and protons.

Table 4.4.1 provides a summary of the properties of the three main types of subatomic particles.

Table 4.4.1: Properties of Subatomic Particles

Particle	Symbol	Mass (kg)	Mass (amu)	Relative Charge	Location
proton	p^+	1.673×10^{-27}	$1.0073 \approx 1$	1+	inside the nucleus
electron	e^-	9.109×10^{-31}	$0.00055 \approx 0$	1−	outside the nucleus
neutron	n^0	1.675×10^{-27}	$1.0087 \approx 1$	0	inside the nucleus

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they are neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

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4.5: Chemical Symbols and the Atomic Number

Learning Objectives

- Define atomic number.
- Determine the number of protons and electrons in an atom.

It is important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up matter, so each one is given a unique name. The names of elements are also represented by unique one- or two-letter symbols, such as H for hydrogen, C for carbon, or He for helium. However, it would be more powerful if these names could be used to identify the numbers of protons and neutrons in the atoms. That's where atomic number and mass number are useful.



Figure **4.5.1**: Each element has a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; Alchemist-hp).

Chemical Symbol

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 in Table 4.5.1.

Table 4.5.1: 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

Atomic Number

Scientists distinguish between different elements by the **atomic number** (Z), which represents the number of protons in the nucleus of one atom of that element. All atoms of an element have the same number of protons, and every element has a different number

of protons in its atoms. If an atom has 26 protons (Figure 4.5.2), we know it is an iron atom. An atom with two protons is always a helium atom, since its atomic number is 2. No other element will have two protons in a single atom.

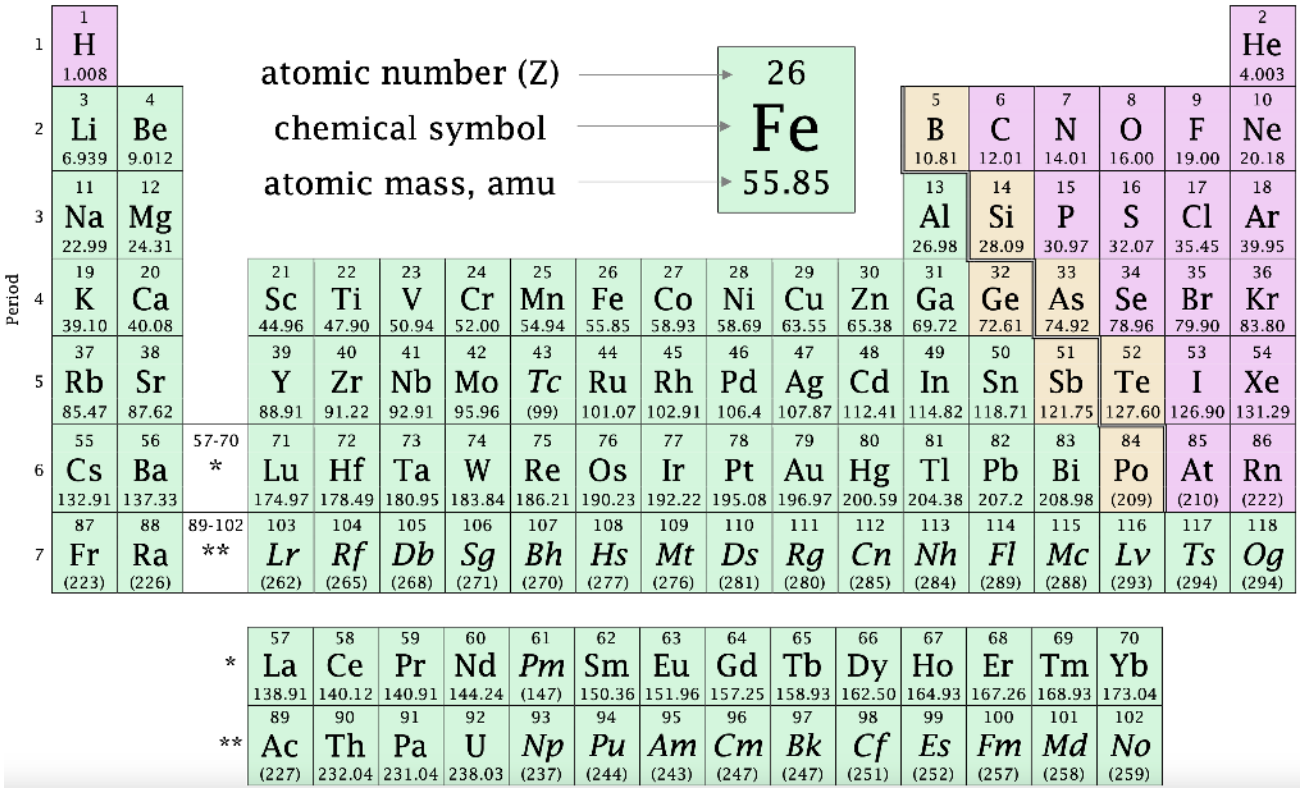


Figure **4.5.2**: Periodic Table showing the atomic numbers and atomic masses.

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

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negatively charged electrons as they have positively charged protons. Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Exercise 4.5.1

How many protons and electrons in one atom of each element, assuming the atom is electrically neutral?

A. Co
B. Na
C. Ca
D. Sr

Answer A

27 protons, 27 electrons

Answer B

11 protons, 11 electrons

Answer C

20 protons, 20 electrons

Answer D

38 protons, 38 electrons

Summary

- Elements are pure substances that make up all matter, so each one is given a unique name.
- The names of elements are also represented by unique one- or two-letter symbols.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.

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4.6: The Periodic Table

Learning Objectives

- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

In the 19th century, many previously unknown elements were discovered, and scientists noted that certain sets of elements had similar chemical properties. For example, chlorine, bromine, and iodine react with other elements (such as sodium) to make similar compounds. Likewise, lithium, sodium, and potassium react with other elements (such as oxygen) to make similar compounds. Why is this so?

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements (see Figure 4.6.1 below), and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.



Figure **4.6.1**: Periodic table by Dmitri Mendeleev, 1871. (NikNaks, Public domain, via Wikimedia Commons)

Mendeleev's Periodic Table

Mendeleev had to list some elements out of the order of their atomic masses to group them with other elements that had similar properties.

The periodic table is one of the cornerstones of chemistry because it organizes all of the known elements on the basis of their chemical properties. A modern version is shown in Figure 4.6.2. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number. The periodic table below also shows how elements may be classified as either a metal, a nonmetal, or a metalloid.

A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be hammered into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. Metals occupy the left three-quarters of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the double-line staircase in the right-hand portion of the periodic table comprise the metalloids.

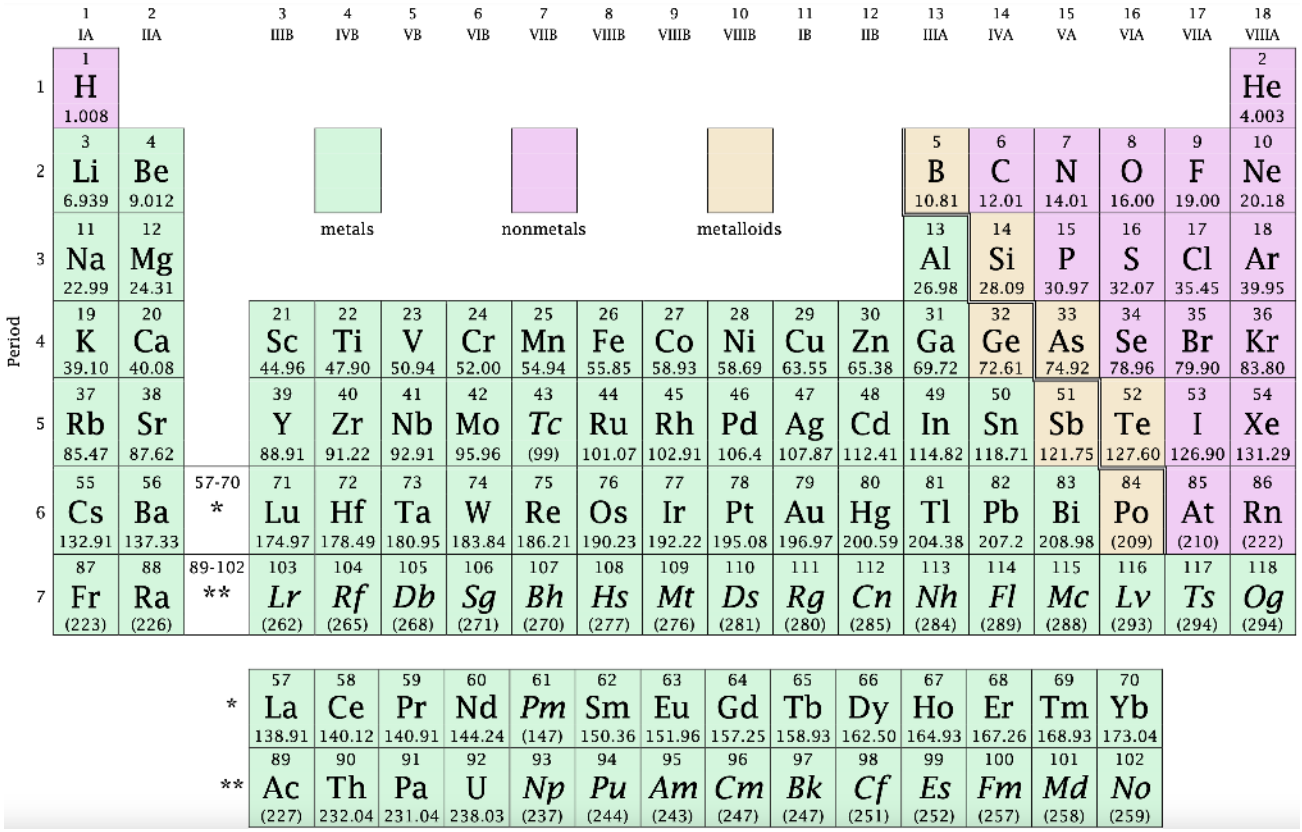


Figure **4.6.2**: Elements may be classified metals, nonmetals, or metalloids. Each group is located in a different part of the periodic table. An interactive periodic table may be found here.

✅ Example 4.6.1

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

A. Mg
B. Br
C. Si

Solution

A. Magnesium lies on the left-hand side of the periodic table, so it should be a metal.
B. Bromine lies on the right-hand side of the periodic table, so it should be a nonmetal.
C. Silicon lies adjacent to the staircase marking the boundary between metals and nonmetals, so it should be a metalloid.

✏️ Exercise 4.6.1

Based on its location in the periodic table, would we expect gold to be a nonmetal, a metal, or a metalloid?

Answer

Gold (Au) is a metal.

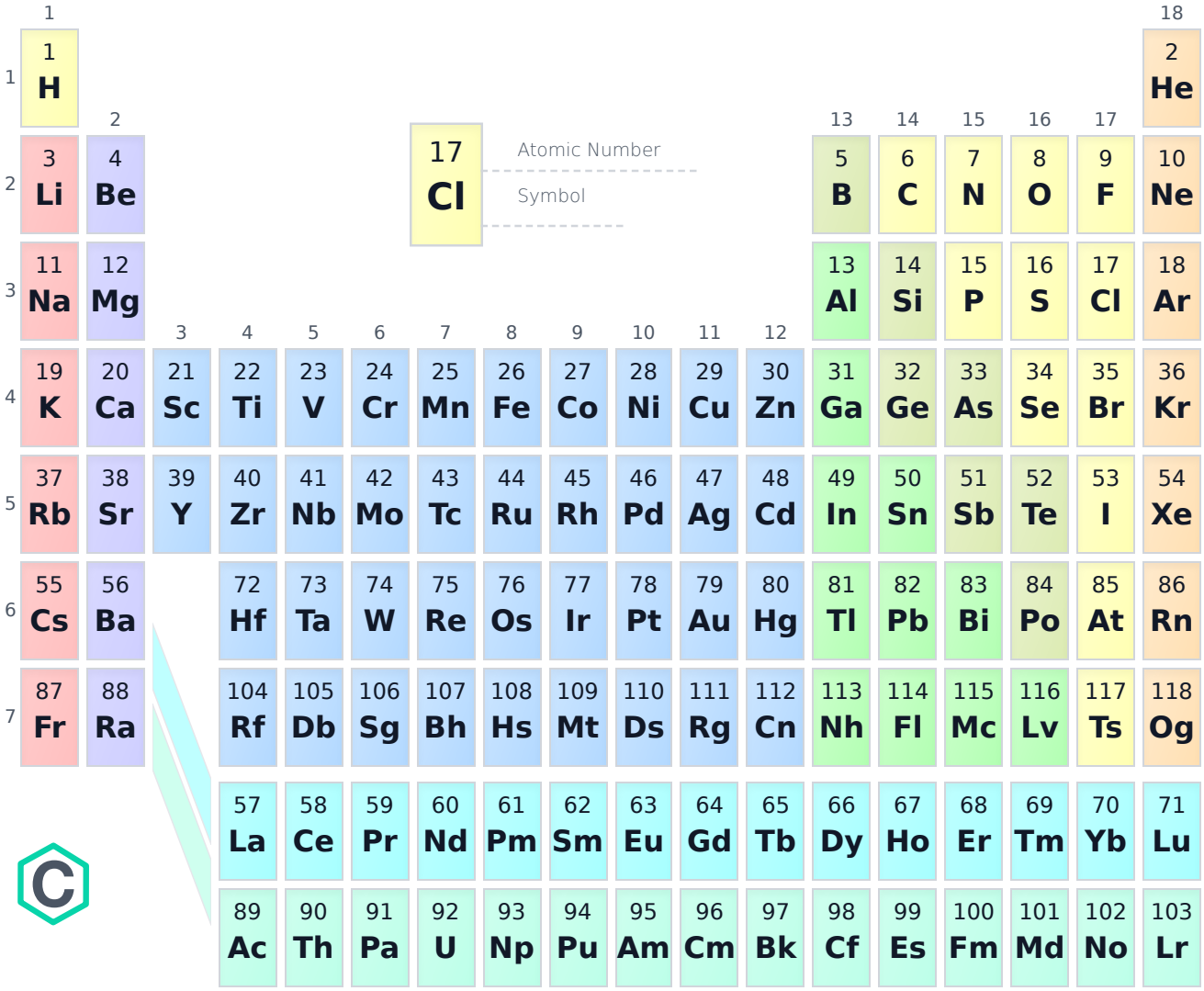
Features of the Periodic Table

Vertical columns in the periodic table are called **groups** while horizontal rows are called **periods**. Elements in the same group have similar chemical properties. Groups in the periodic table are numbered. One convention uses the numbers 1 through 18. A second convention uses Roman numerals followed by either the letter A or B (see Figure 4.6.2 above).

Groups designated with the letter A, those in the first two columns on the left and the last six columns on the right, are called the **main group elements**. The ten-column block in the center designated by the letter B contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. Elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals** (see Figure 4.6.3 below).

As well as being numbered, some of these groups have names – for example, Group IA includes hydrogen and the **alkali metals**, Group IIA holds the **alkaline earth metals**, Group VIIA elements are called the **halogens**, and Group VIIIA elements are the **noble gases** (see Figure 4.6.3 below).

Periodic Table of Elements



https://chem.libretexts.org/@go/page/289366

Figure **4.6.3**: An interactive modern periodic table. The displayed property may be changed by using the dropdown menu to the left of helium, He. (Public Domain; PubChem via NIH)

Group 1 (IA): The Alkali Metals

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

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



Video **4.6.1**: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via https://youtu.be/6ZY6d6jrq-0)

Group 2 (IIA): The Alkaline Earth Metals

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

Group 17 (VIIA): The Halogens

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

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

Group 18 (VIIIA): The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because the noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for

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

✅ Example 4.6.2: Groups

Provide the family or group name of each element.

A. Li
B. Ar
C. Cl

Solution

A. Lithium is an alkali metal (Group 1).
B. Argon is a noble gas (Group 18).
C. Chlorine is a halogen (Group 17).

✏️ Exercise 4.6.2: Groups

Provide the family or group name of each element.

A. F
B. Ca
C. Kr

Answer A

Fluorine is a halogen (Group 17).

Answer B

Calcium is a alkaline earth metal (Group 2).

Answer C

Krypton is a noble gas (Group 18).

Example 4.6.3: Classification of Elements

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

A. Li
B. Ar
C. Am
D. Fe

Solution

A. Lithium is a metal.
B. Argon is a non metal.
C. Americium is an inner transition metal.
D. Iron is a transition metal.

Exercise 4.6.3: Classification of Elements

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

A. F
B. U
C. Cu

Answer A

Fluorine is a nonmetal.

Answer B

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

Answer C

Copper is a metal (and a transition metal too).

To Your Health: Radon

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

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

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

Summary

- The periodic table is an arrangement of the elements in order of increasing atomic number.
- Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods.
- Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.
- The elements can be broadly divided into metals, nonmetals, and metalloids.
- Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of metalloids.
- Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors.
- Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.

2. Sisler, Harry H. Electronic structure, properties, and the periodic law. New york; Reinhold publishing corporation, 1963.
3. Petrucci, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

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4.7: Isotopes and Mass Numbers

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.

As discussed in the the previous section, atoms of the same element always have the same number of protons. However, atoms of the same element may have differing numbers of neutrons. While all carbon atoms have six protons and most have six neutrons, some carbon atoms have seven neutrons, while others have eight. Atoms of the same element that differ in their numbers of neutrons are called **isotopes**. Many isotopes occur naturally. Quite often, there are one or two isotopes of an element that are the most stable and common. Different isotopes of an element generally have the same physical and chemical properties due to the fact they have the same numbers of protons and electrons.

Mass Number

The **mass number** (A) of an atom represents the sum of protons and neutrons in the nucleus of that atom and provides a very close approximation to the mass of the atom in atomic mass units. This is because the mass of the proton and neutron are each about 1 amu, while the mass of the electron is very small in comparison.

$$\text{mass number } (A) = \text{number of protons} + \text{number of neutrons} \tag{4.7.1}$$

Consider oxygen, which has an atomic number (Z) of 8. This means that oxygen atoms have 8 protons. Most atoms of oxygen also have 8 neutrons. Oxygen atoms with 8 neutrons would have a mass number of 16 (8 $\mathbf{p^+}$ + 8 $\mathbf{n^0}$), meaning they would have a mass of about 16 amu.

Some oxygen atoms have 9 neutrons, while others have 10 neutrons. Oxygen atoms with 9 neutrons would have a mass number of 17 (8 $\mathbf{p^+}$ + 9 $\mathbf{n^0}$), meaning they would have a mass of about 17 amu. Oxygen atoms with 10 neutrons would have a mass number of 18 (8 $\mathbf{p^+}$ + 10 $\mathbf{n^0}$), meaning they would have a mass of about 18 amu.

Isotope Notation

The three types of oxygen atoms described above are the three naturally-occurring isotopes of oxygen. While there are several other known isotopes of oxygen, they are not found anywhere in nature and may only be synthesized in the laboratory, which is not necessarily an "easy" thing to do.

Isotopes of elements are designated by their mass number. Oxygen atoms with a mass number of 16 are called oxygen-16, or O-16, while oxygen atoms with a mass number of 17 are called oxygen-17, or O-17, and oxygen atoms with a mass number of 18 are called oxygen-18, or O-18.

Another type of **isotope notation** or **isotope symbol** uses the mass number (A) written as a superscript to the left of the chemical symbol with the atomic number (Z) written below the mass number as a subscript to the left of the chemical symbol.

$$^{A}_{Z}\mathrm{X}$$

Sometimes, the atomic number is left off the isotope notation, since the atomic number and the number of protons for a given element never changes.

$$^{A}\mathrm{X}$$

Assuming the atoms are neutral, the number of protons, electrons, and neutrons in one atom of each of the three naturally-occurring isotopes of oxygen are shown in Table 4.7.1 below.

Table 4.7.1: The Three Naturally-Occurring Isotopes of Oxygen

Isotope	Abbreviation	Isotope Symbol	Isotope Symbol	p^+	e^-	n^0
oxygen-16	O-16	^{16}O	$^{16}_{8}O$	8	8	8
oxygen-17	O-17	^{17}O	$^{17}_{8}O$	8	8	9
oxygen-18	O-18	^{18}O	$^{18}_{8}O$	8	8	10

It is important to note that the mass number is *not* shown on the Periodic Table. You may find all of the known isotopes of a given element by clicking on the chemical symbol of the desired element on the Periodic Table located at https://www.webelements.com/isotopes.html.

✅ Example 4.7.1

A. What is the mass number of an atom of chromium (Cr) that has 30 neutrons?
B. Write the different isotope notations or isotope symbols for this isotope.

Solution

A. Cr has an atomic number (Z) of 24, so it has 24 p^+. With 24 p^+ and 30 n^0, the mass number of this atom is 54.

B. The isotope notations include:

- chromium-54
- Cr-54
- ^{54}Cr
- $^{54}_{24}Cr$

Example 4.7.2

How many protons, electrons, and neutrons are in a neutral atom of $^{204}_{80}Hg$?

Solution

- The atomic number is 80, so there are 80 p^+.
- Since the atom is neutral, there are also 80 e^-.
- Starting with $\textbf{mass number } (A) = \textbf{number of protons} + \textbf{number of neutrons}$, rearrange the equation to solve for the number of neutrons.
  $\textbf{number of neutrons} = \textbf{mass number } (A) - \textbf{number of protons}$ = 204 – 80 = 124 n^0

✏️ Exercise 4.7.1

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

A. $^{60}_{27}Co$
B. Na-24
C. $^{45}_{20}Ca$
D. Sr-90

Answer A

27 p^+, 27 e^-, 33 n^0

Answer B

11 p^+, 11 e^-, 13 n^0

Answer C

20 p^+, 20 e^-, 25 n^0

Answer D

38 p^+, 38 e^-, 52 n^0

Exercise 4.7.2

Write the different isotope notations or isotope symbols for an atom with 42 p^+, 42 e^-, and 52 n^0.

Answer

The isotope notations include:

- molybdenum-94
- Mo-94
- ^{94}Mo
- $^{94}_{42}\text{Mo}$

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.
- Isotopes of a given element contain different numbers of neutrons, therefore, different isotopes have different mass numbers.

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4.8: Atomic Mass

Learning Objectives

- Explain what is meant by the atomic mass of an element.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Since most elements have more than one naturally-occurring isotope that are already naturally premixed, we rarely need to worry about the mass of a specific isotope. Rather, we need to know the average mass of the atoms of an element. We call this mass the **atomic mass**.

Atomic Mass

The atomic mass of an element is defined as the average mass of all the naturally-occurring isotopes for that element. Atomic mass is typically reported in atomic mass units.

Calculating the Atomic Mass

Let's consider the element chlorine. Chlorine has two naturally occurring isotopes, $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$. The mass and natural abundance of these two isotopes was obtained from https://www.webelements.com/chlorine/isotopes.html:

Table 4.8.1: Naturally Occurring Isotopes of Chlorine

Isotope	Mass (amu)	Natural Abundance
$^{35}_{17}\text{Cl}$	34.969	75.78%
$^{37}_{17}\text{Cl}$	36.966	24.22%

One of the first things we may notice is that the mass of one atom of $^{35}_{17}\text{Cl}$ is close to 35 amu and that the mass of $^{37}_{17}\text{Cl}$ is close to 37 amu. Remember that this is because the mass of each proton is about 1 amu and the mass of each neutron is about 1 amu, while each electron has a mass that is so small relative to the proton and neutron that their mass may generally be ignored.

In order to find the atomic mass for chlorine, it may be tempting to simply sum the masses of the two isotopes together and divide the result by 2. However, this does not take into consideration that $^{35}_{17}\text{Cl}$ has almost three times the natural abundance of $^{37}_{17}\text{Cl}$.

When the abundances are not equal, which will always be the case, an average is best accomplished by calculating what is called a **weighted average**, in which the fractional abundance of each naturally occurring isotope is multiplied by its mass and the resulting masses are summed together. We summarize this as follows:

$$\text{Atomic mass} = (\text{fractional abundance}_1)\,(\text{mass}_1) + (\text{fractional abundance}_2)\,(\text{mass}_2) + \cdots \tag{4.8.1}$$

Keep in mind that the fractional abundance represents $\frac{1}{100}$ of the percent abundance. Just like the percent natural abundances of isotopes should sum to 100%, the sum of fractional abundances of isotopes should sum to 1.

Table 4.8.2: Calculating the Atomic Mass of Chlorine

Isotope	Mass (amu)	Natural Abundance	Calculation	Mass (amu)
$^{35}_{17}\mathrm{Cl}$	34.969	75.78%	$\mathbf{34.969\ amu \times 0.7578 =}$	26.5000
$^{37}_{17}\mathrm{Cl}$	36.966	24.22%	$\mathbf{36.966\ amu \times 0.2422 =}$	8.95532
			atomic mass of chlorine =	35.45532

As we can see, the calculated atomic mass for chlorine is 35.45 amu. This matches the atomic mass listed on the Periodic Table for chlorine.

✅ Example 4.8.1: Boron Isotopes 1

Simply knowing that boron has two naturally occurring isotopes, $^{10}_{5}\mathbf{B}$ and $^{11}_{5}\mathbf{B}$, which of the two isotopes is more abundant?

Solution

According to the Periodic Table, the atomic mass of boron is 10.81 amu. Since the mass of one atom of $^{10}_{5}\mathbf{B}$ would be approximately 10 amu and the mass of one atom of $^{11}_{5}\mathbf{B}$ would be about 11 amu, we can conclude that the most abundant isotope is $^{11}_{5}\mathbf{B}$, since the atomic mass is closer to the mass of $^{11}_{5}\mathbf{B}$ than to $^{10}_{5}\mathbf{B}$.

This is supported using natural abundance data available from https://www.webelements.com/boron/isotopes.html, which shows that 19.9% of boron atoms are $^{10}_{5}\mathbf{B}$, while 80.1% are $^{11}_{5}\mathbf{B}$.

✅ Example 4.8.2: Boron Isotopes 2

Given that the atomic mass of boron is 10.81 amu and that there are two naturally occurring isotopes of boron, $^{10}_{5}\mathbf{B}$ and $^{11}_{5}\mathbf{B}$, what is the percent natural abundance of each isotope?

Solution

As noted above, just like the percent natural abundances of isotopes should sum to 100%, the sum of fractional abundances of isotopes should sum to 1. Therefore, if we assume the fractional abundance of $^{10}_{5}\mathrm{B}$ is x, then the fractional abundance of $^{11}_{5}\mathrm{B}$ must be $1-x$. Summarizing the information into the table below shows we have enough information to calculate the value of x, hence the fractional abundance of each.

Isotope	Approximate Mass (amu)	Fractional Abundance	Mass (amu)
$^{10}_{5}\mathrm{B}$	10.0	x	$10.0 \times (x)$
$^{11}_{5}\mathrm{B}$	11.0	$1-x$	$11.0 \times (1-x)$
		atomic mass of boron =	10.81

At this point, we can solve for x:

$$10.0\,x + 11.0\,(1-x) = 10.81$$

$$10.0\,x + 11.0 - 11.0\,x = 10.81$$

$$-x = -0.2 \text{ or } x = 0.2$$

In other words, the fractional abundance of $^{10}_{5}\mathrm{B}$ is x = 0.2 and the fractional abundance of $^{11}_{5}\mathrm{B}$ is $1-x = 1-0.2$ = 0.8. Since a percentage is 100 times its fractional amount,

The natural abundance of $^{10}_{5}\mathrm{B}$ is 100 × 0.2 = 20%

The natural abundance of $^{11}_{5}\mathrm{B}$ is 100 × 0.8 = 80%

This compares very well to the natural abundance data available from https://www.webelements.com/boron/isotopes.html, which shows that 19.9% of boron atoms are $^{10}_{5}\mathrm{B}$, while 80.1% are $^{11}_{5}\mathrm{B}$.

Exercise 4.8.1

Neon has three naturally occurring isotopes. Using the data below, calculate the atomic mass of neon.

Table 4.8.3: Naturally Occurring Isotopes of Neon

Isotope	Mass (amu)	Natural Abundance	Calculation	Mass (amu)
$^{20}_{10}\mathrm{Ne}$	19.992	90.48%		
$^{21}_{10}\mathrm{Ne}$	20.994	0.27%		
$^{22}_{10}\mathrm{Ne}$	21.991	9.25%		

Answer

20.18 amu

Summary

- An element's atomic mass is the weighted average of the masses of the isotopes of an element

- An element's atomic mass can be calculated provided the relative abundance of the element's naturally occurring isotopes and the masses of those isotopes are known.
- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number and the element's atomic mass.

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4.9: Ion Formation

Learning Objectives

- Define the two types of ions.

Up until now, we have only considered atoms that are neutral, where the number of electrons is equal to the number of protons. As we recall, an atom consists of a positively charged nucleus that contains the protons and neutrons that is surrounded by a cloud of negatively charged electrons (Figure 4.9.1). This makes the electrons more accessible, resulting in their ability to more readily be gained or lost. This is also why it is the number of protons in the nucleus that determines the identity of the element – the number of protons does not change.



Figure 4.9.1: Helium atom showing the nucleus surrounded by an electron cloud. Cepheiden, CC BY-SA 3.0, via Wikimedia Commons

As we will learn in Chapter 11, atoms tend to form chemical bonds by having complete outer shells of electrons. In some cases, it is easier for an atom to gain electrons to make an outer shell complete. In other cases, it is easier for an atom to lose outer shell electrons to arrive at a complete shell that lies below.

An atom that has gained or lost one or more electrons is called an **ion**. An ion always has a charge that is positive (+) or negative (−). An ion with a positive charge is called a **cation**, while an ion with a negative charge is called an **anion**. An easy mnemonic to distinguish the two is that the the word *cation* contains a plus sign (in other words, think of it as ca+ion).

Cations

Metals generally have three or fewer outer shell electrons making them more likely to lose electrons. Atoms that lose electrons acquire a positive charge since there are fewer of the negatively charged electrons that remain in the electron cloud compared to the positively charged protons that are present in the nucleus. Consider a sodium atom, which ordinarily loses one electron (Figure 4.9.2). The cation produced in this manner, Na^+, is called a sodium ion to distinguish it from a neutral atom of sodium.

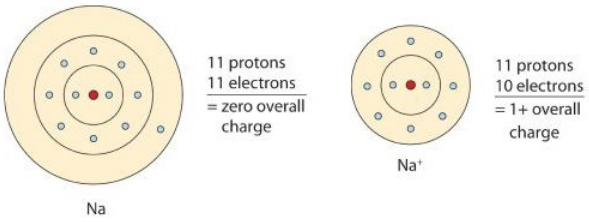


Figure 4.9.2: The formation of a sodium ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion has only 10 electrons and a charge of 1+.

Also notice that the ion charge is written as 1+ rather than +1. The convention for showing charges on particles is to always write the sign *after* the number. Using this convention gives us the "heads up" that we are looking at the charge of a particle rather than an ordinary positive or negative number. If we want to summarize the process of sodium losing an electron to form the ion, we may write it like this:

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + \mathrm{e^-} \tag{4.9.1}$$

In table salt, NaCl, this electron is "donated" to an atom of chlorine.

Anions

Nonmetals generally have four or more outer shell electrons making them more likely to gain electrons. Atoms that gain electrons acquire a negative charge since there would be more of the negatively charged electrons present in the electron cloud compared to the positively charged protons in the nucleus. Consider a chlorine atom, which ordinarily gains one electron (Figure 4.9.3). The resulting anion, Cl^-, is called the chloride ion. Note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion.

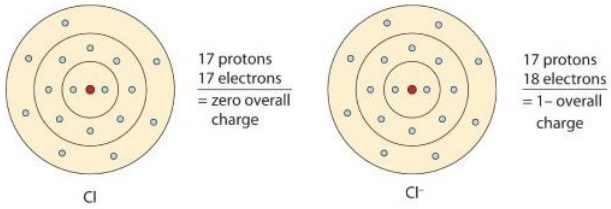


Figure 4.9.3: The formation of a chloride ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1− charge.

If we want to summarize the process of chlorine gaining an electron to form the ion, we may write it like this:

$$Cl + e^- \rightarrow Cl^- \tag{4.9.2}$$

In table salt, this electron comes from the sodium atom.

Location, Location, Location!

In many cases, elements that belong to the same group on the periodic table form ions with the same charge because they have the same number of outer shell 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 4.9.4 shows how the charge on many ions may be predicted its location.

Once again, note the convention of writing the sign after the number. The barium cation is written Ba^{2+}, *not* Ba^{+2}.

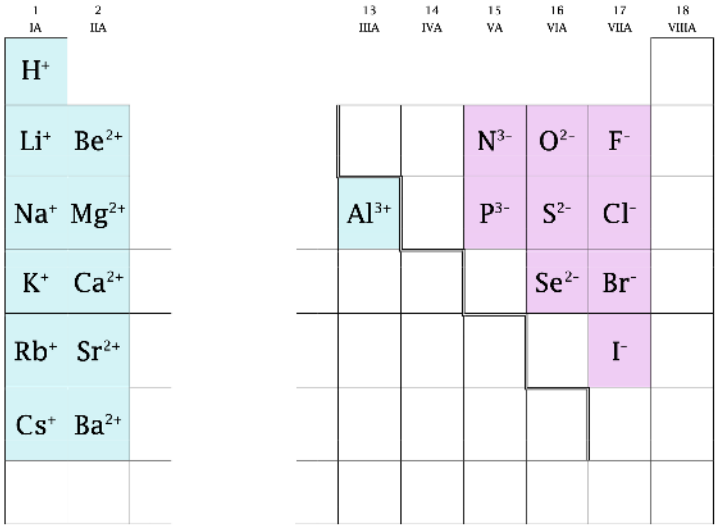


Figure 4.9.4: Predicting ionic charges. The charge an atom acquires when it becomes an ion is often related to its position on the periodic table.

✅ Example 4.9.1

A. How many protons and electrons are in a potassium ion, K^+?

B. What is the chemical symbol and charge of an ion that has 25 p^+ and 23 e^-?

Solution

A. Potassium has an atomic number is 19, so there are 19 p^+.

Since there is a single positive (+) charge on the ion, there must be one more positively charged proton in the nucleus than there are negatively charged electrons in the electron cloud.

Therefore, there are 18 e^-.

B. With 25 p^+, the atomic number (*Z*) is 25. This identifies the element as manganese, Mn.

Since there are two more protons, p^+, than electrons, e^-, there is an excess of two positive (+) charges on the ion, meaning its charge is 2+.

Therefore, we write the symbol and charge as Mn^{2+}.

✏️ Exercise 4.9.1

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

A. How many protons and electrons are in a phosphide ion, P^{3-}?

B. What is the formula of an ion (its chemical symbol and charge) that has 53 p^+ and 54 e^-?

Answer A

15 p^+, 18 e^-

Answer B

I^-

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4.11: Exercises

4.1: Cutting Aluminum Until You Get Atoms

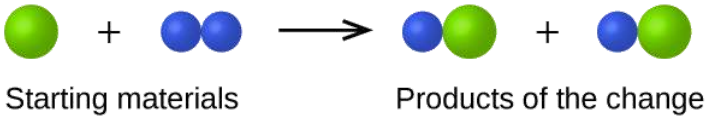
1. Give the proper formula for each diatomic element.

Answer

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

4.2: Early Atomic Theory

2. In the following drawing, the green spheres represent atoms of a certain element. The purple spheres represent atoms of another element. If the spheres of different elements touch, they are part of a single unit of a compound. The following chemical change represented by these spheres may violate one of the ideas of Dalton's atomic theory. Which one?



Answer

The starting materials consist of one green sphere and two purple spheres. The products consist of two green spheres and two purple spheres. This violates Dalton's postulate that that atoms are not created during a chemical change, but are merely redistributed.

4.3: Discovery of the Nucleus

3. Explain how atoms are composed.

Answer

Atoms have a small nucleus containing protons and neutrons. The nucleus is surrounded by "empty" space in which the electrons travel.

4. Where is most of the mass of an atom located?

Answer

in the nucleus

4.4: Protons, Neutrons, and Electrons

5. Define *atomic mass unit*. What is its abbreviation?

Answer

An atomic mass unit is $\frac{1}{12}$ of the mass of a carbon-12 atom. Its abbreviation is amu.

6. Which is larger, a proton or an electron?

Answer

proton

7. Which is larger, a neutron or an electron?

Answer

neutron

8. What are the charges for each of the three subatomic particles?

Answer

proton = 1+
neutron = 0,
electron = 1–

9. How are electrons and protons similar? How are they different?

Answer

They are similar in that they are both charged particles. They are different in that the charges are different and they have different masses.

10. How are protons and neutrons similar? How are they different?

Answer

They are similar in that they both have a mass of approximately 1 amu. They are different in that protons have a 1+ charge and neutrons are neutral.

4.5: Chemical Symbols and the Atomic Number

11. Define *atomic number*.

Answer

The number of protons in the nucleus of one atom of an element.

12. What is the atomic number for a boron atom?

Answer

5

13. What is the atomic number of helium?

Answer

2

14. Give the chemical symbol for each element.

a. sodium
b. argon
c. nitrogen
d. radon

Answer

a. Na
b. Ar
c. N
d. Rn

15. Give the chemical symbol for each element.

a. silver
b. gold
c. mercury
d. iodine

Answer

a. Ag
b. Au
c. Hg
d. I

16. Give the name of the element.

a. Si
b. Mn
c. Fe
d. Cr

Answer

a. silicon
b. manganese
c. iron
d. chromium

17. Give the name of the element.

a. F
b. Cl
c. Br
d. I

Answer

a. fluorine
b. chlorine
c. bromine
d. iodine

4.6: The Periodic Table

18. Distinguish between a metal and a nonmetal.

Answer

Metals are typically shiny, conduct electricity and heat well, and are malleable and ductile; nonmetals are a variety of colors and phases, are brittle in the solid phase, and do not conduct heat or electricity well.

19. What properties do metalloids have?

Answer

Metalloids have properties that are intermediate between metals and nonmetals.

20. Why is iron considered a metal?

Answer

Iron is a metal because it is solid, is shiny, and conducts electricity and heat well.

21. Why is oxygen considered a nonmetal?

Answer

Oxygen is a gas. It is not a good conductor of heat or electricity.

22. Elemental carbon is a black, dull-looking solid that conducts heat and electricity well. It is very brittle and cannot be made into thin sheets or long wires. Of these properties, how does carbon behave as a metal? How does carbon behave as a nonmetal?

Answer

Carbon behaves as a metal because it conducts heat and electricity well. It is a nonmetal because it is black and brittle and cannot be made into sheets or wires.

23. Pure silicon is shiny and silvery but does not conduct electricity or heat well. Of these properties, how does silicon behave as a metal? How does silicon behave as a nonmetal?

Answer

Silicon behaves as a metal because it is silver and shiny. It behaves as a nonmetal because it does not conduct heat or electricity well.

24. Use its place on the periodic table to determine if indium, In, atomic number 49, is a metal or a nonmetal.

Answer

metal

25. Only a few atoms of astatine, At, atomic number 85, have been detected. On the basis of its position on the periodic table, would you expect it to be a metal or a nonmetal?

Answer

It is a metalloid. However, between the two options, it would behave more like a nonmetal than a metal.

26. Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

a. uranium
b. bromine
c. strontium
d. neon
e. gold
f. americium
g. rhodium
h. sulfur
i. carbon
j. potassium

Answer

a. metal, inner transition metal
b. nonmetal, main-group element
c. metal, main-group element
d. nonmetal, main-group element
e. metal, transition metal
f. metal, inner transition metal
g. metal, transition metal
h. nonmetal, main-group element
i. nonmetal, main-group element
j. metal, main-group element

27. Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

a. cobalt
b. europium
c. iodine
d. indium
e. lithium
f. oxygen
g. cadmium
h. terbium
i. rhenium

Answer

a. metal, transition metal
b. metal, inner transition metal
c. nonmetal, main-group element
d. metal, main-group element

e. metal, main-group element
f. nonmetal, main-group element
g. metal, transition metal
h. metal, inner transition metal
i. metal, transition metal

28. Using the periodic table, identify the lightest member of each of the following groups:

a. noble gases
b. alkaline earth metals
c. alkali metals

Answer

a. He
b. Be
c. Li (as H is not technically an alkali metal)

29. Using the periodic table, identify the heaviest member of each of the following groups:

a. alkali metals
b. noble gases
c. alkaline earth metals

Answer

a. Fr
b. Og
c. Ra

30. Use the periodic table to give the name and symbol for each of the following elements:

a. the noble gas in the same period as germanium
b. the alkaline earth metal in the same period as selenium
c. the halogen in the same period as lithium

Answer

a. Kr, krypton
b. Ca, calcium
c. F, fluorine

31. Use the periodic table to give the name and symbol for each of the following elements:

a. the halogen in the same period as the alkali metal with 11 protons
b. the alkaline earth metal in the same period with the neutral noble gas with 18 electrons
c. the noble gas in the same row as an isotope with 25 protons
d. the noble gas in the same period as gold

Answer

a. Cl, chlorine
b. Mg, magnesium
c. Kr, krypton

d. Rn, radon

32. Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

a. the alkali metal with 11 protons and a mass number of 23
b. the noble gas element with 75 neutrons in its nucleus and 54 electrons in the neutral atom
c. the isotope with 33 protons and 40 neutrons in its nucleus
d. the alkaline earth metal with 88 electrons and 138 neutrons
e. the noble gas, used in lighting, with 10 electrons and 10 neutrons
f. the lightest alkali metal with three neutrons

Answer

a. $^{23}_{11}\text{Na}$
b. $^{129}_{54}\text{Xe}$
c. $^{73}_{33}\text{As}$
d. $^{226}_{88}\text{Ra}$
e. $^{20}_{10}\text{Ne}$
f. $^{6}_{3}\text{Li}$

4.7: Isotopes and Mass Numbers

33. Which pair represents isotopes?

a. $^{4}_{2}\text{He}$ and $^{3}_{2}\text{He}$
b. $^{56}_{26}\text{Fe}$ and $^{56}_{25}\text{Mn}$
c. $^{28}_{14}\text{Si}$ and $^{31}_{15}\text{P}$

Answer

pair a

34. Which pair represents isotopes?

a. $^{40}_{20}\text{Ca}$ and $^{40}_{19}\text{K}$
b. $^{56}_{26}\text{Fe}$ and $^{56}_{28}\text{Fe}$
c. $^{238}_{92}\text{U}$ and $^{235}_{92}\text{U}$

Answer

pair c

35. Give complete symbols of each atom, including the atomic number and the mass number.

a. an oxygen atom with 8 protons and 8 neutrons
b. a potassium atom with 19 protons and 20 neutrons
c. a lithium atom with 3 protons and 4 neutrons

Answer

a. $^{16}_{8}\text{O}$
b. $^{39}_{19}\text{K}$
c. $^{7}_{3}\text{Li}$

36. Give complete symbols of each atom, including the atomic number and the mass number.

a. a magnesium atom with 12 protons and 12 neutrons
b. a magnesium atom with 12 protons and 13 neutrons
c. a xenon atom with 54 protons and 77 neutron

Answer

a. $^{24}_{12}\text{Mg}$
b. $^{25}_{12}\text{Mg}$
c. $^{131}_{54}\text{Xe}$

37. Americium-241 is an isotope used in smoke detectors. What is the complete symbol for this isotope?

Answer

$^{241}_{95}\text{Am}$

38. Carbon-14 is an isotope used to perform radioactive dating tests on previously living material. What is the complete symbol for this isotope?

Answer

$^{14}_{6}\text{C}$

39. A rare isotope of helium has a single neutron in its nucleus. Write the complete atomic symbol of this isotope.

Answer

$^{3}_{2}\text{He}$

40. Americium-241 is a crucial part of many smoke detectors. How many neutrons are present in its nucleus?

Answer

146 neutrons

41. Potassium-40 is a radioactive isotope of potassium that is present in the human body. How many neutrons are present in its nucleus?

Answer

21 neutrons

42. Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

a. $^{10}_{5}\text{B}$
b. $^{199}_{80}\text{Hg}$
c. $^{63}_{29}\text{Cu}$
d. $^{13}_{6}\text{C}$

e. $^{77}_{34}\text{Se}$

Answer

a. 5 protons, 5 electrons, 5 neutrons
b. 80 protons, 80 electrons, 119 neutrons
c. 29 protons, 29 electrons, 34 neutrons
d. 6 protons, 6 electrons, 7 neutrons
e. 34 protons, 34 electrons, 43 neutrons

43. Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:

a. $^{7}_{3}\text{Li}$
b. $^{125}_{52}\text{Te}$
c. $^{109}_{47}\text{Ag}$
d. $^{15}_{7}\text{N}$
e. $^{31}_{15}\text{P}$

Answer

a. 3 protons, 3 electrons, 4 neutrons
b. 52 protons, 52 electrons, 73 neutrons
c. 47 protons, 47 electrons, 62 neutrons
d. 7 protons, 7 electrons, 8 neutrons
e. 15 protons, 15 electrons, 16 neutrons

4.8: Atomic Mass

44. Define *atomic mass*. What is its unit?

Answer

Atomic mass is the average mass of atoms of an element. Its unit is atomic mass units (amu).

45. Determine the atomic mass of each element, given the isotopic composition.

a. lithium, which is 92.4% lithium-7 (mass 7.016 amu) and 7.60% lithium-6 (mass 6.015 amu)
b. oxygen, which is 99.76% oxygen-16 (mass 15.995 amu), 0.038% oxygen-17 (mass 16.999 amu), and 0.205% oxygen-18 (mass 17.999 amu)

Answer

a. 6.94 amu
b. 16.00 amu

46. Determine the atomic mass of each element, given the isotopic composition.

a. neon, which is 90.48% neon-20 (mass 19.992 amu), 0.27% neon-21 (mass 20.994 amu), and 9.25% neon-22 (mass 21.991 amu)
b. uranium, which is 99.27% uranium-238 (mass 238.051 amu) and 0.720% uranium-235 (mass 235.044 amu)

Answer

a. 20.18 amu
b. 238.0 amu

47. An element has the following natural abundances and isotopic masses: 90.92% abundance with 19.99 amu, 0.26% abundance with 20.99 amu, and 8.82% abundance with 21.99 amu. Calculate the average atomic mass of this element.

Answer

20.17 amu

48. Average atomic masses listed by IUPAC are based on a study of experimental results. Bromine has two isotopes ^{79}Br and ^{81}Br, whose masses (78.9183 and 80.9163 amu) and abundances (50.69% and 49.31%) were determined in earlier experiments. Calculate the average atomic mass of bromine based on these experiments.

Answer

79.90 amu

49. Variations in average atomic mass may be observed for elements obtained from different sources. Lithium provides an example of this. The isotopic composition of lithium from naturally occurring minerals is 7.5% ^{6}Li and 92.5% ^{7}Li, which have masses of 6.01512 amu and 7.01600 amu, respectively. A commercial source of lithium, recycled from a military source, was 3.75% ^{6}Li (and the rest ^{7}Li). Calculate the average atomic mass values for each of these two sources.

Answer

mineral = 6.94 amu; commercial = 6.978 amu

50. The ^{18}O:^{16}O abundance ratio in some meteorites is greater than that used to calculate the average atomic mass of oxygen on earth. Is the average mass of an oxygen atom in these meteorites greater than, less than, or equal to that of a terrestrial oxygen atom?

Answer

The average mass of an oxygen atom in the meteorites will be greater than the average mass of a terrestrial oxygen atom due to the extra ^{18}O.

4.9: Ion Formation

51. Explain how cations form.

Answer

Cations are formed when an atom loses electrons.

52. Explain how anions form.

Answer

Anions are formed when an atom gains electrons.

53. Give the charge each atom takes when it forms an ion.

a. K

b. O

Answer

a. K^+
b. O^{2-}

54. Give the charge each atom takes when it forms an ion.

a. Ca
b. I

Answer

a. Ca^{2+}
b. I^-

55. Give the charge each atom takes when it forms an ion.

a. Al
b. Br

Answer

a. Al^{3+}
b. Br^-

56. Give the charge each atom takes when it forms an ion.

a. S
b. Na

Answer

a. S^{2-}
b. Na^+

57. Write the symbol for each of the following ions:

a. the ion with a 1+ charge, atomic number 55, and mass number 133
b. the ion with 54 electrons, 53 protons, and 74 neutrons
c. the ion with atomic number 15, mass number 31, and a 3− charge
d. the ion with 24 electrons, 30 neutrons, and a 3+ charge

Answer

a. $^{133}Cs^+$
b. $^{127}I^-$
c. $^{31}P^{3-}$
d. $^{57}Co^{3+}$

58. Write the symbol for each of the following ions:

a. the ion with a 3+ charge, 28 electrons, and a mass number of 71

b. the ion with 36 electrons, 35 protons, and 45 neutrons
c. the ion with 86 electrons, 142 neutrons, and a 4+ charge
d. the ion with a 2+ charge, atomic number 38, and mass number 87

Answer

$^{71}Ga^{3+}$

$^{80}Br^{-}$

$^{232}Th^{4+}$

$^{87}Sr^{2+}$

59. Determine the number of protons, electrons, and neutrons in the following isotopes that are used in medical diagnoses:

a. atomic number 9, mass number 18, charge of 1−
b. atomic number 43, mass number 99, charge of 7+
c. atomic number 53, atomic mass number 131, charge of 1−
d. atomic number 81, atomic mass number 201, charge of 1+

Answer

a. 9 protons, 10 electrons, 9 neutrons
b. 43 protons, 36 electrons, 56 neutrons
c. 53 protons, 54 electrons, 78 neutrons
d. 81 protons, 80 electrons, 120 neutrons

60. The following are properties of isotopes of two elements that are essential in our diet. Determine the number of protons, electrons and neutrons in each and name them.

a. atomic number 26, mass number 58, charge of 2+
b. atomic number 53, mass number 127, charge of 1−

Answer

a. 26 protons, 24 electrons, 32 neutrons; iron(II) ion
b. 53 protons, 54 electrons, 74 neutrons; iodide ion

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CHAPTER OVERVIEW

5: Chemical Nomenclature

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multi-atom combinations are called molecules. Molecules are the smallest part of a substance that have the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

5.1: Elements Combine to Form Compounds
5.2: Chemical Formulas
5.3: A Closer Look at Elements and Compounds
5.4: Classification of Compounds
5.5: Binary Ionic Compounds
5.6: Ions With Variable Charges
5.7: Ionic Compounds Containing Polyatomic Ions
5.8: Binary Molecular Compounds
5.9: Acids
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5.11: For Future Use
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5.1: Elements Combine to Form Compounds

When building a house, the starting point is a blueprint of what the house will look like. The plan states how many windows and what kind, how many doors and what style, how many rooms and what type (bedroom, kitchen, other). The blueprint shows how the different pieces will go together to make the house. As long as the blueprint is followed and exactly the same items are used, the house will be identical to its blueprint.

Compounds

A **compound** is a substance that contains two or more elements chemically combined in a fixed proportion. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen atoms as carbon atoms. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances – carbon and hydrogen.

Recall that the components of a mixture may be separated from one another by physical means. This is not true for compounds. Compounds require chemical means to separate them into elements and/or simpler compounds.

Sodium chloride, also known as table salt, is an ionic compound with the chemical formula NaCl, representing a 1:1 ratio of sodium and chloride ions. It is commonly used as a condiment and food preservative. Salt can be created by adding two very reactive elements together: sodium, Na (s), metal and chlorine, Cl_2 (g), gas.

$$2\,Na\,(s) + Cl_2\,(g) \rightarrow 2\,NaCl\,(s) \tag{5.1.1}$$

The element sodium (Figure 5.1.1a) is a very reactive metal. Given the opportunity, it will react with the sweat on your hands and form sodium hydroxide, which is a very corrosive substance. The element chlorine (Figure 5.1.1b) is a pale yellow, corrosive gas that should not be inhaled due to its poisonous nature. Bring these two hazardous substances together, however, and they react to make the ionic compound sodium chloride (Figure 5.1.1c), known simply as salt.

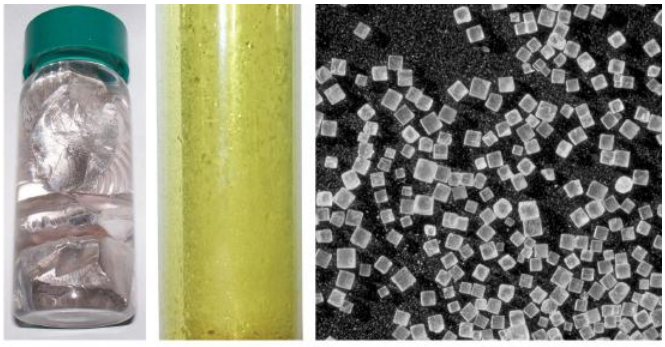


Figure 5.1.1: Sodium + Chlorine → Sodium Chloride (a) Sodium is a very reactive metal. (b) Chlorine is a pale yellow, noxious gas. (c) Together, sodium and chlorine make sodium chloride – salt – which is necessary for our survival. Source: (a) courtesy of reenhorn1, (b) courtesy of Benjah-bmm27, (c) © Thinkstock.

Salt is necessary for life. Na^+ ions are one of the main ions in the human body and are necessary to regulate the fluid balance in the body. Cl^- ions are necessary for proper nerve function and respiration. Both of these ions are supplied by salt. The taste of salt is one of the fundamental tastes; salt is probably the most ancient flavoring known, and one of the few rocks we eat. Clearly when the elemental sodium and chlorine combine (Equation 5.1.1), the resulting salt product has radically different properties (both physical and chemical). This reaction is spectacular to observe (Video 5.1.1).



Video 5.1.1: Making table salt using sodium metal and chlorine gas

Another compound is sugar, which is the generic name for sweet, soluble carbohydrates, many of which are used in food. Sugar has the chemical formula $C_{12}H_{22}O_{11}$ and is constructed from different elements than salt: carbon, hydrogen and oxygen. While sugar qualitatively resembles table salt (often confused in the kitchen), the two have distinctly different physical and chemical properties. There are various types of sugar derived from different sources. While sugar is made with carbon, hydrogen, and oxygen, it is considerably harder to synthesize from its constituent elements than table salt (Equation 5.1.1). However, the thermal decomposition is considerably easier and can be represented as a dehydration of sucrose to pure carbon and water vapor in Equation 5.1.2, and demonstrated in Video 5.1.2.

$$\mathrm{C_{12}H_{22}O_{11}}(s) + \mathrm{heat} \rightarrow 12\,\mathrm{C}(s) + 11\,\mathrm{H_2O}(g) \tag{5.1.2}$$



Video 5.1.2: A science experiment in the kitchen shows what happens to sugar molecules when they are heated. The experiment did not disappoint!

As with salt, sugar has radically different properties (both physical and chemical) than its constituent elements. This difference in properties of constituent elements and compounds is a central feature of chemical reactions.

Summary

- A compound is a substance that contains two or more elements chemically combined in a fixed proportion.
- Compounds may be broken down into simpler substances by chemical means.

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5.2: Chemical Formulas

Learning Objectives

- Determine the number of different atoms in a formula.
- Define chemical formula, molecular formula, and empirical formula.

A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen atoms and oxygen atoms chemically bonded in a 2:1 ratio, which leads to the chemical formula for water being written as H_2O.



Figure 5.2.1: A molecule of water, H_2O.

Molecular Formula

A **molecular formula** is a chemical formula of a molecular compound that represents the numbers of atoms of each element present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

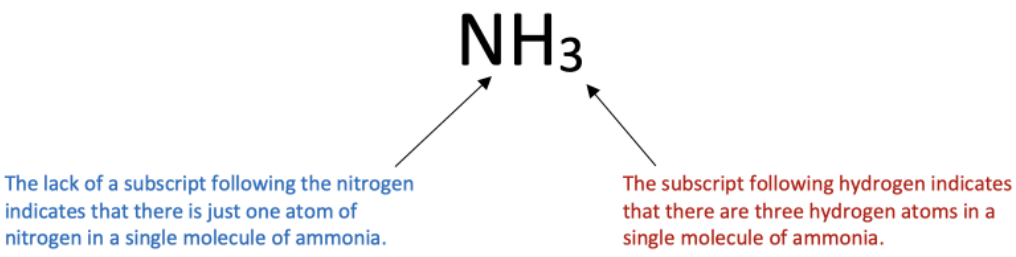


Figure 5.2.2: The molecular formula for ammonia.

The arrangement of the elements depends on the particular structure, which is not of concern at this point. The number of atoms of each kind is indicated by a subscript following the atom. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. A value of "1" is implied when there is no subscript written after an element. We would not write N_3H for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule.

If we consider the molecular formula for sulfuric acid, H_2SO_4, one of the most widely produced chemicals in the United States, we could conclude that one molecule of H_2SO_4 contains two hydrogen atoms, one sulfur atom, and four oxygen atoms.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $\mathbf{C_6H_{12}O_6}$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is $\mathbf{CH_2O}$. When chemists analyze an unknown compound, the first step is often to determine its empirical formula.

- molecular formula: $C_6H_{12}O_6$
- empirical formula: CH_2O

There are a many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.

Polyatomic Ions

Certain groups of atoms often bond together to form what is called a polyatomic ion, a group that often behaves as a single unit in chemical reactions. Polyatomic ions are discussed in further detail in Section 5.7. If more than one of a particular polyatomic ion exists in a chemical formula, it will be enclosed in parentheses followed by a subscript that indicates the quantity of that polyatomic ion present in the formula. To count the total number of atoms for formulas with polyatomic ions enclosed in parentheses, the subscript is used as a multiplier for each atom or number of atoms.

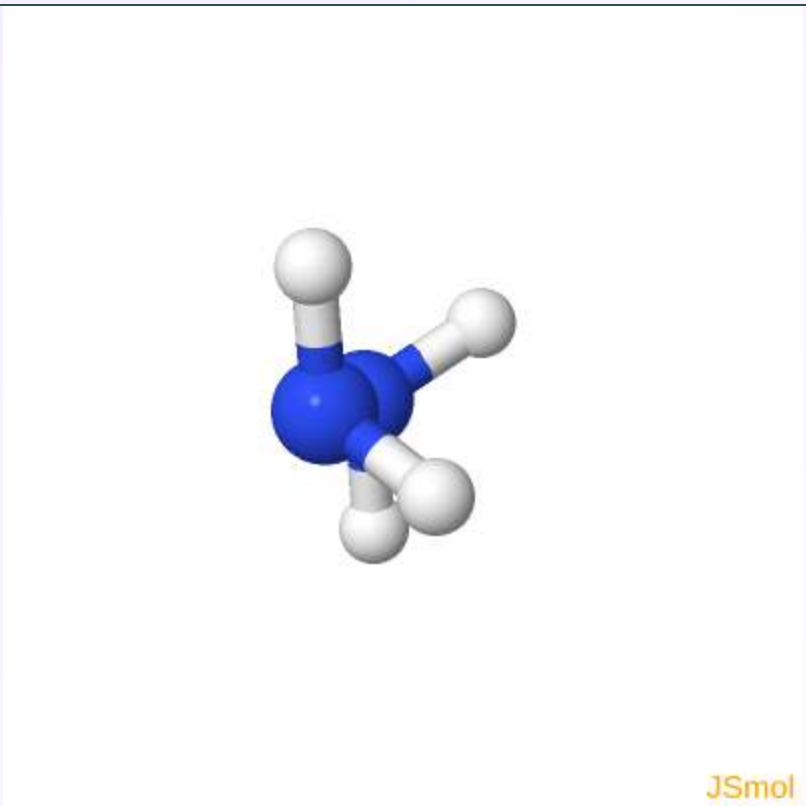
✅ Example 5.2.1

How many atoms of each element are represented by the formula of calcium phosphate, $Ca_3(PO_4)_2$?

Solution

$Ca_3(PO_4)_2$	**Explanation**
3 Ca atoms	The subscript "3" follows the symbol of calcium.
2 P atoms	P has no subscript, which implies there is 1 phosphorus atom within PO_4. The subscript of "2" outside the parentheses serves as a multiplier of all the atoms within the parentheses. Therefore, $2 \times 1 = 2$.
8 O atoms	O has a subscript of "4". The subscript of "2" outside the parentheses serves as a multiplier of all the atoms within the parentheses. Therefore, $2 \times 4 = 8$.

Exercise 5.2.1



Given that the blue atoms represent nitrogen and the white atoms represent hydrogen:

A. What is its molecular formula? (nitrogen is written first in the formula)
B. What is its empirical formula? (again, nitrogen is written first)

Answer A

N_2H_4

Answer B

NH_2

Exercise 5.2.2

How many atoms of each element are represent in each formula?

A. $C_{12}H_{22}O_{11}$
B. Li_2CO_3
C. $(NH_4)_2Cr_2O_7$

Answer A

12 carbon, 22 hydrogen, 11 oxygen

Answer B

2 lithium, 1 carbon, 3 oxygen

Answer C

2 nitrogen, 8 hydrogen, 2 chromium, 7 oxygen

Summary

- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- If only one atom of a specific type is present, no subscript is used.
- For atoms that have two or more of a specific type of atom present, a subscript is written after the symbol for that atom.
- Polyatomic ions in chemical formulas are enclosed in parentheses followed by a subscript if more than one of the same type of polyatomic ion exist.
- Molecular formulas do not indicate how the atoms are arranged in the molecule.
- The empirical formula tells the lowest whole-number ratio of elements in a compound. The empirical formula does not show the actual number of atoms.

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5.3: A Closer Look at Elements and Compounds

Learning Objectives

- Classify substances as atomic elements, molecular elements, molecular compounds, or ionic compounds.

Atomic Elements

Most elements exist with **individual atoms** as their basic unit. It is assumed that there is only one atom in a formula if there is no numerical subscript to the right of a chemical symbol.

Molecular Elements

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multi-atom combinations are called **molecules**. A molecule is the smallest part of a substance that retains the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is comprised of two or more atoms.

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules, called **diatomic molecules**. A complete list of elements that exist naturally as diatomic molecules are shown Table 5.3.1. As with any molecule, these elements are labeled with a **molecular formula**, a formal listing of what and how many atoms are in a single molecule.

Table 5.3.1: Elements That Exist as Diatomic Molecules

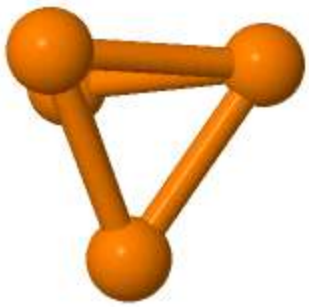
Hydrogen	Oxygen	Nitrogen	Chlorine	Bromine	Iodine	Fluorine
H_2	O_2	N_2	Cl_2	Br_2	I_2	F_2

It would be helpful to memorize the elements that exist naturally as diatomic molecules. One way to do this is to use a mnemonic. The elements in Table 5.3.1 were purposefully arranged to provide such a mnemonic. When arranged in this order, the elements H, O, N, Cl, Br, I, F may be pronounced phonetically as HONK-uhl-briff.

Other elements exist as polyatomic molecules. For example, sulfur naturally exists as an eight-atom molecule, S_8, while phosphorus exists as a four-atom molecule, P_4 , as shown in Figure 5.3.1 below.



JSmol



JSmol

Figure **5.3.1**: Molecular models of S_8 and P_4. (from the PubChem database).

Ionic Compounds

The elements in the periodic table are divided into specific groupings: the metals, the nonmetals, the metalloids, and so on. These groupings are largely based on physical properties and on the tendency of the various elements to bond with other elements by forming either an ionic or a covalent bond. As a general rule of thumb, compounds that involve a metal bonding with either a nonmetal or a metalloid will display ionic bonding.

Figure 5.3.2 shows an atomic level view of sodium chloride. This compound would be expected to be ionic since a metal, Na, is bonded to a nonmetal, Cl. Since the ratio of sodium ions to chloride ions is 1:1, the formula for sodium chloride is NaCl. The basic unit of an ionic compound that retains the identity of the compound is called a **formula unit**. It is incorrect to refer to the basic unit of an ionic compound as a molecule (refer to the section on Molecular Compounds below). One formula unit of NaCl contains one sodium ion and one chloride ion.



Figure **5.3.2**: Atomic level view of sodium chloride, NaCl. (Benjah-bmm27 , Public domain, via Wikimedia Commons)

Molecular Compounds

Compounds that are composed of only nonmetals and/or metalloids with non-metals will display covalent bonding and will be classified as molecular compounds. Examples of molecular compounds include nitrogen monoxide (NO), consisting of two nonmetals covalently bonded together, and silicon dioxide (SiO_2), consisting of a metalloid covalently bonded to a nonmetal. The basic unit of molecular compounds is the **molecule**.

✅ Example 5.3.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

A. Fe
B. PCl_3
C. LiBr
D. P_4
E. oxygen gas

Solution

A. Fe (iron) is an element that is represented with no subscript, so it is an **atomic element**.
B. PCl_3 is made up of two nonmetals, so it is a **molecular compound**.
C. LiBr is made up of lithium, a metal, and bromine, a nonmetal, so it is an **ionic compound**.
D. P_4 is a substance that is made up of four atoms of the same element, so it is a **molecular element**.
E. The formula for oxygen gas is O_2 so it is a **molecular element**.

✏️ Exercise 5.3.1

Provide the classification (i.e. atomic element, molecular element, molecular compound, or ionic compound) of each substance.

A. I_2
B. He
C. H_2O
D. Al
E. CuCl

Answer A

molecular element

Answer B

atomic element

Answer C

molecular compound

Answer D

atomic element

Answer E

ionic compound

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5.4: Classification of Compounds

Learning Objectives

- Classify compounds based on their composition.

Nomenclature is the process of naming chemical compounds so they may be easily identified as separate chemicals. The primary function of **chemical nomenclature** is to ensure that a spoken or written chemical name leaves no ambiguity as to which chemical compound the name refers – each chemical name should refer to a single substance. A less important aim is to ensure that each substance has a single name, although a limited number of alternative names is acceptable in many cases. Preferably, the name also conveys some information about the structure or chemistry of a compound. A common name will often suffice to identify a chemical compound in a particular set of circumstances. To be more generally applicable, the name should indicate the chemical formula at a very minimum. To be more specific still, the three-dimensional arrangement of the atoms may need to be specified.

Ionic Compounds vs. Molecular Compounds

The distinction between ionic compounds and molecular compounds was discussed in Section 5.3 and is summarized in Table 5.4.1 below.

Table 5.4.1: Distinguishing Between Ionic Compounds and Molecular Compounds

Ionic Compounds	Molecular Compounds
Consist of metals bonded to nonmetals with one class of exceptions – ammonium compounds	Consist of nonmetals and/or metalloids bonded to other nonmetals and/or metalloids
Atoms joined by ionic bonds	Atoms joined by covalent bonds

Further Classification of Compounds

Students typically become more successful at writing chemical formulas and chemical names when they are able to properly classify compounds. Many compounds are said to be **binary**, which means they consist of two different elements. Other compounds contain **polyatomic ions**, previously identified as groups of covalently bonded atoms that often act as single units in chemical reactions. **Acids** are compounds that contain the element hydrogen, H, typically written at the front of a chemical formula. **Oxyacids** are acids that also contain the element oxygen.

With this in mind, many compounds may be classified into five different categories, each of which uses a slightly different, but related, set of rules. These five common classes of compounds are:

1. Binary ionic compounds
2. Ionic compounds containing polyatomic ions
3. Binary molecular compounds
4. Binary acids
5. Oxyacids

Exercise 5.4.1

Classify each compound into one of the five categories listed above. Provide a short explanation for your choice.

A. H_3PO_4
B. Ag_2S

C. HI
D. SF_6
E. Na_2SO_4
F. $Ti(C_2H_3O_2)_4$

Answer A

H_3PO_4 is an **oxyacid**. H is written at the front of the formula, making it an acid. Since it also contains O, it is an oxyacid.

Answer B

Ag_2S is a **binary ionic compound**. Ag_2S has two elements, which makes it binary. Ag is a metal and S is a nonmetal, making the compound ionic.

Answer C

HI is a **binary acid**. HI has two elements, which makes it binary. H is written at the front of the formula, making it an acid.

Answer D

SF_6 is a **binary molecular compound**. SF_6 has two elements, which makes it binary. S and F are both nonmetals, making the compound molecular.

Answer E

Na_2SO_4 is an **ionic compound containing a polyatomic ion**. Na_2SO_4 has three elements, so it cannot be binary. Na is a metal and both S and O are nonmetals, making the compound ionic.

Answer F

$Ti(C_2H_3O_2)_4$ is an **ionic compound containing a polyatomic ion**. $Ti(C_2H_3O_2)_4$ has four elements, so it cannot be binary. Ti is a metal, while C, H, and O are all nonmetals, making the compound ionic.

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5.5: Binary Ionic Compounds

Learning Objectives

- Write the correct names and formulas for binary ionic compounds.

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in a crystal lattice containing many ions each of the cation and anion. An ionic formula, like NaCl, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to write the names of binary ionic compounds, as well as to determine the correct ratio of ions, so that you may correctly write a formula.

Writing Chemical Names

As we recall, a binary ionic compound is one that contains two elements, one of which is a metal and the other of which is a nonmetal. The metal is always written first in the name and formula followed by the nonmetal. The convention for naming binary ionic compounds is

- Metals retain their name.
- Nonmetals retain their root name followed by a suffix of *-ide*.

The root name represents the first one or two syllables of an element's name:

Table 5.5.1: Root Names for the Common Nonmetals

Symbol	Name	Root Name	Symbol	Name	Root Name
H	hydrogen	*hydr-*	P	phosphorus	*phosph-*
C	carbon	*carb-*	S	sulfur	*sulf-*
N	nitrogen	*nitr-*	Cl	chlorine	*chlor-*
O	oxygen	*ox-*	Br	bromine	*brom-*
F	fluorine	*fluor-*	I	iodine	*iod-*

✅ Example 5.5.1: Writing Chemical Names

Write the name for each compound.

A. Al_2S_3
B. $BaBr_2$
C. K_3N

Solution

A. Al = aluminum. S = sulfur. The root name of sulfur is *sulf-*. The nonmetal has a suffix of *-ide*. Al_2S_3 is aluminum sulfide.

B. Ba = barium. Br = bromine. The root name of bromine is *brom-*. The nonmetal has a suffix of *-ide*. $BaBr_2$ is barium bromide.

C. K = potassium. N = nitrogen. The root name of nitrogen is *nitr-*. The nonmetal has a suffix of *-ide*. K_3N is potassium nitride.

Writing Chemical Formulas

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charges to balance each other. If the charges of the ions are 2+ and 2–, nothing needs to be done – the charges are already balanced. If the charges of the ions are 3+ and 1–, three of the ions with the 1– charge are needed to balance one of the ions with the 3+ charge.

When balancing charges, many find it easiest to think of the least common multiple between the magnitude of charge. For example, if the charges of the ions are 2+ and 3–, the least common multiple between 2 and 3 is 6. Therefore, the charge will be balanced when the positive charges sum to 6 and the negative charges sum to 6.

The ions shown in Figure 5.5.1 have fixed charges related to their position on the periodic table, as previously discussed in Section 4.8. Note that most main-group elements (groups labeled with Roman numerals followed by the letter *A*) have fixed charges, while most transition metals (groups labeled with Roman numerals followed by the letter *B*) have charges that are variable. Metals that have variable charges are discussed in the next section.

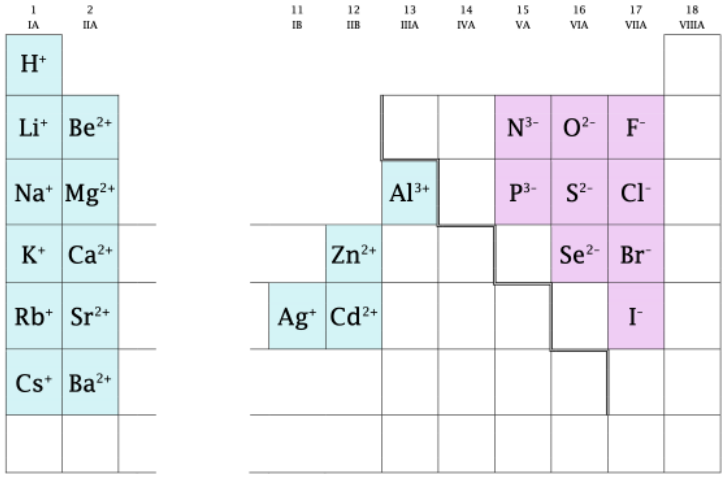


Figure **5.5.1**: Ions with fixed charges related to position on the periodic table.

✅ Example 5.5.2: Writing Chemical Formulas

Write the chemical formulas for:

- lithium fluoride
- sodium oxide
- strontium iodide

Solution

	lithium fluoride
1. Write the symbol and charge of the cation and the anion.	Li^+ F^-
2. Balance the charge of the cations and anions.	Li^+ F^-
3. Explanation.	1+ balances 1–

4. Write the final formula. Charges are left off. Subscripts of 1 are implied.	LiF

✅ Example 5.5.3: Writing Chemical Formulas

Write the chemical formulas for:

- magnesium sulfide
- aluminum oxide

Solution

	magnesium sulfide	aluminum oxide
1. Write the symbol and charge of the cation and anion.	Mg^{2+} S^{2-}	Al^{3+} O^{2-}
2. Balla	Mg^{2+} S^{2-}	Al^{3+} O^{2-} Al^{3+} O^{2-} O^{2-}

	magnesium sulfide	aluminum oxide
nce the charge of the cations and anions.		
3. Explanation.	2+ balances 2–	6+ balances 6–
4. Wri	MgS	Al_2O_3

magnesium sulfide	aluminum oxide
the final formula. Charges are left off. Subscripts of 1 ar	

magnesium sulfide	aluminum oxide
e i m p l i e d .	

Exercise 5.5.1

If the chemical formula is provided, write the name. If the name is provided, write the formula.

A. Li_2S
B. Ca_3N_2
C. potassium phosphide
D. barium oxide

Answer A

lithium sulfide

Answer B

calcium nitride

Answer C

K_3P

Answer D

BaO

Summary

- Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.
- Ionic compounds must be neutral overall.

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5.6: Ions With Variable Charges

Learning Objectives

- Write the correct names and formulas for binary ionic compounds that have ions with variable charges.

Up until now, it has been assumed that elements form ions with fixed charges. However, many metals form ions that have variable charges. In fact, the large majority of transition metals form ions with variable charges. There is also a small block of main group metals that form ions with variable charges as well (see Figure 5.6.1 below). As we will find out, there is no need to memorize the charges of metals that form ions with variable charges, as the charges may be identified by using Roman numerals. When writing the formula of the ion, the name of the metal is immediately followed by parentheses (no space) within which is a Roman numeral equal to the magnitude of charge on the ion.

For example,

- the iron(II) ion has a formula of Fe^{2+}.
- the iron(III) ion has a formula of Fe^{3+}.

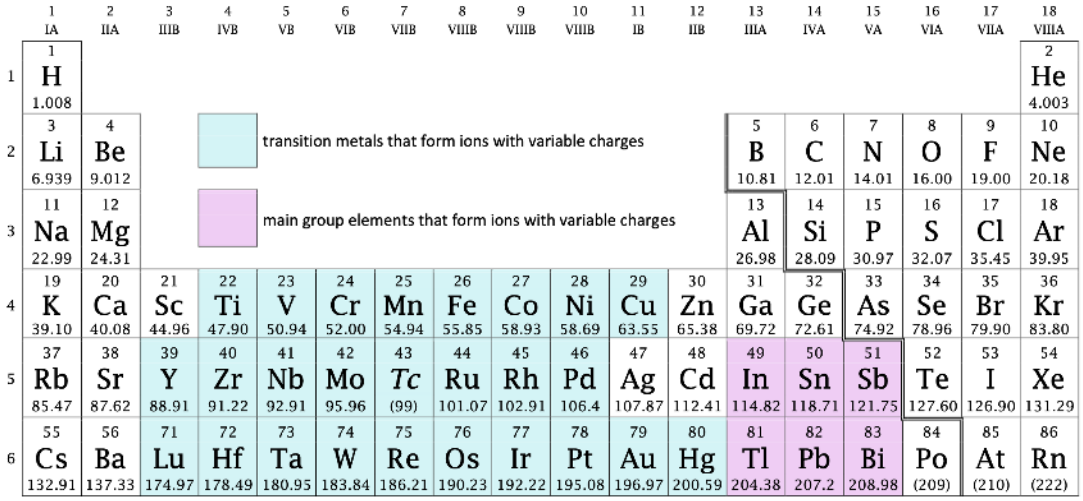


Figure 5.6.1: The periodic table showing elements that form ions with variable charges.

Exercise 5.6.1

If the formula of an ion is provided, write the name. If the name of an ion is provided, write the formula.

A. Mn^{3+}
B. copper(I) ion
C. lead(IV) ion
D. Cu^{2+}

Answer A

manganese(III) ion

Answer B

Cu^{+}

Answer C

Pb^{4+}

Answer D
copper(II) ion

When writing names for compounds containing metals that form ions with variable charges, the name *must include* the Roman numeral that shows the charge on the metal ion. Roman numerals are *not* included when the metal only forms ions with fixed charges. Chemical formulas themselves will never include Roman numerals as the charge on the metal ions present may be determined from the subscripts in the chemical formula.

✅ Example 5.6.1: Writing Chemical Formulas

Write the chemical formula for each compound.

A. iron(II) oxide
B. iron(III) oxide

Solution

	iron(II) oxide	iron(III) oxide
1. Write the symbol and charge of the cation and the anion.	Fe^{2+} O^{2-}	Fe^{3+} O^{2-}
2. Balla	Fe^{2+} O^{2-}	Fe^{3+} O^{2-} Fe^{3+} O^{2-} O^{2-}

	iron(II) oxide	iron(III) oxide
nce the charge of the cation and anions.		
3. Explanation.	2+ balances 2–	6+ balances 6–
4. Wri	FeO	Fe_2O_3

iron(II) oxide	iron(III) oxide
the final formula. Charges are left off. Subscripts of 1 ar	

iron(II) oxide	iron(III) oxide
e i m p l i e d .	

✅ Example 5.6.2: Writing Chemical Names

Give the correct name for each compound.

A. CuO
B. Cu_2O

Solution

	CuO	Cu_2O
1. Write the name leaving room for a Roman numeral, since copper forms ions with variable charges.	copper( ) oxide	copper( ) oxide
2. Based on position on the periodic table, the oxide ion has a charge of 2–. Note that CuO has only one copper ion, while Cu_2O has two copper ions. Both have only one oxide ion.	$Cu^{?+}$ O^{2-}	$Cu^{?+}$ $Cu^{?+}$ O^{2-}
3. Determine the charge of the copper ion ($Cu^{?+}$).	Cu^{2+} must have a charge of 2+ (Cu^{2+}) to balance oxide's charge of 2– (O^{2-}).	Each $Cu^{?+}$ must have a charge of 1+ (Cu^{+}) to balance oxide's charge of 2– (O^{2-}).
4. The charge on the copper ion determines the Roman numeral that belongs inside the parentheses.	Cu^{2+} means copper(II), so the name is: copper(II) oxide	Cu^{+} means copper(I), so the name is: copper(I) oxide

✏️ Exercise 5.6.2

If a formula is provided, write the name of the compound. If a name is provided, write the formula of the compound.

A. titanium(IV) sulfide
B. Co_2O_3
C. $PbCl_2$

Answer A

TiS_2

Answer B

cobalt(III) oxide

Answer C

lead(II) chloride

Stock System vs. Common Names

The approach for naming the ions for metals that have variable charges is called the **Stock system**, in which an ion's positive charge is indicated by a Roman numeral in parentheses after the element name. The Na^+ ion is not called the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

A 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 root name of the element. The *-ic* suffix represents the cation with the larger of the two charges, and the *-ous* suffix represents the cation with the smaller of the two charges. In many cases, the stem of the element name comes from the Latin name of the element. Table 5.6.1 lists the elements that use the common system, along with their respective cation names. *This text uses the Stock system of nomenclature.*

Table 5.6.1: The Modern and Common System of Cation Names

Element	Root Name	Charge	Modern Name	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion

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5.7: Ionic Compounds Containing Polyatomic Ions

Learning Objectives

- Write correct names and formulas for ionic compounds containing polyatomic ions.

Some ions consist of groups of atoms bonded together that carry an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, the nitrate ion has one nitrogen atom and three oxygen atoms with an overall charge of 1−. The formula of the nitrate ion is written as NO_3^-. A list of the most common polyatomic ions appears in Table 5.7.1. Since these ions are regularly used throughout the rest of this text, it generally a good practice to memorize this list of ions.

Table 5.7.1: The Most Common Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$
bicarbonate ion (hydrogen carbonate ion)	HCO_3^-
carbonate ion	CO_3^{2-}
chlorate ion	ClO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
phosphate ion	PO_4^{3-}
sulfate ion	SO_4^{2-}

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

Writing formulas for ionic compounds containing polyatomic ions involves the same steps as for a binary ionic compound. The cation is written first in the chemical formula followed by the anion. Consider potassium nitrate. Potassium is located in Group IA of the periodic table, so it forms ions with a 1+ charge, K^+. As shown above in Table 5.7.1, the formula of the nitrate ion is NO_3^-. Since the charges of the two ions are already balanced, the formula for potassium nitrate may be written as KNO_3.

When more than one polyatomic ion is required to balance charge, the formula of the polyatomic ion is enclosed in parentheses followed by a subscript that indicates the quantity of that polyatomic ion required to balance the charge. This means that the formula of calcium nitrate is written as $Ca(NO_3)_2$, since calcium is located in Group IIA and forms ions with a 2+ charge, Ca^{2+}, and will require two nitrate ions, NO_3^-, to balance the charge of a single calcium ion.

It should be emphasized that enclosing a polyatomic ion in parentheses is only done when more than one is required to balance charge. Parentheses are not included when only one polyatomic ion is required to balance charge. That is why the formula of potassium nitrate is written as KNO_3 and *not* as $K(NO_3)$.

✅ Example 5.7.1: Writing Chemical Formulas

Write the chemical formulas for:

- copper(I) carbonate

- aluminum acetate
- magnesium hydroxide

Solution

	copper(I) carbonate		aluminum acetate		magnesium hydroxide	
1. Write the symbol and charge of the cation and the anion.	Cu^+	CO_3^{2-}	Al^{3+}	$C_2H_3O_2^-$	Mg^{2+}	OH^-
2. Balance the charge of the cations and anions.	Cu^+ Cu^+	CO_3^{2-}	Al^{3+}	$C_2H_3O_2^-$ $C_2H_3O_2^-$ $C_2H_3O_2^-$	Mg^{2+}	OH^- OH^-
3. Explanation.	2+ balances 2–		3+ balances 3–		2+ balances 2–	
4. Write the final formula. Charges are left off. Subscripts of 1 are implied.	Cu_2CO_3		$Al(C_2H_3O_2)_3$		$Mg(OH)_2$	

Writing Names for Ionic Compounds Containing Polyatomic Ions

The reverse process is used when writing names for ionic compounds containing polyatomic ions. Each ion is named in the order that they appear in the chemical formula. The most important part is to recognize that a polyatomic ion is present in the compound. Na_3PO_4 would be called sodium phosphate, since it contains the sodium ion, Na^+, and the phosphate ion, PO_4^{3-}. $Mg(ClO_3)_2$ would be called magnesium chlorate, since it contains the magnesium ion, Mg^{2+}, and the chlorate ion, ClO_3^-.

✅ Example 5.7.2: Writing Chemical Names

Write the name for each compound.

A. $(NH_4)_2S$
B. $FeSO_4$
C. $Fe_2(SO_4)_3$

Solution

A. NH_4^+ is the ammonium ion. S = sulfur. The root name of sulfur is *sulf-*. The nonmetal has a suffix of *-ide*. $(NH_4)_2S$ is ammonium sulfide.

B. SO_4^{2-} is the sulfate ion. Fe is a transition element that forms ions with multiple charges. Since only one iron ion is needed to balance the charge of one sulfate ion, it must have a 2+ charge. An iron ion with a 2+ charge is called the iron(II) ion. $FeSO_4$ is iron(II) sulfate.

C. SO_4^{2-} is the sulfate ion. Fe is a transition element that forms ions with multiple charges. Since two iron ions are needed to balance the charge of three sulfate ions, it must have a 3+ charge. An iron ion with a 3+ charge is called the iron(III) ion. $Fe_2(SO_4)_3$ is iron(III) sulfate.

Oxyanions

You may have noticed that most of the polyatomic ions in Table 5.7.1 carry a negative charge and contain the element oxygen. Anions that contain the element oxygen are called **oxyanions**. Many oxyanions belong to a series where the number of oxygen atoms in the ion varies, as shown in Table 5.7.2.

Table 5.7.2: Oxyanions Belonging to a Series

Name	Formula
chlorate ion	ClO_3^-
chlorite ion	ClO_2^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
sulfate ion	SO_4^{2-}
sulfite ion	SO_3^{2-}

Note that oxyanions may end with either an *-ate* or *-ite* suffix. In other words, if you encounter a compound (or polyatomic ion) that ends with *-ate* or *-ite*, it will *always* contain oxygen. To be clear, simply containing oxygen doesn't mean the compound will always end with an *-ate* or *-ite* suffix. We have seen examples of oxides, where oxygen is part of a binary compound such as lithium oxide, Li_2O, and hydroxides, where oxygen is part of the polyatomic ion OH^-.

When comparing oxyanions that end with *-ate* and *-ite*, the ion that ends with *-ate* will always have one more oxygen than the corresponding ion that ends with *-ite*. The ion in the series that ends with *-ate* is the one found in Table 5.7.1: The Most Common Polyatomic Ions. This is because the *-ate* form is more commonly found in nature than the *-ite* form. It is also the reason why memorizing the form that ends with *-ate* is recommended rather than the one that ends with *-ite*. Furthermore, if you learn the one that ends with *-ate*, the ion that ends with *-ite* can easily be obtained, since it always has one less oxygen than the *-ate* form.

It should also be pointed out that just because the *-ite* form always has one less oxygen than the *-ate* form doesn't mean that an *-ite* form is possible. Even though the carbonate ion, CO_3^{2-}, is quite common, there is no such thing as the carbonite ion. Some of you reading this might be saying right now, "Wait a minute! I remember hearing about carbonite in Star Wars!" That's true. However, Star Wars is science fiction and may have even been the reason for the choice of using the word carbonite...it doesn't really exist!

✅ Example 5.7.3

A. What is the formula of zinc sulfite?

B. Write the chemical name of $Sr(NO_2)_2$?

C. If the bromate ion has a formula of BrO_3^-, what is the formula of the bromite ion?

Solution

A. From Figure 5.5.1, Zn is one of the three transition elements that has a fixed charge. Zinc ions have a charge of 2+, Zn^{2+}. The formula of the sulfite ion is SO_3^{2-}, which has one less oxygen atom than the sulfate ion. The charge on Zn^{2+} balances the charge on SO_3^{2-}, so the formula of zinc sulfite is $ZnSO_3$.

B. Sr^{2+} is the formula of the strontium ion. NO_2^- is the formula of the nitrite ion, which has one less oxygen than the nitrate ion. $Sr(NO_2)_2$ is strontium nitrite.

C. The *-ite* form always contains one less oxygen atom that the ion with an *-ate* suffix. If the bromate ion is BrO_3^-, then the bromite ion has a formula of BrO_2^-.

Exercise 5.7.1

Write the formula for each:

A. silver cyanide
B. calcium bicarbonate
C. lithium chlorite
D. nickel(III) sulfate
E. phosphite ion

Answer A

AgCN

Answer B

$Ca(HCO_3)_2$

Answer C

$LiClO_2$

Answer D

$Ni_2(SO_4)_3$

Answer E

PO_3^{3-}

Exercise 5.7.2

Write the name of each:

A. Na_2SO_3
B. $Cd(C_2H_3O_2)_2$
C. $(NH_4)_3PO_4$
D. KOH
E. $Pb(NO_3)_4$

Answer A

sodium sulfite

Answer B

cadmium acetate

Answer C

ammonium phosphate

Answer D

potassium hydroxide

Answer E

lead(IV) nitrate

Additional Polyatomic Ions

There are many other polyatomic ions that may be encountered in this text besides the ones already listed. Table 5.7.3 is not an all-inclusive list of additional polyatomic ions that exist, but rather ions you may expect to see in this text. Ions are grouped together if they are a part of a series.

Table 5.7.3: Additional Polyatomic Ions

Name	Formula
hypochlorite ion	ClO^-
perchlorate ion	ClO_4^-
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
oxalate ion	$C_2O_4^{2-}$
permanganate ion	MnO_4^-
monohydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
bisulfate ion (hydrogen sulfate ion)	HSO_4^-
bisulfite ion (hydrogen sulfite ion)	HSO_3^-

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5.8: Binary Molecular Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

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

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



Figure **5.8.1**: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Larger molecules can have many, many bonds that serve to hold the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O. A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18}.



Figure 5.8.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas and a prominent air pollutant produced by internal combustion engines.

Naming binary molecular compounds is similar to naming binary ionic compounds. The convention for naming binary molecular compounds:

- The first element in the formula retains its name.
- The second element retains its root name followed by a suffix of *-ide*.
- Greek prefixes (see Table 5.8.1) are added to each name* to indicate the number of atoms of each element present in the formula.

Table 5.8.1: Greek Prefixes for Naming Binary Molecular Compounds

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

*The prefix *mono-* is not used for the first element's name.

Note

- Generally, the element that is above and/or to the left in the Periodic Table is written first in the formula, though there are a few exceptions. The order of common nonmetals in formulas of binary molecular compounds:
  - C, P, N, H, S, I, Br, Cl, O, F
- An *a* or *o* at the end of a prefix is dropped when it precedes oxide. Examples:
  - carbon monoxide (rather than carbon monooxide)
  - dinitrogen pentoxide (rather than dinitrogen pentaoxide)
- The prefix is *mono-* is usually not added to the first element's name even if there is only one atom of the first element in a formula.

Some examples of molecular compounds are listed in Table 5.8.2.

Table 5.8.2: Examples of Molecular Compounds

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

Notice that the *mono-* prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The formula S_2Cl_2 emphasizes that formulas for molecular compounds are not reduced to their lowest ratios. The *o* of the *mono-* and the *a* of *hepta-* are dropped from the name when paired with oxide.

✅ Example 5.8.1: Writing Chemical Names

Write the name for each compound.

A. ClF_3
B. P_2F_4
C. $BaBr_2$

Solution

A. Cl = chlorine. F = fluorine. The root name of fluorine is *fluor-*. The second element has a suffix of *-ide*. Cl does not need a prefix, since the formula shows one atom. The prefix that means "3" is *tri-*. ClF_3 is chlorine trifluoride.

B. P = phosphorus. F = fluorine. The root name of fluorine is *fluor-*. The second element has a suffix of *-ide*. The prefix that means 2 is *di-*. The prefix that means "4" is *tetra-*. P_2F_4 is diphosphorus tetrafluoride.

C. Be careful. $BaBr_2$ is a binary *ionic* compound. Rules for naming binary ionic compound were discussed in Section 5.5. Greek prefixes are *not* used for binary ionic compounds. Ba = barium. Br = bromine. The root name of bromine is *brom-*. The nonmetal has a suffix of *-ide*. $BaBr_2$ is barium bromide.

✅ Example 5.8.2: Writing Chemical Formulas

Write the chemical formula for each compound.

A. sulfur hexafluoride
B. dinitrogen pentoxide
C. aluminum oxide

Solution

Writing formulas for binary molecular compounds is simple. The prefixes tell exactly how many atoms of each element are present in one molecule. There is no need to look up charges based on the position in the periodic table. Nonmetals bond to each other with covalent bonds, where electrons are shared rather than transferred.

A. S = sulfur. F = fluorine. S has no prefix, so there is only one atom. The prefix *hexa-* means "6". Sulfur hexafluoride is SF_6.

B. N = nitrogen. O = oxygen. The prefix *di-* means "2". The prefix *pent-* means "5" (*a* is left off when paired with oxide). Dinitrogen pentoxide is N_2O_5.

C. Be careful. Aluminum oxide is a binary *ionic* compound. Rules for writing formulas for binary ionic compound were discussed in Section 5.5. Greek prefixes are *not* used for binary ionic compounds. The charge on an aluminum ion is 3+, Al^{3+}. The charge on an oxide ion is 2–, O^{2-}. Charge is balanced with two Al^{3+} and three O^{2-}. Aluminum oxide is Al_2O_3.

Exercise 5.8.1

If the chemical formula is provided, write the name. If the name is provided, write the formula.

A. CF_4
B. diarsenic pentaiodide
C. SO_3

Answer A

carbon tetrafluoride

Answer B

As_2I_5

Answer C

sulfur trioxide

Simple Molecular Compounds with Common Names

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
- H_2O_2: hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrogen atom and oxygen atoms in a nitrate ion are held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.

Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix -ide. Numerical prefixes are used to specify the number of atoms in a molecule.

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5.9: Acids

A spot test for gold has been in use for decades. The sample is first treated with nitric acid. Other metals may react or dissolve in this acid, but gold will not. Then the sample is added to a mixture of nitric acid and hydrochloric acid. Gold will only dissolve in this mixture. The term "acid test" arose from the California gold rush in the late 1840's when this combination was used to test for the presence of real gold. It has since come to mean "tested and approved" in a number of fields.

Acids

An **acid** can be defined in several ways. The most straightforward definition is that an acid is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions (H^+) when dissolved in water.



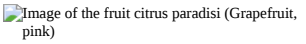


Figure **5.9.1**: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids.

This is a different type of compound than the others we have seen thus far. Acids are molecular, which means that in their pure state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$\mathrm{HCl} \rightarrow \mathrm{H^+ + Cl^-} \tag{5.9.1}$$

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula. The remainder of the acid (other than the H) is the anion after the acid dissolves.

Binary Acids

The name of an acid is based on the other elements with which it is combined. A **binary acid** is an acid that consists of hydrogen combined with one other element. The most common binary acids contain a halogen. The convention for naming binary acids is

hydro- + root name of the other element + *-ic acid*

In other words, the acid name begins with the prefix *hydro-*, followed immediately by the root name of the other element, which is then followed by the suffix *-ic acid*. It should also be noted that binary acids are only assigned their acid name when they are dissolved in water, which we designate with a phase label of (*aq*), which stands for aqueous (pronounced AY-quee-us).

Therefore, the name of HCl (*aq*) is hydrochloric acid, since the root name of chlorine is *chlor-* and we are showing that the HCl is dissolved in water. As a gas, HCl (*g*) is called hydrogen chloride. While this distinction is important, the phase label is usually left off and the correct name is usually based on the context in which it is used. In the absence of a phase label, one is usually safe to assume that a binary acid is named as an acid, in which the *hydro-* prefix and *-ic acid* suffix are used.

When writing the formula of an acid, the formula of a hydrogen ion, H^+, is combined with the formula and charge of the nonmetal anion in the acid. To write the formula of hydroiodic acid, H^+ would be combined with I^-, since iodine is located in Group VIIA of the periodic table (see Figure 5.5.1). The charges H^+ and I^- are already balanced, so the formula of hydroiodic acid is HI (*aq*). The formula could also be simply written as HI, since we know the context in which it was used (as an acid, according to its name).

✅ Example 5.9.1: Binary Acids

If the chemical formula is provided, write the name. If the name is provided, write the formula.

A. HF (*aq*)
B. hydrosulfuric acid

Solution

A. Since there are just two elements and the first element in the formula is H, this a binary acid. Binary acids have a *hydro-* prefix and an *-ic acid* suffix. The root name of fluorine is *fluor-*. HF (*aq*) is hydrofluoric acid.

B. The name has a *hydro-* prefix and an *-ic acid* suffix, making it a binary acid. H^+ would be combined with S^{2-}. The charge on sulfide is 2– based on its position (Group VIA) on the periodic table. Charge is balanced with two H^+ for every one S^{2-}. Hydrosulfuric acid is H_2S (*aq*) or simply H_2S.

Oxyacids

An acid always contains the element hydrogen and the H is always written at the front of the chemical formula. An **oxyacid** is an acid that also contains the element oxygen, along with a third (and occasionally fourth) element. The third element is almost always a nonmetal. Oxyacids may also be thought of as a compound where H^+ has combined with an oxyanion. As we recall, the most common oxyanions end with an *-ate* suffix, while other oxyanions end with an *-ite* suffix.

The convention for naming oxyacids:

- Identify and name the oxyanion present in the acid.
- If the oxyanion name ends with *-ate*, drop the *-ate* suffix from the anion name and replace with an *-ic acid* suffix.
- If the oxyanion name ends with *-ite*, drop the *-ite* suffix from the anion name and replace with an *-ous acid* suffix.

✅ Example 5.9.2: Naming Oxyacids

Write the name for each acid.

A. HNO_3
B. HNO_2

Solution

	HNO_3	HNO_2
1. Identify the oxyanion present in the acid.	NO_3^-	NO_2^-
2. Name the oxyanion	nitrate	nitrite

	HNO_3	HNO_2
n.		
3. Drop-*ate* or -*ite* from the anion name.	nitr-	nitr-
4. Replace -*ate* with -*ic*	nitric acid	nitrous acid

HNO_3	HNO_2
c a c i d . R e p l a c e - *i t e* w i t h - *o u s a c i d* .	

✅ Example 5.9.3: Writing Formulas for Oxyacids

Write the chemical formula for each acid.

A. chlorous acid
B. chloric acid

Solution

The process of arriving at a name now needs to be reversed.

	chlorous acid	chloric acid
1. Identify suffix as *-ic acid* or *-ous acid*.	chlorous acid	chloric acid
2. Replace *-ic acid*	chlorite	chlorate

	chlorous acid	chloric acid
with *-ate*. Replace *-ous acid* with *-ite*.		
3. Since it is an acid	H^+ with ClO_2^-	H^+ with ClO_3^-

	chlorous acid	chloric acid
	d , c o m b i n e H + w i t h t h e f o r m u l a o f t h e o x y a n i o n .	
4.	B a l a n c e c h a	one H^+ balances one ClO_2^-	one H^+ balances one ClO_3^-

chlorous acid	chloric acid
r g e s .	
5. W r i t e t h e f i n a l f o r m u l a . C h a r g e s a r e l e f t o f f .	$HClO_2$	$HClO_3$

⚡ Sulfur and Phosphorus

When sulfur or phosphorus are present in acids, their root names are different than when they are present in anions. Everything else about naming acids that contain sulfur or phosphorus remains the same.

Element	Root Name in Anions	Root Name in Acids
sulfur	sulf-	sulfur-
phosphorus	phosph-	phosphor-

✅ Example 5.9.4: Naming Oxyacids

Write the name for H_2SO_4.

Solution

	H_2SO_4
1. Identify the oxyanion present in the acid.	SO_4^{2-}
2. Name the oxyanion.	sulfate
3. Drop *-ate* or *-ite* from the anion name.	sulf-
4. Replace *-ate* with *-ic acid*. Replace *-ite* with *-ous acid*. Sulfur has a root name of sulfur- in acids.	sulfuric acid

✏️ Exercise 5.9.1

If the chemical formula is provided, write the name. If the name is provided, write the formula.

A. sulfurous acid
B. HBr (*aq*)
C. acetic acid
D. H_2CO_3

Answer A

H_2SO_3

Answer B

hydrobromic acid

Answer C

$HC_2H_3O_2$

Answer D

carbonic acid

Summary

- Acids are molecular compounds that release hydrogen ions when dissolved in water.
- A binary acid consists of hydrogen and one other element.
- Oxyacids contain hydrogen, oxygen, and at least one other element.
- The name of the acid is based on the anion attached to the hydrogen.

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5.10: Nomenclature Summary

To summarize, many compounds may be classified into five different categories, each of which uses a slightly different, but related, set of rules. These five common classes of compounds are

1. Binary ionic compounds
2. Ionic compounds containing polyatomic ions
3. Binary molecular compounds
4. Binary acids
5. Oxyacids

Binary Ionic Compounds

A binary ionic compound is one that contains two elements, one of which is a metal and the other of which is a nonmetal. The metal is always written first in the name and formula followed by the nonmetal. The convention for naming binary ionic compounds is

- Metals retain their name.
- Nonmetals retain their root name followed by a suffix of *-ide*.

Most main-group elements (groups labeled with Roman numerals followed by the letter *A*) have fixed charges, while most transition metals (groups labeled with Roman numerals followed by the letter *B*) have charges that are variable.

Examples:

- SrF_2
- copper(II) nitride

Ionic Compounds Containing Polyatomic Ions

An ionic compound with polyatomic ions is one that contains more than two elements. These compounds always involve metals combined with nonmetals, unless it contains the ammonium ion, NH_4^+. Polyatomic ions encountered in this text may be found in Table 5.7.1, Table 5.7.2, and Table 5.7.3. It is generally a good practice to memorize the polyatomic ions found in Table 5.7.1.

In an ionic compound containing polyatomic ions:

- The cation name comes first.
- The anion name comes second.

The same holds true for their position in the chemical formula.

Examples:

- $Cr(NO_3)_3$
- sodium carbonate

Binary Molecular Compounds

A binary molecular compound is one that contains two elements, both of which are nonmetals and/or metalloids. Naming binary molecular compounds is similar to naming binary ionic compounds. The convention for naming binary molecular compounds is

- The first element in the formula retains its name.
- The second element retains its root name followed by a suffix of *-ide*.
- Greek prefixes (see Table 5.8.1) are added to each name* to indicate the number of atoms of each element present in the formula. *The prefix *mono-* is not used for the first element's name.

When writing formulas, prefixes tell exactly how many atoms of each element are present in one molecule.

Examples:

- dinitrogen tetrachloride
- IBr_5

Binary Acids

A **binary acid** is an acid that consists of hydrogen combined with one other element. The most common binary acids contain a halogen. The convention for naming binary acids is

hydro- + root name of the other element + *-ic acid*

Binary acids only carry their acid name when they are dissolved in water, indicated by a phase label of (*aq*). When writing a formula for a binary acid, H is always written at the front of the formula. If an acid name has a *hydro-* prefix, it will always be a binary acid.

Examples:

- hydrochloric acid
- HBr (*aq*)

Oxyacids

An **oxyacid** is an acid that also contains the element oxygen, along with a third element (or more). The third element is almost always a nonmetal. Oxyacids may also be thought of as a compound where H^+ has combined with an oxyanion. The convention for naming oxyacids:

- Identify and name the oxyanion present in the acid.
- If the oxyanion name ends with *-ate*, drop the *-ate* suffix from the anion name and replace with an *-ic acid* suffix.
- If the oxyanion name ends with *-ite*, drop the *-ite* suffix from the anion name and replace with an *-ous acid* suffix.

Oxyacids carry their acid name whether they are dissolved in water or not. If there is **not** a *hydro-* prefix in the acid name, it will always be an oxyacid.

Examples:

- H_2CO_3
- sulfurous acid

Exercise 5.10.1

If the chemical formula is provided, write the name. If the name is provided, write the formula.

A. $LiNO_3$
B. P_2S_5
C. Au_2O
D. $Mg_3(PO_4)_2$
E. H_2CO_3
F. hydrochloric acid
G. potassium sulfate
H. silver bromide
I. boron trifluoride
J. chromium(III) acetate

Answer A

lithium nitrate

Answer B

diphosphorus pentasulfide

Answer C

gold(I) oxide

Answer D

magnesium phosphate

Answer E

carbonic acid

Answer F

HCl (*aq*) or simply HCl

Answer G

K_2SO_4

Answer H

AgBr

Answer I

BF_3

Answer J

$Cr(C_2H_3O_2)_3$

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5.12: Exercises

5.2: Chemical Formulas

1. Classify each of the following as an atomic element, a molecular element, an ionic compound, or a molecular compound. State how many atoms are in each formula unit or molecule.
   a. Fe
   b. PCl_3
   c. P_4
   d. $CaBr_2$

Answer

   a. atomic element
   b. molecular compound, 4 atoms
   c. molecular element, 4 atoms
   d. ionic compound, 3 atoms

2. Classify each of the following as an atomic element, a molecular element, an ionic compound, or a molecular compound. State how many atoms are in each formula unit or molecule.
   a. I_2
   b. $Fe(NO_3)_2$
   c. H_2O
   d. Al

Answer

   a. molecular element, 2 atoms
   b. ionic compound, 9 atoms
   c. molecular compound, 3 atoms
   d. atomic element

3. What is the difference between CO and Co?

Answer

CO is a molecular compound containing carbon and oxygen. Co is the atomic element cobalt.

4. What is the difference between H_2O and H_2O_2 (hydrogen peroxide)?

Answer

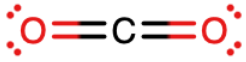
H_2O has two hydrogen atoms and one oxygen atom. H_2O_2 has two hydrogen atoms and two oxygen atoms.

5. Explain why the formula for an atom of the element oxygen and the formula for a molecule of oxygen differ.

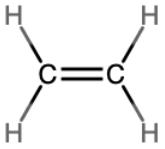
Answer

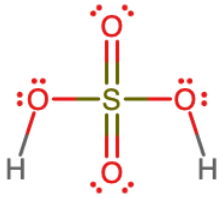
The symbol for the element oxygen, O, represents both the element and one atom of oxygen. A molecule of oxygen, O_2, contains two oxygen atoms; the subscript 2 in the formula must be used to distinguish the diatomic molecule from two single oxygen atoms.

6. Write the molecular and empirical formulas of the following compounds:

a. 

b. 

c. 

d. 

Answer

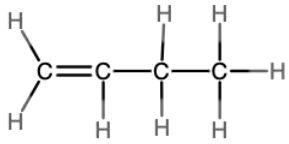
a. molecular = CO_2, empirical = CO_2

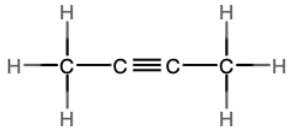
b. molecular = C_2H_2, empirical = CH

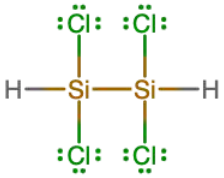
c. molecular = C_2H_4, empirical = CH_2

d. molecular = H_2SO_4, empirical = H_2SO_4

7. Write the molecular and empirical formulas of the following compounds:

a. 

b. 

c. 

d. 

Answer

a. molecular = C_4H_8; empirical = CH_2
b. molecular = C_4H_6; empirical = C_2H_3
c. molecular = $Si_2H_2Cl_4$; empirical = $SiHCl_2$
d. molecular = H_3PO_4; empirical = H_3PO_4

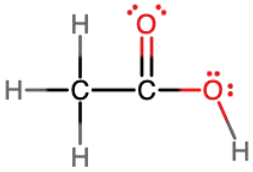
8. Determine the empirical formulas for the following compounds:

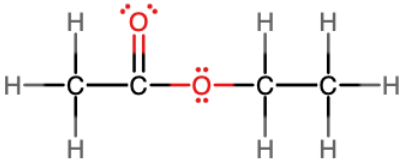
a. caffeine, $C_8H_{10}N_4O_2$
b. sucrose, $C_{12}H_{22}O_{11}$
c. hydrogen peroxide, H_2O_2
d. glucose, $C_6H_{12}O_6$
e. ascorbic acid (vitamin C), $C_6H_8O_6$

Answer

a. $C_4H_5N_2O$
b. $C_{12}H_{22}O_{11}$
c. HO
d. CH_2O
e. $C_3H_4O_3$

9. Write the empirical formulas for the following compounds:

a. 

b. 

Answer

a. CH_2O
b. C_2H_4O

10. Determine the empirical formulas for the following compounds:

a. acetic acid, $C_2H_4O_2$

b. citric acid, $C_6H_8O_7$
c. hydrazine, N_2H_4
d. nicotine, $C_{10}H_{14}N_2$
e. butane, C_4H_{10}

Answer

a. CH_2O
b. $C_6H_8O_7$
c. NH_2
d. C_5H_7N
e. C_2H_5

5.3: A Closer Look at Elements and Compounds

11. From their positions on the periodic table, will Cu and I form a molecular compound or an ionic compound?

Answer

ionic compound

12. From their positions on the periodic table, will N and S form a molecular compound or an ionic compound?

Answer

molecular compound

13. Using the periodic table, predict whether the following chlorides are ionic or covalent: KCl, NCl_3, ICl, $MgCl_2$, PCl_5, and CCl_4.

Answer

ionic = KCl, $MgCl_2$; covalent = NCl_3, ICl, PCl_5, CCl_4

14. Using the periodic table, predict whether the following chlorides are ionic or covalent: $SiCl_4$, PCl_3, $CaCl_2$, CsCl, $CuCl_2$, and $CrCl_3$.

Answer

ionic = $CaCl_2$, CsCl, $CuCl_2$, $CrCl_3$; covalent = $SiCl_4$, PCl_3

5.4: Chemical Nomenclature and the Classification of Compounds

15. Classify each of the following compounds as binary ionic, ionic containing polyatomic ions, binary molecular, binary acid, or oxyacid.

a. NF_3
b. HI
c. $(NH_4)_2CO_3$
d. $Sr_3(PO_4)_2$
e. BaO
f. H_2SO_4

Answer

a. binary molecular
b. binary acid
c. ionic containing polyatomic ions
d. ionic containing polyatomic ions
e. binary ionic
f. oxyacid

16. Classify each of the following compounds as binary ionic, ionic containing polyatomic ions, binary molecular, binary acid, or oxyacid.

a. $HClO_3$
b. $Mg(C_2H_3O_2)_2$
c. H_2S
d. Ag_2S
e. N_2Cl_4
f. $CoBr_2$

Answer

a. oxyacid
b. ionic containing polyatomic ions
c. binary acid
d. binary ionic
e. binary molecular
f. binary ionic

5.5: Binary Ionic Compounds

17. Give the formula and name for each ionic compound formed between the two listed ions.

a. Mg^{2+} and Cl^-
b. Na^+ and O^{2-}
c. Cd^{2+} and O^{2-}

Answer

a. $MgCl_2$, magnesium chloride
b. Na_2O, sodium oxide
c. CdO, cadmium oxide

18. Give the formula and name for each ionic compound formed between the two listed ions.

a. K^+ and S^{2-}
b. Ag^+ and Br^-
c. Sr^{2+} and N^{3-}

Answer

a. K_2S, potassium sulfide
b. AgBr, silver bromide
c. Sr_3N_2, strontium nitride

19. Give the formula and name for each ionic compound formed between the two listed ions.

a. Zn^{2+} and F^{-}
b. Ca^{2+} and O^{2-}
c. Li^{+} and P^{3-}

Answer

a. ZnF_2, zinc fluoride
b. CaO, calcium oxide
c. Li_3P, lithium phosphide

20. Give the formula and name for each ionic compound formed between the two listed ions.

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

Answer

a. Na_3N, sodium nitride
b. SrI_2, strontium iodide
c. Al_2S_3, aluminum sulfide

5.6: Ions With Variable Charges

21. Give the formula and name for each ionic compound formed between the two listed ions.

a. Ni^{2+} and Cl^{-}
b. Fe^{2+} and O^{2-}
c. Fe^{3+} and O^{2-}

Answer

a. $NiCl_2$, nickel(II) chloride
b. FeO, iron(II) oxide
c. Fe_2O_3, iron(III) oxide

22. Give the formula and name for each ionic compound formed between the two listed ions.

a. Cu^{+} and S^{2-}
b. Pt^{2+} and Br^{-}
c. Cr^{2+} and N^{3-}

Answer

a. Cu_2S, copper(I) sulfide
b. $PtBr_2$, platinum(II) bromide
c. Cr_3N_2, chromium(II) nitride

23. Give the formula and name for each ionic compound formed between the two listed ions.

a. Cu^{2+} and F^{-}
b. U^{6+} and O^{2-}
c. Au^{+} and P^{3-}

Answer

a. CuF_2, copper(II) fluoride
b. UO_3, uranium(VI) oxide
c. Au_3P, gold(I) phosphide

24. Give the formula and name for each ionic compound formed between the two listed ions.

a. Sn^{2+} and N^{3-}
b. Co^{2+} and I^-
c. Au^{3+} and S^{2-}

Answer

a. Sn_3N_2, tin(II) nitride
b. CoI_2, cobalt(II) iodide
c. Au_2S_3, gold(III) sulfide

25. Name each of the following compounds:

a. Cr_2O_3
b. $FeCl_2$
c. CrO_3
d. $TiCl_4$
e. CoO
f. SnS_2

Answer

a. chromium(III) oxide
b. iron(II) chloride
c. chromium(VI) oxide
d. titanium(IV) chloride
e. cobalt(II) oxide
f. tin(IV) sulfide

26. Name each of the following compounds:

a. $NiBr_2$
b. WO_3
c. CoS
d. V_2O_5
e. MnO_2
f. Fe_2O_3

Answer

a. nickel(II) bromide
b. tungsten(VI) oxide
c. cobalt(II) sulfide
d. vanadium(V) oxide
e. manganese(IV) oxide
f. iron(III) oxide

5.7: Ionic Compounds Containing Polyatomic Ions

27. Give the formula and name for each ionic compound formed between the two listed ions.
    a. K^+ and SO_4^{2-}
    b. NH_4^+ and S^{2-}
    c. NH_4^+ and PO_4^{3-}

Answer

    a. K_2SO_4, potassium sulfate
    b. $(NH_4)_2S$, ammonium sulfide
    c. $(NH_4)_3PO_4$, ammonium phosphate

28. Give the formula and name for each ionic compound formed between the two listed ions.
    a. Ca^{2+} and NO_3^-
    b. Ca^{2+} and NO_2^-
    c. W^{3+} and $C_2H_3O_2^-$

Answer

    a. $Ca(NO_3)_2$, calcium nitrate
    b. $Ca(NO_2)_2$, calcium nitrite
    c. $W(C_2H_3O_2)_3$, tungsten(III) acetate

29. Give the formula and name for each ionic compound formed between the two listed ions.
    a. Pb^{4+} and SO_4^{2-}
    b. Na^+ and NO_3^-
    c. Li^+ and CO_3^{2-}

Answer

    a. $Pb(SO_4)_2$, lead(IV) sulfate
    b. $NaNO_3$, sodium nitrate
    c. Li_2CO_3, lithium carbonate

30. Give the formula and name for each ionic compound formed between the two listed ions.
    a. NH_4^+ and N^{3-}
    b. Mg^{2+} and CN^-
    c. Al^{3+} and OH^-

Answer

    a. $(NH_4)_3N$, ammonium nitride
    b. $Mg(CN)_2$, magnesium cyanide
    c. $Al(OH)_3$, aluminum hydroxide

31. Give the formula and name for each ionic compound formed between the two listed ions.

a. Ag^{+} and SO_3^{2-}
b. Na^{+} and HCO_3^{-}
c. Fe^{3+} and ClO_3^{-}

Answer

a. Ag_2SO_3, silver sulfite
b. $NaHCO_3$, sodium bicarbonate or sodium hydrogen carbonate
c. $Fe(ClO_3)_3$, iron(III) chlorate

32. Give the formula and name for each ionic compound formed between the two listed ions.

a. Sn^{2+} and $C_2H_3O_2^{-}$
b. Au^{3+} and SO_3^{2-}
c. Sr^{2+} and NO_2^{-}

Answer

a. $Sn(C_2H_3O_2)_2$, tin(II) acetate
b. $Au_2(SO_3)_3$, gold(III) sulfite
c. $Sr(NO_2)_2$, strontium nitrite

33. The uranyl cation has the formula UO_2^{2+}. Propose formulas and names for the ionic compounds between the uranyl cation and F^{-}, SO_4^{2-}, and PO_4^{3-}.

Answer

UO_2F_2, uranyl fluoride; UO_2SO_4, uranyl sulfate; $(UO_2)_3(PO_4)_2$, uranyl phosphate

34. The permanganate anion has the formula MnO_4^{-}. Propose formulas and names for the ionic compounds between the permanganate ion and K^{+}, Ca^{2+}, and Fe^{3+}.

Answer

$KMnO_4$, potassium permangante; $Ca(MnO_4)_2$, calcium permanganate; $Fe(MnO_4)_3$, iron(III) permanganate

35. For each of the following pairs of ions, write the formula of the compound they will form:

a. Ca^{2+}, S^{2-}
b. $\mathbf{NH_4^{+}}$, $\mathbf{SO_4^{2-}}$
c. Al^{3+}, Br^{-}
d. Na^{+}, $\mathbf{SO_4^{2-}}$
e. Mg^{2+}, $\mathbf{PO_4^{3-}}$

Answer

a. CaS
b. $(NH_4)_2SO_4$
c. $AlBr_3$
d. Na_2SO_4
e. $Mg_3(PO_4)_2$

36. For each of the following pairs of ions, write the formula of the compound they will form:

a. K^+, O^{2-}
b. NH_4^+, PO_4^{3-}
c. Al^{3+}, O^{2-}
d. Li^+, CO_3^{2-}
e. Ba^{2+}, PO_4^{3-}

Answer

a. K_2O
b. $(NH_4)_3PO_4$
c. Al_2O_3
d. Li_2CO_3
e. $Ba_3(PO_4)_2$

37. Write the formulas for each of the following compounds:

a. potassium phosphate
b. copper(II) sulfate
c. calcium chloride
d. titanium(IV) oxide
e. ammonium nitrate

Answer

a. K_3PO_4
b. $CuSO_4$
c. $CaCl_2$
d. TiO_2
e. NH_4NO_3

38. Name each of the following compounds:

a. $Ca(ClO_3)_2$
b. $FeSO_4$
c. $CaCO_3$
d. $AgNO_2$
e. $Ni(CN)_3$

Answer

a. calcium chlorate
b. iron(II) sulfate
c. calcium carbonate
d. silver nitrite
e. nickel(III) cyanide

5.8: Binary Molecular Compounds

39. What is the stem of fluorine used in molecule names? CF_4 is one example.

Answer

fluor–

40. What is the stem of bromine used in molecule names? $SiBr_4$ is an example.

Answer

brom–

41. Give the proper name for each molecule.

a. PF_3
b. SCl_2
c. N_2O_3

Answer

a. phosphorus trifluoride
b. sulfur dichloride
c. dinitrogen trioxide

42. Give the proper name for each molecule.

a. NO
b. CS_2
c. As_2O_3

Answer

a. nitrogen monoxide
b. carbon disulfide
c. diarsenic trioxide

43. Give the proper name for each molecule.

a. XeF_2
b. O_2F_2
c. SF_6

Answer

a. xenon difluoride
b. dioxygen difluoride
c. sulfur hexafluoride

44. Give the proper name for each molecule.

a. P_4O_{10}
b. B_2O_3
c. P_2S_3

Answer

a. tetraphosphorus decoxide
b. diboron trioxide
c. diphosphorus trisulfide

45. Give the proper name for each molecule.

a. N_2O
b. N_2O_4
c. N_2O_5

Answer

a. dinitrogen monoxide
b. dinitrogen tetroxide
c. dinitrogen pentoxide

46. Give the proper name for each molecule.

a. SO_2
b. Cl_2O
c. XeF_6

Answer

a. sulfur dioxide
b. dichlorine monoxide
c. xenon hexafluoride

47. Give the proper formula for each name.

a. dinitrogen pentoxide
b. tetraboron tricarbide
c. phosphorus pentachloride

Answer

a. N_2O_5
b. B_4C_3
c. PCl_5

48. Give the proper formula for each name.

a. nitrogen triiodide
b. diarsenic trisulfide
c. iodine trichloride

Answer

a. NI_3
b. As_2S_3
c. ICl_3

49. Give the proper formula for each name.

a. dioxygen difluoride
b. dinitrogen trisulfide
c. xenon tetrafluoride

Answer

a. O_2F_2
b. N_2S_3
c. XeF_4

50. Give the proper formula for each name.

a. chlorine dioxide
b. sulfur dibromide
c. dinitrogen trioxide

Answer

a. ClO_2
b. SBr_2
c. N_2O_3

51. Give the proper formula for each name.

a. iodine trifluoride
b. xenon trioxide
c. disulfur decafluoride

Answer

a. IF_3
b. XeO_3
c. S_2F_{10}

52. Give the proper formula for each name.

a. silicon dioxide
b. carbon disulfide
c. disulfur dibromide

Answer

a. SiO_2
b. CS_2
c. S_2Br_2

5.9: Acids

53. Give the formula for each acid.

a. chloric acid
b. hydroiodic acid

Answer

a. $HClO_3$
b. HI

54. Give the formula for each acid.

a. hydrosulfuric acid
b. phosphorous acid

Answer

a. H_2S
b. H_3PO_3

55. Name each acid.

a. HF(aq)
b. HNO_3(aq)
c. $HC_2H_3O_2$(aq)

Answer

a. hydrofluoric acid
b. nitric acid
c. acetic acid

56. Name each acid.

a. H_2SO_4(aq)
b. H_3PO_4(aq)
c. HCl(aq)

Answer

a. sulfuric acid
b. phosphoric acid
c. hydrochloric acid

5.10: Nomenclature Summary

57. Name the following compounds:

a. $HClO_2$
b. CuO
c. K_2SO_4
d. $BrCl_3$
e. HBr
f. AlF_3

Answer

a. chlorous acid
b. copper(II) oxide
c. potassium sulfate
d. bromine trichloride
e. hydrobromic acid
f. aluminum fluoride

58. Name the following compounds:

a. NaF
b. $(NH_4)_2O$
c. $MnCl_3$
d. H_2S
e. P_4O_6
f. HNO_3

Answer

a. sodium fluoride
b. ammonium oxide
c. manganese(III) chloride
d. hydrosulfuric acid
e. tetraphosphorus hexoxide
f. nitric acid

59. Write the formulas of the following compounds:

a. strontium bromide
b. acetic acid
c. lead(II) oxide
d. gold(III) hydroxide
e. hydrofluoric acid
f. diphosphorus pentasulfide
g. aluminum sulfite
h. krypton tetrafluoride

Answer

a. $SrBr_2$
b. $HC_2H_3O_2$
c. PbO
d. $Au(OH)_3$
e. HF
f. P_2S_5
g. $Al_2(SO_3)_3$
h. KrF_4

60. Write the formulas of the following compounds:

a. lithium carbonate
b. silver sulfide
c. molybdenum(III) nitride
d. diboron hexahydride
e. sulfuric acid
f. trisulfur heptachloride
g. manganese(IV) chlorate
h. hydrobromic acid

Answer

a. Li_2CO_3

b. Ag_2S
c. MoN
d. B_2H_6
e. H_2SO_4
f. S_3Cl_7
g. $Mn(ClO_3)_4$
h. HBr

61. Write the formulas of the following compounds:

a. chlorine dioxide
b. carbonic acid
c. potassium phosphide
d. nickel(III) nitrite
e. aluminum chlorite
f. carbon disulfide

Answer

a. ClO_2
b. H_2CO_3
c. K_3P
d. $Ni(NO_2)_3$
e. $Al(ClO_2)_3$
f. CS_2

62. Write the formulas of the following compounds:

a. barium chloride
b. xenon dioxide
c. iron(III) nitrate
d. acetic acid
e. dinitrogen trioxide
f. tin(IV) chloride

Answer

a. $BaCl_2$
b. XeO_2
c. $Fe(NO_3)_3$
d. $HC_2H_3O_2$
e. N_2O_3
f. $SnCl_4$

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CHAPTER OVERVIEW

6: Chemical Composition

6.1: The Mole
6.2: Molecular Mass and Formula Mass
6.3: Counting Objects By Weighing
6.4: Molar Mass
6.5: Mole Calculations
6.6: Counting Particles By Weighing
6.7: Using Chemical Formulas as Conversion Factors
6.8: Percent-By-Mass Composition
6.9: Mass Percent Composition from a Chemical Formula
6.10: For Future Use
6.11: Exercises

Template:HideTOC

Thumbnail Chapter 6: A mole and Avogadro's number (adaptation by ChemLancer from NobleDame via tenor).

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6.1: The Mole

When objects are small or come in larger quantities, it is often inconvenient, inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with small objects or large quantities in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like donuts or pencils. Other items such as bottle rockets – a type of fireworks – are sold in packages of 144, which is called a "gross", another name for a dozen dozen ($12^2 = 144$).



Figure **6.1.1**: One dozen donuts. (Racool; FreeRange Stock Photos)

Avogadro's Number and the Mole

Chemists use a quantity called the **mole** to count items. One mole of any item contains 602,200,000,000,000,000,000,000, or 6.022×10^{23}, of those items. When writing numbers in scientific notation, or when seeing a bunch of zeroes at the end of a number, it is easy to lose the context of a number's sheer size.

For example, one mole of eggs (6.022×10^{23} eggs), the kind purchased at a grocery store, would be enough to cover the entire planet 60 miles deep with eggs! Or, perhaps you prefer to think of quantities in terms of cold, hard cash. According to the United States Census Bureau Population Clock, there were over 334,630,000 people living in the U.S. in April 2023. With one mole dollars ($\$6.022\times10^{23}$) available, it would be possible to give every person living in the U.S. $1,000,000 per *every single second* of their lives for a duration of 57 years and 1 month! Indeed, this number is quite large!

What could possibly be measured with such an enormous number? If you said atoms or molecules, you would be correct! Individual atoms and molecules are too small to be seen, counted, or weighed on their own. However, if one were to gather together enough atoms or molecules, they may be weighed and even counted.

The **mole**, abbreviated as **mol**, is the base SI unit for the amount of substance, defined as exactly 6.02214076×10^{23} particles (atoms, molecules, ions, or electrons). The current definition was adopted in November 2018, revising its old definition based on the number of atoms in 12 grams of carbon-12 ($^{12}_{6}\mathrm{C}$). For most purposes, 6.022×10^{23} provides an adequate number of significant figures. Just as 1 mole of atoms contains 6.022×10^{23} atoms, 1 mole of eggs contains 6.022×10^{23} eggs.

6.022×10^{23} is also called **Avogadro's number**, after the 19th-century Italian scientist who first proposed a relationship between the volumes of gases and the numbers of particles they contain. There is a particular reason that this number was chosen and this reason will become clear as we proceed.

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6.2: Molecular Mass and Formula Mass

Learning Objectives

- Define formula mass and molecular mass.
- To determine the formula mass of an ionic compound or molecular mass of a molecular compound.

Why is knowledge of composition important? Everything in nature is either chemically or physically combined with other substances. To find the amount of a material in a sample, you need to know what fraction of the sample it is. Some simple applications of composition are: the amount of sodium in sodium chloride for a diet, the amount of iron in iron ore for steel production, the amount of hydrogen in water for hydrogen fuel, and the amount of chlorine in freon to estimate ozone depletion.



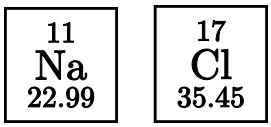
Figure 6.2.1: How much sodium is in this spoonful of salt? (via Flickr; Gianni pepenera).

Molecular Mass and Formula Mass

The composition of a compound may be determined from its chemical formula and the atomic masses of the elements that make up the compound. The mass of a molecule is called its **molecular mass**. The molecular mass is calculated summing the masses of all the atoms shown in the chemical formula.

Ionic compounds are not comprised of discrete molecules, but rather, ions. The smallest unit of an ionic compound that retains the identity of an ionic compound is called a **formula unit**. The mass of a formula unit is called its **formula mass**. The formula mass is calculated summing the masses of all the atoms shown in the chemical formula. In other words, the molecular mass and formula mass are both calculated in an identical manner, by summing the masses of all the atoms shown in the chemical formula.

As you may recall, the periodic table shows the average mass of the atoms for each element. We called this mass the atomic mass. The atomic mass of sodium is 22.99 amu and the atomic mass of chlorine is 35.45 amu.



Since NaCl is an ionic compound, we could calculate its formula mass. 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:

1 Na: 22.99 amu

1 Cl: +35.45 amu

Total: 58.44 amu

In other words, the formula mass of NaCl is 58.44 amu. We may also say that one formula unit of NaCl has a mass of 58.44 amu.

✅ Example 6.2.1

Calculate the molecular mass or formula mass for each compound. (Periodic Table)

A. $C_{12}H_{22}O_{11}$
B. $Ca(NO_3)_2$

Solution

	$C_{12}H_{22}O_{11}$	$Ca(NO_3)_2$
1. Determine how many atoms of each element are present in the compound. Multiply the number of atoms of each element by its atomic mass.	12 C = 12(12.01) = 144.1 amu 22 H = 22(1.008) = 22.18 amu 11 O = 11(16.00) = 176.0 amu	1 Ca = 1(40.08) = 40.08 amu 2 N = 2(14.01) = 28.02 amu 6 O = 6(16.00) = 96.00 amu
2. Sum the masses to find the molecular mass or formula mass.	molecular mass = 342.3 amu	formula mass = 164.10 amu
3. Explanation.	It is called a molecular mass since $C_{12}H_{22}O_{11}$ is a molecular compound.	It is called a formula mass since $Ca(NO_3)_2$ is an ionic compound.
4. Interpretation.	One molecule of $C_{12}H_{22}O_{11}$ has a mass of 342.3 amu.	One formula unit of $Ca(NO_3)_2$ has a mass of 164.10 amu.

✏️ Exercise 6.2.1

Calculate the molecular mass or formula mass for each compound. (Periodic Table)

A. TiO_2
B. $C_6H_{12}O_6$
C. $Fe_3(PO_4)_2$

Answer A

The formula mass is 79.90 amu.

Answer B

The molecular mass is 180.16 amu.

Answer C

The formula mass is 357.5 amu.

Summary

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

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6.3: Counting Objects By Weighing

The sizes of atoms and molecules are so small that it is physically difficult, if not impossible, to directly count them out (Figure 6.3.1). However, atoms and molecules may be counted indirectly by using a common trick of "counting by weighing."

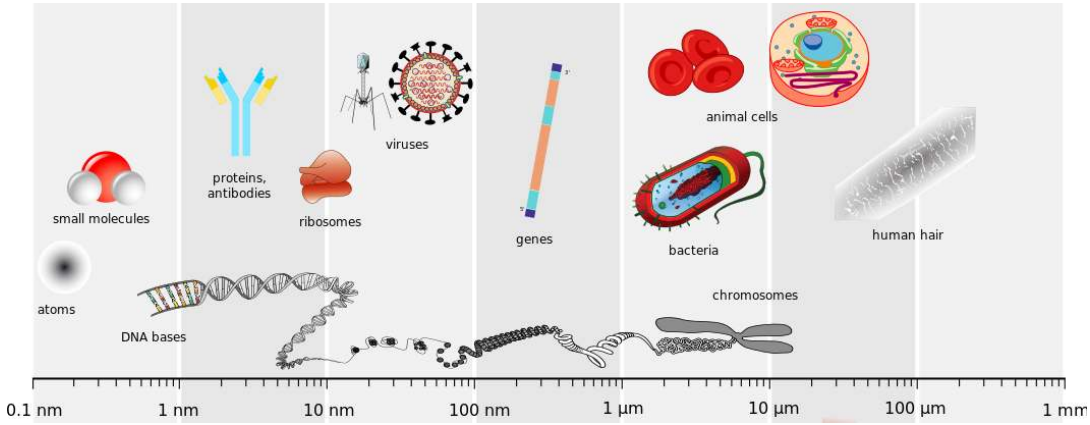


Figure **6.3.1**: A comparison of scale. (CC BY-SA 3.0; Wikipedia)

Counting by Weighing

Consider the example of counting nails in a large box at a hardware store. You need to estimate the number of nails in a box. The weight of an empty box is 113 g and the weight of the box plus a bunch of big nails is 1340 g. Assume that we know that the weight of one nail is 0.450 g. Hopefully it's not necessary to tear open the package and count the nails. We agree that

$$\textbf{mass of nails} = 1340\,\text{g} - 113\,\text{g} = 1227\,\text{g}$$

Therefore

$$\textbf{nails in box} = \frac{1227\,\text{g}}{0.450\ \text{g/nail}} = 2,726.7\,\textbf{nails} = \boxed{2,730\,\textbf{nails}}$$

You have just counted the number of nails in the box by weighing them (rather than by counting them individually).

Exercise 6.3.1

Suppose you wanted to know how many pennies dated between 1983 and today were in your coin collection. To save time, you placed them onto a balance with a plastic pan that had already been tared. If it is known that all pennies minted since 1983 have a mass of 2.5 g, how many pennies dated between 1983 and today are shown on the balance below if they have a combined mass of 124.9 g?



Answer

50 pennies

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6.4: Molar Mass

Learning Objectives

- Calculate the molar mass of an element or compound.
- Determine the number of atoms, formula units, or molecules in one mole of a substance.

Molar Mass of an Element

According to the periodic table, the atomic mass of aluminum is 26.98 amu, copper is 63.55 amu, and carbon is 12.01 amu. Since 1 amu is only 1.674×10^{-24} g, these masses would be way too small to measure on ordinary laboratory equipment. However, if we have 6.022×10^{23} atoms, or 1 mole of each of these elements, we would have 26.98 g Al, 63.55 g Cu, and 12.01 g C. In other words, one mole of any element will have a mass that is numerically equal to its atomic mass and expressed in units of grams. This mass is called the **molar mass**.

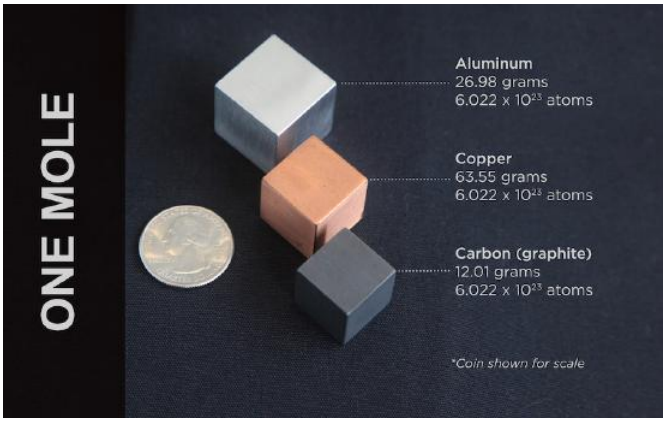


Figure **6.4.1**: One mole of aluminum, copper, and carbon. (R. Press/N. Hanacek/NIST)

We might find these easier to use if the relationships are written out like this (recall that mole is abbreviated as mol):

- 1 mol Al = 26.98 g Al = 6.022×10^{23} atoms Al
- 1 mol Cu = 63.55 g Cu = 6.022×10^{23} atoms Cu
- 1 mol C = 12.01 g C = 6.022×10^{23} atoms C

Molar Mass of a Compound

One mole of any compound will have a mass that is numerically equal to its molecular mass or formula mass and expressed in units of grams. This mass is also called the **molar mass**.

In Section 6.2, the formula mass for NaCl was calculated to be 58.44 amu. This means that the molar mass of NaCl is 58.44 g and it will contain 6.022×10^{23} formula units of NaCl, or one mole of NaCl. We also calculated the molecular mass for $C_{12}H_{22}O_{11}$ to be 342.3 amu. This means that the molar mass of $C_{12}H_{22}O_{11}$ is 342.3 g and it will contain 6.022×10^{23} molecules of $C_{12}H_{22}O_{11}$, or one mole of $C_{12}H_{22}O_{11}$.

We might find these easier to use if the relationships are written out like this (recall that mole is abbreviated as mol):

- 1 mol NaCl = 58.44 g NaCl = 6.022×10^{23} formula units NaCl
- 1 mol $C_{12}H_{22}O_{11}$ = 342.3 g $C_{12}H_{22}O_{11}$ = 6.022×10^{23} molecules $C_{12}H_{22}O_{11}$



Figure **6.4.2**: One mole table salt (NaCl) and sugar ($C_{12}H_{22}O_{11}$). (Lance S. Lund)

Summary of Relevant Terminology

Keeping the terminology straight (see Table 6.4.1) is essential to understanding the mole concept and correctly performing mole calculations.

Table 6.4.1: Relevant Terminology

Substance	Fundamental Unit	Atomic, Molecular, or Formula Mass	Molar Mass
aluminum (Al)	atom	atomic mass = 26.98 amu	molar mass = 26.98 g
sucrose ($C_{12}H_{22}O_{11}$)	molecule	molecular mass = 342.3 amu	molar mass = 342.3 g
sodium chloride (NaCl)	formula unit	formula mass = 58.44 amu	molar mass = 58.44 g

✅ Example 6.4.1

Find the molar mass of each element. Then determine how many atoms are contained in the calculated molar mass. (Periodic Table)

A. K
B. Pb

Solution

A. Since the atomic mass of K is 39.10 amu, it has a molar mass of 39.10 g. There are 6.022×10^{23} atoms K, which is also 1 mole of K, in 39.10 g K.
B. Since the atomic mass of Pb is 207.2 amu, it has a molar mass of 207.2 g. There are 6.022×10^{23} atoms Pb, which is also 1 mole of Pb, in 207.2 g Pb.

✅ Example 6.4.2

Calculate the molar mass for each compound. Then determine how many formula units or molecules are contained in the calculated molar mass. (Periodic Table)

A. H_2O

B. $Ca(NO_3)_2$

Solution

	H_2O	$Ca(NO_3)_2$
1. Determine how many atoms of each element are present in the compound. Multiply the number of atoms of each element by its atomic mass and express in units of grams.	2 H = 2(1.008) = 2.016 g H 1 O = 1(16.00) = 16.00 g O	1 Ca = 1(40.08) = 40.08 g Ca 2 N = 2(14.01) = 28.02 g N 6 O = 6(16.00) = 96.00 g O
2. Sum the masses to find the molar mass.	molar mass = 18.02 g	molar mass = 164.10 g
3. Determine whether the compound is ionic or molecular.	H_2O is a molecular compound.	$Ca(NO_3)_2$ is an ionic compound.
4. Formula units or molecules contained in the molar mass.	6.022×10^{23} molecules H_2O, which is also 1 mol H_2O, are in 18.02 g H_2O.	6.022×10^{23} formula units $Ca(NO_3)_2$, which is also 1 mol $Ca(NO_3)_2$, are in 164.10 g $Ca(NO_3)_2$.

Exercise 6.4.1

Calculate the molar mass for each substance. Then show the number of grams in 1 mol of the substance, as well as the number of atoms, molecules, or formula units that are contained 1 mol of the substance. (Periodic Table)

A. Na
B. Na_2S
C. N_2O

Answer A

molar mass = 22.99 g; Therefore, 1 mol Na = 22.99 g Na = 6.022×10^{23} atoms Na.

Answer B

molar mass = 78.05 g; Therefore, 1 mol Na_2S = 78.05 g Na_2S = 6.022×10^{23} formula units Na_2S.

Answer C

molar mass = 44.02 g; Therefore, 1 mol N_2O = 44.02 g N_2O = 6.022×10^{23} molecules N_2O.

Summary

- A mole is defined as exactly $6.02214076 \times 10^{23}$ particles, e.g., atoms, molecules, ions or electrons.
- There are $6.02214076 \times 10^{23}$ particles in 1 mole. This number is called Avogadro's number.
- The molar mass of an element can be found by referring to the atomic mass on a periodic table with units of grams.
- The molar mass of compounds can be determined by the molar masses of the atoms in their formulas.

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6.5: Mole Calculations

Learning Objectives

- Perform conversions between mass and moles of a compound.

In the previous section, several relationships were written, including:

- 1 mol Al = 26.98 g Al = 6.022×10^{23} atoms Al
- 1 mol $C_{12}H_{22}O_{11}$ = 342.3 g $C_{12}H_{22}O_{11}$ = 6.022×10^{23} molecules $C_{12}H_{22}O_{11}$
- 1 mol NaCl = 58.44 g NaCl = 6.022×10^{23} formula units NaCl

These relationships may be used to convert from grams to moles or vice versa; or from moles to atoms, molecules, or formula units or vice versa. In the next section, we will show how these relationships may also be used to count atoms, molecules, or formula units by weighing.

Conversion Factors

The above relationships allow for a number of possible conversions. Let's start with aluminum, since it provides the simplest conversion. Here are the possible conversions:

mol Al ⇄ g Al	mol Al ⇄ atoms Al	g Al ⇄ atoms Al
$\frac{26.98\ \text{g Al}}{1\ \text{mol Al}}$ or $\frac{1\ \text{mol Al}}{26.98\ \text{g Al}}$	$\frac{6.022\times10^{23}\ \text{atoms Al}}{1\ \text{mol Al}}$ or $\frac{1\ \text{mol Al}}{6.022\times10^{23}\ \text{atoms Al}}$	$\frac{6.022\times10^{23}\ \text{atoms Al}}{26.98\ \text{g Al}}$ or $\frac{26.98\ \text{g Al}}{6.022\times10^{23}\ \text{atoms Al}}$

If converting from moles of aluminum to grams of aluminum, a conversion that places grams of aluminum in the numerator and moles of aluminum in the denominator would be used to enure the proper cancellation of units:

$$\text{mol Al} \xrightarrow{\frac{26.98\ \text{g Al}}{1\ \text{mol Al}}} \text{g Al}$$

This means that the mass of aluminum in 8.36 mol Al may be calculated like this:

$$8.36\ \cancel{\text{mol Al}} \times \frac{26.98\ \text{g Al}}{1\ \cancel{\text{mol Al}}} = \boxed{226\ \text{g Al}}$$

✅ Example 6.5.1

Calculate the mass of 0.220 mol $Cu(NO_3)_2$.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 0.220 mol $Cu(NO_3)_2$ Find: g $Cu(NO_3)_2$
List known relationship(s).	1 mol $Cu(NO_3)_2$ = 187.57 g $Cu(NO_3)_2$ = 6.022×10^{23} formula units $Cu(NO_3)_2$
Prepare a concept map and use the proper conversion factor.	$\text{mol Cu(NO}_3\text{)}_2 \xrightarrow{\frac{187.57\ \text{g Cu(NO}_3\text{)}_2}{1\ \text{mol Cu(NO}_3\text{)}_2}} \text{g Cu(}$
Calculate the answer.	$0.220\ \cancel{\text{mol Cu(NO}_3\text{)}_2} \times \frac{187.57\ \text{g Cu(NO}_3\text{)}_2}{1\ \cancel{\text{mol Cu(NO}_3\text{)}_2}} = \boxed{41.3\ \text{g Cu(NO}_3\text{)}_2}$
Think about your result.	0.220 mol is a little less than ¼ of a mole. Since 1 mole has a mass of 187.57 g, the answer should be a little less than ¼ of this amount. The answer is rounded to three significant figures since the fewest significant figures in the calculation is three.

✅ Example 6.5.2

How many moles are in 0.0114 g Hg?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 0.0114 g Hg Find: mol Hg
List known relationship(s).	1 mol Hg = 200.59 g Hg = 6.022×10^{23} atoms Hg
Prepare a concept map and use the proper conversion factor.	$\text{g Hg} \xrightarrow{\frac{1\ \text{mol Hg}}{200.59\ \text{g Hg}}} \text{mol Hg}$
Calculate the answer.	$0.0114\ \cancel{\text{g Hg}} \times \frac{1\ \text{mol Hg}}{200.59\ \cancel{\text{g Hg}}} = \boxed{5.68 \times 10^{-5}\ \text{mol Hg}}$
Think about your result.	Since there are 200.59 g Hg in 1 mol Hg, 0.0114 g Hg must be a very small fraction of a mole. The answer is rounded to three significant figures since the fewest significant figures in the calculation is three.

✏️ Exercise 6.5.1

A. What is the mass of 0.27 mol Ne?
B. How many grams are in 386 mol $Fe_3(PO_4)_2$?
C. How many moles are present in 7.0 g of NH_3?

D. Calculate the number of moles in 555 g $CaCO_3$.

Answer A

5.4 g Ne

Answer B

1.38×10^5 g $Fe_3(PO_4)_2$

Answer C

0.41 mol NH_3

Answer D

5.55 mol $CaCO_3$

Summary

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

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6.6: Counting Particles By Weighing

The following relationships have previously been established:

- 1 mol Al = 26.98 g Al = 6.022×10^{23} atoms Al
- 1 mol $C_{12}H_{22}O_{11}$ = 342.3 g $C_{12}H_{22}O_{11}$ = 6.022×10^{23} molecules $C_{12}H_{22}O_{11}$
- 1 mol NaCl = 58.44 g NaCl = 6.022×10^{23} formula units NaCl

These relationships may be used to count atoms, molecules, or formula units by weighing just like we counted pennies by weighing.

Counting Atoms, Molecules, or Formula Units By Weighing

In the case of aluminum, here are the possible conversions:

mol Al ⇄ g Al	mol Al ⇄ atoms Al	g Al ⇄ atoms Al
$\frac{26.98\ \text{g Al}}{1\ \text{mol Al}}$ or $\frac{1\ \text{mol Al}}{26.98\ \text{g Al}}$	$\frac{6.022\times10^{23}\ \text{atoms Al}}{1\ \text{mol Al}}$ or $\frac{1\ \text{mol Al}}{6.022\times10^{23}\ \text{atoms Al}}$	$\frac{6.022\times10^{23}\ \text{atoms Al}}{26.98\ \text{g Al}}$ or $\frac{26.98\ \text{g Al}}{6.022\times10^{23}\ \text{atoms Al}}$

When converting from grams of aluminum to atoms of aluminum, a conversion that places atoms of aluminum in the numerator and grams of aluminum in the denominator would be used to enure the proper cancellation of units:

$$\text{g Al} \xrightarrow{\frac{6.022\times10^{23}\ \text{atoms Al}}{26.98\ \text{g Al}}} \text{atoms Al}$$

How many atoms are required to make the average empty aluminum beverage can that has a mass of 14.9 g?

$$14.9\ \cancel{\text{g Al}} \times \frac{6.022 \times 10^{23}\text{atoms Al}}{26.98\ \cancel{\text{g Al}}} = \boxed{3.33 \times 10^{23}\ \text{atoms Al}}$$

This calculation shows how it is possible to to count atoms by weighing using the definitions for molar mass and Avogadro's number.

✅ Example 6.6.1

If a glass contains 9.64×10^{24} molecules H_2O, how many moles of water are in the glass?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 9.64×10^{24} molecules H_2O Find: mol H_2O
List known relationship(s).	1 mol H_2O = 18.02 g H_2O = 6.022×10^{23} molecules H_2O
Prepare a concept map and use the proper conversion factor.	$\boxed{\text{molecules } H_2O} \xrightarrow[6.022\times10^{23}\ \text{molecules } H_2O]{1\ \text{mol } H_2O} \boxed{\text{mol } H_2O}$
Calculate the answer.	$9.64 \times 10^{24}\ \cancel{\text{molecules } H_2O} \times \frac{1\ \text{mol } H_2O}{6.022 \times 10^{23}\ \cancel{\text{molecules } H_2O}} = \boxed{16.0\ \text{mol } H_2O}$
Think about your result.	The given number of water molecules was greater than Avogadro's number, so the number of moles of water molecules must be greater than 1 mole. The answer is rounded to three significant figures since the fewest significant figures in the calculation is three.

✅ Example 6.6.2

⅛ tsp of table salt (NaCl) has a mass of 0.69 g. How many formula units of NaCl are in this mass of salt?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 0.69 g NaCl Find: formula units NaCl
List known relationship(s).	1 mol NaCl = 58.44 g NaCl = 6.022×10^{23} formula units NaCl
Prepare a concept map and use the proper conversion factor.	g NaCl $\xrightarrow{\frac{6.022\times10^{23}\ \text{formula units NaCl}}{58.44\ \text{g NaCl}}}$ formula units
Calculate the answer.	$0.69\ \text{g NaCl} \times \frac{6.022 \times 10^{23}\ \text{formula units NaCl}}{58.44\ \text{g NaCl}}$ $= \boxed{7.1 \times 10^{21}\ \text{formula units NaCl}}$
Think about your result.	There should be a huge number of formula units of NaCl. However, 0.69 g NaCl is quite a bit less than the mass of one mole NaCl (58.44 g), so there should be less than Avogadro's number of formula units. The answer is rounded to two significant figures since the fewest significant figures in the calculation is two.

Exercise 6.6.2

A. What is the mass of one molecule of water, expressed in units of grams? Provide your answer to 3 significant figures.

B. Which of the following amounts contains the fewest atoms?

12.01 g C 26.00 g Cr 40.08 g Ca 58.93 g Co 63.55 g Cu

Answer A

2.99×10^{-23} g H_2O

Answer B

26.00 g Cr

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6.7: Using Chemical Formulas as Conversion Factors

Learning Objectives

- Use chemical formulas as conversion factors.

Figure 6.7.1 shows that 2 hydrogen atoms and 1 oxygen atom are required to make one water molecule, 4 hydrogen atoms and 2 oxygen atoms for two water molecules, and 6 hydrogen atoms and 3 oxygen atoms for three water molecules. To make any number of water molecules, the ratio is always the same: 2 hydrogen atoms to 1 oxygen atom.



JSmol



JSmol



Figure 6.7.1: The ratio of hydrogen atoms to oxygen atoms used to make water molecules, H_2O is always 2:1, regardless of the number of water molecules.

By using formulas to indicate how many atoms of each element we have in a substance, we can relate the number of molecules to the number of atoms of each element. The possible relationships for 1 mole of water are shown in last column of Table 6.7.1.

Table 6.7.1: Molecular Relationships for Water

1 molecule H_2O has:	1 dozen H_2O molecules has:	1 mol H_2O has:	Mole Relationships
2 H atoms	2 dozen H atoms	2 mol H	$\frac{2\text{ mol H}}{1\text{ mol H}_2\text{O}}$ or $\frac{1\text{ mol H}_2\text{O}}{2\text{ mol H}}$
1 O atom	1 dozen O atoms	1 mol O	$\frac{1\text{ mol O}}{1\text{ mol H}_2\text{O}}$ or $\frac{1\text{ mol H}_2\text{O}}{1\text{ mol O}}$

Therefore, with 30.2 mol H_2O, there would be:

$$30.2 \cancel{\text{ mol H}_2\text{O}} \times \frac{2\text{ mol H}}{1\ \cancel{\text{mol H}_2\text{O}}} = \boxed{60.4\text{ mol H}}$$

$$30.2 \cancel{\text{ mol H}_2\text{O}} \times \frac{1\text{ mol O}}{1\ \cancel{\text{mol H}_2\text{O}}} = \boxed{30.2\text{ mol O}}$$

The following example illustrates how we can use the relationships in Table 6.7.1 as conversion factors.

✅ Example 6.7.1: Ethanol

How many moles of hydrogen atoms are present in 2.5 mol of ethanol (C_2H_6O)?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 2.5 mol C_2H_6O Find: mol H atoms
List known relationships.	1 mol C_2H_6O = 6 mol H
Prepare a concept map and use the proper conversion factor.	$\text{mol } C_2H_6O \xrightarrow{\frac{6 \text{ mol H}}{1 \text{ mol } C_2H_6O}} \text{mol H}$
Calculate the answer.	$2.5 \cancel{\text{ mol } C_2H_6O} \times \frac{6 \text{ mol H}}{1 \cancel{\text{ mol } C_2H_6O}} = \boxed{15 \text{ mol H}}$
Think about your result.	There 6 H atoms per C_2H_6O molecule, so the final answer should be 6 times as large.

Exercise 6.7.1

How many moles of sodium, sulfur, and oxygen atoms are in a sample containing 6.75 mol of Na_2SO_4, sodium sulfate?

Answer

13.5 mol Na atoms, 6.75 mol S atoms, and 27.0 mol O atoms

Summary

- In any given formula, the ratio of the number of moles of molecules (or formula units) to the number of moles of atoms can be used as a conversion factor.

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6.8: Percent-By-Mass Composition

Learning Objectives

- Determine percent composition of each element in a compound based on mass.

Packaged foods that you eat typically have nutritional information provided on the label. The label on a jar of peanut butter reveals that one serving size is considered to be **32 g**. The label also gives the masses of various types of compounds that are present in each serving. One serving contains 7 g of protein, 15 g of fat, and 3 g of sugar. By calculating the fraction of protein, fat, or sugar in one serving size of peanut butter and converting to percent values, we can determine the composition of peanut butter on a percent by mass basis.

Percent Composition

Chemists often need to know what elements are present in a compound and in what percentage. In general, percentages may be found by taking the part, dividing it by the whole and multiplying by 100:

$$\% = \frac{\text{part}}{\text{whole}} \times 100 \tag{6.8.1}$$

The **percent-by-mass composition** is the percent by mass of each element in a compound. It is calculated in a similar way to that of the composition of the peanut butter.

$$\%\ \text{by mass} = \frac{\text{mass of element in compound}}{\text{mass of compound}} \times 100\% \tag{6.8.2}$$

The sample problem below shows the calculation of the percent composition of a compound based on mass data.

✅ Example 6.8.1: Percent Composition from Mass Data

A certain newly synthesized compound is known to contain the elements zinc and oxygen. When a 20.00 g sample of this compound is decomposed, 16.07 g of zinc remains. Determine the percent-by-mass composition of the compound.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given : Mass of compound = 20.00 g Mass of Zn = 16.07 g Find: % Composition (%Zn and %O)
List known relationships.	Subtract to find the mass of oxygen in the compound. Mass of oxygen = 20.00 g - 16.07 g = 3.93 g O
Calculate the answers.	Calculate the percent by mass of each element by dividing the mass of that element by the mass of the compound and multiplying by 100%. $\%\ \mathrm{Zn} = \frac{16.07\ \mathrm{g\ Zn}}{20.00\ \mathrm{g}} \times 100\% = \boxed{80.35\%\ \mathrm{Zn}}$ $\%\ \mathrm{O} = \frac{3.93\ \mathrm{g\ O}}{20.00\ \mathrm{g}} \times 100\% = \boxed{19.65\%\ \mathrm{O}}$
Think about your result.	The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Exercise 6.8.1

Sulfuric acid, H_2SO_4 is a very useful chemical in industrial processes. If 196.0 g of sulfuric acid contained 64.0 g of sulfur and 4.0 g of hydrogen, what is the percent composition of the compound?

Answer

2.0% H, 32.7% S, and 65.31% O

Summary

- Processes are described for calculating the percent composition of a compound based on mass.

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6.9: Mass Percent Composition from a Chemical Formula

Learning Objectives

- Determine the percent composition of each element in a compound from the chemical formula.

Nitrogen is a nutrient essential for plant growth. Two common fertilizers used by farmers to supply nitrogen are ammonia (NH_3) and ammonium nitrate (NH_4NO_3). Ammonia is applied as a gas under pressure, called anhydrous ammonia. Ammonium nitrate is applied as a solid. Both have special handling precautions. Ammonia is highly soluble in water and will quickly dissolve in any water that is exposed. In large concentrations, one may quickly be overcome by ammonia, since it will dissolve in the water in your mouth, nose, lungs, and eyes. Ammonium nitrate, on the other hand, is highly explosive and has been the cause of several major deadly explosions over the past few decades.





Figure **6.9.1**: Anhydrous ammonia tanks in a farm field (Andrea Gantz via Flickr). Port of Beirut days after a massive ammonium nitrate explosion in August 2020 (Rashid Khreiss via Unsplash).

Despite the dangers, both fertilizers are still widely used since they each deliver a substantial percentage of nitrogen by mass. But, which of the two supplies the greatest nitrogen by mass? The easiest way to do this is to calculate the molar mass of each compound and to use the portion of that mass contributed by nitrogen to calculate the percentage.

$$= \frac{\text{mass of element in compound}}{\text{molar mass of compound}} \times 100\% \tag{6.9.1}$$

✅ Example 6.9.1

Which substance supplies the greatest percentage of nitrogen, by mass, NH_3 or NH_4NO_3? Provide calculations to support your answer.

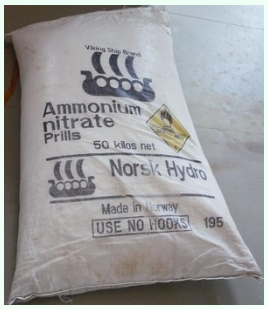
Solution

	NH_3	NH_4NO_3
1. Determine how many atoms of each element are present in the compound. Multiply the number of atoms of each element by its atomic mass and express in units of grams.	1 N = 1(14.01) = 14.01 g N 3 H = 3(1.008) = 3.024 g H	2 N = 2(14.01) = 28.02 g N 4 H = 4(1.008) = 4.032 g H 3 O = 3(16.00) = 48.00 g O
2. Sum the masses to find the molar mass.	molar mass = 17.03 g NH_3	molar mass = 80.05 g NH_4NO_3
3. Calculate the percentage of nitrogen.	$\%N = \frac{14.01\ \cancel{g}}{17.03\ \cancel{g}} \times 100\% = \boxed{82.27\%\ N}$	$\frac{28.02\ \cancel{g}}{80.05\ \cancel{g}} \times 100\% = \boxed{35.00\%\ N}$
4. Which supplies the greatest percentage of N?	From the calculations above, we can see that NH_3 supplies the greatest percentage of nitrogen, by mass.	

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound.

In Example 6.9.1, it was found that ammonium nitrate is 35.00% N. This means that there are 35.00 g N for every 100.00 g NH_4NO_3. Since the mass ratios remain the same regardless of sample size, this also means that there are 35.00 kg N for every 100.00 kg NH_4NO_3 or even 35.00 mg N for every 100.00 mg NH_4NO_3.

✅ Example 6.9.2



Using the results from Example 6.9.1, what mass of nitrogen is in a 50.0 kg bag of ammonium nitrate fertilizer? (Photo: Mike Spiller via Flickr)

Solution

Ammonium nitrate is 35.00% N. This means that there are 35.00 g N for every 100.00 g NH_4NO_3. Since the mass ratios remain the same regardless of sample size, this also means that there are 35.00 kg N for every 100.00 kg NH_4NO_3. Therefore,

$$50.0\ \cancel{\text{kg } NH_4NO_3} \times \frac{35.00\ \text{kg N}}{100.00\ \cancel{\text{kg } NH_4NO_3}} = \boxed{17.5\ \text{kg N}}$$

The solution may also be found using the masses from the molar mass calculation for NH_4NO_3 in Example 6.9.1, in which it was found that there are 28.02 g N for every 80.05 g NH_4NO_3. Since the mass ratios remain the same regardless of sample size, this also means that there are 28.02 kg N for every 80.05 kg NH_4NO_3. Therefore,

$$50.0\ \cancel{\text{kg } NH_4NO_3} \times \frac{28.02\ \text{kg N}}{80.05\ \cancel{\text{kg } NH_4NO_3}} = \boxed{17.5\ \text{kg N}}$$

Note that the answers agree.

Exercise 6.9.1

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer

78.327% Ba and 21.67% F

Exercise 6.9.2

An iron supplement contains 25.0 mg $FeSO_4$. What is the mass of iron in this sample?

Answer

9.19 mg Fe

Summary

- Processes are described for calculating the percent composition of a compound based on the chemical formula.

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6.11: Exercises

6.2: Molecular Mass and Formula Mass

1. What is the total mass (amu) of carbon in each of the following molecules?
    a. CH_4
    b. $CHCl_3$
    c. $C_{12}H_{10}O_6$
    d. $CH_3CH_2CH_2CH_2CH_3$

Answer

    a. 12.01 amu
    b. 12.01 amu
    c. 144.1 amu
    d. 60.05 amu

2. What is the total mass of hydrogen in each of the molecules?
    a. CH_4
    b. $CHCl_3$
    c. $C_{12}H_{10}O_6$
    d. $CH_3CH_2CH_2CH_2CH_3$

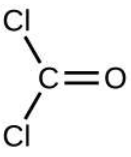
Answer

    a. 4.032 amu
    b. 1.008 amu
    c. 10.08 amu
    d. 12.10 amu

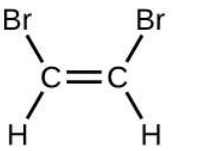
3. Calculate the molecular or formula mass of each of the following:
    a. P_4
    b. H_2O
    c. $Ca(NO_3)_2$
    d. CH_3CO_2H (acetic acid)
    e. $C_{12}H_{22}O_{11}$ (sucrose, cane sugar)

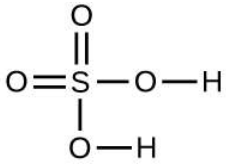
Answer

    a. 123.9 amu
    b. 18.02 amu
    c. 164.10 amu
    d. 60.05 amu
    e. 342.3 amu

4. Determine the molecular mass of the following compounds:
    a. 

b. 

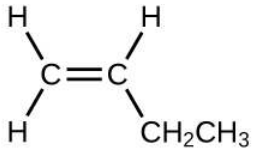
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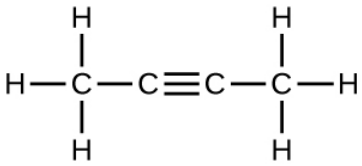
d. 

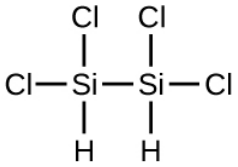
Answer

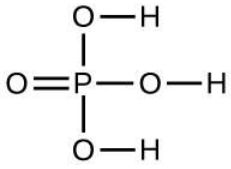
a. 98.91 amu
b. 26.04 amu
c. 185.8 amu
d. 98.09 amu

5. Determine the molecular mass of the following compounds:

a. 

b. 

c. 

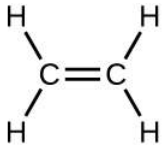
d. 

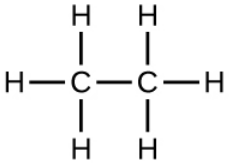
Answer

a. 56.10 amu
b. 54.09 amu
c. 200.0 amu
d. 97.99 amu

6. Which molecule has a molecular mass of 28.05 amu?

a. 

b. 

c. 

Answer

molecule b

6.3: Counting Objects By Weighing

7. Compare 1 mole of H_2, 1 mole of O_2, and 1 mole of F_2.
   a. Which has the largest number of molecules? Explain why.
   b. Which has the greatest mass? Explain why.

Answer

a. They all have the same number of molecules. One mole is 6.022×10^{23} particles, regardless of the identity of the particle.
b. The F_2 will have the greatest mass because a molecule of F_2 weighs more than a molecule of O_2 or H_2.

6.4: Molar Mass

8. How are the molecular mass and the molar mass of a compound similar and how are they different?

Answer

The two masses have the same numerical value, but the units are different: the molecular mass is the mass of 1 molecule in amu while the molar mass is the mass of 6.022×10^{23} molecules in grams.

9. Determine the molar mass of each substance.
   a. Si
   b. SiH_4
   c. K_2O

Answer

a. 28.09 g/mol
b. 32.12 g/mol
c. 94.20 g/mol

10. Determine the molar mass of each substance.
    a. Cl_2
    b. $SeCl_2$
    c. $Ca(C_2H_3O_2)_2$

Answer

a. 70.90 g/mol
b. 149.86 g/mol
c. 158.17 g/mol

11. Determine the molar mass of each substance.

a. Al
b. Al_2O_3
c. $CoCl_3$

Answer

a. 26.98 g/mol
b. 101.96 g/mol
c. 165.3 g/mol

12. Calculate the molar mass of each of the following compounds:

a. hydrofluoric acid, HF
b. ammonia, NH_3
c. nitric acid, HNO_3
d. silver sulfate, Ag_2SO_4
e. boric acid, $B(OH)_3$

Answer

a. 20.01 g/mol
b. 17.03 g/mol
c. 63.02 g/mol
d. 311.81 g/mol
e. 61.83 g/mol

13. Calculate the molar mass of each of the following minerals:

a. limestone, $CaCO_3$
b. halite, NaCl
c. beryl, $Be_3Al_2Si_6O_{18}$
d. malachite, $Cu_2(OH)_2CO_3$
e. turquoise, $CuAl_6(PO_4)_4(OH)_8(H_2O)_4$

Answer

a. 100.09 g/mol
b. 58.44 g/mol
c. 537.5 g/mol
d. 221.1 g/mol
e. 813.4 g/mol

14. Calculate the molar mass of each of the following:

a. the anesthetic halothane, $C_2HBrClF_3$
b. the herbicide paraquat, $C_{12}H_{14}N_2Cl_2$

c. caffeine, $C_8H_{10}N_4O_2$
d. urea, $CO(NH_2)_2$
e. a typical soap, $C_{17}H_{35}CO_2Na$

Answer

a. 197.38 g/mol
b. 257.2 g/mol
c. 194.20 g/mol
d. 60.06 g/mol
e. 306.5 g/mol

6.5: Mole Calculations

15. What is the mass of 4.44 mol of Mn?

Answer

244 g Mn

16. What is the mass of 0.311 mol of Xe?

Answer

40.8 g Xe

17. What is the mass of 12.34 mol of $Al_2(SO_4)_3$?

Answer

4,222 g $Al_2(SO_4)_3$

18. What is the mass of 0.0656 mol of $PbCl_2$?

Answer

18.2 g $PbCl_2$

19. How many moles are present in 45.6 g of CO?

Answer

1.63 mol CO

20. How many moles are present in 0.00339 g of LiF?

Answer

1.31×10^{-4} mol LiF

21. How many moles are present in 1.223 g of SF_6?

Answer

8.373×10^{-3} mol SF_6

22. How many moles are present in 48.8 g of $BaCO_3$?

Answer

0.247 mol $BaCO_3$

23. Determine the mass of each of the following:

a. 2.345 mol LiCl
b. 0.0872 mol acetylene, C_2H_2
c. 3.3×10^{-2} mol Na_2CO_3
d. 1.23×10^{3} mol fructose, $C_6H_{12}O_6$
e. 0.5758 mol $FeSO_4(H_2O)_7$

Answer

a. 99.40 g LiCl
b. 2.27 g C_2H_2
c. 3.5 g Na_2CO_3
d. 2.22×10^{5} g $C_6H_{12}O_6$
e. 160.1 g $FeSO_4(H_2O)_7$

24. The approximate minimum daily dietary requirement of the amino acid leucine, $C_6H_{13}NO_2$, is 1.1 g. What is this requirement in moles?

Answer

8.4×10^{-3} mol $C_6H_{13}NO_2$

25. Determine the mass in grams of each of the following:

a. 0.600 mol of oxygen atoms
b. 0.600 mol of oxygen molecules, O_2
c. 0.600 mol of ozone molecules, O_3

Answer

a. 9.60 g oxygen atoms
b. 19.2 g O_2
c. 28.8 g O_3

26. How many moles are present in 54.8 mL of mercury if the density of mercury is 13.6 g/mL?

Answer

3.72 mol mercury

27. How many moles are present in 56.83 mL of O_2 if the density of O_2 is 0.00133 g/mL?

Answer

2.36×10^{-3} mol O_2

6.6: Counting Particles By Weighing

28. How many atoms are present in 4.55 mol of Fe?

Answer

2.74×10^{24} atoms of Fe

29. How many atoms are present in 0.0665 mol of K?

Answer

4.00×10^{22} atoms of K

30. How many molecules are present in 2.509 mol of H_2S?

Answer

1.511×10^{24} molecules of H_2S

31. How many molecules are present in 0.336 mol of acetylene (C_2H_2)?

Answer

2.02×10^{23} molecules of C_2H_2

32. How many moles are present in 3.55×10^{24} Pb atoms?

Answer

5.90 mol Pb

33. How many moles are present in 2.09×10^{22} Ti atoms?

Answer

0.0347 mol Ti

34. How many moles are present in 1.00×10^{23} PF_3 molecules?

Answer

0.166 mol PF_3

35. How many moles are present in 5.52×10^{25} penicillin molecules?

Answer

91.7 mol penicillin

36. How many moles are present in 6.411 kg of CO_2? How many molecules is this?

Answer

145.7 mol CO_2; 8.772×10^{25} molecules of CO_2

37. How many moles are present in 2.998 mg of SCl_4? How many molecules is this?

Answer

1.724×10^{-5} mol SCl_4; 1.038×10^{19} molecules of SCl_4

38. What is the mass in milligrams of 7.22×10^{20} molecules of CO_2?

Answer

52.8 mg CO_2

39. What is the mass in kilograms of 3.408×10^{25} molecules of SiS_2?

Answer

5.220 kg SiS_2

40. What is the mass in grams of 1 molecule of H_2O?

Answer

2.992×10^{-23} g H_2O

41. What is the mass in grams of 1 atom of Al?

Answer

4.480×10^{-23} g Al

42. What is the volume in mL of 3.44 mol of Ga if the density of Ga is 6.08 g/mL?

Answer

39.4 mL Ga

43. What is the volume in L of 0.662 mol of He if the density of He is 0.1785 g/L?

Answer

14.8 L He

44. A 55-kg woman has 7.5×10^{-3} mol of hemoglobin (molar mass = 64,456 g/mol) in her blood. How many hemoglobin molecules is this? What is this quantity in grams?

Answer

4.5×10^{21} molecules of hemoglobin; 4.8×10^{2} g hemoglobin

45. Diamond is one form of elemental carbon. An engagement ring contains a diamond weighing 1.25 carats (1 carat = 200 mg). How many atoms are present in the diamond?

Answer

1.25×10^{22} atoms of elemental carbon

46. The Cullinan diamond was the largest natural diamond ever found (January 25, 1905). It weighed 3104 carats (1 carat = 200 mg). How many carbon atoms were present in the stone?

Answer

3.113×10^{25} atoms of carbon

47. One 55.0-gram serving of a particular cereal supplies 2.70 x 10^{2} mg of sodium, 11.0% of the recommended daily allowance. How many moles and atoms of sodium are in the recommended daily allowance?

Answer

0.0117 mol sodium; 7.07×10^{21} atoms of sodium

48. A certain nut crunch cereal contains 11.0 grams of sugar (sucrose, $C_{12}H_{22}O_{11}$) per serving size of 60.0 grams. How many servings of this cereal must be eaten to consume 0.0278 moles of sugar?

Answer

0.865 servings of cereal

49. Which of the following represents the least number of molecules?
    a. 20.0 g of H_2O (18.02 g/mol)
    b. 77.0 g of CH_4 (16.06 g/mol)
    c. 68.0 g of CaH_2 (42.09 g/mol)
    d. 100.0 g of N_2O (44.02 g/mol)
    e. 84.0 g of HF (20.01 g/mol)

Answer

20.0 g H_2O

6.7: Using Chemical Formulas as Conversion Factors

50. Which contains the greatest mass of oxygen: 0.75 mol of ethanol (C_2H_5OH), 0.60 mol of formic acid (HCO_2H), or 1.0 mol of water (H_2O)? Explain why.

Answer

Formic acid. Its formula has twice as many oxygen atoms as the other two compounds (one each). Therefore, 0.60 mol of formic acid would be equivalent to 1.20 mol of a compound containing a single oxygen atom.

51. Which contains the greatest number of moles of oxygen atoms: 1 mol of methanol (CH_3OH), 1 mol of acetic acid (CH_3CO_2H), or 1 mol of glycerol ($C_3H_8O_3$)? Explain why.

Answer

Glycerol. Each molecule of glycerol has three atoms of oxygen as compared to one atom of oxygen in a molecule of methanol and two atoms of oxygen in acetic acid.

52. Determine the number of moles of the compound and determine the number of moles of each type of atom in each of the following:
    a. 2.12 g of potassium bromide, KBr
    b. 0.1488 g of phosphoric acid, H_3PO_4
    c. 23 kg of calcium carbonate, $CaCO_3$
    d. 78.452 g of aluminum sulfate, $Al_2(SO_4)_3$
    e. 0.1250 mg of caffeine, $C_8H_{10}N_4O_2$

Answer

a. 0.0178 mol KBr; 0.0178 mol K; 0.0178 mol Br
b. 1.518×10^{-3} mol H_3PO_4; 4.555×10^{-3} mol H; 1.518×10^{-3} mol P; 6.074×10^{-3} mol O
c. 2.3×10^{2} mol $CaCO_3$; 2.3×10^{2} mol Ca; 2.3×10^{2} mol C; 6.9×10^{2} mol O
d. 0.2293 mol $Al_2(SO_4)_3$; 0.4586 mol Al; 0.6878 mol S; 2.751 mol O
e. 6.437×10^{-7} mol caffeine; 5.149×10^{-6} mol C; 6.437×10^{-6} mol H; 2.575×10^{-6} mol N; 1.287×10^{-6} mol O

53. Determine the number of atoms and the mass of zirconium, silicon, and oxygen found in 0.3384 mol of zircon, $ZrSiO_4$, a semiprecious stone.

Answer

zirconium: 2.038×10^{23} atoms; 30.87 g; silicon: 2.038×10^{23} atoms; 9.506 g; oxygen: 8.151×10^{23} atoms; 21.66 g

54. Determine which of the following contains the greatest mass of hydrogen: 1 mol of CH_4, 0.6 mol of C_6H_6, or 0.4 mol of C_3H_8.

Answer

1 mol CH_4

55. Determine which of the following contains the greatest mass of aluminum: 122 g of $AlPO_4$, 266 g of Al_2Cl_6, or 225 g of Al_2S_3.

Answer

225 g Al_2S_3

56. A tube of toothpaste contains 0.76 g of sodium monofluorophosphate (Na_2PO_3F) in 100 mL.
    a. What mass of fluorine atoms in mg was present?
    b. How many fluorine atoms were present?

Answer

a. 1.0×10^2 mg F
b. 3.2×10^{21} atoms F

6.8: Percent-By-Mass Composition

57. A 125 g sample of a compound known to contain only phosphorus and oxygen was shown to contain 54.5 g of phosphorus. What is the percent by mass composition of the compound?

Answer

43.6% P; 56.4% O

58. A compound containing only carbon and hydrogen was submitted for analysis. A 1.237 g sample of the compound was found to contain 1.140 g of carbon. What is the percent composition of the compound?

Answer

92.16% C; 7.84% H

59. Dichloroethane, a compound that is often used for dry cleaning, contains carbon, hydrogen, and chlorine. A 14.83 g sample of dichloroethane contained 3.60 g of carbon and 0.61 g of hydrogen. What is the percent by mass composition of dichloroethane?

Answer

24.3% C; 4.1% H; 71.6% Cl

60. A major textile dye manufacturer developed a new yellow dye containing carbon, nitrogen, and hydrogen. When 16.94 g of the dye was analyzed, it was found to contain 3.00 g of nitrogen and 1.07 g of hydrogen. What is the percent composition of the dye?

Answer

75.98% C; 17.7% N; 6.32% H

6.9: Mass Percent Composition from a Chemical Formula

61. Calculate the following to four significant figures:
    a. the percent composition of ammonia, NH_3

b. the percent composition of photographic “hypo,” $Na_2S_2O_3$
c. the percent of calcium ion in $Ca_3(PO_4)_2$

Answer

a. 82.25% N; 17.75% H
b. 29.08% Na; 40.56% S; 30.36% O
c. 38.76% Ca^{2+}

62. Determine the following to four significant figures:

a. the percent composition of hydrazoic acid, HN_3
b. the percent composition of TNT, $C_6H_2(CH_3)(NO_2)_3$
c. the percent of SO_4^{2-} in $Al_2(SO_4)_3$

Answer

a. 2.342% H; 97.66% N
b. 37.01% C; 2.219% H; 18.50% N; 42.26% O
c. 84.23% SO_4^{2-}

63. Determine the percent ammonia, NH_3, in $Co(NH_3)_6Cl_3$, to three significant figures.

Answer

38.2% NH_3

64. Determine the percent water in $CuSO_4 \cdot 5H_2O$ to three significant figures.

Answer

36.1% H_2O

65. Copper(I) iodide (CuI) is often added to table salt as a dietary source of iodine. How many moles of CuI are contained in 1.00 lb (454 g) of table salt containing 0.0100% CuI by mass?

Answer

2.38×10^{-4} mol CuI

66. A cough syrup contains 5.0% ethyl alcohol, C_2H_5OH, by mass. If the density of the solution is 0.9928 g/mL, how many grams of ethyl alcohol are contained in a 20.0 mL dose?

Answer

0.99 g ethyl alcohol

67. D5W is a solution used as an intravenous fluid. It is a 5.0% by mass solution of dextrose, $C_6H_{12}O_6$, in water. If the density of D5W is 1.029 g/mL, calculate the mass of dextrose in a 1.00 L IV fluid bag.

Answer

51 g dextrose

68. Find the mass of sulfuric acid in 325 mL of a 40.0% by mass aqueous solution of sulfuric acid, H_2SO_4, for which the density is 1.3057 g/mL.

Answer

1.70×10^2 g H_2SO_4

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CHAPTER OVERVIEW

7: Chemical Reactions

7.1: Whiz Bang!
7.2: Evidence of a Chemical Reaction
7.3: Chemical Equations
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7.5: Classifying Chemical Reactions
7.6: Combustion Reactions
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Thumbnail Chapter 7: Slow-motion capture of a Mylar balloon filled with hydrogen and oxygen gas exploding upon ignition during a test conducted by Anoka-Ramsey Community College chemistry faculty.

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7.1: Whiz Bang!

Children often get cuts and scrapes while playing. When this happens, it is important to cleanse the wound. If this happened to you as a kid, you may have had hydrogen peroxide poured on the cut. One thing that you may have noticed, in addition to any stinging, were the tiny bubbles forming in the cut. This was due to a chemical reaction known as a decomposition reaction which breaks hydrogen peroxide down to oxygen gas and water. Hydrogen peroxide undergoes this decomposition naturally all of the time. However, when added to a cut, enzymes in the body make the reaction go faster and the changes can be visualized. Enzymes are not the only substance which can speed up this reaction.

What happens when we add one of these other substances to speed up the decomposition of a *lot* of hydrogen peroxide, along with some soap to form bubbles with the oxygen gas? A reaction commonly known as elephant toothpaste, is shown in Video **7.1.1** below.



Video **7.1.1**: Elephant toothpaste reaction. Hydrogen peroxide decomposes to form lots of oxygen gas and water very quickly.

Chemistry is not only concerned with the composition of matter. It is also interested in how different types of matter interact with each other. These interactions are known as **chemical reactions** and are what make chemistry such an interesting field of science. Chemical reactions are occurring all of the time and all around us. Some of them, like the elephant toothpaste, can be very exciting to observe as they are accompanied by changes in colors, phases, heat, and light that grab our attention. Others are much more subtle and we tend not to notice them. In either case, these reactions involve changes in how atoms and molecules are arranged leading to the formation of new substances. In doing so, chemical reactions drive everything that happens to us on a daily basis and are therefore important to study and understand.

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7.2: Evidence of a Chemical Reaction

Learning Objectives

- Identify the evidence for chemical reactions.

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

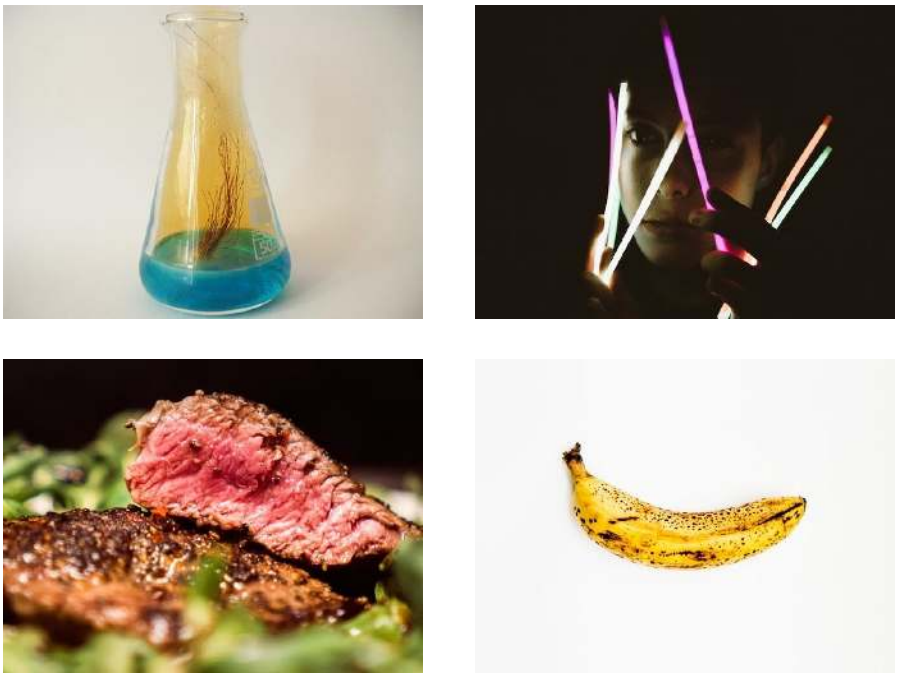


Figure **7.2.1**: Clockwise from upper left. (A) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide (unknown via Wikipedia Commons). (B) A chemical reaction that gives off light, a process called chemiluminescence, occurs when a glow stick is bent, a crack is heard and the inner solutions mix together (Anton Darius via Unsplash). (C) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change (Justus Menke via Unsplash). (D) A banana turning brown is a chemical change as new, darker substances that have a different taste form (Bruno Emmanuelle via Unsplash).

To identify a chemical reaction, we look for a **chemical change**. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance.

Various signs of a chemical reaction include:

1. Change in Color
2. Formation of a Gas
3. Formation of a Precipitate
4. Change in Odor
5. Change in Temperature
6. Something is Burning

7. Light is Being Produced



Video **7.2.1**: Evidence of a Chemical Reaction

✅ Example 7.2.1: Evidence of a Chemical Reaction

Which of the following is a chemical reaction?

A. Freezing liquid mercury.
B. Adding yellow to blue to make green.
C. Cutting a piece of paper into two pieces.
D. Burning natural gas.
E. Dissolving salt in water.

Solution

A, B, C, & E involve only physical changes. Burning involves a reaction with oxygen where the reacting materials have been consumed and the products have a different chemical composition, so the answer is D.

✏️ Exercise 7.2.1

Which of these involve a chemical reaction?

A. Painting a wall blue.
B. Mixing baking soda and vinegar to produce a gas.
C. Ice cream melting.
D. Boiling water.
E. Roasting a marshmallow over a flame until it turns brown.

Answer

B and E

Summary

- Chemical reactions can be identified via a wide range of different observable factors including change in color, energy change (temperature change or light produced), gas production, something burning, and the formation of a precipitate.

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7.3: Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, such as (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

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

Reactants and Products

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

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that – what is produced – or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about making tacos, our reactants would be tortillas, meat, cheese, lettuce, tomatoes, sour cream, and salsa. What would be the products? Tacos! The reaction vessel would be our plate or countertop.

$$\underbrace{\text{tortillas} + \text{meat} + \text{cheese} + \text{lettuce} + \text{tomatoes} + \text{sour cream} + \text{salsa}}_{\text{Ingredients = Reactants}} \rightarrow \underbrace{\text{tacos}}_{\text{Product}}$$

Writing Chemical Equations

Let's consider the explosive reaction between hydrogen and oxygen gases (see Figure 7.3.1 below). One might describe this reaction using words like this:

- Hydrogen gas reacts with oxygen gas to yield water in the gaseous phase.



Figure 7.3.1: Slow-motion capture of a Mylar balloon filled with hydrogen and oxygen gas exploding upon ignition during a test conducted by Anoka-Ramsey Community College chemistry faculty.

Word descriptions are typically simplified by writing out chemical reactions in the form of a chemical equation using chemical formulas to represent the substances that are reacting. The reaction between hydrogen and oxygen may be simplified by the chemical equation:

$$H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(g)$$

Hydrogen and oxygen are the reactants and water is the product.

$$\underbrace{H_2\,(g) + O_2\,(g)}_{\text{Reactants}} \rightarrow \underbrace{H_2O\,(g)}_{\text{Product}}$$

In chemical reactions, reactants are shown to the left of the arrow ($\rightarrow$) and products are shown after the arrow. The general equation for a chemical reaction is:

$$\textbf{Reactants} \rightarrow \textbf{Products}$$

As to why water is produced as a gas, it may be observed that the water is produced as a result of an explosion as shown in Figure 7.3.1. At the elevated temperature of the explosion, water exists as a gas and not as a liquid. This leads to a few special symbols that we need to know in order to "talk" in chemical shorthand. Table 7.3.1 provides a summary of the major symbols used in chemical equations.

Table 7.3.1: Symbols Used in Chemical Equations

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

The symbols (*s*), (*l*), (*g*), and (*aq*) are called phase labels. Phase labels aren't always necessary. However, they add clarity when the physical states of each substance involved are unclear or when they add important context.

A reaction between aqueous solutions of calcium nitrate and sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate may be written as:

$$\mathrm{Ca(NO_3)_2\,(aq) + NaOH\,(aq) \rightarrow Ca(OH)_2\,(s) + NaNO_3\,(aq)}$$

Many would agree that the chemical equation is much easier to read.

We should also be able to read a chemical equation and put into words a description for a chemical reaction. Consider,

$$\mathrm{Cu\,(s) + AgNO_3\,(aq) \rightarrow Cu(NO_3)_2\,(aq) + Ag\,(s)}$$

This equation may be described something like

- Solid copper reacts with an aqueous solution of silver nitrate to produce an aqueous solution of copper(II) nitrate and solid silver metal.

To turn word equations into chemical equations, it may be helpful to think of the process something like this:

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs belong.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (making sure all ionic compounds are charge balanced).
3. Write the correct formulas for all elements, remembering that that there are seven diatomic elements (H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2), meaning that they are always found in pairs in nature. Diatomic elements were discussed in Section 5.3.

✅ Example 7.3.1

Convert the chemical equations into word equations and the word equations into chemical equations.

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

Solution

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

B. Reactants: propane ($\mathbf{C_3H_8}$) and oxygen ($\mathbf{O_2}$)

Product: carbon dioxide ($\mathbf{CO_2}$) and water ($\mathbf{H_2O}$)

$$\boxed{\mathbf{C_3H_8\ (g)\ +\ O_2\ (g)\ \rightarrow\ CO_2\ (g)\ +\ H_2O\ (l)}}$$

C. Reactants: hydrogen fluoride ($\mathbf{HF}$) and potassium carbonate ($\mathbf{K_2CO_3}$)

Products: potassium fluoride ($\mathbf{KF}$), water ($\mathbf{H_2O}$), and carbon dioxide ($\mathbf{CO_2}$)

$$\boxed{\mathbf{HF\ (g)\ +\ K_2CO_3\ (aq)\ \rightarrow\ KF\ (aq)\ +\ H_2O\ (l)\ +\ CO_2\ (g)}}$$

Exercise 7.3.1

Convert the chemical equations into word equations and the word equations into chemical equations.

A. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.

B. $\mathbf{HCl\,(aq) + LiOH\,(aq) \rightarrow LiCl\,(aq) + H_2O\,(l)}$

C. Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer A

$\mathbf{H_2\ (g) + N_2\ (g) \rightarrow NH_3\ (g)}$

Answer B

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

Answer C

$\mathbf{Cu\ (s) + O_2\ (g) \rightarrow CuO\ (s)}$

Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

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7.4: Balancing Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the law of conservation of matter in a chemical reaction.

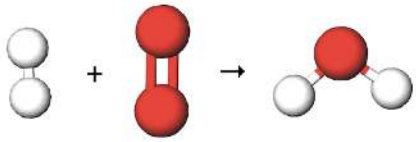
Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are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

Chemical Equations and the Law of Conservation of Matter

In the previous section, the reaction between hydrogen gas and oxygen gas to produce water in the gaseous phase was shown as a chemical equation:

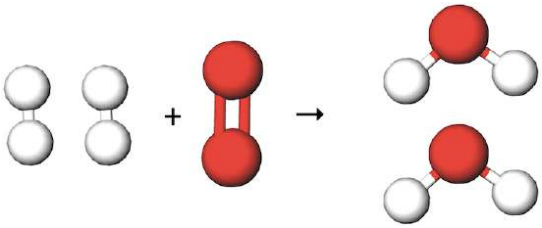
$$H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(g)$$

At the molecular level, the reaction would look something like this:



Notice that there are two oxygen atoms on the left hand side of the equation and only one oxygen atom on the right. Where did that one oxygen atom go? Matter can't just disappear. Once again, atoms are never created or destroyed in chemical reactions.

Now, let's consider what happens if we start with two hydrogen molecules and one oxygen molecule:



This combination results in two water molecules being formed – there are four hydrogen atoms and two oxygen atoms on the left hand side of the equation and four hydrogen atoms and two oxygen atoms on the right. Any violation of the law of conservation of matter has been avoided. The chemical equation for this reaction above would be written as:

$$2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g)$$

This form of the chemical equation is called a **balanced chemical equation**. A balanced chemical equation is one that conforms to the law of conservation of matter.

Coefficients and Subscripts

There are two types of numbers that appear in chemical equations. There are **subscripts**, which are part of the chemical formulas of the reactants and products:

$$\mathbf{2\,H_2}\,(g) + \mathbf{O_2}\,(g) \rightarrow \mathbf{2\,H_2O}\,(g)$$

Once the formulas for the reactants and products are determined, the *subscripts may not be changed*.

There are also **coefficients** that are placed in front of chemical formulas to indicate how many molecules of that substance are used or produced:

$$\boxed{2}\,H_2\,(g) + O_2\,(g) \rightarrow \boxed{2}\,H_2O\,(g)$$

Coefficients may be changed in order to balance the equation. The equation above indicates that two molecules of hydrogen gas are reacting with one molecule of oxygen gas to produce two molecules of water.

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described. For example, had the equation for the reaction:

$$H_2\,(g) + O_2\,(g) \rightarrow H_2O\,(g)$$

been balanced by changing the subscript for the oxygen on H_2O to become H_2O_2, the reaction no longer shows the production of water. The equation is indeed balanced. However, it shows the production of hydrogen peroxide, a very different substance:

$$\cancel{H_2\,(g) + O_2\,(g) \rightarrow H_2O_2\,(g)}$$

Therefore, once again, *chemical equations are balanced by changing the coefficients*:

$$\boxed{2}\,H_2\,(g) + O_2\,(g) \rightarrow \boxed{2}\,H_2O\,(g)$$

Just like subscripts of "1" are implied in chemical formulas (the formula for water is written as H_2O and not H_2O_1), coefficients of "1" are also implied. In other words, the coefficient for O_2 (*g*) in the balanced equation above is "1", even though the number "1" is not written.

It may also be helpful to think of coefficients in chemical equations as if they were coefficients in algebra. Coefficients in algebra are written at the front of a term or variable, not in the middle. Just like in algebra, a coefficient in a chemical equation means to multiply. In other words, by writing "2 H_2O", we are indicating a total of 2×2, or 4, hydrogen atoms altogether.

Balancing Chemical Equations

The simplest and most generally useful method for balancing chemical equations is called “balancing by inspection,” also known as trial and error. There are some strategies that may be used to reduce the number of trials and errors to wind up with an equation that is ultimately balanced correctly.

⚓ Strategies for Balancing Chemical Equations

1. The most complex formula should usually be balanced first.
2. The simplest formula should usually be balanced last.
3. The least common multiple between two numbers may be used to determine the coefficients.
4. Polyatomic ions (if present on both sides of the chemical equation) may be balanced as a unit.
5. Fractional coefficients are sometimes necessary to balance the equation when balancing the simplest formula that remains. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole number coefficients.
6. Count the numbers of atoms of each element on each side of the equation to ensure the chemical equation is balanced.

✅ Example 7.4.1: Oxidation of Aluminum

Balance the chemical equation for the reaction of aluminum metal with oxygen gas to produce solid aluminum oxide:

$$Al\,(s) + O_2\,(g) \rightarrow Al_2O_3\,(s)$$

Solution

1. The most complex formula is Al_2O_3 (s) and should be balanced first, while the simplest formula is Al (s) and should be balanced last.

2. There are 2 oxygen atoms on the left and 3 oxygen atoms on the right. The least common multiple between 2 and 3 is 6. To make 6 oxygen atoms on each side, a coefficient of "3" is placed in front of O_2 (g) and a coefficient of "2" is placed in front of Al_2O_3 (s):

$$Al\ (s) + 3\ O_2\ (g) \rightarrow 2\ Al_2O_3\ (s)$$

3. This still leaves aluminum atoms unbalanced. At the point, there is 1 aluminum atom on the left and 4 aluminum atoms on the right. To balance the aluminum atoms, a coefficient of "4" is placed in front of Al (s):

$$\boxed{4\ Al\ (s) + 3\ O_2\ (g) \rightarrow 2\ Al_2O_3\ (s)}$$

4. A final check should be done to ensure the equation remains balanced for all elements. Sometimes it is useful placing the information into a table and tallying each atom of an element as the equation is being balanced:

Element	Reactant Side	Product Side
Al	𝍸	𝍸
O	𝍸 |	𝍸 |

✅ Example 7.4.2: Decomposition of Potassium Chlorate

Balance the chemical equation for the decomposition of solid potassium chlorate in the presence of heat to yield solid potassium chloride and oxygen gas. Note that a triangle appears above the arrow to indicate that heat is necessary for this reaction to occur:

$$KClO_3\ (s) \xrightarrow{\triangle} KCl\ (s) + O_2\ (g)$$

Solution

1. The most complex formula is $KClO_3$ (s) and should be balanced first. KCl (s) or O_2 (g) seem to be equally simple. However, it appears that potassium atoms and chlorine atoms are already balanced, so one may begin by balancing the oxygen atoms. Let's keep an ongoing tally this time.

Element	Reactants	Products
K	|	|
Cl	|	|
O	𝍷	||

2. There are 3 oxygen atoms on the left and 2 oxygen atoms on the right. The least common multiple between 2 and 3 is 6. To make 6 oxygen atoms on each side, a coefficient of "2" is placed in front of $KClO_3$ (s) and a coefficient of "3" is placed in front of O_2 (g):

$$2\ KClO_3\ (s) \xrightarrow{\triangle} KCl\ (s) + 3\ O_2\ (g)$$

Element	Reactants	Products
K	||	|
Cl	||	|
O	𝍸 |	𝍸 |

3. This now makes both potassium and chlorine atoms unbalanced. However, this is easily remedied by placing a coefficient of "2" is placed in front of KCl (s):

$$\boxed{2\ KClO_3\ (s) \xrightarrow{\triangle} 2\ KCl\ (s) + 3\ O_2\ (g)}$$

4. A final check should be done to ensure the equation remains balanced for all elements.

Element	Reactants	Products
K	||	||
Cl	||	||
O	𝍸 |	𝍸 |

✅ Example 7.4.3: Combustion of Butane

Balance the chemical equation for the combustion of butane liquid (C_4H_{10}) in the presence of oxygen gas to yield carbon dioxide and water in the gaseous phase:

$$C_4H_{10}\,(l) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)$$

Solution

1. The most complex formula is $C_4H_{10}\,(l)$ and should be balanced first, while the simplest formula is $O_2\,(g)$ and should be balanced last. Let's keep an ongoing tally once again, this time using numbers.

Element	Reactant Side	Product Side
C	4	1
H	10	2
O	2	3

2. There are 4 carbon atoms on the left and 1 carbon atom on the right. The least common multiple between 4 and 1 is 4. To make 4 carbon atoms on the right, a coefficient of "4" is placed in front of $CO_2\,(g)$:

$$C_4H_{10}\,(l) + O_2\,(g) \rightarrow 4\,CO_2\,(g) + H_2O\,(g)$$

Element	Reactant Side	Product Side
C	4	4
H	10	2
O	2	9

3. There are 10 hydrogen atoms on the left and 2 hydrogen atoms on the right. The least common multiple between 10 and 2 is 10. To make 10 hydrogen atoms on the right, a coefficient of "5" is placed in front of $H_2O\,(g)$:

$$C_4H_{10}\,(l) + O_2\,(g) \rightarrow 4\,CO_2\,(g) + 5\,H_2O\,(g)$$

Element	Reactant Side	Product Side
C	4	4
H	10	10
O	2	13

4. Only oxygen atoms remain unbalanced, as originally desired. At the point, there are 2 oxygen atoms on the left and 13 oxygen atoms on the right. This means it should be possible to place a coefficient (x) in front of $O_2\,(g)$ to balance the equation:

$$C_4H_{10}\,(l) + x\,O_2\,(g) \rightarrow 4\,CO_2\,(g) + 5\,H_2O\,(g)$$

5. It appears this is the time to employ Strategy #5 listed above, using a fractional coefficient to balance the oxygen atoms, since we need to ask ourselves, "What number times two would equal 13?" In algebra, this question would be written as

$$2x = 13$$

Dividing each side by 2, a solution for x may be found:

$$x = \frac{13}{2}$$

In other words, a coefficient of "$\frac{13}{2}$" may placed in front of $O_2\,(g)$:

$$C_4H_{10}\,(l) + \tfrac{13}{2}\,O_2\,(g) \rightarrow 4\,CO_2\,(g) + 5\,H_2O\,(g)$$

Element	Reactant Side	Product Side
C	4	4
H	10	10
O	13	13

6. As noted in Strategy #5, both sides of the equations should be multiplied by the denominator when fractional coefficients are used in order to obtain whole number coefficients.

$$2 \times \left[C_4H_{10}\,(l) + \tfrac{13}{2}\,O_2\,(g) \rightarrow 4\,CO_2\,(g) + 5\,H_2O\,(g)\right]$$

$$\boxed{2\,C_4H_{10}\,(l) + 13\,O_2\,(g) \rightarrow 8\,CO_2\,(g) + 10\,H_2O\,(g)}$$

7. A final check should be done to ensure the equation remains balanced for all elements.

Element	Reactant Side	Product Side
C	8	8
H	20	20
O	26	26

✏️ Exercise 7.4.1

Balance the following chemical equations.

A. $V\,(s) + O_2\,(g) \rightarrow V_2O_5\,(s)$
B. $Mg\,(s) + O_2\,(g) \rightarrow MgO\,(s)$
C. $C_6H_{12}O_6\,(s) + O_2\,(g) \rightarrow CO_2\,(s) + H_2O\,(l)$
D. $H_2O_2\,(aq) \xrightarrow{MnO_2} H_2O\,(l) + O_2\,(g)$

Answer A

$4\,V\,(s) + 5\,O_2\,(g) \rightarrow 2\,V_2O_5\,(s)$

Answer B

$2\,Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$

Answer C

$C_6H_{12}O_6\,(s) + 6\,O_2\,(g) \rightarrow 6\,CO_2\,(s) + 6\,H_2O\,(l)$

Answer D

$2\,H_2O_2\,(aq) \xrightarrow{MnO_2} 2\,H_2O\,(l) + O_2\,(g)$

✅ Example 7.4.4: Precipitation of Lead(II) Chloride

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

Solution

1. Start by writing out a chemical equation from the word equation. Recall that the nitrate ion (NO_3^-) is a polyatomic ion and may be balanced as a unit.

$$Pb(NO_3)_2\,(aq) + NaCl\,(aq) \rightarrow PbCl_2\,(s) + NaNO_3\,(aq)$$

Unit	Reactant Side	Product Side
Pb	|	|
NO_3	||	|
Na	|	|
Cl	|	||

2. By looking at the tallies, it appears nitrate ions are easily balanced by placing a "2" in front of $NaNO_3$ (*aq*).

$$Pb(NO_3)_2\ (aq) + NaCl\ (aq) \rightarrow PbCl_2\ (s) + 2\ NaNO_3\ (aq)$$

Unit	Reactant Side	Product Side
Pb	|	|
NO_3	||	||
Na	|	||
Cl	|	||

3. At this point, both sodium and chlorine atoms unbalanced. However, this is easily remedied by placing a coefficient of "2" is placed in front of NaCl (*aq*):

$$\boxed{Pb(NO_3)_2\ (aq) + 2\ NaCl\ (aq) \rightarrow PbCl_2\ (s) + 2\ NaNO_3\ (aq)}$$

4. A final check shows the equation is now balanced for all elements.

Unit	Reactant Side	Product Side
Pb	|	|
NO_3	||	||
Na	||	||
Cl	||	||

Exercise 7.4.2

Write out a balanced chemical equation for these word equations.

A. Nitrogen gas reacts with hydrogen gas to form gaseous ammonia.
B. Aqueous solutions of aluminum nitrate and lithium carbonate are mixed to yield a solid aluminum carbonate precipitate and an aqueous solution of lithium nitrate.
C. Liquid hexyne (C_6H_{10}) reacts with hydrogen gas in the presence of a platinum catalyst to yield liquid hexane (C_6H_{14}).

Answer A

$$N_2\ (g) + 3\ H_2\ (g) \rightarrow 2\ NH_3\ (g)$$

Answer B

$$2\ Al(NO_3)_3\ (aq) + 3\ Li_2CO_3\ (aq) \rightarrow Al_2(CO_3)_3\ (s) + 6\ LiNO_3\ (aq)$$

Answer C

$$C_6H_{10}\ (l) + 2\ H_2\ (g) \xrightarrow{Pt} C_6H_{14}\ (l)$$

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

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7.5: Classifying Chemical Reactions

Learning Objectives

- Classify a chemical reaction as a synthesis, decomposition, combustion, single replacement, double replacement reaction.
- Predict the products of simple reactions.

There are an infinite number of chemical reactions that are possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Many chemical reactions may be classified into one or more of five basic types: combination (or synthesis), decomposition, combustion, single replacement, and double replacement. It is important to note, however, that many reactions may classified in more than one way.

Combination Reactions

A **combination reaction** is a reaction in which two or more substances combine to form a single new substance. Combination reactions are also called synthesis reactions. The general form of a combination reaction is:

$$A + B \rightarrow AB$$

One type of combination reaction that occurs frequently is the reaction of an element with oxygen to form an oxide. Metals and nonmetals both react readily with oxygen under most conditions. Magnesium reacts rapidly and dramatically when ignited, combining with oxygen from the air to produce a fine powder of magnesium oxide:

$$2\ Mg\ (s) + O_2\ (g) \rightarrow 2\ MgO\ (s)$$

Notice that in order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2). When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form either sulfur dioxide or sulfur trioxide:

$$S\ (s) + O_2\ (g) \rightarrow SO_2\ (g)$$

$$2\ S\ (s) + 3\ O_2\ (g) \rightarrow 2\ SO_3\ (g)$$

Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron(II) oxide and iron(III) oxide:

$$2\ Fe\ (s) + O_2\ (g) \rightarrow 2\ FeO\ (s)$$

$$4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$$

Combination reactions can also take place when two compounds react to form a more complex compound. A very common example is the reactions of oxides with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide:

$$CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$$

Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$H_2O\ (l) + SO_3\ (g) \rightarrow H_2SO_4\ (aq)$$




Figure 7.5.1: Acid rain has severe consequences on both natural and manmade objects. The trees in the forest on the left have been killed by acid rain. Acid rain also degrades marble statues like the one on the right. (OpenStax, CC BY 4.0, via Wikimedia Commons)

Decomposition Reactions

A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$$AB \rightarrow A + B$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury(II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas:

$$2\ HgO\ (s) \rightarrow 2\ Hg\ (l) + O_2\ (g)$$



Video 7.5.1: Mercury(II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.

A reaction is also considered to be a decomposition reaction even when one or more of the products are still compounds. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide. Carbonic acid decomposes easily at room temperature into carbon dioxide and water:

$$CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$$

$$H_2CO_3\ (aq) \rightarrow CO_2\ (g) + H_2O\ (l)$$

Combustion Reactions

A **combustion reaction** is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat. The products of a combustion reaction depend on the substance being burned. If the substance being burned contains carbon, one of the products will be carbon dioxide. If the substance being burned contains hydrogen, one of the products will be water. If the substance contains sulfur, one of the products will be sulfur dioxide.

The primary component of natural gas is methane, CH_4. The complete combustion of methane yields carbon dioxide and water:

$$CH_4\,(g) + 2\,O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

When the supply of oxygen becomes limited, incomplete combustion typically occurs causing carbon monoxide, CO, to be produced in addition to carbon dioxide. In this text, complete combustion will always be assumed to occur unless otherwise specified. Combustion reactions are covered more extensively in Section 7.6.

Single Replacement Reactions

A fourth type of reaction is the single replacement reaction, in which one element replaces a similar element in a compound. One general form of a single-replacement reaction (also called a single-displacement reaction) is:

$$AB + C \rightarrow CB + A$$

In this general reaction, element C is a metal and replaces element A, also a metal, in the compound AB. One example is when a copper wire is placed into an aqueous solution of silver nitrate (see Figure 7.5.2 below). Silver leaves the solution forming crystals of silver metal, while copper metal goes into solution yielding the characteristic blue color of an aqueous copper(II) nitrate solution.

$$2\,AgNO_3\,(aq) + Cu\,(s) \rightarrow Cu(NO_3)_2\,(aq) + 2\,Ag\,(s)$$



Figure 7.5.2: Crystals of silver metal form on a coil of copper wire that has been immersed into an aqueous solution of silver nitrate. The solution becomes progressively darker bluish-green as copper metal goes into solution forming copper(II) nitrate. (Tony Hudson via Wikipedia)

When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$AB + C \rightarrow AC + B$$

where C is a nonmetal and replaces the nonmetal B in the compound AB. One example is when an aqueous solution of sodium bromide reacts with the element chlorine to produce aqueous sodium chloride and elemental bromine:

$$2\,NaBr\,(aq) + Cl_2\,(g) \rightarrow 2\,NaCl\,(aq) + Br_2\,(l)$$

The reactivity of the halogens (Group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

Double Replacement Reactions

A **double-replacement reaction** (also called double-displacement reaction) is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a double-replacement reaction is:

$$AB + CD \rightarrow AD + CB$$

In this reaction, A and C represent positively-charged cations, while B and D represent negatively-charged anions. Double-replacement reactions generally occur between substances in aqueous solution. In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water.

A precipitate forms in a double-replacement reaction when the cations from one of the reactants combine with the anions from the other reactant to form an insoluble ionic compound. When aqueous solutions lead(II) nitrate and potassium iodide are mixed, the following reaction occurs:

$$Pb(NO_3)_2\,(aq) + 2\,KI\,(aq) \rightarrow PbI_2\,(s) + 2\,KNO_3\,(aq)$$

There are very strong attractive forces that occur between Pb^{2+} and I^- ions and the result is a brilliant yellow precipitate (Figure 7.5.3). The other product of the reaction, potassium nitrate, remains soluble.



Figure **7.5.3**: Lead(II) iodide precipitates when potassium iodide is mixed with lead(II) nitrate. (CC BY-SA 3.0; PRHaney).

Determining the products of precipitation reactions and predicting whether or not a precipitation reaction occurs is discussed more extensively in Section 7.8.

Table 7.5.1: A Summary of Chemical Reaction Classification

Name of Reaction	General Form	Examples
1. Combination (Synthesis)	A + B → AB	$N_2\,(g) + 2\,O_2\,(g) \rightarrow 2\,NO_2\,(g)$ $CO_2\,(g) + H_2O\,(l) \rightarrow H_2CO_3\,(aq)$
2. Decomposition	AB → A + B	$CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$
3. Combustion	burning in the presence of $O_2\,(g)$	$C_7H_{16}\,(l) + 11\,O_2\,(g) \rightarrow 7\,CO_2\,(g) + 8\,H_2O\,(g)$
4. Single Replacement	AB + C → CB + A AB + C → AC + B	$ZnCl_2\,(aq) + Mg\,(s) \rightarrow MgCl_2\,(aq) + Zn\,(s)$ $2\,KI\,(aq) + Cl_2\,(aq) \rightarrow 2\,KCl\,(aq) + I_2\,(aq)$
5. Double Replacement	AB + CD → AD + CB	$BaCl_2\,(aq) + Na_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + 2\,NaCl\,(aq)$

The classification scheme is by no means complete. Other reaction classifications include oxidation-reduction reactions (also called redox reactions), acid-base reactions, and condensation reactions, to name a few. In this text, determining the products of combustion reactions and double replacement reacts (specifically precipitation reactions) will be most emphasized. These reactions are covered in the next few sections.

Exercise 7.5.1

Classify the following reactions as combination, decomposition, combustion, single replacement, or double replacement.

A. $\mathbf{Mg\,(s) + Cu(NO_3)_2\,(aq) \rightarrow Mg(NO_3)_2\,(aq) + Cu\,(s)}$
B. $\mathbf{4\,V\,(s) + 5\,O_2\,(g) \rightarrow 2\,V_2O_5\,(s)}$
C. $\mathbf{(NH_4)_2SO_4\,(aq) + Ba(NO_3)_2\,(aq) \rightarrow BaSO_4\,(s) + 2\,NH_4NO_3\,(aq)}$
D. $\mathbf{C_6H_{12}O_6\,(s) + 6\,O_2\,(g) \rightarrow 6\,CO_2\,(s) + 6\,H_2O\,(l)}$
E. $\mathbf{2\,H_2O_2\,(aq) \xrightarrow{MnO_2} 2\,H_2O\,(l) + O_2\,(g)}$

Answer A

single replacement

Answer B

combination

Answer C

double replacement

Answer D

combustion

Answer E

decomposition

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7.6: Combustion Reactions

Learning Objectives

- Recognize a reaction as a combustion reaction.
- Complete and balance chemical equations for combustion reactions.

A **combustion reaction** is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat. Combustion reactions involve O_2 as one reactant. The combustion of hydrogen gas producing water vapor qualifies as a combustion reaction:

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

The *Hindenberg* was a hydrogen-filled airship that suffered an accident upon its attempted landing in New Jersey in 1937. The hydrogen immediately combusted in a huge fireball, destroying the airship and killing 36 people. The chemical reaction was a simple one: hydrogen combining with oxygen to produce water.

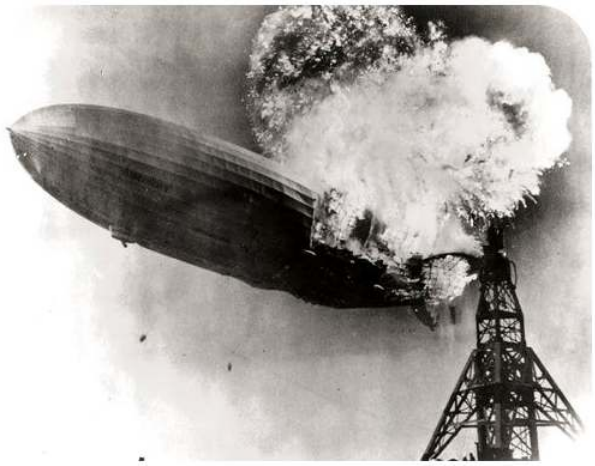


Figure **7.6.1**: Explosion of the Hindenberg.

Many combustion reactions occur with a hydrocarbon, a compound comprised solely of carbon and hydrogen. The products of the complete combustion of hydrocarbons are carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amounts of heat energy. Propane, C_3H_8, is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$C_3H_8\ (g) + 5\ O_2\ (g) \rightarrow 3\ CO_2\ (g) + 4\ H_2O\ (g)$$

As a general rule, combustion of a reactant that contains:

- carbon will yield carbon dioxide, CO_2, as one of the products.
- hydrogen will yield water, H_2O, as one of the products.
- sulfur will yield sulfur dioxide, SO_2, as one of the products.

When the supply of oxygen becomes limited, incomplete combustion typically occurs. The incomplete combustion carbon or carbon-containing compounds leads to the production of carbon monoxide, CO, in addition to carbon dioxide. In this text, however, complete combustion will always be assumed to occur unless otherwise specified.

✅ Example 7.6.1: Combustion of Ethanol

Ethanol (C_2H_5OH) may be used as a fuel source in an alcohol lamp. Write the balanced chemical equation that shows the combustion of ethanol.

Solution

Combustion is a reaction in which a substance reacts with oxygen gas.

$$C_2H_5OH\ (l) + O_2\ (g) \rightarrow$$

Since C_2H_5OH contains carbon, CO_2 will be one of the products. Since it also contains hydrogen, H_2O will be another product. H_2O is produced in the gaseous phase due to the high temperatures that accompany combustion reactions.

$$C_2H_5OH\ (l) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g)$$

Carbon atoms are balanced by placing a coefficient of "2" in front of CO_2 (g). Hydrogen atoms are balanced by placing a "3" in front of H_2O (g).

$$C_2H_5OH\ (l) + O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3\ H_2O\ (g)$$

The simplest formula, O_2 (g), is balanced last using a coefficient of "3".

$$C_2H_5OH\ (l) + 3\ O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3\ H_2O\ (g)$$

Exercise 7.6.1: Combustion of Hexane

Write the balanced equation for the combustion of hexane, C_6H_{14}, which is a liquid at room temperature.

Answer

$2\ C_6H_{14}\ (l) + 19\ O_2\ (g) \rightarrow 12\ CO_2\ (g) + 14\ H_2O\ (g)$

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7.7: Solubility Rules for Ionic Compounds

Learning Objectives

- Define and give examples of electrolytes.
- Determine the solubility of common ionic compounds.

When some substances are dissolved in water, they undergo either a physical or a chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not yield ions when dissolved are called nonelectrolytes. If the physical or chemical process that generates the ions is essentially 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a weak electrolyte.

Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 7.7.1).

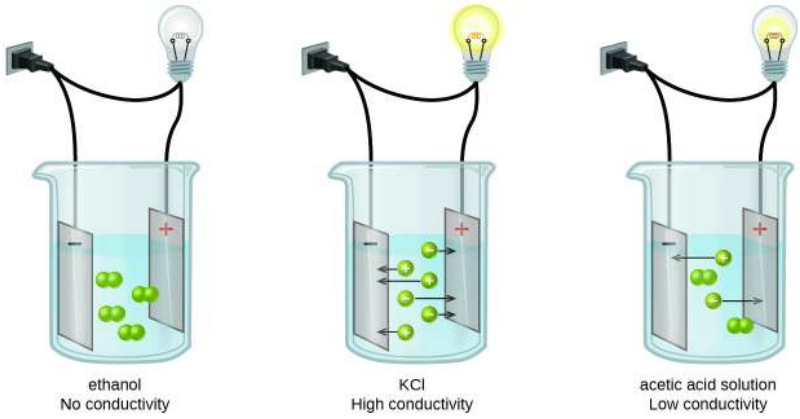


Figure **7.7.1**: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Ionic Electrolytes

Water and other polar molecules are characterized by a slightly positive region and a slightly negative region and are therefore attracted to ions, as shown in Figure 7.7.2. The electrostatic attraction between an ion and a molecule with a dipole is called an ion-dipole attraction. These attractions play an important role in the dissolution of ionic compounds in water, which will be later discussed in Chapter 14.

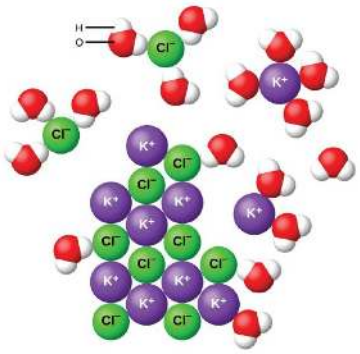


Figure **7.7.2**: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

Let us consider what happens at the microscopic level when we add solid KCl to water. Ion-dipole forces attract the slightly positive (hydrogen) end of the polar water molecules to the negative chloride ions at the surface of the solid, and they attract the slightly negative (oxygen) end to the positive potassium ions. The water molecules penetrate between individual K^+ and Cl^- ions and surround them, reducing the strong interionic forces that bind the ions together and letting them move off into solution as solvated ions, as Figure 7.7.2 shows. The reduction of the electrostatic attraction permits the independent motion of each hydrated ion in a dilute solution, resulting in an increase in the disorder of the system, as the ions change from their fixed and ordered positions in the crystal to mobile and much more disordered states in solution. This increased disorder is responsible for the dissolution of many ionic compounds, including KCl, which dissolve with absorption of heat.

In other cases, the electrostatic attractions between the ions in a crystal are so large, or the ion-dipole attractive forces between the ions and water molecules are so weak, that the increase in disorder cannot compensate for the energy required to separate the ions, and the crystal is insoluble. Such is the case for compounds such as calcium carbonate (limestone), calcium phosphate (the inorganic component of bone), and iron oxide (rust).

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write an equation showing an exchange of ions; but both products, sodium chloride and ammonium nitrate, are soluble and remain in the solution as ions. It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called **solubility rules** (Tables 7.7.1 and 7.7.2).

Table 7.7.1: Solubility Rules for Soluble Substances

Soluble in Water	Important Exceptions (Insoluble)
All Group IA (alkali metals) and ammonium compounds (containing Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+)	none
All nitrates, chlorates, and acetates (containing NO_3^-, ClO_3^-, $C_2H_3O_2^-$)	none
Most sulfates (containing SO_4^{2-})	Sulfates of Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, and Pb^{2+}
Most chlorides, bromides, and iodides (containing Cl^-, Br^-, I^-)	Chlorides, bromides, and iodides of Ag^+, Pb^{2+}, and Hg_2^{2+}

Table 7.7.2: Solubility Rules for Insoluble Substances

Insoluble in Water	Important Exceptions (Soluble)
Most carbonates, oxalates, and phosphates (containing CO_3^{2-}, $C_2O_4^{2-}$, PO_4^{3-})	Group IA (alkali metals) and ammonium compounds
Most hydroxides (containing OH^-)	Group IA (alkali metals) and ammonium compounds Hydroxides of Ca^{2+}, Sr^{2+}, Ba^{2+} sparingly soluble – though usually considered insoluble
Most sulfides (containing S^{2-})	Group IA (alkali metals) and ammonium compounds Sulfides of Ca^{2+}, Sr^{2+}, Ba^{2+}

✅ Example 7.7.1: Solubility

Classify each compound as soluble or insoluble

A. $Zn(NO3)2$
B. $PbBr_2$
C. $Sr3(PO4)2$

Solution

A. Nitrates are soluble in water with no exceptions, so $Zn(NO_3)_2$ is soluble.
B. Most bromides are soluble in water. However, combinations with Pb^{2+} are an exception, so $PbBr_2$ is insoluble.
C. Most phosphates are insoluble and there is no exception when combined with Sr^{2+}, so $Sr_3(PO_4)_2$ is insoluble.

✏️ Exercise 7.7.1: Solubility

Classify each compound as soluble or insoluble.

A. $Mg(OH)_2$

B. KBr
C. $Pb(NO_3)_2$

Answer A

insoluble

Answer B

soluble

Answer C

soluble

Summary

- Substances that dissolve in water to yield ions are called electrolytes.
- Nonelectrolytes are substances that do not produce ions when dissolved in water.
- Solubility rules allow prediction of what products will be insoluble in water.

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7.8: Precipitation Reactions

Learning Objectives

- To identify a precipitation reaction and predict solubility.

A precipitation reaction is a reaction that yields an insoluble product – a precipitate – when two solutions are mixed. When a colorless solution of silver nitrate is mixed with a yellow-orange solution of potassium dichromate, a reddish precipitate of silver dichromate is produced (see Video 7.8.1).

$$2\ AgNO_3\ (aq) + K_2Cr_2O_7\ (aq) \rightarrow Ag_2Cr_2O_7\ (s) + 2\ KNO_3\ (aq)$$

This chemical equation has the general form of an exchange reaction:

$$\mathbf{AC + BD \rightarrow \underset{insoluble}{AD} + BC} \tag{7.8.1}$$

Thus precipitation reactions are a subclass of exchange reactions that occur between ionic compounds when one of the products is insoluble. Because both components of each compound change partners, such reactions are sometimes called double-replacement reactions. Precipitation reactions are used to isolate metals that have been extracted from their ores, and to recover precious metals for recycling.



Video **7.8.1**: Mixing potassium dichromate and silver nitrate together to initiate a precipitation reaction.

Just as important as predicting the product of a reaction is knowing when a chemical reaction will *not* occur. Simply mixing solutions of two different chemical substances does *not* guarantee that a reaction will take place. For example, if 500 mL of an aqueous sodium chloride solution is mixed with 500 mL of an aqueous potassium bromide solution, the final solution has a volume of 1000 mL and contains a mixture of Na^+ (*aq*), Cl^- (*aq*), K^+ (*aq*), and Br^- (*aq*) ions. As you will see in (Figure 7.8.1), none of these species reacts with any of the others. When these solutions are mixed, the only effect is to dilute each solution with the other.

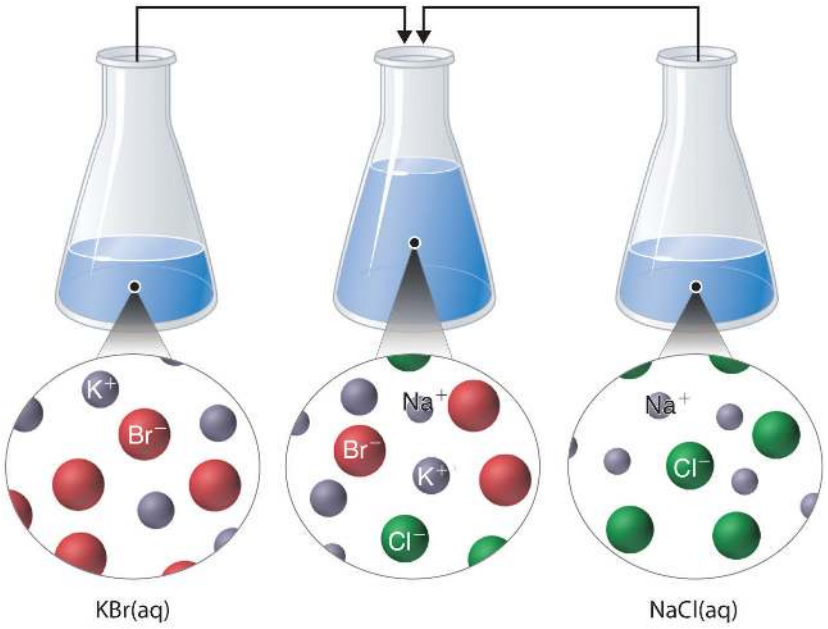


Figure **7.8.1**: The effect of mixing aqueous KBr and NaCl solutions. Because no net reaction occurs, the only effect is to dilute each solution with the other. (Water molecules are omitted from molecular views of the solutions for clarity.)

Predicting Precipitation Reactions

A precipitation reaction occurs when a solid precipitate forms after mixing two strong electrolyte solutions. As stated previously, if none of the species in the solution reacts then no net reaction occurs.

⚡ A Case for No Reaction

When two aqueous ionic compounds are mixed together, there will be no reaction when both of the potential products would also be aqueous ionic compounds.

Let's see what happens when aqueous solutions of barium chloride and lithium sulfate are mixed:

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow$$

Since barium chloride and lithium sulfate are both strong electrolytes, each dissociates completely in water resulting in a solution that contains all of the constituent anions and cations. Barium chloride is comprised of barium ions, Ba^{2+}, and chloride ions, Cl^-, while lithium sulfate is comprised of lithium ions, Li^+, and sulfate ions, SO_4^{2-}. The anions exchange partners.



Barium ions pair up with sulfate ions forming barium sulfate, $BaSO_4$ (Ba^{2+} balances the charge of SO_4^{2-}). Lithium ions pair up with chloride ions forming LiCl (Li^+ balances the charge of Cl^-).

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4 + LiCl$$

Referring to the solubility rules found in the previous section, $BaSO_4$ is insoluble in water resulting in a phase label of (*s*), while LiCl is soluble in water resulting in a phase label of (*aq*).

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + LiCl\,(aq)$$

At this point, the equation may be balanced by placing a coefficient of "2" in front of LiCl (*aq*).

$$BaCl_2\,(aq) + Li_2SO_4\,(aq) \rightarrow BaSO_4\,(s) + 2\,LiCl\,(aq)$$

Although soluble barium salts are toxic, $BaSO_4$ is so insoluble that it can be used to diagnose stomach and intestinal problems without being absorbed into tissues. An outline of the digestive organs appears on x-rays of patients who have been given a "barium milkshake" or a "barium enema", a suspension of very fine $BaSO_4$ particles in water.



Figure **7.8.2**: An x-ray of the digestive organs of a patient who has swallowed a "barium milkshake." A barium milkshake is a suspension of very fine $BaSO_4$ particles in water; the high atomic mass of barium makes it opaque to x-rays. (OpenStax, CC BY 4.0, via Wikimedia Commons)

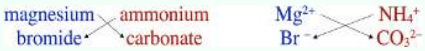
✅ Example 7.8.1

Complete and balance the following equation. If no reaction takes place, write "No Reaction" after the arrow.

$$MgBr_2\,(aq) + (NH_4)_2CO_3\,(aq) \rightarrow$$

Solution

$MgBr_2$ is called magnesium bromide and is comprised of magnesium ions, Mg^{2+}, and bromide ions, Br^-, while $(NH_4)_2CO_3$ is called ammonium carbonate and is comprised of ammonium ions, NH_4^+, and carbonate ions, CO_3^{2-}. The anions exchange partners.



Magnesium ions pair up with carbonate ions forming magnesium carbonate, $MgCO_3$ (Mg^{2+} balances the charge of CO_3^{2-}). Ammonium ions pair up with bromide ions forming NH_4Br (NH_4^+ balances the charge of Br^-).

$$MgBr_2\,(aq) + (NH_4)_2CO_3\,(aq) \rightarrow MgCO_3 + NH_4Br$$

Referring to the solubility rules, $MgCO_3$ is insoluble in water resulting in a phase label of (*s*), while NH_4Br is soluble in water resulting in a phase label of (*aq*).

$$MgBr_2\,(aq) + (NH_4)_2CO_3\,(aq) \rightarrow MgCO_3\,(s) + NH_4Br\,(aq)$$

At this point, the equation may be balanced by placing a coefficient of "2" in front of NH_4Br (*aq*).

$$MgBr_2\,(aq) + (NH_4)_2CO_3\,(aq) \rightarrow MgCO_3\,(s) + 2\,NH_4Br\,(aq)$$

✅ Example 7.8.2

Write a balanced chemical equation that shows what happens when aqueous solutions of sodium carbonate and potassium iodide are mixed. If there is no reaction, write "No Reaction" after the arrow.

Solution

Sodium carbonate is comprised of sodium ions, Na^+, and carbonate ions, CO_3^{2-}, while potassium iodide is comprised of potassium ions, K^+, and iodide ions, I^-. The anions exchange partners.



Sodium ions pair up with iodide ions forming sodium iodide, NaI (Na^+ balances the charge of I^-). Potassium ions pair up with carbonate ions forming K_2CO_3 (two K^+ balances the charge of one CO_3^{2-}).

$$Na_2CO_3\,(aq) + KI\,(aq) \rightarrow NaI + K_2CO_3$$

Referring to the solubility rules, both NaI and K_2CO_3 are soluble in water resulting in a phase label of (*aq*).

$$Na_2CO_3\,(aq) + KI\,(aq) \rightarrow NaI\,(aq) + K_2CO_3\,(aq)$$

Since this equation shows two aqueous ionic compounds being mixed together to potentially yield two aqueous ionic compounds, no reaction will take place.

$$Na_2CO_3\ (aq) + KI\ (aq) \rightarrow \text{No Reaction}$$

✅ Example 7.8.3

Write a balanced chemical equation that shows what happens when aqueous solutions of iron(III) nitrate and sodium hydroxide are mixed. If there is no reaction, write "No Reaction" after the arrow.

Solution

Iron(III) nitrate is comprised of iron(III) ions, Fe^{3+}, and nitrate ions, NO_3^-, while sodium hydroxide is comprised of sodium ions, Na^+, and hydroxide ions, OH^-. The anions exchange partners.



Iron(III) ions pair up with hydroxide ions forming iron(III) hydroxide, $Fe(OH)_3$ (one Fe^{3+} balances the charge of three OH^-). Sodium ions pair up with nitrate ions forming $NaNO_3$ (Na^+ balances the charge of NO_3^-).

$$Fe(NO_3)_3\ (aq) + NaOH\ (aq) \rightarrow Fe(OH)_3 + NaNO_3$$

Referring to the solubility rules, $Fe(OH)_3$ is insoluble in water resulting in a phase label of (*s*), while $NaNO_3$ is soluble in water resulting in a phase label of (*aq*).

$$Fe(NO_3)_3\ (aq) + NaOH\ (aq) \rightarrow Fe(OH)_3\ (s) + NaNO_3\ (aq)$$

At this point, the equation may be balanced by placing a coefficient of "3" in front of NaOH (*aq*) and $NaNO_3$ (*aq*).

$$Fe(NO_3)_3\ (aq) + 3\ NaOH\ (aq) \rightarrow Fe(OH)_3\ (s) + 3\ NaNO_3\ (aq)$$

✏️ Exercise 7.8.1

Complete and balance each equation. If no reaction takes place, write "No Reaction" after the arrow.

A. $Li_3PO_4\ (aq) + KOH\ (aq) \rightarrow$
B. $MgSO_4\ (aq) + SrI_2\ (aq) \rightarrow$

Answer A

$Li_3PO_4\ (aq) + KOH\ (aq) \rightarrow$ No Reaction

Answer B

$MgSO_4\ (aq) + SrI_2\ (aq) \rightarrow SrSO_4\ (s) + MgI_2\ (aq)$

✏️ Exercise 7.8.2

Write a balanced chemical equation that shows what happens when aqueous solutions of each pair are mixed. If no reaction occurs, write "No Reaction" after the arrow.

A. Potassium phosphate and calcium chloride
B. Zinc nitrate and sodium carbonate

Answer A

$2\ K_3PO_4\ (aq) + 3\ CaCl_2\ (aq) \rightarrow Ca_3(PO_4)_2\ (s) + 6\ KCl\ (aq)$

Answer B

$Zn(NO_3)_2\ (aq) + Na_2CO_3\ (aq) \rightarrow ZnCO_3\ (s) + 2\ NaNO_3\ (aq)$

Summary

- In a **precipitation reaction**, a subclass of exchange reactions, an insoluble material (a **precipitate**) forms when two electrolyte solutions are mixed.
- To predict the product of a precipitation reaction, all species initially present in the solutions are identified, as are any combinations likely to produce an insoluble salt.

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7.9: Acid-Base and Gas Evolution Reactions

Learning Objectives

- Identify when a reaction will evolve a gas.

Neutralization Reactions

Acids and bases react chemically with each other to form **salts**. A salt is a general chemical term for any ionic compound formed from an acid and a base. In reactions where the acid is a hydrogen-ion-containing compound and the base is a hydroxide-ion-containing compound, water is also a product. The general reaction is as follows:

$$\text{acid} + \text{base} \rightarrow \text{salt} + \text{water}$$

The reaction of acid and base to make water and a salt is called **neutralization**. Like any chemical equation, a neutralization chemical equation must be properly balanced. For example, the neutralization reaction between sodium hydroxide and hydrochloric acid is as follows:

$$\text{NaOH}\,(aq) + \text{HCl}\,(aq) \rightarrow \text{NaCl}\,(aq) + \text{H}_2\text{O}\,(l)$$

with coefficients all understood to be one. The neutralization reaction between sodium hydroxide and sulfuric acid is as follows:

$$2\ \text{NaOH}\,(aq) + \text{H}_2\text{SO}_4\,(aq) \rightarrow \text{Na}_2\text{SO}_4\,(aq) + 2\ \text{H}_2\text{O}\,(l)$$

✅ Example 7.9.1: Neutralizing Nitric Acid

Nitric acid, $\text{HNO}_3\,(aq)$, can be neutralized by calcium hydroxide, $\text{Ca(OH)}_2\,(aq)$. Write a balanced chemical equation for the reaction between these two compounds and identify the salt that it produces.

Solution

Write the unbalanced equation. This is a double replacement reaction, so the cations and anions swap to create new products. Referring to the solubility rules, $\text{Ca(NO}_3)_2$ is soluble in water resulting in a phase label of (aq). Water, H_2O, is a liquid at room temperature resulting in a phase label of (l).

$$\text{Ca(OH)}_2\,(aq) + \text{HNO}_3\,(aq) \rightarrow \text{Ca(NO}_3)_2\,(aq) + \text{H}_2\text{O}\,(l)$$

At this point, the equation may be balanced by placing a coefficient of "2" in front of $\text{HNO}_3\,(aq)$ and $\text{H}_2\text{O}\,(l)$.

$$\text{Ca(OH)}_2\,(aq) + 2\ \text{HNO}_3\,(aq) \rightarrow \text{Ca(NO}_3)_2\,(aq) + 2\ \text{H}_2\text{O}\,(l)$$

The salt formed is calcium nitrate, $\text{Ca(NO}_3)_2\,(aq)$.

✏️ Exercise 7.9.1

Hydrobromic acid, $\text{HBr}\,(aq)$, may be neutralized with potassium hydroxide, $\text{KOH}\,(aq)$. Write a balanced chemical equation for the reaction between these two compounds and identify the salt that it produces.

Answer

$$\text{KOH}\,(aq) + \text{HBr}\,(aq) \rightarrow \text{KBr}\,(aq) + \text{H}_2\text{O}\,(l)$$

The salt is potassium bromide, $\text{KBr}\,(aq)$.

Gas Evolving Reactions

Figure 7.9.1 shows an apparatus that may be used for collecting gases in the laboratory.

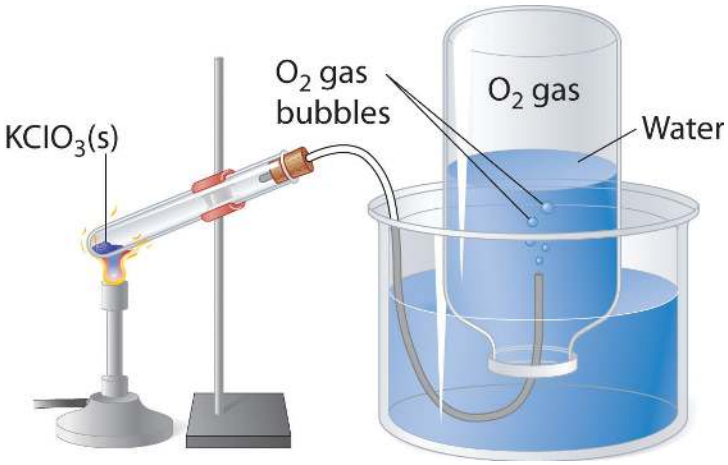


Figure **7.9.1**: An apparatus for collecting gases by the displacement of water.

A gas evolution reaction is a chemical process that produces a gas, such as hydrogen or carbon dioxide. In the following examples, an acid reacts with a carbonate, producing salt, carbon dioxide, and water, respectively. For example, nitric acid reacts with sodium carbonate to form sodium nitrate, carbon dioxide, and water (Table 7.9.1):

$$2\ HNO_3\ (aq) + Na_2CO_3\ (aq) \rightarrow 2\ NaNO_3\ (aq) + CO_2\ (g) + H_2O\ (l)$$

Hydrochloric acid reacts with calcium carbonate to form calcium chloride, carbon dioxide, and water:

$$2\ HCl\ (aq) + CaCO_3\ (aq) \rightarrow CaCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$$

Table 7.9.1: Types of Compounds That Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas Evolved	Example
sulfide	none	H_2S	$2\ HCl\ (aq) + K_2S\ (aq) \rightarrow H_2S\ (g) + 2\ KCl\ (aq)$
carbonates and bicarbonates	H_2CO_3	CO_2	$2\ HCl\ (aq) + K_2CO_3\ (aq) \rightarrow H_2CO_3\ (aq) + 2\ KCl\ (aq)$ Final Equation: $2\ HCl\ (aq) + K_2CO_3\ (aq) \rightarrow H_2O\ (l) + CO_2\ (g) + 2\ KCl\ (aq)$
sulfites and bisulfites	H_2SO_3	SO_2	$2\ HCl\ (aq) + K_2SO_3\ (aq) \rightarrow H_2SO_3\ (aq) + 2\ KCl\ (aq)$ Final Equation: $2\ HCl\ (aq) + K_2SO_3\ (aq) \rightarrow H_2O\ (l) + SO_2\ (g) + 2\ KCl\ (aq)$
ammonia	NH_4OH	NH_3	$NH_4Cl\ (aq) + KOH\ (aq) \rightarrow NH_4OH\ (aq) + 2\ KCl\ (aq)$ Final Equation: $NH_4Cl\ (aq) + KOH\ (aq) \rightarrow H_2O\ (l) + NH_3\ (g) + 2\ KCl\ (aq)$

The oxidation of metals by strong acids is another common example of a gas evolution reaction. This reaction will yield a metal salt and hydrogen gas.

$$2\ HCl\ (aq) + Zn\ (s) \rightarrow ZnCl_2\ (aq) + H_2\ (g)$$

Here, hydrochloric acid oxidizes zinc to produce an aqueous metal salt and hydrogen gas bubbles.

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7.11: Exercises

7.3: Chemical Equations

1. From the statement “nitrogen and hydrogen react to produce ammonia,” identify the reactants and the products.

Answer

reactants = nitrogen (N_2) and hydrogen (H_2); products = ammonia (NH_3)

2. From the statement “sodium metal reacts with water to produce sodium hydroxide and hydrogen,” identify the reactants and the products.

Answer

reactants = sodium (Na) and water (H_2O); products = sodium hydroxide (NaOH) and hydrogen (H_2)

3. From the statement “magnesium hydroxide reacts with nitric acid to produce magnesium nitrate and water,” identify the reactants and the products.

Answer

reactants = magnesium hydroxide ($Mg(OH)_2$) and nitric acid (HNO_3); products = magnesium nitrate ($Mg(NO_3)_2$) and water (H_2O)

4. From the statement “propane reacts with oxygen to produce carbon dioxide and water,” identify the reactants and the products.

Answer

reactants = propane (C_3H_8) and oxygen (O_2); products = carbon dioxide (CO_2) and water (H_2O)

5. Chemical equations can also be used to represent physical processes. Write a chemical equation for the boiling of water, including the proper phase labels.

Answer

$H_2O(l) \rightarrow H_2O(g)$

6. Chemical equations can also be used to represent physical processes. Write a chemical equation for the freezing of water, including the proper phase labels.

Answer

$H_2O(l) \rightarrow H_2O(s)$

7.4: Balancing Chemical Equations

7. What does it mean to say an equation is balanced? Why is it important for an equation to be balanced?

Answer

An equation is balanced when the same number of each element is represented on the reactant and product sides. Equations must be balanced to accurately reflect the law of conservation of matter.

8. Explain why $4\ Na(s) + 2\ Cl_2(g) \rightarrow 4\ NaCl(s)$ should not be considered a proper chemical equation.

Answer

It does not have the lowest possible whole number coefficients.

9. Explain why $H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(l)$ should not be considered a proper chemical equation.

Answer

It has a fractional coefficient.

10. Write and balance the chemical equation described by Exercise 1.

Answer

$N_2 + 3\ H_2 \rightarrow 2\ NH_3$

11. Write and balance the chemical equation described by Exercise 2.

Answer

$2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$

12. Write and balance the chemical equation described by Exercise 3.

Answer

$Mg(OH)_2 + 2\ HNO_3 \rightarrow Mg(NO_3)_2 + 2\ H_2O$

13. Write and balance the chemical equation described by Exercise 4. The formula for propane is C_3H_8.

Answer

$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

14. Balance the following equations:

a. $NaClO_3 \rightarrow NaCl + O_2$
b. $N_2 + H_2 \rightarrow N_2H_4$
c. $Al + O_2 \rightarrow Al_2O_3$
d. $C_2H_4 + O_2 \rightarrow CO_2 + H_2$

Answer

a. $2\ NaClO_3 \rightarrow 2\ NaCl + 3\ O_2$
b. $N_2 + 2\ H_2 \rightarrow N_2H_4$

c. $4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3$
d. $C_2H_4 + 2\ O_2 \rightarrow 2\ CO_2 + 2\ H_2$

15. Balance the following equations:

a. $PCl_5(s) + H_2O(l) \rightarrow POCl_3(l) + HCl(aq)$
b. $Cu(s) + HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + H_2O(l) + NO(g)$
c. $H_2(g) + I_2(s) \rightarrow HI(s)$
d. $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$
e. $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
f. $(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + H_2O(g)$
g. $P_4(s) + Cl_2(g) \rightarrow PCl_3(l)$
h. $PtCl_4(s) \rightarrow Pt(s) + Cl_2(g)$

Answer

a. $PCl_5(s) + H_2O(l) \rightarrow POCl_3(l) + 2\ HCl(aq)$
b. $3\ Cu(s) + 8\ HNO_3(aq) \rightarrow 3\ Cu(NO_3)_2(aq) + 4\ H_2O(l) + 2\ NO(g)$
c. $H_2(g) + I_2(s) \rightarrow 2HI(s)$
d. $4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$
e. $2\ Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$
f. $(NH_4)_2Cr_2O_7(s) \rightarrow Cr_2O_3(s) + N_2(g) + 4\ H_2O(g)$
g. $P_4(s) + 6\ Cl_2(g) \rightarrow 4\ PCl_3(l)$
h. $PtCl_4(s) \rightarrow Pt(s) + 2\ Cl_2(g)$

16. Balance the following equations:

a. $Ag(s) + O_2(g) + H_2S(g) \rightarrow Ag_2S(s) + H_2O(l)$
b. $P_4(s) + O_2(g) \rightarrow P_4O_{10}(s)$
c. $Pb(s) + H_2O(l) + O_2(g) \rightarrow Pb(OH)_2(s)$
d. $Fe(s) + H_2O(l) \rightarrow Fe_3O_4(s) + H_2(g)$
e. $Sc_2O_3(s) + SO_3(l) \rightarrow Sc_2(SO_4)_3(s)$
f. $Ca_3(PO_4)_2(s) + H_3PO_4(aq) \rightarrow Ca(H_2PO_4)_2(aq)$
g. $Al(s) + H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(s) + H_2(g)$
h. $TiCl_4(s) + H_2O(g) \rightarrow TiO_2(s) + HCl(g)$

Answer

a. $4\ Ag(s) + O_2(g) + 2\ H_2S(g) \rightarrow 2\ Ag_2S(s) + 2\ H_2O(l)$
b. $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$
c. $2\ Pb(s) + 2\ H_2O(l) + O_2(g) \rightarrow 2\ Pb(OH)_2(s)$
d. $3\ Fe(s) + 4\ H_2O(l) \rightarrow Fe_3O_4(s) + 4\ H_2(g)$
e. $Sc_2O_3(s) + 3\ SO_3(l) \rightarrow Sc_2(SO_4)_3(s)$
f. $Ca_3(PO_4)_2(s) + 4\ H_3PO_4(aq) \rightarrow 3\ Ca(H_2PO_4)_2(aq)$
g. $2\ Al(s) + 3\ H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(s) + 3\ H_2(g)$
h. $TiCl_4(s) + 2\ H_2O(g) \rightarrow TiO_2(s) + 4\ HCl(g)$

17. Write a balanced molecular equation describing each of the following chemical reactions.

a. Solid calcium carbonate is heated and decomposes to solid calcium oxide and carbon dioxide gas.
b. Gaseous butane, C_4H_{10}, reacts with diatomic oxygen gas to yield gaseous carbon dioxide and water vapor.

c. Aqueous solutions of magnesium chloride and sodium hydroxide react to produce solid magnesium hydroxide and aqueous sodium chloride.
d. Water vapor reacts with sodium metal to produce solid sodium hydroxide and hydrogen gas.

Answer

a. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
b. $2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$
c. $MgCl_2(aq) + 2\ NaOH(aq) \rightarrow Mg(OH)_2(s) + 2\ NaCl(aq)$
d. $2\ H_2O(g) + 2\ Na(s) \rightarrow 2\ NaOH(s) + H_2(g)$

18. Write a balanced equation describing each of the following chemical reactions.

a. Solid potassium chlorate, $KClO_3$, decomposes to form solid potassium chloride and diatomic oxygen gas.
b. Solid aluminum metal reacts with solid diatomic iodine to form solid Al_2I_6.
c. When solid sodium chloride is added to aqueous sulfuric acid, hydrogen chloride gas and aqueous sodium sulfate are produced.
d. Aqueous solutions of phosphoric acid and potassium hydroxide react to produce aqueous potassium dihydrogen phosphate ($H_2PO_4^-$)and liquid water.

Answer

a. $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$
b. $2\ Al(s) + 3\ I_2(s) \rightarrow Al_2I_6(s)$
c. $2\ NaCl(s) + H_2SO_4(aq) \rightarrow 2\ HCl(g) + Na_2SO_4(aq)$
d. $H_3PO_4(aq) + KOH(aq) \rightarrow KH_2PO_4(aq) + H_2O(l)$

19. Colorful fireworks often involve the decomposition of barium nitrate and potassium chlorate and the reaction of the metals magnesium, aluminum, and iron with oxygen.

a. Write the formulas of barium nitrate and potassium chlorate.
b. The decomposition of solid potassium chlorate leads to the formation of solid potassium chloride and diatomic oxygen gas. Write an equation for the reaction.
c. The decomposition of solid barium nitrate leads to the formation of solid barium oxide, diatomic nitrogen gas, and diatomic oxygen gas. Write an equation for the reaction.
d. Write separate equations for the reactions of the solid metals magnesium, aluminum, and iron with diatomic oxygen gas to yield the corresponding metal oxides. (Assume the iron oxide contains Fe^{3+} ions.)

Answer

a. $Ba(NO_3)_2$; $KClO_3$
b. $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$
c. $2\ Ba(NO_3)_2(s) \rightarrow 2\ BaO(s) + 2\ N_2(g) + 5\ O_2(g)$
d. $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$; $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$; $4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$

20. Aqueous hydrogen fluoride (hydrofluoric acid) is used to etch glass and to analyze minerals for their silicon content. Hydrogen fluoride will also react with sand (silicon dioxide). Write a balanced equation for the reaction of solid silicon dioxide with hydrofluoric acid to yield gaseous silicon tetrafluoride and liquid water.

Answer

$SiO_2(s) + 4\ HF(aq) \rightarrow SiF_4(g) + 2\ H_2O(l)$

21. The mineral fluorite (calcium fluoride) occurs extensively in Illinois. Solid calcium fluoride can also be prepared by the reaction of aqueous solutions of calcium chloride and sodium fluoride, yielding aqueous sodium chloride as the other product. Write a balanced equation for this reaction.

Answer

$CaCl_2(aq) + 2\ NaF(aq) \rightarrow CaF_2(s) + 2\ NaCl(aq)$

22. A novel process for obtaining magnesium from sea water involves several reactions. Write a balanced chemical equation for each step of the process.
    a. The first step is the decomposition of solid calcium carbonate from seashells to form solid calcium oxide and gaseous carbon dioxide.
    b. The second step is the formation of solid calcium hydroxide as the only product from the reaction of the solid calcium oxide with liquid water.
    c. Solid calcium hydroxide is then added to the seawater, reacting with dissolved magnesium chloride to yield solid magnesium hydroxide and aqueous calcium chloride.
    d. The solid magnesium hydroxide is added to a hydrochloric acid solution, producing dissolved magnesium chloride and liquid water.
    e. Finally, the magnesium chloride is melted and electrolyzed to yield liquid magnesium metal and diatomic chlorine gas.

Answer

a. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
b. $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$
c. $Ca(OH)_2(s) + MgCl_2(aq) \rightarrow Mg(OH)_2(s) + CaCl_2(aq)$
d. $Mg(OH)_2(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + 2\ H_2O(l)$
e. $MgCl_2(l) \rightarrow Mg(l) + Cl_2(g)$

7.5: Classifying Chemical Reactions

23. What are the general characteristics that help you recognize single-replacement reactions?

Answer

The reactants have an element and a compound which combine to form a new element and a new compound.

24. What are the general characteristics that help you recognize double-replacement reactions?

Answer

Two compounds (typically ionic) switch partners to form two new compounds.

25. Classify each of the following reactions as synthesis, decomposition, combustion, single-replacement, or double-replacement.
    a. $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$
    b. $CaO + CO_2 \rightarrow CaCO_3$
    c. $2\ HBr + Cl_2 \rightarrow 2\ HCl + Br_2$
    d. $C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$
    e. $CaCO_3 \rightarrow CaO + CO_2$

Answer

a. double-replacement
b. synthesis
c. single-replacement
d. combustion
e. decomposition

26. Classify each of the following reactions as synthesis, decomposition, combustion, single-replacement, or double-replacement.

a. $HCl + NaOH \rightarrow NaCl + H_2O$
b. $C_5H_{12} + 8\ O_2 \rightarrow 5\ CO_2 + 6\ H_2O$
c. $CH_4 + 2\ F_2 \rightarrow CF_4 + 2\ H_2$
d. $Na_2O + CO_2 \rightarrow Na_2CO_3$
e. $H_2SO_3 \rightarrow H_2O + SO_2$

Answer

a. double-replacement
b. combustion
c. single-replacement
d. synthesis
e. decomposition

27. Classify each of the following reactions as synthesis, decomposition, combustion, single-replacement, or double-replacement.

a. $H_2 + Cl_2 \rightarrow 2\ HCl$
b. $Na_2CO_3 \rightarrow Na_2O + CO_2$
c. $Pb(NO_3)_2 + H_2SO_4 \rightarrow PbSO_4 + 2\ HNO_3$
d. $2\ HCl + Zn \rightarrow ZnCl_2 + H_2$
e. $C_2H_4 + O_2 \rightarrow C_2H_4O_2$

Answer

a. synthesis
b. decomposition
c. double-replacement
d. single-replacement
e. synthesis

28. Classify each of the following reactions as synthesis, decomposition, combustion, single-replacement, or double-replacement.

a. $Ca + Br_2 \rightarrow CaBr_2$
b. $C_6H_{12} + 9\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$
c. $H_2O + C \rightarrow CO + H_2$
d. $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$
e. $Al(OH)_3 + 3\ HBr \rightarrow AlBr_3 + 3\ H_2O$

Answer

a. synthesis
b. combustion
c. single-replacement

d. decomposition
e. double-replacement

29. Use the following equations to answer the next three questions:

i. $H_2O(s) \rightarrow H_2O(l)$
ii. $2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$
iii. $CH_3OH(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

a. Which equation describes a physical change?
b. Which equation identifies the reactants and products of a combustion reaction?
c. Which equation is not balanced?

Answer

a. equation i
b. equation iii
c. equation iii

7.6: Combustion Reactions

30. Complete and balance each combustion equation.

a. $C_4H_9OH + O_2 \rightarrow ?$
b. $CH_3SH + O_2 \rightarrow ?$

Answer

a. $C_4H_9OH + 6\ O_2 \rightarrow 4\ CO_2 + 5\ H_2O$
b. $CH_3SH + 3\ O_2 \rightarrow CO_2 + 2\ H_2O + SO_2$

7.7: Solubility Rules for Ionic Compounds

31. Classify each of the following compounds as soluble or insoluble in water.

a. AgBr
b. Na_3PO_4
c. CaC_2O_4
d. K_2SO_4

Answer

a. insoluble
b. soluble
c. insoluble
d. soluble

32. Classify each of the following compounds as soluble or insoluble in water.

a. $Pb(OH)_2$
b. SrS
c. $AgC_2H_3O_2$
d. $Fe(NO_3)_2$

Answer

a. insoluble
b. soluble
c. soluble
d. soluble

33. Classify each of the following compounds as soluble or insoluble in water.

a. $BaSO_4$
b. Li_2CO_3
c. NaOH
d. Hg_2I_2

Answer

a. insoluble
b. soluble
c. soluble
d. insoluble

7.8: Precipitation Reactions

34. Explain what is wrong with this double-replacement reaction: $NaCl(aq) + KBr(aq) \rightarrow NaK(aq) + ClBr(aq)$

Answer

The cations are combined in one product and the anions are combined in the other. The products each need to have one cation and one anion.

35. Predict the products of and balance this double-replacement reaction: $Na_2SO_4(aq) + SrCl_2(aq) \rightarrow ?$

Answer

$Na_2SO_4(aq) + SrCl_2(aq) \rightarrow SrSO_4(s) + 2\ NaCl(aq)$

36. Predict the products of and balance this double-replacement reaction: $BaCl_2(aq) + AgNO_3(aq) \rightarrow ?$

Answer

$BaCl_2(aq) + 2\ AgNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + 2\ AgCl(s)$

37. For each of the following reactions, predict the products and write the balanced chemical equation. If no reaction occurs, write "no reaction."

a. $Zn(NO_3)_2(aq) + NaOH(aq) \rightarrow ?$
b. $Ca(C_2H_3O_2)_2(aq) + HNO_3(aq) \rightarrow ?$
c. $Na_2CO_3(aq) + Sr(NO_3)_2(aq) \rightarrow ?$
d. $Pb(NO_3)_2(aq) + KBr(aq) \rightarrow ?$

Answer

a. $Zn(NO_3)_2(aq) + 2\ NaOH(aq) \rightarrow Zn(OH)_2(s) + 2\ NaNO_3(aq)$
b. no reaction

c. $Na_2CO_3(aq) + Sr(NO_3)_2(aq) \rightarrow 2\ NaNO_3(aq) + SrCO_3(s)$
d. $Pb(NO_3)_2(aq) + 2\ KBr(aq) \rightarrow PbBr_2(s) + 2\ KNO_3(aq)$

38. For each of the following reactions, predict the products and write the balanced chemical equation. If no reaction occurs, write "no reaction."

a. $AgClO_3(aq) + KI(aq) \rightarrow ?$
b. $K_2O(aq) + Na_2CO_3(aq) \rightarrow ?$
c. $Na_2CO_3(aq) + Sr(NO_3)_2(aq) \rightarrow ?$
d. $(NH_4)_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow ?$

Answer

a. $AgClO_3(aq) + KI(aq) \rightarrow AgI(s) + KClO_3(aq)$
b. no reaction
c. $Na_2CO_3(aq) + Sr(NO_3)_2(aq) \rightarrow 2\ NaNO_3(aq) + SrCO_3(s)$
d. $(NH_4)_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow 2\ NH_4NO_3(aq) + BaSO_4(s)$

39. For each of the following reactions, predict the products and write the balanced chemical equation. If no reaction occurs, write "no reaction."

a. $K_3PO_4(aq) + SrCl_2(aq) \rightarrow ?$
b. $NaOH(aq) + MgCl_2(aq) \rightarrow ?$
c. $KC_2H_3O_2(aq) + Li_2CO_3(aq) \rightarrow ?$
d. $KOH(aq) + AgNO_3(aq) \rightarrow ?$

Answer

a. $2\ K_3PO_4(aq) + 3\ SrCl_2(aq) \rightarrow 6\ KCl(aq) + Sr_3(PO_4)_2(s)$
b. $2\ NaOH(aq) + MgCl_2(aq) \rightarrow 2\ NaCl(aq) + Mg(OH)_2(s)$
c. no reaction
d. $KOH(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + AgOH(s)$

40. For each of the following reactions, predict the products and write the balanced chemical equation. If no reaction occurs, write "no reaction."

a. $FeCl_2(aq) + AgNO_3(aq) \rightarrow ?$
b. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow ?$
c. $KCl(aq) + NaC_2H_3O_2(aq) \rightarrow ?$
d. $Fe_2(SO_4)_3(aq) + Sr(NO_3)_2(aq) \rightarrow ?$

Answer

a. $FeCl_2(aq) + 2\ AgNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + 2\ AgCl(s)$
b. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaCl(aq)$
c. no reaction
d. $Fe_2(SO_4)_3(aq) + 3\ Sr(NO_3)_2(aq) \rightarrow 2\ Fe(NO_3)_3(aq) + 3\ SrSO_4(s)$

7.9: Acid-Base and Gas Evolution Reactions

41. Write balanced chemical equations for the following neutralization reactions.

a. HCl and KOH

b. H_2SO_4 and KOH
c. H_3PO_4 and $Ni(OH)_2$

Answer

a. $HCl + KOH \rightarrow KCl + H_2O$
b. $H_2SO_4 + 2\ KOH \rightarrow K_2SO_4 + 2\ H_2O$
c. $2\ H_3PO_4 + 3\ Ni(OH)_2 \rightarrow Ni_3(PO_4)_2 + 6\ H_2O$

42. Write balanced chemical equations for the following neutralization reactions.

a. HBr and $Fe(OH)_3$
b. HNO_2 and $Al(OH)_3$
c. $HClO_3$ and $Mg(OH)_2$

Answer

a. $3\ HBr + Fe(OH)_3 \rightarrow FeBr_3 + 3\ H_2O$
b. $3\ HNO_2 + Al(OH)_3 \rightarrow Al(NO_2)_3 + 3\ H_2O$
c. $2\ HClO_3 + Mg(OH)_2 \rightarrow Mg(ClO_3)_2 + 2\ H_2O$

43. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.

a. $HI(aq) + KOH(aq) \rightarrow\ ?$
b. $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow\ ?$

Answer

a. $HI(aq) + KOH(aq) \rightarrow KI(aq) + H_2O(l)$
b. $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow BaSO_4(s) + 2\ H_2O(l)$

44. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.

a. $HNO_3(aq) + Fe(OH)_3(s) \rightarrow\ ?$
b. $H_3PO_4(aq) + LiOH(aq) \rightarrow\ ?$

Answer

a. $3\ HNO_3(aq) + Fe(OH)_3(s) \rightarrow Fe(NO_3)_3(aq) + 3\ H_2O(l)$
b. $H_3PO_4(aq) + 3\ LiOH(aq) \rightarrow Li_3PO_4(aq) + 3\ H_2O(l)$

45. Write a balanced chemical equation for the neutralization reaction between each given acid and base. Include the proper phase labels.

a. $HClO_3(aq) + Zn(OH)_2(s) \rightarrow\ ?$
b. $H_2C_2O_4(s) + Sr(OH)_2(aq) \rightarrow\ ?$

Answer

a. $2\ HClO_3(aq) + Zn(OH)_2(s) \rightarrow Zn(ClO_3)_2(aq) + 2\ H_2O(l)$
b. $H_2C_2O_4(s) + Sr(OH)_2(aq) \rightarrow SrC_2O_4(s) + 2\ H_2O(l)$

46. Write a balanced chemical equation for the following acid-base equations:

a. HCl gas reacts with solid $Ca(OH)_2(s)$.
b. A solution of $Sr(OH)_2$ is added to a solution of HNO_3.
c. A solution of $HClO_4$ is added to a solution of LiOH.
d. Aqueous H_2SO_4 reacts with NaOH.
e. $Ba(OH)_2$ reacts with HF gas.

Answer

a. $2\ HCl(g) + Ca(OH)_2(s) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$
b. $Sr(OH)_2(aq) + 2\ HNO_3(aq) \rightarrow Sr(NO_3)_2(aq) + 2\ H_2O(l)$
c. $HClO_4(aq) + LiOH(aq) \rightarrow LiClO_4(aq) + H_2O(l)$
d. $H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$
e. $Ba(OH)_2(aq) + 2\ HF(g) \rightarrow BaF_2(aq) + 2\ H_2O(l)$

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CHAPTER OVERVIEW

8: Stoichiometry

So far, we have talked about chemical reactions in terms of individual atoms and molecules. Although this works, most of the reactions occurring around us involve much larger amounts of chemicals. Even a tiny sample of a substance will contain millions, billions, or a hundred billion billions of atoms and molecules. How do we compare amounts of substances to each other in chemical terms when it is so difficult to count to a hundred billion billion? Actually, there are ways to do this, which we will explore in this chapter. In doing so, we will increase our understanding of stoichiometry, which is the study of the numerical relationships between the reactants and the products in a balanced chemical reaction.

8.1: Climate Science And Carbon Dioxide
8.2: An Automobile Factory
8.3: Stoichiometry and the Molar Interpretation
8.4: Molar Ratios and Mole-to-Mole Conversions
8.5: Mass-to-Mass Conversions
8.6: Limiting Reactants and Excess Reactants
8.7: Theoretical Yield and Percent Yield
8.8: For Future Use
8.9: Exercises

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8.1: Climate Science And Carbon Dioxide

Carbon dioxide (CO_2) is an important heat-trapping (greenhouse) gas, which is released through human activities such as deforestation and burning fossil fuels, as well as natural processes such as respiration and volcanic eruptions. Figure 8.1.1 shows CO_2 levels during the last three glacial cycles, as reconstructed from ice cores, as well as direct measurements from 2000-2025.

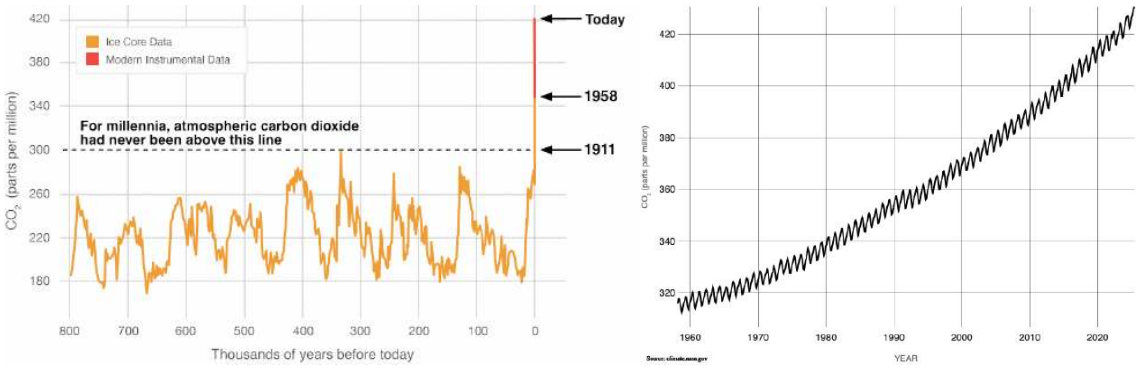


Figure **8.1.1**: Atomopheric CO_2 concentrations based on the comparison of atmospheric samples contained in ice cores and more recent direct measurements, provides evidence that atmospheric CO_2 has increased since the Industrial Revolution. Data as of April 2025. (Credit NOAA via NASA)

Carbon dioxide, CO_2, is the primary greenhouse gas emitted through human activities. In 2015, CO_2 accounted for about 82.2% of all U.S. greenhouse gas emissions from human activities. Carbon dioxide is naturally present in the atmosphere as part of the Earth's carbon cycle (the natural circulation of carbon among the atmosphere, oceans, soil, plants, and animals). Human activities are altering the carbon cycle, both by adding more CO_2 to the atmosphere and by influencing the ability of natural sinks, like forests, to remove CO_2 from the atmosphere. While CO_2 emissions come from a variety of natural sources, human-related emissions are responsible for the increase that has occurred in the atmosphere since the industrial revolution.

The main human activity that emits CO_2 is the combustion of fossil fuels (coal, natural gas, and oil) for energy and transportation, although certain industrial processes and land-use changes also emit CO_2. As an example of how CO_2 can be generated, consider the combustion of octane, a component of gasoline:

$$2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)$$

This balanced reaction demonstrates that for every two molecules of octane that are burned, 16 molecules of CO_2 are generated.

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8.2: An Automobile Factory

When making anything, whether it is at home or in a factory, it is important to make sure there are enough raw materials on hand. Likewise, if one is working with a limited supply of raw materials, it would be helpful to know how much of a particular product may be made. At home, this could very well apply to making cookies or sandwiches or knowing how many doses of medication may be left.



Figure 8.2.1: A classic '57 Chevy. (Doug Watanabe via Unsplash)

How Much Product Can We Make?

For our introduction, let's assume we are working in an automobile factory. The objective in an automobile factory, of course, is to make cars. There are many, many parts that go into making an automobile. For simplicity, let's look at the number of frames, bumpers, and wheels that go into making one car. In order to make one car, we need one frame, two bumpers, and four wheels. In equation form, this could be written as :

$$1 \text{ frame} + 2 \text{ bumpers} + 4 \text{ wheels} \rightarrow 1 \text{ car}$$

Next, let's consider a scenario where there were 50 frames, 80 bumpers, and 100 wheels available. A reasonable question might be, what is the maximum number of cars that could be made from the available supply?

Looking at the equation, we should be able to identify the following relationships and possible conversions:

Relationship	Possible Conversions	
1 frame: 1 car	$\frac{1 \text{ frame}}{1 \text{ car}}$	$\frac{1 \text{ car}}{1 \text{ frame}}$
2 bumpers: 1 car	$\frac{2 \text{ bumpers}}{1 \text{ car}}$	$\frac{1 \text{ car}}{2 \text{ bumpers}}$
4 wheels: 1 car	$\frac{4 \text{ wheels}}{1 \text{ car}}$	$\frac{1 \text{ car}}{4 \text{ wheels}}$

The appropriate conversion factor is used in such a manner that units cancel when calculating the number of cars.

$$50 \cancel{\text{ frames}} \times \frac{1 \text{ car}}{1 \cancel{\text{ frame}}} = \cancel{50 \text{ cars}}$$

$$80 \cancel{\text{ bumpers}} \times \frac{1 \text{ car}}{2 \cancel{\text{ bumpers}}} = \cancel{40 \text{ cars}}$$

$$100 \cancel{\text{ wheels}} \times \frac{1 \text{ car}}{4 \cancel{\text{ wheels}}} = \boxed{25 \text{ cars}}$$

These calculations show that 25 cars is the maximum number that may be made from a supply of 50 frames, 80 bumpers, and 100 wheels. Even though there are enough frames to make 50 cars and enough bumpers to make 40 cars, we run out of wheels once 25 cars are made. It is impossible to make any more than 25 cars due to the limited supply of wheels.

This analogy allows the introduction of several terms that will be used throughout this chapter and remainder of this text:

⚡ Terms Associated with Stoichiometry

- The **limiting reactant** is the reactant that limits the amount of product that can be made.
- An **excess reactant** is one that is not entirely consumed.
- The **theoretical yield** is the maximum amount of product that may be made when all of the limiting reactant is converted to product.
- The **actual yield** is the actual amount of product that is made. The actual yield may never exceed the theoretical yield.
- The **percent yield** is the percentage of the theoretical yield that is actually obtained. The percent yield may never exceed 100%.

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

In the example above, *wheels are the limiting reactant* since they limited the number of cars that could be made. *Frames and bumpers are both excess reactants* since more cars could have been made from the frames and bumpers had there been more wheels available. *The theoretical yield is 25 cars*, since this is the maximum number of cars that could be made from the available supply of frames, bumpers, and wheels.

Let's go one step further. Let's say that some of the wheels were found to be defective. As a result, only 23 cars were made. In this case, *there was a 92% yield of cars*. To calculate the percent yield:

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{23\ \cancel{\text{cars}}}{25\ \cancel{\text{cars}}} \times 100 = \boxed{92\%\ \text{yield}}$$

How Much Raw Materials Are Needed?

What if we wanted to go the other direction? In other words, how could we find out how much raw materials are needed when a specific amount of product is needed. Using the automobile factory and simplified example and equation once again:

$$1\ \text{frame} + 2\ \text{bumpers} + 4\ \text{wheels} \rightarrow 1\ \text{car}$$

How many frames, bumpers, and wheels would be needed to make 126 cars? The same relationships previously used still apply and each relationship may be used as a conversion factor in such a manner that units cancel.

$$126\ \cancel{\text{cars}} \times \frac{1\ \text{frame}}{1\ \cancel{\text{car}}} = \boxed{126\ \text{frames}}$$

$$126\ \cancel{\text{cars}} \times \frac{2\ \text{bumpers}}{1\ \cancel{\text{car}}} = \boxed{252\ \text{bumpers}}$$

$$126\ \cancel{\text{cars}} \times \frac{4\ \text{wheels}}{1\ \cancel{\text{car}}} = \boxed{504\ \text{wheels}}$$

As we can see, it would take 126 frames, 252 bumpers, and 504 wheels to make 126 cars.

✅ Example 8.2.1

If 30 bumpers were available, how many frames and wheels would be needed to make complete cars?

Solution

The equation is:

$$1\ \text{frame} + 2\ \text{bumpers} + 4\ \text{wheels} \rightarrow 1\ \text{car}$$

The relevant relationships from the equation are:

$$1\ \text{frame: } 2\ \text{bumpers}$$

$$2\ \text{bumpers: } 4\ \text{wheels}$$

Each relationship may be used as a conversion factor in such a manner that units cancel.

$$30 \cancel{\text{ bumpers}} \times \frac{1 \text{ frame}}{2 \cancel{\text{ bumpers}}} = \boxed{15 \text{ frames}}$$

$$30 \cancel{\text{ bumpers}} \times \frac{4 \text{ wheels}}{2 \cancel{\text{ bumpers}}} = \boxed{60 \text{ wheels}}$$

These answers make sense, since the number of frames should be half of the number of bumpers and the number of wheels should be twice the number of bumpers.

Exercise 8.2.1

How many cars could be made using 300 wheels, assuming there were excess frames and bumpers available?

Answer

75 cars

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8.3: Stoichiometry and the Molar Interpretation

Learning Objectives

- Explain the meaning of the term "stoichiometry."
- Determine the relative amounts of each substance in chemical equations.

The previous section used an everyday example of an automobile factory to show the relationships between raw materials and products. In Chapter 7, we learned that chemical equations provide us with information about the types of particles that react to form products. Chemical equations also provide us with the relative number of particles and moles that react to form products. In this section, we will apply this knowledge to explore the quantitative relationships that exist between the amounts of reactants and products in a balanced chemical equation, also known as **stoichiometry**.

Stoichiometry, by definition, is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in a balanced chemical equation. The word stoichiometry is actually from two Greek words: *στοικηιον*, which means "element", and *μετρον*, which means "measure".

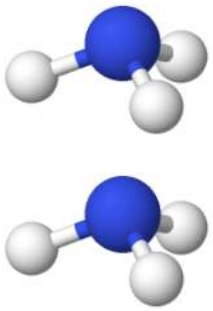
Interpreting Chemical Equations

The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Consider the production of ammonia, NH_3, from nitrogen and hydrogen gases, an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

The coefficients used, as we have learned, tell us the relative amounts of each substance in the equation. So for every 1 unit of nitrogen, N_2, we have, we need to have 3 units of hydrogen, H_2. For every 1 dozen nitrogen molecules, we need 3 dozen hydrogen molecules. Because the "mole" is also a counting unit, we can interpret this equation in terms of moles, as well. For every 1 mole of nitrogen, we need 3 moles of hydrogen. Table 8.3.1 shows several ways of interpreting this chemical equation.

Table 8.3.1: Several Ways to Interpret a Chemical Equation

$N_2\ (g)$	+	$3\ H_2\ (g)$	→	$2\ NH_3\ (g)$
	+		→	
1 molecule N_2	+	3 molecules H_2	→	2 molecules NH_3
1 dozen molecules N_2	+	3 dozen molecules H_2	→	2 dozen molecules NH_3
1 mol N_2	+	3 mol H_2	→	2 mol NH_3
28.02 g N_2	+	6.048 g H_2	→	34.07 g NH_3

The most convenient unit for counting particles is the mole, since mole-sized quantities are measured in units of grams. This is the conventional method of interpreting any balanced chemical equation and is called the **molar interpretation** of a chemical equation.

⚡ Molar Interpretation

A method for interpreting balanced chemical equations in which the coefficients represent the number of moles of each substance.

Finally, if each mole quantity is converted to grams by using the molar mass, we can see that the law of conservation of mass is followed. 1 mole of nitrogen has a mass of 28.02 g, while 3 moles of hydrogen has a mass of 6.048 g, and 2 moles of ammonia has a mass of 34.07 g (slight difference due to rounding all masses to four significant figures). Mass and the number of atoms must be conserved in any chemical reaction.

✅ Example 8.3.1

The balanced chemical equation for the combustion of ethane, C_2H_6, is

$$2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$$

A. Describe the equation in terms of the number of formula units or molecules.
B. Use a molar interpretation to describe the equation.

Solution

A. 2 molecules C_2H_6 react with 7 molecules O_2 to form 4 molecules CO_2 and 6 molecules H_2O.
B. 2 mol C_2H_6 react with 7 mol O_2 to form 4 mol CO_2 and 6 mol H_2O.

Exercise 8.3.1

Given the balanced chemical equation:

$$KBrO_3 + 6\ KI + 6\ HBr \rightarrow 7\ KBr + 3\ I_2 + 3\ H_2O$$

A. Describe the equation in terms of the number of formula units or molecules.
B. Use a molar interpretation to describe the equation.

Answer A

1 formula unit $KBrO_3$ reacts with 6 formula units KI and 6 molecules HBr to yield 7 formula units KBr, 3 molecules I_2, and 3 molecules H_2O.

Answer B

1 mol $KBrO_3$, 6 mol KI, and 6 mol HBr react together to yield 7 mol KBr, 3 mol I_2, and 3 mol H_2O.

Summary

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in a balanced chemical equation.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- The coefficients of a balanced equation can be used to determine the ratio of moles of all substances in the reaction.

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8.4: Molar Ratios and Mole-to-Mole Conversions

Learning Objectives

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

The previous section described the molar interpretation of a balanced chemical equation. The balanced chemical equation:

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

tells us that 1 mol N_2 reacts with 3 mol H_2 to yield 2 mol NH_3. Referring back to the automobile factory analogy, this also means that several relationships and possible conversions may be identified. When the coefficients from the balanced chemical equation are used to represent conversion factors in terms of the number of moles, the relationships are called **molar ratios**.

Relationships	Molar Ratios	
1 mol N_2: 3 mol H_2	$\frac{1\text{ mol }N_2}{3\text{ mol }H_2}$	$\frac{3\text{ mol }H_2}{1\text{ mol }N_2}$
1 mol N_2: 2 mol NH_3	$\frac{1\text{ mol }N_2}{2\text{ mol }NH_3}$	$\frac{2\text{ mol }NH_3}{1\text{ mol }N_2}$
3 mol H_2: 2 mol NH_3	$\frac{3\text{ mol }H_2}{2\text{ mol }NH_3}$	$\frac{2\text{ mol }NH_3}{3\text{ mol }H_2}$

The molar ratios may then be used to convert between moles of one substance and moles of a different substance, in much the same way that we were able to convert between wheels and cars, bumpers and wheels, or cars and bumpers.

Suppose we wanted to know how many moles of H_2 were needed to completely react with 17.6 mol N_2. The appropriate molar ratio is used in such a manner that units cancel when calculating the the moles of H_2.

$$17.6\ \cancel{\text{mol N}_2} \times \frac{3\text{ mol H}_2}{1\ \cancel{\text{mol N}_2}} = \boxed{52.8\text{ mol H}_2}$$

Suppose we wanted to know how many moles of NH_3 could be made when 0.773 mol H_2 reacted with an excess of N_2. Once again, the appropriate molar ratio is used in such a manner that units cancel when calculating the the moles of NH_3. Note that the amount of N_2 is irrelevant, since it is present in excess.

$$0.773\ \cancel{\text{mol H}_2} \times \frac{2\text{ mol NH}_3}{3\ \cancel{\text{mol H}_2}} = \boxed{0.515\text{ mol NH}_3}$$

Example 8.4.1

Identify all of the possible molar ratios for the balanced chemical equation:

$$2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g)$$

Solution

Relationships	Molar Ratios	
2 mol H_2: 1 mol O_2	$\frac{2\ \text{mol H}_2}{1\ \text{mol O}_2}$	$\frac{1\ \text{mol O}_2}{2\ \text{mol H}_2}$
2 mol H_2: 2 mol H_2O	$\frac{\cancel{2}\ \text{mol H}_2}{\cancel{2}\ \text{mol H}_2\text{O}} = \frac{1\ \text{mol H}_2}{1\ \text{mol H}_2\text{O}}$	$\frac{\cancel{2}\ \text{mol H}_2\text{O}}{\cancel{2}\ \text{mol H}_2} = \frac{1\ \text{mol H}_2\text{O}}{1\ \text{mol H}_2}$
1 mol O_2: 2 mol H_2O	$\frac{1\ \text{mol O}_2}{2\ \text{mol H}_2\text{O}}$	$\frac{2\ \text{mol H}_2\text{O}}{1\ \text{mol O}_2}$

✅ Example 8.4.2

Calculate the number of moles of hydrogen and oxygen required to make 3.34 mol of H_2O according to the balanced chemical equation:

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

Solution

The appropriate molar ratios (see Example 8.4.1) are used in such a manner that units cancel when calculating moles of H_2 and moles of O_2.

$$3.34\ \cancel{\text{mol H}_2\text{O}} \times \frac{1\ \text{mol H}_2}{1\ \cancel{\text{mol H}_2\text{O}}} = \boxed{3.34\ \text{mol H}_2}$$

$$3.34\ \cancel{\text{mol H}_2\text{O}} \times \frac{1\ \text{mol O}_2}{2\ \cancel{\text{mol H}_2\text{O}}} = \boxed{1.67\ \text{mol O}_2}$$

✏️ Exercise 8.4.1

A. Write a balanced chemical equation that shows the reaction between aluminum metal and oxygen gas to form an aluminum oxide solid.

B. Identify all of the possible molar ratios.

Answer A

$4\ Al\ (s) + 3\ O_2\ (g) \rightarrow 2\ Al_2O_3\ (s)$

Answer B

Relationships	Molar Ratios	
4 mol Al: 3 mol O_2	$\frac{4\ \text{mol Al}}{3\ \text{mol O}_2}$	$\frac{3\ \text{mol O}_2}{4\ \text{mol Al}}$
4 mol Al: 2 mol Al_2O_3	$\frac{\cancel{4}\ \text{mol Al}}{\cancel{2}\ \text{mol Al}_2\text{O}_3} = \frac{2\ \text{mol Al}}{1\ \text{mol Al}_2\text{O}_3}$	$\frac{\cancel{2}\ \text{mol Al}_2\text{O}_3}{\cancel{4}\ \text{mol Al}} = \frac{1\ \text{mol Al}_2\text{O}_3}{2\ \text{mol Al}}$
3 mol O_2: 2 mol Al_2O_3	$\frac{3\ \text{mol O}_2}{2\ \text{mol Al}_2\text{O}_3}$	$\frac{2\ \text{mol Al}_2\text{O}_3}{3\ \text{mol O}_2}$

Exercise 8.4.2

Given the balanced chemical equation:

$$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

A. How many moles of O_2 are needed to completely react with 3.08 mol C_8H_{18}?
B. How many moles of C_8H_{18} will react when 7.7 mol CO_2 are produced?

Answer A

38.5 mol O_2

Answer B

0.96 mol C_8H_{18}

Summary

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

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8.5: Mass-to-Mass Conversions

Learning Objectives

- Convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.

Mole-to-Mass Conversions

We have established that a balanced chemical equation is balanced in terms of moles, as well as atoms or molecules. We have used balanced equations to set up ratios, in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions – such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that relation to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:

$$\boxed{\text{moles of A}} \xrightarrow[\text{ratio}]{\text{molar}} \boxed{\text{moles of B}} \xrightarrow[\text{mass}]{\text{molar}} \boxed{\text{grams of B}}$$

Collectively, these conversions are called *mole-to-mass* calculations. As an example, consider the balanced chemical equation

$$Fe_2O_3 + 3\ SO_3 \rightarrow Fe_2(SO_4)_3$$

Suppose we wanted to know how many grams of SO_3 will react with 3.59 mol of Fe_2O_3. Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate *molar ratio*, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. From here, the mass of SO_3 may be calculated using the *molar mass* of SO_3 as a conversion factor.

Relationship		Possible Conversions
molar ratio	1 mol Fe_2O_3: 3 mol SO_3	$\frac{1\text{ mol }Fe_2O_3}{3\text{ mol }SO_3}$ or $\frac{3\text{ mol }SO_3}{1\text{ mol }Fe_2O_3}$
molar mass	1 mol SO_3 = 80.07 g SO_3	$\frac{80.07\text{ g }SO_3}{1\text{ mol }SO_3}$ or $\frac{1\text{ mol }SO_3}{80.07\text{ g }SO_3}$

As before, start with the quantity given (3.59 mol Fe_2O_3):

$$3.59\ \cancel{\text{mol Fe}_2\text{O}_3} \times \frac{3\text{ mol SO}_3}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} = 10.77\text{ mol SO}_3$$

Moles of Fe_2O_3 cancel, leaving moles of SO_3 as the units. This answer may then be converted to grams of SO_3:

$$10.77\ \cancel{\text{mol SO}_3} \times \frac{80.07\text{ g SO}_3}{1\ \cancel{\text{mol SO}_3}} = \boxed{862\text{ g SO}_3}$$

The final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner. It is generally easier if the problem is set up in a single series of steps:

$$3.59\ \cancel{\text{mol Fe}_2\text{O}_3} \times \frac{3\ \cancel{\text{mol SO}_3}}{1\ \cancel{\text{mol Fe}_2\text{O}_3}} \times \frac{80.07\text{ g SO}_3}{1\ \cancel{\text{mol SO}_3}} = \boxed{862\text{ g SO}_3}$$

Notice that the identical answer is obtained, in much the same manner as when multi-step dimensional analysis was introduced in Section 2.7.

✅ Example 8.5.1: Generation of Aluminum Oxide

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2\ AlCl_3 + 3\ H_2O \rightarrow Al_2O_3 + 6\ HCl$$

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 249 g AlCl3 Find: moles HCl
List other known quantities.	1 mol AlCl3 = 133.33 g $AlCl_3$ 2 mol $AlCl_3$: 6 mol HCl
Prepare a concept map using the proper conversion factor(s).	$\boxed{\text{g AlCl}_3} \xrightarrow{\frac{1\text{ mol AlCl}_3}{133.33\text{ g AlCl}_3}} \boxed{\text{mol AlCl}_3} \xrightarrow{\frac{6\text{ mol HCl}}{2\text{ mol AlCl}_3}} \boxed{\text{mol HCl}}$
Cancel units and calculate.	$249\ \cancel{\text{g AlCl}_3} \times \frac{1\ \cancel{\text{mol AlCl}_3}}{133.33\ \cancel{\text{g AlCl}_3}} \times \frac{6\text{ mol HCl}}{2\ \cancel{\text{mol AlCl}_3}} = \boxed{5.60\text{ mol HCl}}$
Think about your result.	249 g $AlCl_3$ is less than 266.66 g (the mass for 2 mol $AlCl_3$). Since the relationship is 6 mol HCl to 2 mol $AlCl_3$, the answer should be less than 6 mol HCl. The answer of 5.60 mol HCl is indeed less than 6 mol HCl.

✏️ Exercise 8.5.1: Generation of Aluminum Oxide

How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to this chemical equation?

$$2\ AlCl_3 + 3\ H_2O \rightarrow Al_2O_3 + 6\ HCl$$

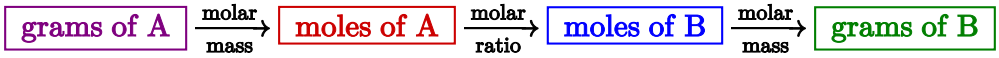
Answer

0.442 mol Al_2O_3

Mass-to-Mass Conversions

In the real world, the number of molecules or moles are not measured directly. Rather, it is the masses of the reactants and/or products that will be measured directly. Consequently, it is much more practical to make all measurements in terms of the number of grams. It is a small step to move from *mole-to-mass* calculations to *mass-to-mass* calculations.

If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then – and only then – the balanced chemical equation is used to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. This text will generally approach this type of problem in a single series of steps.

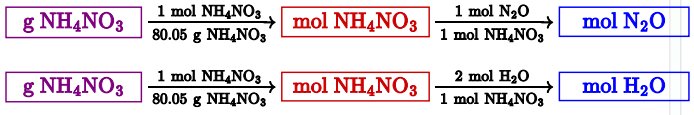
✅ Example 8.5.2: Decomposition of Ammonium Nitrate

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

$$NH_4NO_3\ (s) \rightarrow N_2O\ (g) + 2\ H_2O\ (l)$$

Find the mass of each product formed when 45.7 g of ammonium nitrate is decomposed.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 45.7 g NH_4NO_3 Find: Mass N_2O, Mass H_2O
List other known quantities.	1 mol NH_4NO_3 = 80.05 g NH_4NO_3 1 mol N_2O = 44.02 g N_2O 1 mol H_2O = 18.02 g H_2O 1 mol NH_4NO_3: 1 mol N_2O 1 mol NH_4NO_3: 2 mol H_2O
Prepare concept maps using the proper conversion factor(s).	
Cancel units and calculate.	$45.7\ \cancel{g\ NH_4NO_3} \times \frac{1\ \cancel{mol\ NH_4NO_3}}{80.05\ \cancel{g\ NH_4NO_3}} \times \frac{1\ \cancel{mol\ N_2O}}{1\ \cancel{mol\ NH_4NO_3}} \times \frac{44.02\ g\ N_2O}{1\ \cancel{mol\ N_2O}} = \boxed{25.1\ g\ N_2O}$ $45.7\ \cancel{g\ NH_4NO_3} \times \frac{1\ \cancel{mol\ NH_4NO_3}}{80.05\ \cancel{g\ NH_4NO_3}} \times \frac{2\ \cancel{mol\ H_2O}}{1\ \cancel{mol\ NH_4NO_3}} \times \frac{18.02\ g\ H_2O}{1\ \cancel{mol\ H_2O}} = \boxed{20.6\ g\ H_2O}$
Think about your result.	The sum of the masses of the two products is equal to the mass of ammonium nitrate decomposed. This demonstrates the Law of Conservation of Mass. Each answer has three significant figures.

✏️ Exercise 8.5.2: Carbon Tetrachloride

Methane reacts with elemental chlorine to make carbon tetrachloride, CCl_4. The balanced chemical equation for this reaction is:

$$CH_4\ (g) + 4\ Cl_2\ (g) \rightarrow CCl_4\ (l) + 4\ HCl\ (g)$$

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

Answer

909.1 g HCl

✅ Example 8.5.3: Production of Ammonia

Taking into consideration the balanced chemical equation:

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

what mass of hydrogen gas is needed to prepare 404 g of ammonia?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 404 g NH_3 Find: Mass H_2
List other known quantities.	1 mol NH3 = 17.03 g NH_3 1 mol H_2 = 2.016 g H_2 3 mol H_2: 2 mol NH_3
Prepare concept maps using the proper conversion factor(s).	$\text{g NH}_3 \xrightarrow{\frac{1\text{ mol NH}_3}{17.03\text{ g NH}_3}} \text{mol NH}_3 \xrightarrow{\frac{3\text{ mol H}_2}{2\text{ mol NH}_3}} \text{mol H}_2 \xrightarrow{\frac{2.016\text{ g H}_2}{1\text{ mol H}_2}}$
Cancel units and calculate.	$404\ \text{g NH}_3 \times \frac{1\ \text{mol NH}_3}{17.03\ \text{g NH}_3} \times \frac{3\ \text{mol H}_2}{2\ \text{mol NH}_3} \times \frac{2.016\ \text{g H}_2}{1\ \text{mol H}_2} = \boxed{71.7\ \text{g H}_2}$
Think about your result.	404 g NH_3 is a bit more than 20 mol N_3. Since the molar ratio is 3 mol H_2: 2 mol NH_3, a bit more than 30 mol H_2 are needed. 71.7 g H_2 is a bit more than 30 mol H_2.

✏️ Exercise 8.5.3: Vanadium(V) Oxide

Vanadium metal reacts with oxygen gas to yield solid vanadium(V) oxide:

$$4\,V\,(s) + 5\,O_2\,(g) \rightarrow 2\,V_2O_5\,(s)$$

A. What is the molar ratio between V and O_2?
B. How many moles of V will react with 9.25 mol O_2?
C. What mass of vanadium is needed to produce 0.301 g V_2O_5?

Answer A

4 mol V: 5 mol O_2

Answer B

7.40 mol V

Answer C

0.169 g V

Summary

- Calculations involving conversions between moles of a substance and the mass of that substance can be done using conversion factors.
- A balanced chemical reaction can be used to determine molar and mass relationships between substances.

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8.6: Limiting Reactants and Excess Reactants

Learning Objectives

- Identify the limiting reactant (limiting reagent) in a given chemical reaction.
- Calculate how much product will be produced from the limiting reactant.
- Calculate how much reactant(s) remains when the reaction is complete.

Section 8.2 used an automobile factory to introduce terminology that extends to the stoichiometry associated with chemical reactions. We learned that the **limiting reactant** is the reactant that limits the amount of product that can be made, while an **excess reactant** is one that is not entirely consumed. We also learned that the **theoretical yield** is the maximum amount of product that may be made when all of the limiting reactant is converted to product.

Check out the simulation below for examples that review the concept of limiting reactants and excess reactants and applies the concept to making sandwiches and molecules. Then test your understanding with a game.

PhET Simulation: Limiting and Excess Reactants

View this interactive simulation illustrating the concepts of limiting and excess reactants.



Consider the reaction between hydrogen and chlorine to make hydrogen chloride:

$$H_2\,(g) + Cl_2\,(g) \rightarrow 2\,HCl\,(g)$$

The balanced equation shows that hydrogen and chlorine react in a 1:1 stoichiometric ratio. If these reactants are provided in any other amounts, one of the reactants will nearly always be entirely consumed, thus limiting the amount of product that may be generated. This substance is the limiting reactant, and the other substance is the excess reactant. Identifying the limiting and excess reactants for a given situation requires computing the molar amounts of each reactant provided and comparing them to the stoichiometric amounts represented in the balanced chemical equation.

For example, imagine combining 3 moles of H_2 and 2 moles of Cl_2. This represents a 3:2 (or 1.5:1) ratio of hydrogen to chlorine present for reaction, which is greater than the stoichiometric ratio of 1:1. Hydrogen, therefore, is present in excess, and chlorine is the limiting reactant. Reaction of all of the provided chlorine (2 moles) will consume 2 of the 3 moles of hydrogen provided, leaving 1 mole of hydrogen that is unconsumed (see Figure 8.6.1).

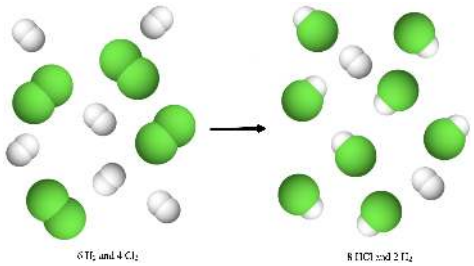


Figure **8.6.1**: When H_2 and Cl_2 are combined in non-stoichiometric amounts, one of these reactants will limit the amount of HCl that can be produced. This illustration shows a reaction in which hydrogen is present in excess and chlorine is the limiting reactant.

An alternative approach to identifying the limiting reactant involves comparing the amount of product expected for the complete reaction of each reactant. Each reactant amount is used to separately calculate the amount of product that would be formed per the reaction's stoichiometry. The reactant yielding the lesser amount of product is the limiting reactant.

If we were to again imagine combining 3 moles of H_2 and 2 moles of Cl_2, complete reaction of the provided hydrogen would yield:

$$3 \text{ mol } H_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } H_2} = 6 \text{ mol HCl produced}$$

Complete reaction of the provided chlorine would produce:

$$2 \text{ mol } Cl_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } Cl_2} = \boxed{4 \text{ mol HCl produced}}$$

The chlorine will be completely consumed once 4 moles HCl have been produced. Since enough hydrogen was provided to yield 6 moles HCl, there will be hydrogen that remains unconsumed once this reaction is complete.

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 that is entirely consumed is called the limiting reactant; the other reactant or reactants are present *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reactant and which is/are the excess reactant(s).

The key to recognizing which reactant is the limiting reactant is to determine the reactant that yields the *least* amount of product is the limiting reactant. It doesn't matter which product is used for the calculation, as long as the same product is used in for the comparison. It also doesn't matter whether the number of moles or grams of that product is calculated. However, knowing the final mass of product is generally more useful.

✅ Example 8.6.1: Limiting Reactants and Mole-to-Mole Calculations

Consider the balanced equation:

$$4\ C_2H_3Br_3 + 11\ O_2 \rightarrow 8\ CO_2 + 6\ H_2O + 6\ Br_2$$

A. How many moles of Br_2 could be produced when 36 mol $C_2H_3Br_3$ are reacted with 33 mol O_2?
B. Identify the limiting reactant(s) and excess reactant(s).

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 36 mol $C_2H_3Br_3$ reacted; 33 mol O_2 reacted Find: theoretical yield Br_2; identify limiting reactant(s) and excess reactant(s)
List other known quantities.	4 mol $C_2H_3Br_3$: 6 mol Br_2 11 mol O_2: 6 mol Br_2
Prepare concept maps using the proper conversion factor(s).	$\text{mol } C_2H_3Br_3 \xrightarrow{\frac{6 \text{ mol } Br_2}{4 \text{ mol } C_2H_3Br_3}} \text{mol } Br_2$ $\text{mol } O_2 \xrightarrow{\frac{6 \text{ mol } Br_2}{11 \text{ mol } O_2}} \text{mol } Br_2$
Calculate the theoretical yield. Select the smallest answer.	$36 \text{ mol } C_2H_3Br_3 \times \frac{6 \text{ mol } Br_2}{4 \text{ mol } C_2H_3Br_3} = 54 \text{ mol } Br_2$ $33 \text{ mol } O_2 \times \frac{6 \text{ mol } Br_2}{11 \text{ mol } O_2} = \boxed{18 \text{ mol } Br_2}$
Identify the limiting reactant(s) and excess reactant(s).	The limiting reactant is O_2 since it would yield the least amount of product (18 mol Br_2). The excess reactant is $C_2H_3Br_3$ since its complete reaction would have yielded up to 54 mol Br_2.
Think about your result.	It usually is not possible to determine the limiting reactant looking at the initial amounts, since the reactants have different coefficients.

✅ Example 8.6.2: Limiting Reactants and Mass-to-Mass Calculations

Consider once again the balanced equation:

$$4\ C_2H_3Br_3 + 11\ O_2 \rightarrow 8\ CO_2 + 6\ H_2O + 6\ Br_2$$

A. What is the theoretical yield of CO_2, in grams, when 76.4 grams of $C_2H_3Br_3$ are reacted with 49.1 grams of O_2?
B. Identify the limiting reactant(s) and excess reactant(s).

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 76.4 g $C_2H_3Br_3$ reacted; 49.1 g O_2 reacted Find: theoretical yield CO_2; identify limiting reactant(s) and excess reactant(s)
List other known quantities.	1 mol $C_2H_3Br_3$ = 266.74 g $C_2H_3Br_3$ 1 mol O_2 = 32.00 g O_2 1 mol CO_2 = 44.01 g CO_2 4 mol $C_2H_3Br_3$: 8 mol CO_2 11 mol O_2: 8 mol CO_2
Prepare concept maps using the proper conversion factor(s).	$\text{g } C_2H_3Br_3 \xrightarrow{\frac{1\text{ mol } C_2H_3Br_3}{266.74\text{ g } C_2H_3Br_3}} \text{mol } C_2H_3Br_3 \xrightarrow{\frac{8\text{ mol } CO_2}{4\text{ mol } C_2H_3Br_3}} \text{mol } CO_2$ $\text{g } O_2 \xrightarrow{\frac{1\text{ mol } O_2}{32.00\text{ g } O_2}} \text{mol } O_2 \xrightarrow{\frac{8\text{ mol } CO_2}{11\text{ mol } O_2}} \text{mol } CO_2 \xrightarrow{\frac{44.01\text{ g } CO_2}{1\text{ mol } CO_2}}$
Calculate the theoretical yield. Select the smallest answer.	$76.4\text{ g } C_2H_3Br_3 \times \frac{1\text{ mol } C_2H_3Br_3}{266.74\text{ g } C_2H_3Br_3} \times \frac{8\text{ mol } CO_2}{4\text{ mol } C_2H_3Br_3} \times \frac{44.01\text{ g } CO_2}{1\text{ mol } CO_2} = \boxed{25.2\text{ g } CO_2}$ $49.1\text{ g } O_2 \times \frac{1\text{ mol } O_2}{32.00\text{ g } O_2} \times \frac{8\text{ mol } CO_2}{11\text{ mol } O_2} \times \frac{44.01\text{ g } CO_2}{1\text{ mol } CO_2} = 49.1\text{ g } CO_2$
Identify the limiting reactant(s) and excess reactant(s).	The limiting reactant is $C_2H_3Br_3$ since it would yield the least amount of product (25.2 g CO_2). The excess reactant is O_2 since its complete reaction would have yielded up to 49.1 g CO_2.
Think about your result.	It usually is not possible to determine the limiting reactant looking at the initial masses, since the reactants have different molar masses and different coefficients.

✅ Example 8.6.3: Limiting Reactant and Mass of Excess Reactant

5.00 g Rb are combined with 3.44 g $MgCl_2$ according to the chemical reaction:

$$2\text{ Rb }(s) + MgCl_2\ (s) \rightarrow \text{Mg }(s) + 2\text{ RbCl }(s)$$

A. What mass of Mg is formed?
B. What mass of which reactant is left over?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 5.00 g Rb reacted; 3.44 g $MgCl_2$ reacted Find: theoretical yield Mg; identify excess reactant and its mass
List other known quantities.	1 mol Rb = 85.47 g Rb 1 mol $MgCl_2$ = 95.21 g $MgCl_2$ 1 mol Mg = 24.31 g Mg 2 mol Rb: 1 mol Mg 1 mol $MgCl_2$: 1 mol Mg
Prepare concept maps using the proper conversion factor(s).	g Rb $\xrightarrow{\frac{1\text{ mol Rb}}{85.47\text{ g Rb}}}$ mol Rb $\xrightarrow{\frac{1\text{ mol Mg}}{2\text{ mol Rb}}}$ mol Mg $\xrightarrow{\frac{24.31\text{ g Mg}}{1\text{ mol Mg}}}$ g Mg g $MgCl_2$ $\xrightarrow{\frac{1\text{ mol }MgCl_2}{95.21\text{ g }MgCl_2}}$ mol $MgCl_2$ $\xrightarrow{\frac{1\text{ mol Mg}}{2\text{ mol }MgCl_2}}$ mol Mg $\xrightarrow{\frac{24.31\text{ g Mg}}{1\text{ mol Mg}}}$
Calculate the theoretical yield. Select the smallest answer.	$5.00\ \text{g Rb} \times \frac{1\ \text{mol Rb}}{85.47\ \text{g Rb}} \times \frac{1\ \text{mol Mg}}{2\ \text{mol Rb}} \times \frac{24.31\ \text{g Mg}}{1\ \text{mol Mg}} = \boxed{0.711\ \text{g Mg}}$ $3.44\ \text{g }MgCl_2 \times \frac{1\ \text{mol }MgCl_2}{95.21\ \text{g }MgCl_2} \times \frac{1\ \text{mol Mg}}{1\ \text{mol }MgCl_2} \times \frac{24.31\ \text{g Mg}}{1\ \text{mol Mg}} = 0.878\ \text{g Mg}$
Identify the limiting reactant(s) and excess reactant(s).	The limiting reactant is Rb since it would yield the least amount of product (0.711 g Mg). The excess reactant is $MgCl_2$ since its complete reaction would have yielded up to 0.878 g Mg.
Calculate the mass of excess reactant that reacts.	Start with 5.00 g Rb, since it is the limiting reactant and known to completely react. $5.00\ \text{g Rb} \times \frac{1\ \text{mol Rb}}{85.47\ \text{g Rb}} \times \frac{1\ \text{mol }MgCl_2}{2\ \text{mol Rb}} \times \frac{95.21\ \text{g }MgCl_2}{1\ \text{mol }MgCl_2} = \boxed{2.78\ \text{g }MgCl_2\ \text{reacted}}$
Calculate the mass of excess reactant that remains and think about the result.	We started with 3.44 g $MgCl_2$ and found that 2.78 g $MgCl_2$ reacted. This makes sense, since $MgCl_2$ is the excess reactant. Less should react than we started with. 3.44 g $MgCl_2$ initially – 2.78 g $MgCl_2$ reacted = 0.66 g $MgCl_2$ remain

Exercise 8.6.1

Given the balanced chemical equation:

$$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$$

A. If 24.5 moles of C_8H_{18} are reacted with 245 moles of O_2, how many moles of CO_2 are produced assuming complete reaction?
B. How many moles of which reactant remain unconsumed?

Answer A

157 mol CO_2 are produced.

Answer B

4.9 mol C_8H_{18} remain unconsumed.

Exercise 8.6.2

Potassium superoxide, KO_2, is used in rebreathing gas masks to generate oxygen according to the equation below.

$$4\ KO_2\ (s) + 2\ H_2O\ (l) \rightarrow 4\ KOH\ (s) + 3\ O_2\ (g)$$

A. What is the theoretical yield of oxygen, in units of grams, when 108 g of KO_2 reacts with 16.8 g of H_2O?
B. Which reactant is the limiting reactant?
C. Which reactant is present in excess?

Answer A

36.5 g O_2 are produced.

Answer B

The limiting reactant is KO_2.

Answer C

H_2O is present in excess.

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8.7: Theoretical Yield and Percent Yield

Learning Objectives

- Calculate percentage or actual yields from known amounts of reactants.

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

Percent Yield

Chemical reactions in the real world don't always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than predicted. Besides spills and other experimental errors, there are usually losses due to an incomplete reaction, undesirable side reactions, etc. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. As we have previously learned, this is called the **theoretical yield**, the maximum amount of product that can be formed from the given amounts of reactants. The **actual yield** is the amount of product that is actually formed when the reaction is carried out in the laboratory. The **percent yield** is the ratio of the actual yield to the theoretical yield, expressed as a percentage. Once again,

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and unnecessary expense.

The percent yield of a desired product may never exceed 100% for the reasons indicated earlier. Should a percent yield ever be greater than 100%, this would be an indication of some sort of experimental error or that the reaction contains impurities that cause its mass to be greater than if the product was pure. When chemists synthesize a desired chemical, they are always careful to purify the products of the reaction. Example **8.7.1** illustrates the steps for determining percent yield.

✅ Example 8.7.1: Theoretical Yield and Percent Yield

Potassium chlorate decomposes upon heating, according to the reaction below:

$$2\ KClO_3\ (s) \xrightarrow{\Delta} 2\ KCl\ (s) + 3\ O_2\ (g)$$

If 40.0 g $KClO_3$ is heated until it completely decomposes, resulting in the collection of 14.3 g of oxygen gas,

A. What is the theoretical yield of oxygen gas?
B. What is the percent yield of oxygen gas?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 40.0 g $KClO_3$ reacted; 14.3 g O_2 produced Find: theoretical yield O_2; percent yield O_2
List other known quantities.	1 mol $KClO_3$ = 122.55 g $KClO_3$ 1 mol O_2 = 32.00 g O_2 2 mol $KClO_3$: 3 mol O_2
Prepare concept maps using the proper conversion factor(s).	$\text{g } KClO_3 \xrightarrow{\frac{1\text{ mol } KClO_3}{122.55\text{ g } KClO_3}} \text{mol } KClO_3 \xrightarrow{\frac{3\text{ mol } O_2}{2\text{ mol } KClO_3}} \text{mol } O_2 \xrightarrow{\frac{32.0}{1\text{ m}}}$
Calculate the theoretical yield.	$40.0\ \text{g } KClO_3 \times \frac{1\ \text{mol } KClO_3}{122.55\ \text{g } KClO_3} \times \frac{3\ \text{mol } O_2}{2\ \text{mol } KClO_3} \times \frac{32.00\ \text{g } O_2}{1\ \text{mol } O_2} = \boxed{15.7\ \text{g } O_2}$
Calculate the percent yield.	$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{14.3\ \text{g } O_2}{15.7\ \text{g } O_2} \times 100\% = \boxed{91.1\%\text{ yield}}$
Think about your result.	The percent yield is less than 100%, but still quite high (over 90%). This seems like a reasonable yield for a chemical reaction.

✅ Example 8.7.2: Limiting Reactants, Theoretical Yield, and Percent Yield

Vanadium metal reacts with oxygen gas to yield solid vanadium(V) oxide:

$$4\ V\ (s) + 5\ O_2\ (g) \rightarrow 2\ V_2O_5\ (s)$$

A. What mass of V_2O_5 can be made when 41.3 g V reacts with 35.0 g O_2?
B. If 62.3 g V_2O_5 were actually produced, what is the percent yield?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 41.3 g V reacted; 35.0 g O_2 reacted; 62.3 g V_2O_5 produced Find: theoretical yield V_2O_5; percent yield V_2O_5
List other known quantities.	1 mol V = 50.94 g V 1 mol O_2 = 32.00 g O_2 1 mol V_2O_5 = 181.88 g V_2O_5 4 mol V: 2 mol V_2O_5 5 mol O_2: 2 mol V_2O_5
Prepare concept maps using the proper conversion factor(s).	$\text{g V} \xrightarrow{\frac{1\ \text{mol V}}{50.94\ \text{g V}}} \text{mol V} \xrightarrow{\frac{2\ \text{mol } V_2O_5}{4\ \text{mol V}}} \text{mol } V_2O_5 \xrightarrow{\frac{181.88\ \text{g } V_2O_5}{1\ \text{mol } V_2O_5}} \text{g V}$ $\text{g } O_2 \xrightarrow{\frac{1\ \text{mol } O_2}{32.00\ \text{g } O_2}} \text{mol } O_2 \xrightarrow{\frac{2\ \text{mol } V_2O_5}{5\ \text{mol } O_2}} \text{mol } V_2O_5 \xrightarrow{\frac{181.88\ \text{g } V_2O_5}{1\ \text{mol } V_2O_5}} \text{g}$
Calculate the theoretical yield. Select the smallest answer.	$41.3\ \text{g V} \times \frac{1\ \text{mol V}}{50.94\ \text{g V}} \times \frac{2\ \text{mol } V_2O_5}{4\ \text{mol V}} \times \frac{181.88\ \text{g } V_2O_5}{1\ \text{mol } V_2O_5} = \boxed{73.7\ \text{g } V_2O_5}$ $35.0\ \text{g } O_2 \times \frac{1\ \text{mol } O_2}{32.00\ \text{g } O_2} \times \frac{2\ \text{mol } V_2O_5}{5\ \text{mol } O_2} \times \frac{181.88\ \text{g } V_2O_5}{1\ \text{mol } V_2O_5} = 79.6\ \text{g } V_2O_5$
Calculate the percent yield.	$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{62.3\ \text{g } V_2O_5}{73.7\ \text{g } V_2O_5} \times 100\% = \boxed{84.5\%\ \text{yield}}$
Think about your result.	The percent yield is less than 100%, but still quite high (almost 85%). This seems like a reasonable yield for a chemical reaction.

Exercise 8.7.1

What is the percent yield of a reaction that produces 12.5 g of the Freon CF_2Cl_2 from 32.9 g of CCl_4 and an excess of HF?

$$CCl_4 + 2\ HF \rightarrow CF_2Cl_2 + 2\ HCl$$

Answer

48.3% yield

Summary

- Theoretical yield is calculated based on the stoichiometry of the chemical equation.
- The actual yield is experimentally determined.
- The percent yield is determined by calculating the ratio of actual yield to theoretical yield.

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8.9: Exercises

8.3: Stoichiometry and the Molar Interpretation

1. What are all the conversion factors that can be constructed from the balanced chemical reaction: $2\ S(s) + 3\ O_2(g) \rightarrow 2\ SO_3(g)$?

Answer

$$\frac{2\ mol\ S}{3\ mol\ O_2};\ \frac{3\ mol\ O_2}{2\ mol\ S};\ \frac{1\ mol\ S}{1\ mol\ SO_3};\ \frac{1\ mol\ SO_3}{1\ mol\ S};\ \frac{3\ mol\ O_2}{2\ mol\ SO_3};\ \frac{2\ mol\ SO_3}{3\ mol\ O_2}$$

2. Construct the three independent conversion factors possible for these two reactions:
   a. $2\ H_2 + O_2 \rightarrow 2\ H_2O$
   b. $H_2 + O_2 \rightarrow H_2O_2$

   Why are the ratios between H_2 and O_2 different?

Answer

The conversion factors are different because the stoichiometries of the balanced chemical reactions are different.

8.4: Molar Ratios and Mole-to-Mole Conversions

3. Given the chemical equation: $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
   a. Balance the equation.
   b. How many moles of H_2 are produced when 332 moles of Na react?

Answer

a. $2\ Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$
b. 166 mol H_2

4. Given the chemical equation: $S(s) + O_2(g) \rightarrow SO_3(g)$
   a. Balance the equation.
   b. How many moles of O_2 are needed when 38 moles of S react?

Answer

a. $2\ S(s) + 3\ O_2(g) \rightarrow 2\ SO_3(g)$
b. 57 mol O_2

5. For the balanced chemical equation:

$$6\ H^+(aq) + 2\ MnO_4^-(aq) + 5\ H_2O_2(l) \rightarrow 2\ Mn^{2+}(aq) + 5\ O_2(g) + 8\ H_2O(l)$$

   how many moles of H_2O are produced when 75 moles of H_2O_2 react?

Answer

120 mol H_2O

6. For the balanced chemical reaction:

$$2\ C_6H_6(l) + 15\ O_2(g) \rightarrow 12\ CO_2(g) + 6\ H_2O(l)$$

how many moles of CO_2 are produced when 56 moles of C_6H_6 react?

Answer

3.4×10^2 mol CO_2

7. For the balanced chemical equation:

$$Fe_2O_3 + 3\ SO_3 \rightarrow Fe_2(SO_4)_3$$

suppose we need to make 145 moles of $Fe_2(SO_4)_3$. How many moles of SO_3 do we need?

Answer

435 mol SO_3

8. One way to make sulfur hexafluoride is to react thioformaldehyde, CH_2S, with elemental fluorine:

$$CH_2S + 6\ F_2 \rightarrow CF_4 + 2\ HF + SF_6$$

If 45.8 moles of SF_6 are needed, how many moles of F_2 are required?

Answer

275 mol F_2

9. For the balanced chemical equation:

$$3\ Cu(s) + 2\ NO_3^-(aq) + 8\ H^+(aq) \rightarrow 3\ Cu^{2+}(aq) + 4\ H_2O(l) + 2\ NO(g)$$

how many moles of Cu^{2+} are formed when 55.7 mol of H^+ are reacted?

Answer

20.9 mol Cu^{2+}

10. For the balanced chemical equation:

$$Al(s) + 3\ Ag^+(aq) \rightarrow Al^{3+}(aq) + 3\ Ag(s)$$

how many moles of Ag are produced when 0.661 mol of Al are reacted?

Answer

1.98 mol Ag

11. For the balanced chemical reaction:

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(l)$$

how many moles of H_2O are produced when 0.669 mol of NH_3 react?

Answer

1.00 mol H_2O

12. For the balanced chemical reaction:

$$4\ NaOH(aq) + 2\ S(s) + 3\ O_2(g) \rightarrow 2\ Na_2SO_4(aq) + 2\ H_2O(l)$$

how many moles of Na_2SO_4 are formed when 1.22 mol of O_2 react?

Answer

0.813 mol Na_2SO_4

13. For the balanced chemical reaction:

$$4\ KO_2(s) + 2\ CO_2(g) \rightarrow 2\ K_2CO_3(s) + 3\ O_2(g)$$

determine the number of moles of both products formed when 6.88 mol of KO_2 react.

Answer

3.44 mol K_2CO_3; 5.16 mol O_2

14. For the balanced chemical reaction

$$2\ AlCl_3 + 3\ H_2O \rightarrow Al_2O_3 + 6\ HCl$$

determine the number of moles of both products formed when 0.0552 mol of $AlCl_3$ react.

Answer

0.0276 mol Al_2O_3; 0.166 mol HCl

8.5: Mass-to-Mass Conversions

15. What mass of CO_2 is produced by the combustion of 1.00 mol of CH_4?

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

Answer

44.0 g CO_2

16. What mass of H_2O is produced by the combustion of 1.00 mol of CH_4?

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$$

Answer

36.0 g H_2O

17. What mass of HgO is required to produce 0.692 mol of O_2?

$$2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$$

Answer

3.00×10^2 g HgO

18. What mass of $NaHCO_3$ is needed to produce 2.659 mol of CO_2?

$$2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

Answer

446.8 g $NaHCO_3$

19. How many moles of Al can be produced from 10.87 g of Ag?

$$Al(NO_3)_3(aq) + 3\ Ag(s) \rightarrow Al(s) + 3\ AgNO_3(aq)$$

Answer

0.03359 mol Al

20. How many moles of HCl can be produced from 0.226 g of $SOCl_2$?

$$SOCl_2(l) + H_2O(l) \rightarrow SO_2(g) + 2\ HCl(g)$$

Answer

3.80×10^{-3} mol HCl

21. How many moles of O_2 are needed to prepare 1.00 g of $Ca(NO_3)_2$?

$$Ca(s) + N_2(g) + 3\ O_2(g) \rightarrow Ca(NO_3)_2(s)$$

Answer

0.0183 mol O_2

22. How many moles of C_2H_5OH are needed to generate 106.7 g of H_2O?

$$C_2H_5OH(l) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(l)$$

Answer

1.974 mol C_2H_5OH

23. What mass of O_2 can be generated by the decomposition of 100.0 g of $NaClO_3$?

$$2\ NaClO_3(s) \rightarrow 2\ NaCl(s) + 3\ O_2(g)$$

Answer

45.10 g O_2

24. What mass of Li_2O is needed to react with 1,060 g of CO_2?

$$Li_2O(aq) + CO_2(g) \rightarrow Li_2CO_3(aq)$$

Answer

7.20×10^2 g Li_2O

25. What mass of Fe_2O_3 must be reacted to generate 324 g of Al_2O_3?

$Fe_2O_3(s) + 2\ Al(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$

Answer

507 g Fe_2O_3

26. What mass of Fe is generated when 100.0 g of Al are reacted?

$Fe_2O_3(s) + 2\ Al(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$

Answer

207.0 g Fe

27. What mass of MnO_2 is produced when 445 g of H_2O are reacted?

$H_2O(l) + 2\ MnO_4^-(aq) + Br^-(aq) \rightarrow BrO_3^-(aq) + 2\ MnO_2(s) + 2\ OH^-(aq)$

Answer

4290 g MnO_2

28. What mass of $PbSO_4$ is produced when 29.6 g of H_2SO_4 are reacted?

$Pb(s) + PbO_2(s) + 2\ H_2SO_4(aq) \rightarrow 2\ PbSO_4(s) + 2\ H_2O(l)$

Answer

91.5 g $PbSO_4$

29. If 83.9 g of ZnO are formed, what mass of Mn_2O_3 is formed with it?

$Zn(s) + 2\ MnO_2(s) \rightarrow ZnO(s) + Mn_2O_3(s)$

Answer

163 g Mn_2O_3

30. If 14.7 g of NO_2 are reacted, what mass of H_2O is reacted with it?

$3\ NO_2(g) + H_2O(l) \rightarrow 2\ HNO_3(aq) + NO(g)$

Answer

1.92 g H_2O

31. If 88.4 g of CH_2S are reacted, what mass of HF is produced?

$CH_2S + 6\ F_2 \rightarrow CF_4 + 2\ HF + SF_6$

Answer

76.7 g HF

32. If 100.0 g of Cl_2 are needed, what mass of NaOCl must be reacted?

$$NaOCl + HCl \rightarrow NaOH + Cl_2$$

Answer

105.0 g NaOCl

33. Calculate the mass of each product when 100.0 g of CuCl react according to the reaction

$$2\ CuCl(aq) \rightarrow CuCl_2(aq) + Cu(s)$$

What do you notice about the sum of the masses of the products? What concept is being illustrated here?

Answer

67.90 g $CuCl_2$; 32.10 g Cu; the sum is 100.0 g; this agrees with the Law of Conservation of Mass

34. Calculate the mass of each product when 500.0 g of $SnCl_2$ react according to the reaction

$$2\ SnCl_2(aq) \rightarrow SnCl_4(aq) + Sn(s)$$

What do you notice about the sum of the masses of the products? What concept is being illustrated here?

Answer

343.5 g $SnCl_4$; 156.5 g Sn; the sum is 500.0 g; this agrees with the Law of Conservation of Mass

35. Write the balanced equation and then determine the number of moles and the mass of chlorine, Cl_2, required to react with 10.0 g of sodium metal, Na, to produce sodium chloride, NaCl.

Answer

$2\ Na + Cl_2 \rightarrow 2\ NaCl$; 0.217 mol Cl_2; 15.4 g Cl_2

36. Write the balanced equation and then determine the number of moles and the mass of oxygen formed by the decomposition of 1.252 g of mercury(II) oxide.

Answer

$2\ HgO \rightarrow 2\ Hg + O_2$; 2.890×10^{-3} mol O_2; 0.09249 g O_2

37. Write the balanced equation and then determine the number of moles and the mass of sodium nitrate, $NaNO_3$, required to produce 128 g of oxygen. ($NaNO_2$ is the other product.)

Answer

$2\ NaNO_3 \rightarrow O_2 + 2\ NaNO_2$; 8.00 mol $NaNO_3$; 6.80×10^2 g $NaNO_3$

38. Write the balanced equation and then determine the number of moles and the mass of carbon dioxide formed by the combustion of 20.0 g of carbon in an excess of oxygen.

Answer

$C + O_2 \rightarrow CO_2$; 1.67 mol CO_2; 73.3 g CO_2

39. Write the balanced equation and then determine the number of moles and the mass of Mg required to react with 5.00 g of HCl and produce $MgCl_2$ and H_2.

Answer

$Mg + 2\ HCl \rightarrow MgCl_2 + H_2$; 0.0686 mol Mg; 1.67 g Mg

40. Write the balanced equation and then determine the number of moles and the mass of oxygen formed by the decomposition of 1.252 g of silver(I) oxide.

Answer

$2\ Ag_2O \rightarrow O_2 + 4\ Ag$; 2.701×10^{-3} mol O_2; 0.08644 g O_2

41. Write the balanced equation and then determine the number of moles and the mass of water formed by the combustion of 20.0 g of acetylene, C_2H_2, in an excess of oxygen.

Answer

$2\ C_2H_2 + 5\ O_2 \rightarrow 4\ CO_2 + 2\ H_2O$; 0.768 g H_2O; 13.8 g H_2O

42. Write the balanced equation and then determine the number of moles and the mass of barium peroxide, BaO_2, needed to produce 2.500 g of barium oxide, BaO (O_2 is the other product.)

Answer

$2\ BaO_2 \rightarrow 2\ BaO + O_2$; 0.01630 mol BaO_2; 2.761 g BaO_2

43. I_2 is produced by the reaction of 0.4235 mol of $CuCl_2$ according to the following equation: $2\ CuCl_2 + 4\ KI \rightarrow 2\ CuI + 4\ KCl + I_2$
    a. How many moles of I_2 are produced?
    b. What mass of I_2 is produced?

Answer

    a. 0.2118 mol I_2
    b. 53.74 g I_2

44. Silver is often extracted from ores as $K[Ag(CN)_2]$ and then recovered by the reaction: $2\ K[Ag(CN)_2] + Zn \rightarrow 2\ Ag + Zn(CN)_2 + 2\ KCN$
    a. How many moles of $Zn(CN)_2$ are produced by the reaction of 35.27 g of $K[Ag(CN)_2]$?
    b. What mass of $Zn(CN)_2$ is produced?

Answer

a. 0.08861 mol $Zn(CN)_2$
b. 10.41 g $Zn(CN)_2$

8.6: Limiting Reactants and Excess Reactants

45. Given the statement “20.0 g of methane is burned in excess oxygen,” is it obvious which reactant is the limiting reactant?

Answer

Yes. If oxygen is in excess, the methane must be the limiting reactant.

46. Given the statement “the metal is heated in the presence of excess hydrogen,” is it obvious which substance is the limiting reactant despite not specifying any quantity of reactant?

Answer

Yes. If the hydrogen is in excess, the metal must be the limiting reactant.

47. Acetylene (C_2H_2) is formed by reacting 7.08 g of C and 4.92 g of H_2.

$$2\ C(s) + H_2(g) \rightarrow C_2H_2(g)$$

What is the limiting reactant? How many grams of C_2H_2 are produced?

Answer

Limiting reactant = C; 7.67 g C_2H_2

48. Ethane (C_2H_6) is formed by reacting 7.08 g of C and 4.92 g of H_2.

$$2\ C(s) + 3\ H_2(g) \rightarrow C_2H_6(g)$$

What is the limiting reactant? How many grams of C_2H_6 are produced?

Answer

Limiting reactant = C; 8.86 g C_2H_6

49. If 35.6 g of P_4O_6 is reacted with 4.77 g of H_2O, what is the limiting reactant, and what mass of H_3PO_3 is produced?

$$P_4O_6(s) + 6\ H_2O(l) \rightarrow 4\ H_3PO_3(aq)$$

Answer

Limiting reactant = H_2O; 14.5 g H_3PO_3

50. If 377 g of NO_2 is reacted with 244 g of H_2O, what is the limiting reactant, and how many grams of HNO_3 and NO are produced?

$$3\ NO_2(g) + H_2O(l) \rightarrow 2\ HNO_3(aq) + NO(g)$$

Answer

Limiting reactant = NO_2; 344 g HNO_3 and 82.0 g NO

51. To form the precipitate $PbCl_2$, 2.88 g of NaCl and 7.21 g of $Pb(NO_3)_2$ are mixed in solution. How much precipitate is formed? How much of which reactant is in excess?

$$2\ NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\ NaNO_3(aq)$$

Answer

6.06 g $PbCl_2$ formed; 0.34 g excess NaCl

52. In a neutralization reaction, 18.06 g of KOH are reacted with 13.43 g of HNO_3. What mass of H_2O is produced, and what mass of which reactant is in excess?

$$KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$$

Answer

3.839 g H_2O formed; 6.10 g excess KOH

53. What is the limiting reactant in a reaction that produces sodium chloride from 8 g of sodium and 8 g of diatomic chlorine?

$$2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$$

Answer

Cl_2

54. What is the limiting reactant when 30.0 g of propane, C_3H_8, is burned with 75.0 g of oxygen.

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

Answer

O_2

55. What is the limiting reactant when 0.50 g of Cr and 0.75 g of H_3PO_4 react according to the following chemical equation?

$$2\ Cr(s) + 2\ H_3PO_4(aq) \rightarrow 2\ CrPO_4(s) + 3\ H_2(g)$$

Answer

H_3PO_4

56. What is the limiting reactant when 1.50 g of lithium and 1.50 g of nitrogen combine to form lithium nitride, a component of advanced batteries, according to the following *unbalanced* equation?

$$Li + N_2 \rightarrow Li_3N$$

Answer

Li

8.7: Theoretical Yield and Percent Yield

57. What is the difference between the theoretical yield and the actual yield?

Answer

Theoretical yield is what you expect stoichiometrically from a chemical reaction; actual yield is what you actually get from a chemical reaction.

58. What is the difference between the actual yield and the percent yield?

Answer

Actual yield is the measured amount of product obtained from a chemical reaction; percent yield is the ratio of the actual yield to the amount predicted from stoichiometry.

59. A worker isolates 2.675 g of SiF_4 after reacting 2.339 g of SiO_2 with HF. What are the theoretical yield and the actual yield?

$$SiO_2(s) + 4\ HF(g) \rightarrow SiF_4(g) + 2\ H_2O(l)$$

Answer

Theoretical yield = 4.052 g; actual yield = 2.675 g

60. A worker synthesizes aspirin, $C_9H_8O_4$, according to this chemical equation. If 12.66 g of $C_7H_6O_3$ are reacted and 12.03 g of aspirin are isolated, what are the theoretical yield and the actual yield?

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + HC_2H_3O_2$$

Answer

Theoretical yield = 16.51 g; actual yield = 12.03 g

61. A chemist decomposes 1.006 g of $NaHCO_3$ and obtains 0.0334 g of Na_2CO_3. What are the theoretical yield and the actual yield?

$$2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

Answer

Theoretical yield = 0.6346 g; actual yield = 0.0334 g

62. A chemist combusts a 3.009 g sample of C_5H_{12} and obtains 3.774 g of H_2O. What are the theoretical yield and the actual yield?

$$C_5H_{12}(l) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(l)$$

Answer

Theoretical yield = 4.508 g; actual yield = 3.774 g

63. What is the percent yield in Exercise 59?

Answer

66.02%

64. What is the percent yield in Exercise 60?

Answer

72.86%

65. What is the percent yield in Exercise 61?

Answer

5.26%

66. What is the percent yield in Exercise 62?

Answer

83.72%

67. A student isolated 25 g of a compound following a procedure that would theoretically yield 81 g. What was his percent yield?

Answer

31%

68. A sample of 0.53 g of carbon dioxide was obtained by heating 1.31 g of calcium carbonate. What is the percent yield for this reaction?

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Answer

92%

69. The phosphorus pentoxide used to produce phosphoric acid for cola soft drinks is prepared by burning phosphorus in oxygen.
    a. What is the limiting reactant when 0.200 mol of P_4 and 0.200 mol of O_2 react according to the reaction: $P_4 + 5\ O_2 \rightarrow P_4O_{10}$?
    b. Calculate the percent yield if 10.0 g of P_4O_{10} is isolated from the reaction.

Answer

a. O_2
b. 87.7%

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CHAPTER OVERVIEW

9: Gases

Of the three basic phases of matter – solids, liquids, and gases – only one of them has predictable physical properties: gases. In fact, the study of the properties of gases was the beginning of the development of modern chemistry from its alchemical roots. The interesting thing about some of these properties is that they are independent of the identity of the gas. That is, it doesn't matter if the gas is helium gas, oxygen gas, or sulfur vapors; some of their behavior is predictable and, as we will find, very similar. In this chapter, we will review some of the common behaviors of gases. Gases have no definite shape or volume; they tend to fill whatever container they are in. They can compress and expand, sometimes to a great extent. Gases have extremely low densities, one-thousandth or less the density of a liquid or solid. Combinations of gases tend to mix together spontaneously; that is, they form solutions. Air, for example, is a solution of mostly nitrogen and oxygen. Any understanding of the properties of gases must be able to explain these characteristics.

9.1: Straws!
9.2: Kinetic Molecular Theory
9.3: Gas Pressure
9.4: Pressure and Volume
9.5: Volume and Temperature
9.6: Pressure and Temperature
9.7: The Combined Gas Law
9.8: For Future Use
9.9: Exercises

Template:HideTOC

Thumbnail Chapter 9: Movement of gas molecules. (Julio Miguel A Enriquez and Monica Muñoz via Wikimedia Commons)

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9.1: Straws!

A drinking straw is a tube for transferring a beverage from its container to the mouth of the drinker and is typically a thin tube of plastic (such as polypropylene and polystyrene) or other material. Many people believe that when they drink a liquid they are sucking the liquid up, however the liquid is really being pushed up. A straw works because when you suck the air out of the straw, it creates a vacuum. This causes a decrease in air pressure on the inside of the straw. Since the atmospheric pressure is greater on the outside of the straw, liquid is forced into and up the straw and into your mouth (Figure 9.1.1).



Figure **9.1.1**: Since the pressure acting on the surface of the liquid inside the straw is less than atmospheric pressure acting on the surface of the liquid inside the cup, the liquid gets pushed up the straw. (image generated by Lance S. Lund using Microsoft Designer)

⛓ How Long of a Straw is Possible?

With the straw just sitting in the glass, the pressure on the surface of the beverage is the same all over, including on the little bit of surface inside the straw. When you suck the air out of the straw, you decrease the pressure inside the straw, allowing the higher pressure on the rest of the surface to push the drink up the straw and into your mouth. Because it is really the atmosphere that is doing the pushing, the atmospheric pressure limits how high a liquid will go up a straw.

If you formed a perfect vacuum within the straw, the pressure outside of the straw at sea level would be enough to push water to a total height of about 10.3 m or 33.8 feet. A 10.3-m column of water exerts the same pressure – 101,325 N/m^2 or 14.7 lb/in^2 (psi) – as do the gas molecules in our atmosphere. At sea level, the air pressure is enough to support a column of water about thirty feet high. This means that even if you could suck all the air out of a 40-foot straw, the water would not rise more than 33.8 feet.

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9.2: Kinetic Molecular Theory

Learning Objectives

- State the major concepts behind the kinetic molecular theory of gases.
- Relate the general properties of gases to the kinetic molecular theory.

Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic molecular theory of gases**. It is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average speed of gas particles is dependent on the temperature of the gas.

Figure 9.2.1 provides a representation of how the gas phase may be mentally pictured.

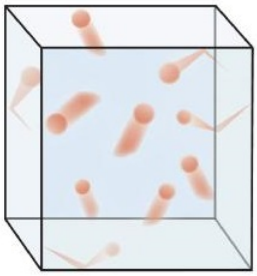


Figure **9.2.1**: The kinetic molecular theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because gases are mostly empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic molecular theory. Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

Summary

- The physical behavior of gases is explained by the kinetic molecular theory of gases.
- An ideal gas adheres exactly to the kinetic molecular theory of gases.

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9.3: Gas Pressure

Learning Objectives

- Define pressure.
- Learn the units of pressure and how to convert between them.

The kinetic molecular theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces distributed over the area of the walls of the container causes something we call pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$\text{pressure} = \frac{\text{force}}{\text{area}} \Rightarrow P = \frac{F}{A} \tag{9.3.1}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure. In this case, the gas is being "held in" by the earth's gravity, rather than the gas being in a container. The pressure of the atmosphere at sea level is about 14.7 pounds of force for every square inch of surface area: 14.7 lb/in^2.

Pressure has a variety of units. The formal, SI unit of pressure is the **pascal** (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square meter). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere** (atm), which was originally defined as the average atmospheric pressure at sea level.

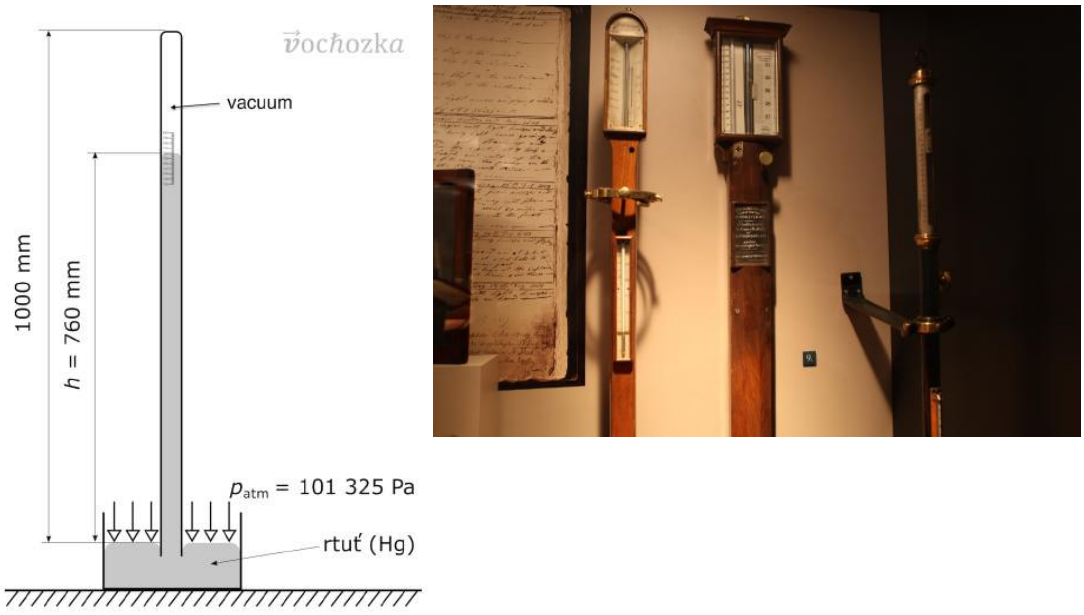


Figure 9.3.1: Left: Measurement of barometric pressure with a mercury barometer (Vladimir658 via Wikimedia Commons). Right: Mercury barometers (Kelly Michals via Flickr).

However, "average atmospheric pressure at sea level" is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimeters of mercury** (mm Hg), which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer (see Figure 9.3.1).

With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mm Hg, or 760 torr. We thus have the following equivalents:

1 atm = 760 mm Hg = 760 torr

Additional pressure conversions include:

1 atm = 29.92 in. Hg = 14.7 psi (14.7 lb/in^2) = 101,325 Pa = 1.01325 bars

We can use these equivalents as with any equivalence – to perform conversions from one unit to another. A more complete set of conversions used throughout this text may be found in Section 17.2.

✅ Example 9.3.1: Pressure Conversion

Express a pressure of 11.7 psi in units of mm Hg and Pa.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 11.7 psi Find: ? mm Hg, ? Pa
List other known quantities.	1 atm = 760 mm Hg 1 atm = 14.7 psi = 101,325 Pa Therefore, 760 mm Hg = 14.7 psi = 101,325 Pa
Prepare concept maps using the proper conversion factor(s).	psi $\xrightarrow[14.7\ \text{psi}]{760\ \text{mm Hg}}$ mm Hg psi $\xrightarrow[14.7\ \text{psi}]{101{,}325\ \text{Pa}}$ Pa
Calculate and cancel units.	$11.7\ \cancel{\text{psi}} \times \frac{760\ \text{mm Hg}}{14.7\ \cancel{\text{psi}}} = 605\ \text{mm Hg}$ $11.7\ \cancel{\text{psi}} \times \frac{101{,}325\ \text{Pa}}{14.7\ \cancel{\text{psi}}} = 8.06 \times 10^4\ \text{Pa}$
Think about your results.	11.7 psi torr is less than 1 atm, as are the final answers. (605 mm Hg, 8.06×10^4 Pa)

✏️ Exercise 9.3.1

A. How many atmospheres are there in 1,022 torr?

B. Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer A

1.345 atm

Answer B

653 torr

Summary

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted.

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9.4: Pressure and Volume

Learning Objectives

- Learn what is meant by the term gas laws.
- Learn and apply Boyle's Law.

When seventeenth-century scientists began studying the physical properties of gases, they noticed some simple relationships between some of the measurable properties of the gas. Consider the relationship between pressure (P) and volume (V), as shown in Figure 9.4.1.

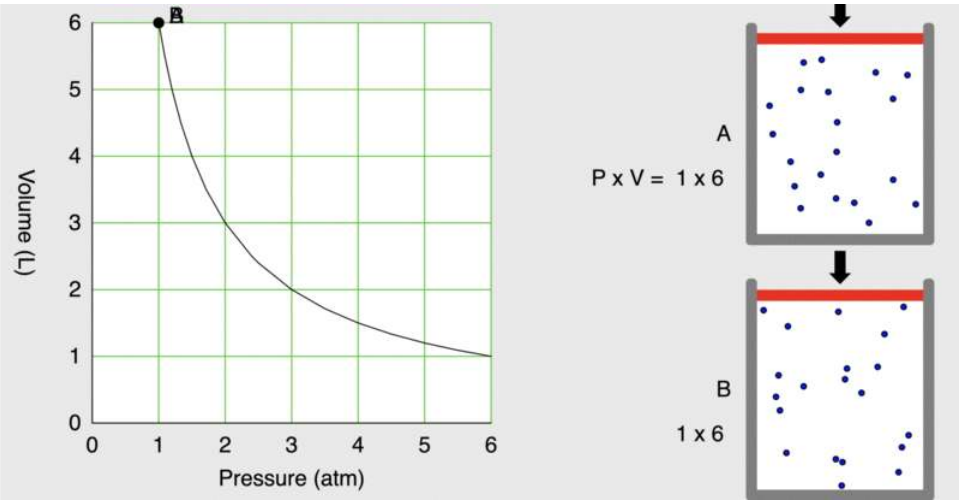


Figure 9.4.1: As the pressure on a piston increases, the volume inside the piston decreases. Note that the product, $P \times V$, remains constant.

Boyle's Law

Assuming that the amount of a gas and the temperature (T) of the gas were kept constant, scientists found that pressure and volume were *inversely proportional*, i.e. as one increased, the other decreased in a proportionate manner. Or, as one decreased, the other increased in a proportionate manner.

This means that when two variables are inversely proportional, doubling one variable causes the other to be cut in half. Reducing one variable to ⅕ of its original value will cause the other variable to increase by a factor of 5. As a result, the product of the two variables remains constant. In the case of pressure and volume,

$$P \times V = \text{constant} \tag{9.4.1}$$

Once again, this is for a given amount of gas at a constant temperature. If either volume or pressure changes while amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labeled P_1 and V_1 and the new conditions are labeled P_2 and V_2, we obtain

$$P_1V_1 = P_2V_2 \tag{9.4.2}$$

Equation 9.4.2 is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behavior of a gas. This particular gas law is called **Boyle's Law**, after the English scientist Robert Boyle, who first announced it in 1662.

Tactics for working with mathematical formulas are different from tactics for working with conversion factors.

- First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables.

- Second, in most formulas, some mathematical rearrangements that require basic algebra must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation.

- Finally, units must be consistent. For example, in Boyle's Law there are two pressure variables; they must have the same unit. There are also two volume variables; they also must have the same units. In most cases, it won't matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

✅ Example 9.4.1

4.01 L of nitrogen gas has an initial pressure of 2.44 atm. What is the volume of the gas if the pressure changes to 1.93 atm? Assume a constant temperature.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 2.44 atm, V_1 = 4.01 L, P_2 = 1.93 atm Find: V_2 = ? L
List other known quantities.	none
Plan the problem.	Rearrange the equation to solve for V_2. $V_2 = \frac{P_1 \times V_1}{P_2}$
Calculate and cancel units.	Substitute the given quantities into the equation and solve. $V_2 = \frac{2.44 \cancel{\text{ atm}} \times 4.01 \text{ L}}{1.93 \cancel{\text{ atm}}} = 5.07 \text{ L}$
Think about your result.	Pressure and volume are inversely proportional. The pressure decreased from 2.44 atm to 1.93 atm, so volume should increase to compensate, which it did (from 4.01 L to 5.07 L).

✅ Example 9.4.2:

An 88.8 mL sample of carbon dioxide has an initial pressure of 722 mm Hg. What is the pressure of the gas if its volume changes to 0.663 L.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 722 mm Hg, V_1 = 88.8 mL, V_2 = 0.633 L Find: P_2 = ? mm Hg
List other known quantities.	1 L = 1000 mL
Plan the problem.	Rearrange the equation to solve for P_2. $$P_2 = \frac{P_1 \times V_1}{V_2}$$
Calculate and cancel units.	$$0.663 \cancel{L} \times \frac{1000 \text{ mL}}{1 \cancel{L}} = 663 \text{ mL}$$ Substitute the known quantities into the equation and solve. $$P_2 = \frac{722 \text{ mm Hg} \times 88.8 \cancel{\text{mL}}}{663 \cancel{\text{mL}}} = 96.7 \text{ mm Hg}$$
Think about your result.	The volume increased, so the pressure should have decreased, which it did (from 722 mm Hg to 96.7 mm Hg).

Exercise 9.4.1

A. What pressure is required to compress a 124 mL of oxygen gas at 102 torr to a volume of 37.8 mL?

B. Suppose that the pressure of 456 mL of helium at 5.96 psi was changed to 1.55 atm. What is the new volume of the helium, in units of milliliters?

Answer A

335 torr (Note: temperature may be assumed to be constant in instances where temperature data is not provided.)

Answer B

119 mL (Note: temperature may be assumed to be constant in instances where temperature data is not provided.)

Summary

- The behavior of gases can be modeled with gas laws.
- Boyle’s Law relates the pressure and volume of a gas at constant temperature and amount.

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9.5: Volume and Temperature

Learning Objectives

- Learn and apply Charles's Law.

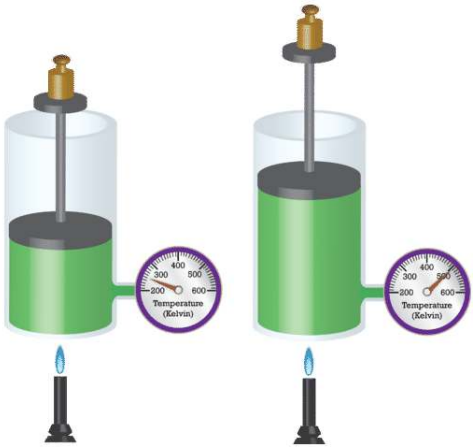


Figure **9.5.1**: Baking bread. (Petty Officer 3rd Class Charles Oki, US Navy via Wikimedia)

Everybody enjoys the smell and taste of freshly-baked bread. It is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which causes the dough to expand. The end result is an enjoyable treat, especially when covered with melted butter.

Charles's Law

Consider the relationship between the volume of a gas (V) and its temperature (T), as shown in Figure **9.5.2**. Change the temperature on the simulator on the right side of the figure by dragging the liquid level in the thermometer up and down. Notice the effect of changing the temperature on the volume.



(a)

403 ERROR

The request could not be satisfied.

Request blocked. We can't connect to the server for this app or website at this time. There might be too much traffic or a configuration error. Try again later, or contact the app or website owner.
If you provide content to customers through CloudFront, you can find steps to troubleshoot and help prevent this error by reviewing the CloudFront documentation.

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Generated by cloudfront (CloudFront)
Request ID: F-U6c41qEdV-rCyiwbgcT33PRJRlFELbM
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(b)

Figure **9.5.2**: (a) As a container of confined gas is heated, its molecules increase in kinetic energy, resulting in an increase in volume. (b) *V* and *T* Simulator: change the temperature by dragging the orange dot at the top of the liquid level in the thermometer.

French physicist Jacques Charles (1746-1823) studied the effect of temperature on the volume of a gas at constant pressure. Charles found that the volume of a given amount of gas increased in proportion with the absolute temperature of the gas. This means volume and absolute temperature are *directly proportional*, i.e. as one increases, the other increases in a proportionate manner. Or, as one decreases, the other decreases in a proportionate manner.

When two variables are directly proportional, doubling one variable causes the other to double as well. Reducing one variable to ⅕ of its original value will cause the other variable to decrease to ⅕ of its original. As a result, the ratio of the two variables remains constant. In the case of volume and absolute temperature,

$$\frac{V}{T} = \text{constant} \tag{9.5.1}$$

Once again, this is for a given amount of gas at a constant pressure. If either volume or temperature changes while amount and pressure stay the same, then the other property must change so that the ratio of the two properties still equals that same constant. That is, if the original conditions are labeled V_1 and T_1 and the new conditions are labeled V_2 and T_2, we obtain

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \tag{9.5.2}$$

This relationship is known as **Charles's Law**. It is important to note that the temperatures are absolute temperatures. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stop of molecular motion.

When volume and absolute temperature data are graphed, the result is a straight line, indicative of a direct proportion, shown in the figure below.

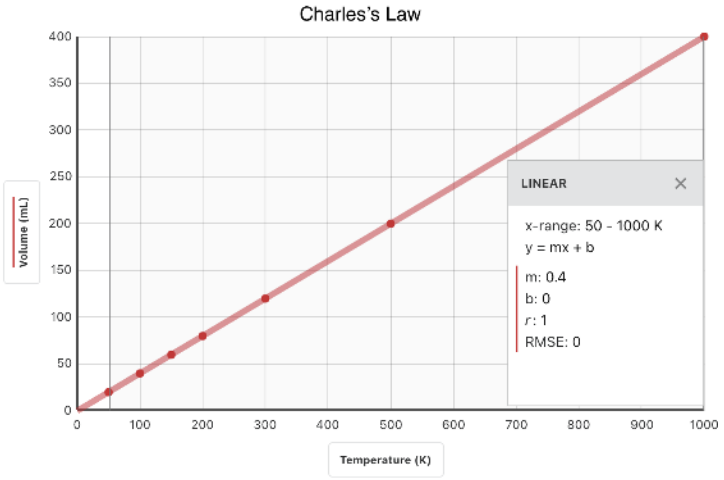


Figure 9.5.3:The volume of a gas increases as the Kelvin temperature increases.

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

The direct relationship will only hold if the temperatures are expressed in kelvins; Celsius temperatures will not work. Recall that $T_K = T_{°C} + 273.15$.

✅ Example 9.5.1:

A balloon is filled to a volume of 2.20 L at a temperature of 22°C. The balloon is then heated to a temperature of 71°C. Find the new volume of the balloon.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: V_1 = 2.20 L, T_1 = 22°C = 295 K, T_2 = 71°C = 344 K Find: V_2 = ? L
List other known quantities.	The temperatures were first converted to kelvins. $T_K = T_{°C} + 273.15$
Plan the problem.	Rearrange the equation to solve for V_2. $V_2 = \frac{V_1 \times T_2}{T_1}$
Calculate and cancel units.	Substitute the given quantities into the equation and solve. $V_2 = \frac{2.20\ \text{L} \times 344\ \cancel{\text{K}}}{295\ \cancel{\text{K}}} = 2.57\ \text{L}$
Think about your result.	The volume increases as the temperature increases. Note: Pressure was assumed to be constant.

✅ Example 9.5.2:

An air sample has a volume of 34.8 L and at −67ºC. What must be the temperature of the air, in ºC, for the volume to decrease to 25.0 L?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: V_1 = 34.8 L, T_1 = –67°C = 206 K, V_2 = 25.0 L Find: T_2 = ? K
List other known quantities.	The temperatures were first converted to kelvins. $T_K = T_{°C} + 273.15$
Plan the problem.	Rearrange the equation to solve for T_2. $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ Cross multiply to yield $V_1 \times T_2 = V_2 \times T_1$ $\frac{\cancel{V_1} \times T_2}{\cancel{V_1}} = \frac{V_2 \times T_1}{V_1} \Rightarrow \boxed{T_2 = \frac{V_2 \times T_1}{V_1}}$
Calculate and cancel units.	Substitute the known quantities into the equation and solve. $T_2 = \frac{V_2 \times T_1}{V_1} = \frac{(25.0 \cancel{\text{ L}})(206 \text{ K})}{34.8 \cancel{\text{ L}}} = 148 \text{ K}$ $T_{°C} = T_K - 273.15 = 148 - 273.15 = -125°C$
Think about your result.	The temperature decreased, meaning the volume decreased (which it did). Note: Pressure was assumed to be constant.

Exercise 9.5.1

A. What is the volume of a balloon if it was initially filled with helium to a volume of 3.77 L at –18°C and cooled much further to a temperature of –150°C?

B. A sample of nitrogen gas has a volume of 277 mL at −38°C. If the sample originally had a volume of 623 mL, what was its temperature at that volume, in °C?

Answer A

1.82 L

Answer B

256°C

Summary

- Charles’s Law relates the volume and temperature of a gas at constant pressure and amount.

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9.6: Pressure and Temperature

Learning Objectives

- Learn and apply Gay-Lussac's Law.

Propane tanks are widely used with barbeque grills. But it's not fun to find out halfway through grilling that you have run out of gas. You can buy gauges that measure the pressure inside the tank to see how much is left. The gauge measures pressure and will register a higher pressure on a hot day than it will on a cold day. So you need to take the air temperature into account when you decide whether or not to refill the tank before your next cook-out.

Gay-Lussac's Law

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. This is because the increase in kinetic energy at higher temperatures results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778-1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given amount of gas varies in direct proportion to the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} = \text{constant} \tag{9.6.1}$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \tag{9.6.2}$$

On the simulator in Figure 9.6.1 below, click the $\boxed{\text{Gay} - \text{Lussac's Law}}$ button. Then click on the $\Uparrow$ and $\Downarrow$ arrows to vary the temperature and the pressure. Then click the $\boxed{\text{Add Data}}$ button to add up to 5 points on the graph.



Figure 9.6.1: A gas laws simulator courtesy of teachchemistry.org.

Notice that the graph of pressure vs. temperature also illustrates a direct proportion. As a gas is cooled at constant volume, its pressure continually decreases until the gas condenses to a liquid.

✅ Example 9.6.1

The gas in an aerosol can is under a pressure of 3.00 atm at a temperature of 25ºC. It is dangerous to dispose of an aerosol can by incineration. Calculate the pressure in the aerosol can at a temperature of 845ºC.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: P_1 = 3.00 atm, T_1 = 25°C = 298 K, T_2 = 845°C = 1118 K Find: P_2 = ? atm
List other known quantities.	The temperatures were first converted to kelvins. $T_K = T_{°C} + 273.15$
Plan the problem.	Rearrange the equation to solve for P_2. $$P_2 = \frac{P_1 \times T_2}{T_1}$$
Calculate and cancel units.	Now substitute the known quantities into the equation and solve. $$P_2 = \frac{3.00\ \text{atm} \times 1118\ \cancel{\text{K}}}{298\ \cancel{\text{K}}} = 11.3\ \text{atm}$$
Think about your result.	The pressure increases dramatically due to a large increase in temperature.

Exercise 9.6.1

An incandescent lightbulb has an internal pressure of 13.9 psi at 20°C. It is designed to withstand an internal pressure of up to 41.7 psi. To what temperature, in °C, may the bulb be heated before it exceeds the maximum pressure it was designed to withstand?

Answer

606°C

Summary

- Pressure and temperature at constant volume are directly proportional.

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9.7: The Combined Gas Law

Learning Objectives

- Learn and apply the Combined Gas Law.

One thing that may be noticed about all the gas laws is that, collectively, volume and pressure are always in the numerator, and temperature is always in the denominator. This suggests that a gas law that combines pressure, volume, and temperature for a given amount of gas may be proposed. This gas law is known as the **Combined Gas Law**, and its mathematical form is:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \tag{9.7.1}$$

This allows changes in all three major properties of a gas to be followed. Again, the usual warnings apply when algebraically solving for an unknown variable (it must be isolated on one side of the equation in the numerator), units (P_1 and P_2 must each the same units, as must V_1 and V_2), and units of temperature must be in kelvins.

✅ Example 9.7.1:

An 8.33-L sample of gas at 1.82 atm and 286 K simultaneously changes its temperature to 355 K and volume to 5.72 L. What is the final pressure of the gas?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: V_1 = 8.33 L, P_1 = 1.82 atm, T_1 = 286 K, V_2 = 5.72 L, T_2 = 355 K Find: P_2 = ? atm
List other known quantities.	none
Plan the problem.	Rearrange the equation to solve for P_2. $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$ $$\frac{P_1V_1}{T_1} \times \frac{T_2}{V_2} = \frac{P_2 \cancel{V_2}}{\cancel{T_2}} \times \frac{\cancel{T_2}}{\cancel{V_2}} \Rightarrow \boxed{P_2 = \frac{P_1V_1T_2}{T_1V_2}}$$
Calculate and cancel units.	Substitute the known quantities into the equation and solve. $$P_2 = \frac{(1.82 \text{ atm})(8.33 \cancel{\text{L}})(355 \cancel{\text{K}})}{(286 \cancel{\text{K}})(5.72 \cancel{\text{L}})} = 3.29 \text{ atm}$$
Think about your result.	The pressure increased, which is a little more difficult to predict because two properties of the gas were changing.

Exercise 9.7.1

A weather balloon was filled with helium to a volume of 256 L at 760 mm Hg and 0°C. The weather balloon rose to an altitude where the volume of the balloon expanded slightly to 272 L and the pressure inside the balloon decreased to 684 mm Hg. If no gas was added to or removed from the balloon, what was the new temperature of the helium inside the balloon, in °C?

Answer

–12°C

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \xrightarrow{\text{cross multiply}} P_1V_1T_2 = P_2V_2T_1 \xrightarrow{\text{solve for } T_2} T_2 = \frac{P_2V_2T_1}{P_1V_1} = \frac{(684 \cancel{\text{ mm Hg}})(272 \cancel{\text{L}})(273.15 \text{ K})}{(760 \cancel{\text{ mm Hg}})(256 \cancel{\text{L}})} = 261 \text{ K} = -12°\text{C}$$

As with other gas laws, when calculating the value of a variable in the denominator, you may either cross-multiply all the terms (as shown above) or to first take the reciprocal to get the variable in the numerator. As always, the variable for which you are solving for must be in the numerator and all by itself on one side of the equation.

As with other gas laws, if you need to determine the value of a variable in the denominator of the combined gas law, you can either cross-multiply all the terms or just take the reciprocal of the combined gas law. Remember, the variable you are solving for must be

in the numerator and all by itself on one side of the equation.

Summary

- The Combined Gas Law relates pressure, volume, and temperature of a gas.

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9.9: Exercises

9.3: Gas Pressure

1. Why are sharp knives more effective than dull knives (Hint: think about the definition of pressure)?

Answer

The cutting edge of a knife that has been sharpened has a smaller surface area than a dull knife. Since pressure is force per unit area, a sharp knife will exert a higher pressure with the same amount of force and cut through material more effectively.

2. Why do some small bridges have weight limits that depend on how many wheels or axles the crossing vehicle has?

Answer

More axles or wheels will increase the surface area of the vehicle with the bridge, decreasing the pressure.

3. Why should you roll or belly-crawl rather than walk across a thinly-frozen pond?

Answer

Lying down distributes your weight over a larger surface area, exerting less pressure on the ice compared to standing up. If you exert less pressure, you are less likely to break through thin ice.

4. A typical barometric pressure in Redding, California, is about 750 mm Hg. Calculate this pressure in atm and kPa.

Answer

0.99 atm; 1.0×10^2 kPa

5. A typical barometric pressure in Denver, Colorado, is 615 mm Hg. What is this pressure in atmospheres and kilopascals?

Answer

0.809 atm; 82.0 kPa

6. A typical barometric pressure in Kansas City is 740 torr. What is this pressure in atmospheres, in millimeters of mercury, and in kilopascals?

Answer

0.97 atm; 740 mm Hg; 99 kPa

7. Canadian tire pressure gauges are marked in units of kilopascals. What reading on such a gauge corresponds to 32 psi?

Answer

2.2×10^2 kPa

8. During the Viking landings on Mars, the atmospheric pressure was determined to be on the average about 6.50 millibars. What is that pressure in torr and kPa?

Answer

4.88 torr; 0.650 kPa

9. The pressure of the atmosphere on the surface of the planet Venus is about 88.8 atm. Compare that pressure in psi to the normal pressure on earth at sea level in psi.

Answer

Earth: 14.7 psi; Venus: 1.31×10^3 psi

10. A medical laboratory catalog describes the pressure in a cylinder of a gas as 14.82 MPa. What is the pressure of this gas in atmospheres and torr?

Answer

146.3 atm; 1.112×10^5 torr

11. Consider this scenario and answer the following questions: On a mid-August day in the northeastern United States, the following information appeared in the local newspaper: atmospheric pressure at sea level 29.97 in. Hg (1015 mbar).
    a. What was the pressure in kPa?
    b. The pressure near the seacoast in the northeastern United States is usually reported near 30.0 in. Hg. During a hurricane, the pressure may fall to near 28.0 in. Hg. Calculate the drop in pressure in torr.

Answer

a. 101.5 kPa
b. 51 torr drop

9.4: Pressure and Volume

12. If the volume of a fixed amount of a gas is tripled at constant temperature, what happens to the pressure?

Answer

The pressure decreases by a factor of 3.

13. A 2.0 mL sample of air in a syringe has a pressure of 743 mm Hg. What is the pressure of the air when the volume is increased to 16.0 mL at constant temperature?

Answer

93 mm Hg

14. A container of nitrogen gas has a volume of 34.1 L at a pressure of 0.643 atm and 298 K. What will be the volume of the nitrogen if the pressure is increased to 0.917 atm at 298 K?

Answer

23.9 L

15. A sealed bag of chips contains air with a pressure of 755 mm Hg and a volume of 1.00 L. What will be the volume of the air in the bag when the chips are shipped over mountains where the pressure is 635 mm Hg?

Answer

1.19 L

16. One mol of N_2 gas has a volume of 6.318 L at a pressure of 1.000 atm at −196°C. What pressure is needed to compress the N_2 gas to 34.65 mL at the same temperature?

Answer

182.3 atm

17. A large scuba tank with a volume of 18 L is rated for a pressure of 220 bar. The tank is filled at 20°C and contains enough air to supply 1860 L of air to a diver at a pressure of 2.37 atm (a depth of 45 feet). Was the tank filled to capacity at 20°C?

Answer

Yes. The pressure of the 18 L of air in the scuba tank would be 240 atm or 240 bar.

9.5: Volume and Temperature

18. Sometimes leaving a bicycle in the sun on a hot day will cause a blowout. Why?

Answer

The temperature of the air in the bicycle tire will increase, increasing the volume of the air. If the volume of the air exceeds the capacity of the tire, it will cause a blowout.

19. What is the temperature of an 11.2 L sample of carbon monoxide, CO, at 744 torr if it occupies 13.3 L at 55°C and 744 torr?

Answer

276 K or 3°C

20. A 3.00 L sample of air at 25 C is cooled to –196°C using liquid nitrogen at constant pressure. What is the new volume of the air?

Answer

0.78 L

21. What temperature is required to cause a 1.15 L balloon at 35°C and 1.1 atm to occupy a volume of 2.55 L at 1.1 atm?

Answer

683 K or 410.°C

22. A 2.50 L volume of hydrogen measured at –100.°C is warmed to 100.°C. Calculate the volume of the gas at the higher temperature, assuming no change in pressure.

Answer

5.39 L

9.6: Pressure and Temperature

23. If the temperature of a fixed amount of a gas is doubled at constant volume, what happens to the pressure?

Answer

The pressure is also doubled.

24. A 2.00 L steel container holding 456 torr of oxygen gas at 25°C is heated. What will be the temperature of the gas when the pressure is 775 torr?

Answer

506 K or 233°C

25. What will be the pressure of a gas that is cooled to –75°C from a starting temperature and pressure of 35°C and 784 mm Hg? Assume that the volume remains constant.

Answer

504 mm Hg

26. A pressure of a sample of neon gas initially at 35°C and 1.24 atm drops to 0.872 atm. Assuming that the volume does not change, what is the new temperature of the gas?

Answer

217 K or –56°C

27. A spray can is used until it is empty except for the propellant gas, which has a pressure of 1344 torr at 23°C. If the can is thrown into a fire (T = 475°C), what will be the pressure in the hot can?

Answer

3.40×10^3 torr

9.7: The Combined Gas Law

28. What is the final temperature of a gas whose initial conditions are 667 mL, 822 torr, and 67°C and whose final volume and pressure are 1.334 L and 2.98 atm, respectively? Assume the amount remains constant.

Answer

1870 K or 1.60×10^3°C

29. What is the final pressure of a gas whose initial conditions are 1.407 L, 2.06 atm, and −67°C and whose final volume and temperature are 608 mL and 449 K, respectively? Assume the amount remains constant.

Answer

10.4 atm

30. At an altitude of 40 km above the earth's surface, the atmospheric pressure is 5.00 torr, and the surrounding temperature is −20°C. If a weather balloon filled with He at 760. torr and 22°C has a volume of 24.2 L, what is its final volume when it reaches 40 km in altitude? (Assume the pressure of the gas equals the surrounding pressure.)

Answer

3.15×10^3 L

31. If a balloon is filled with He at 760. torr and 22°C has a volume of 24.2 L, what is its final volume if it descends to the bottom of the Mariana Trench, where the surrounding temperature is 1.4°C and the pressure is 1,060 atm?

Answer

0.0212 L

32. A high altitude balloon is filled with 1.41×10^4 L of hydrogen at a temperature of 21 °C and a pressure of 745 torr. What is the volume of the balloon at a height of 20 km, where the temperature is –48 °C and the pressure is 63.1 torr?

Answer

1.27×10^5 L

33. A cylinder of medical oxygen has a volume of 35.4 L, and contains O_2 at a pressure of 151 atm and a temperature of 25 °C. What volume of O_2 does this correspond to at normal body conditions, that is, 1 atm and 37 °C?

Answer

5.56×10^3 L

34. A balloon that is 100.21 L at 21 °C and 0.981 atm is released and just barely clears the top of Mount Crumpet in British Columbia. If the final volume of the balloon is 144.53 L at a temperature of 5.24 °C, what is the pressure experienced by the balloon as it clears Mount Crumpet?

Answer

0.644 atm

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CHAPTER OVERVIEW

10: Electrons in Atoms

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Thumbnail Chapter 10: Experiments with a plasma ball. (friedpxels via Giphy)

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10.1: Waves

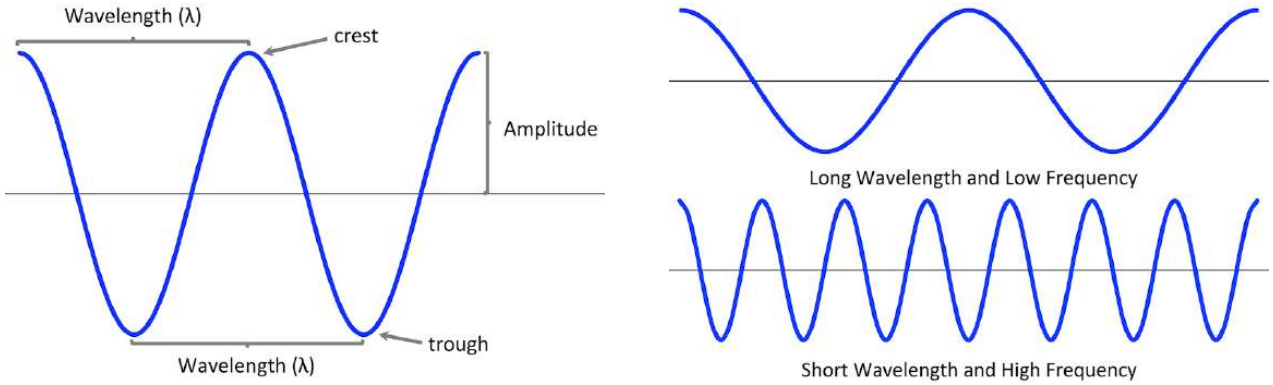
Learning Objectives

- Define the terms wavelength and frequency with respect to wave-form energy.
- State the relationship between wavelength and frequency with respect to electromagnetic radiation.

During the summer, many people enjoy going to the beach. Beach-goers can swim, have picnics, and work on their tans. If a person gets too much sun, a case of sunburn is often the result. A particular set of solar wavelengths are especially harmful to the skin. This portion of the solar spectrum is known as UV-B, with wavelengths of 280 to 320 nm. Sunscreens can be effective in protecting skin against both the immediate skin damage and the long-term possibility of skin cancer due to UV-B radiation.

Waves

Waves are characterized by their repetitive motion. Imagine a toy boat riding the waves in a wave pool. As the water wave passes under the boat, it moves up and down in a regular and repeated fashion. While the wave travels horizontally, the boat only travels vertically up and down. The figure below shows two examples of waves.



(a) (b)

Figure **10.1.1**: A wave consists of alternating crests and troughs. The wavelength (λ) is defined as the distance between any two identical points on adjacent waves. The amplitude is the height of the wave. A wave with a longer wavelength has a low frequency because fewer waves pass a given point in a certain amount of time. A wave with a shorter wavelength has a higher frequency.

A wave cycle consists of one complete wave – starting at the zero point, going up to a wave **crest**, going back down to a wave **trough**, and back to the zero point again. The **wavelength** of a wave is the distance between any two corresponding points on adjacent waves. It is easiest to visualize the wavelength of a wave as the distance from one wave crest to the next. In an equation, wavelength is represented by the Greek letter lambda (λ). Depending on the type of wave, wavelength can be measured in meters, centimeters, or nanometers (1 m = 10^9 nm). The **frequency**, represented by the Greek letter nu (ν), is the number of waves that pass a certain point in a specified amount of time. Typically, frequency is measured in units of cycles per second or waves per second. One wave per second is also called a hertz (Hz) and in SI units is a reciprocal second (s^{-1}).

Figure 10.1.1 (b) above shows an important relationship between the wavelength and frequency of a wave. The top wave has a longer wavelength than the bottom wave. However, if you picture yourself at a stationary point watching these waves pass by, fewer of the top waves would pass by in a given amount of time. Thus the frequency of the top wave is less than that of the bottom wave. Wavelength and frequency are therefore inversely related. As the wavelength of a wave increases, its frequency decreases. The equation that relates the two is:

$$c = \lambda\nu \tag{10.1.1}$$

where c is the speed of light and has a value of 2.998×10^8 m/s. For the relationship to hold mathematically, if the speed of light is used in m/s, the wavelength is measured in units of meters and the frequency in units of hertz.

Summary

- All waves can be defined in terms of their wavelength and frequency.
- $c = \lambda\nu$ expresses the relationship between wavelength and frequency.

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10.2: The Electromagnetic Spectrum

Learning Objectives

- Know the properties of different types of electromagnetic radiation.

The light given off by the sun which allows us to see is a type of electromagnetic radiation and travels in electromagnetic waves. Electromagnetic waves have an extremely wide range of wavelengths, frequencies, and energies. The highest energy form of electromagnetic waves are gamma (γ) rays and the lowest energy form are radio waves. The figure below shows the **electromagnetic spectrum**, which includes all forms of electromagnetic radiation.

Electromagnetic radiation that lies to the left in Figure 10.2.1 are non-ionizing forms of radiation. Those furthest to the left have the lowest energy and are called **radio waves**. This portion of the electromagnetic spectrum is where we find the broadcast frequencies for AM and FM radio. **Microwaves** have the next higher energy of electromagnetic radiation. It should be most obvious that microwave ovens use radiation in the microwave portion of the spectrum. However, cellphones, wireless routers, and radar also use the microwave portion of the spectrum.

Infrared radiation has the next higher energy. Many of the things we associate with giving off heat are the result of infrared radiation. Believe it or not, humans give off abundant infrared radiation. So do heat lamps used to keep food warm. Night vision goggles and thermal imagining cameras detect infrared radiation and are thus used to "see" things that give off heat. The tiny section in the middle is called the **visible spectrum** or **visible light**. This portion of the spectrum is what is visible to humans and is described in further detail below.



Figure **10.2.1**: The electromagnetic spectrum, with its various regions labeled. The borders of each region are approximate. (NASA, Public domain, via Wikimedia Commons).

Electromagnetic radiation that lies to the right of visible light in Figure 10.2.1 are ionizing forms of radiation, i.e. they cause matter to form ions when they absorb these forms of radiation. Lying just above visible light, in terms of energy, are **ultraviolet rays**. Because ultraviolet rays are an ionizing form of radiation, they may cause skin damage and a long term risk of skin cancer. Ultraviolet rays are only able to penetrate the outer layers of skin. On the other hand, **x-rays** have a higher energy and are able to penetrate the human body. This makes them useful for internal imaging of the human body. However, due to the fact they are an ionizing form of radiation, humans are advised to limit as much as possible the number of medical x-rays they have per year. **Gamma rays** have the highest energy and can be quite dangerous to living systems. Exposure is often a result of radioactive decay and should be greatly limited.

The colors of visible spectrum, shown below, may be learned using the mnemonic, ROY G BIV, for red orange yellow green blue indigo violet. Indigo is a color that lies between blue and violet on the spectrum and is simply used to make a handy mnemonic. When learning the colors in this order, you are learning them in order from low energy to high energy. Note that red lies next to the infrared portion of the electromagnetic spectrum and violet lies next to the ultraviolet portion of the spectrum.



The light energies that are in the visible range are electromagnetic waves that cause the human eye to respond when those frequencies enter the eye. The eye sends a signal to the brain and the individual "sees" various colors. The highest energy waves in the visible region cause the brain to see violet and as the energy decreases, the colors change to blue, green, yellow, orange, and red. When the energy of the wave is above or below the visible range, the eye does not respond to them. When the eye receives several different frequencies at the same time, the colors are blended by the brain. If all frequencies of light strike the eye together, the brain sees white. If there are no visible frequencies striking the eye, the brain sees black.

The objects that you see around you are light absorbers – that is, the chemicals on the surface of the object will absorb certain frequencies and not others. Your eyes detect the frequencies that strike your eye. Therefore, if your friend is wearing a red shirt, it means the dye in that shirt absorbs every frequency except red and the red frequencies are reflected. If your only light source was

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

✅ Example 10.2.1

A. Which type of radiation has the longest wavelength, microwaves or ultraviolet rays?
B. Which type of radiation has the greatest frequency, microwaves or ultraviolet rays?
C. Which type of radiation has the greatest energy, microwaves or ultraviolet rays?

Solution

Referring to Figure 10.2.1,

A. Wavelengths decrease as one moves to the right on the diagram, so microwaves have the longest wavelength.
B. Frequencies increase as one moves to the right on the diagram, so ultraviolet rays have the greatest frequency.
C. Energy increases as one moves to the right on the diagram, so ultraviolet rays have the greatest energy.

✏️ Exercise 10.2.1

A. Which color of the visible spectrum has the longest wavelength?
B. Which color of the visible spectrum has the greatest frequency?
C. Which color of the visible spectrum has the greatest energy?

Answer A

red

Answer B

violet

Answer C

violet

Summary

- Electromagnetic radiation has a wide spectrum, including gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves, and radio waves.
- The different colors of light differ in their frequencies (or wavelengths).

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10.3: Emission Spectra

The incandescent light bulbs that have been phasing out over the past decade or so contain a very thin wire filament that emits light when heated. The filament in these bulbs is typically made of tungsten, since it can be heated to high temperatures without melting. The light emitted by a hot tungsten filament contains virtually every frequency and therefore, appears white. Such a spectrum of frequencies is called a **continuous spectrum** (no gaps or lines stand out). Every element emits light when energized by heating or passing electric current through it. Solid elements begin to glow when they are heated sufficiently, while elements in gaseous form emit light when electricity passes through them or when they are atomized in a flame. This is the source of light emitted by neon signs and is also the source of light in a fire.

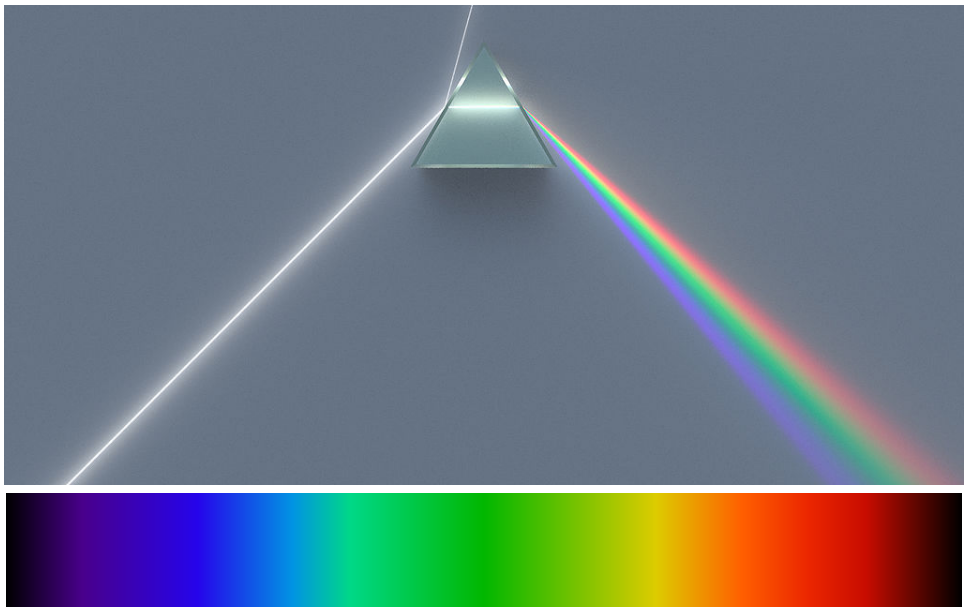


Figure **10.3.1**: White light passing through a prism to produce a continuous spectrum (top – Spigget, CC BY-SA 3.0, via Wikimedia Commons). A continuous spectrum (bottom – Stkl, Public domain, via Wikimedia Commons).

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 (see Figure 10.3.1).

The **emission spectrum** (or **line spectrum**) of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity. When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. However, when separated using a prism or diffraction grating, the hydrogen spectrum reveals four individual frequencies (see top image in Figure 10.3.2). The pink color of the tube is the result of our eyes blending the four colors.

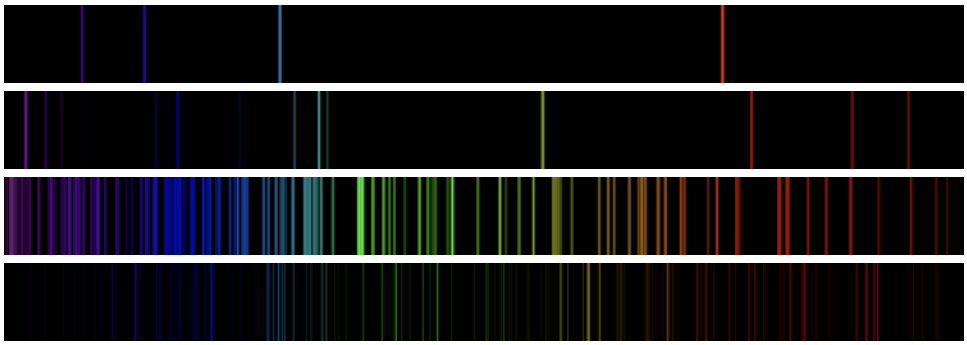


Figure **10.3.2**: From top to bottom, the atomic emission spectra for H, He, Mg, and Xe. (Christopher Thomas via Wikimedia Commons)

Every atom has its own characteristic spectrum; no two emission **spectra** (spectra is the plural form of spectrum) are alike. The images in Figure 10.3.2 above show the emission spectra of hydrogen, helium, magnesium, and xenon, respectively. Video 10.3.1 shows how line spectra may be produced and observed. Since each element has a unique emission spectrum, they are often referred to as the "fingerprints" of the elements.



Video **10.3.1**: Spectrum Demo: Continuous and Emission

Applications of Emission Spectra

You may have heard about scientists discussing which elements are present in the sun or a more distant star, and after hearing that, wondered how we could know what elements were present in a place no one has ever been. The elements present in distant stars may be determined by analyzing the spectrum that comes from the stars and observing that there are lines that represent wavelengths that have been absorbed by star's stellar atmosphere. If the exact four lines that compose hydrogen's emission spectrum are present in the lines that had been absorbed by the star's atmosphere then the star contains hydrogen.

While there is some variation in the composition of stars, hydrogen is by far the most abundant element found in stars, followed by helium, and these two elements are found in virtually all stars. Knowing this to be true, it has also been shown that missing lines in the spectra of stars in distant galaxies are not quite identical to the spectra for hydrogen and helium found here on Earth. Instead, they are almost always shifted more toward the reddish portion of the spectrum. It turns out that the more distant the galaxy, the more redshifted the spectra tend to be. Shifts in the spectra are caused by distant galaxies moving away from the Earth. The greater the redshift appears, the greater the speed at which the distant galaxy is moving away from the Earth. This is how scientists know the universe is expanding. See Video 10.3.2 for a short animation.



Video **10.3.2**: Redshift of distant galaxies.

Additional applications of emission spectra or line spectra include different color "neon" lights, not all of which contain the element neon. Fireworks of different colors are also the result of using different elements, since each element has its own unique spectrum. Video **10.3.3** shows different color flames that are the result of using salts that contain different elements.



Video **10.3.3**: Salts containing different elements yield different color flames.

Summary

- Each element has its own unique emission spectrum.

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10.4: The Bohr Model

Learning Objectives

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

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.

Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moved in a curved path, it gave 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 orbits 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.

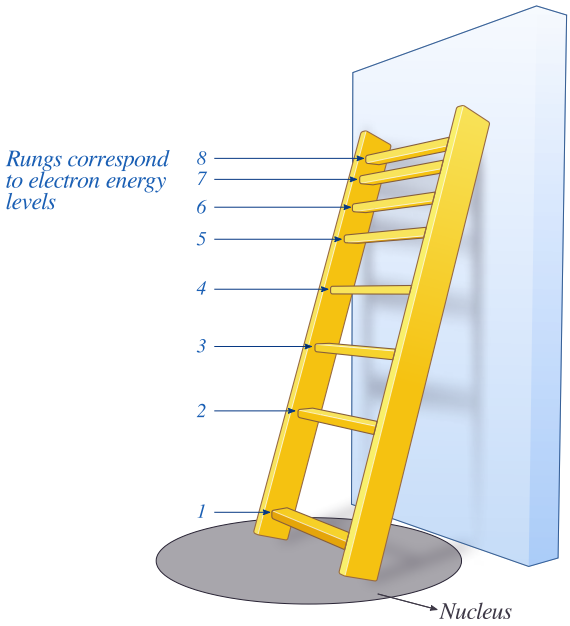


Figure **10.4.1**: The energy levels of the electrons can be viewed as rungs on a ladder. Note that the spacing between rungs gets smaller at higher energies (CC BY-NC; Ümit Kaya)

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 8 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 three people on it, they would be on the lowest three 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.

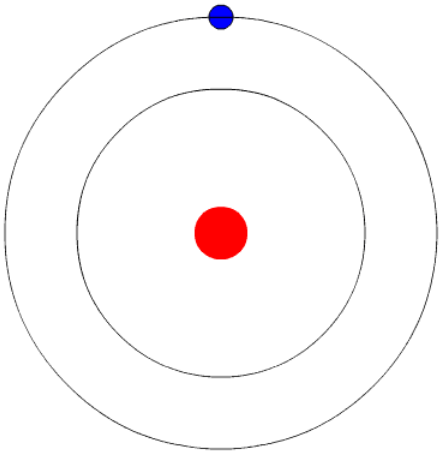


Figure 10.4.2: In Bohr's Model of the atom, electrons absorb energy to move to a higher level and release energy to move to lower levels. (CC BY-SA 3.0; Kurzon).

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 changing energy levels. The electrons are typically in their lowest state of potential energy possible, called the **ground state** (compare it to a ball on the ground in Figure 10.4.3). If the electrons are provided energy (through heat, electricity, light, etc.), the electrons in an atom could absorb energy by jumping to an energy level with a higher state of potential energy, or **excited state** (compare it to a ball at the platform level near the top in Figure 10.4.3). 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.

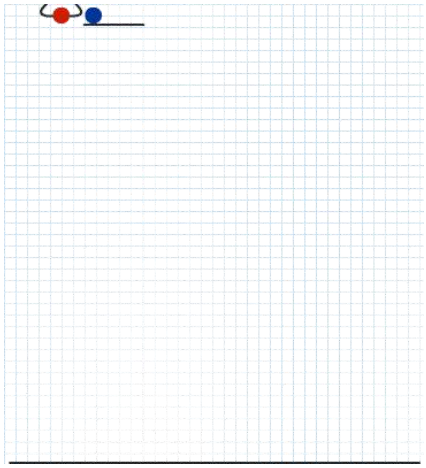


Figure 10.4.3: Two balls moving from a state of higher potential energy to a lower state of potential energy. (OilerLagrangian, CC BY-SA 4.0, via Wikimedia Commons)

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. In the case of hydrogen's emission spectrum, the wavelengths in the visible portion of the spectrum all involve transitions where an electron drops from a higher energy level down to the second energy level ($n = 2$). See Figure 10.4.4.



Figure 10.4.4: Hydrogen atom series for visible wavelengths: λ_3 = 656 nm (red), λ_4 = 486 nm (cyan) and λ_5 = 433 nm (blue). (MikeRun, CC BY-SA 4.0, via Wikimedia Commons)

As an interesting side note, these jumps between quantized energy levels are called quantum jumps or quantum leaps. Since the electron had to undergo an immediate transition from one energy level to another instantaneously, in effect, the electron could not exist in between. This expression took on a new meaning to represent huge leaps or advancements in knowledge, power, or technology. In science fiction, a quantum leap is typically thought of as being a huge jump across space and or time, as was popularized in the TV show, Quantum Leap.

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.

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10.5: Quantum Mechanics and Atomic Orbitals

Learning Objectives

- Define quantum mechanics.
- Differentiate between an orbit and an orbital.

How do you study something that seemingly makes no sense? We talk about electrons being in orbits and it sounds like we can tell where that electron is at any moment. We can draw pictures of electrons in orbit, but the reality is that we don't know exactly where they are. We are going to take a look at an area of science that even leaves scientists puzzled. When asked about quantum mechanics, Niels Bohr (who proposed the Bohr model of the atom) said: "Anyone who is not shocked by quantum theory has not understood it." Richard Feynman (one of the founders of modern quantum theory) stated: "I think I can safely say that nobody understands quantum theory." So, let's take a short trip into a land that challenges our everyday world.

Quantum Mechanics

The study of motion of large objects such as baseballs is called mechanics, or more specifically, classical mechanics. When a baseball is thrown, we can predict the exact path of the baseball at any point in time based on its velocity and the angle at which it is thrown. Because of the quantum nature of the electron and other tiny particles moving at high speeds, classical mechanics is inadequate to accurately describe their motion. **Quantum mechanics** is the study of the motion of objects that are atomic or subatomic in size and thus demonstrate wave-particle duality. In classical mechanics, the size and mass of the objects involved effectively obscures any quantum effects, so that such objects appear to gain or lose energies in any amounts. Particles whose motion is described by quantum mechanics gain or lose energy in small pieces called **quanta**.

One of the fundamental (and hardest to understand) principles of quantum mechanics is that the electron is both a particle and a wave. In the everyday macroscopic world of things we can see, something cannot be both. However, this duality exists in the quantum world of the submicroscopic on the atomic scale. At the heart of quantum mechanics is the idea that we cannot accurately specify the location of an electron. If we learn something about its location, we know less about its speed. If we learn something about its speed, we know less about its location. This observation became known as the **Heisenberg uncertainty principle** in recognition of the mathematical observation by Werner Heisenberg in 1927. All we can say is that there is a probability that it exists within a certain volume of space.

The scientist, Erwin Schrödinger, developed an equation that deals with these calculations, often referred to as the Schrödinger **wave equation**, in which electrons were treated as if they were waves. The wave equation defined **orbitals** to depict the locations of electrons in atoms to replace the circular orbits that were central to the Bohr model. These orbitals were defined as regions of space where there is a high probability of finding an electron.

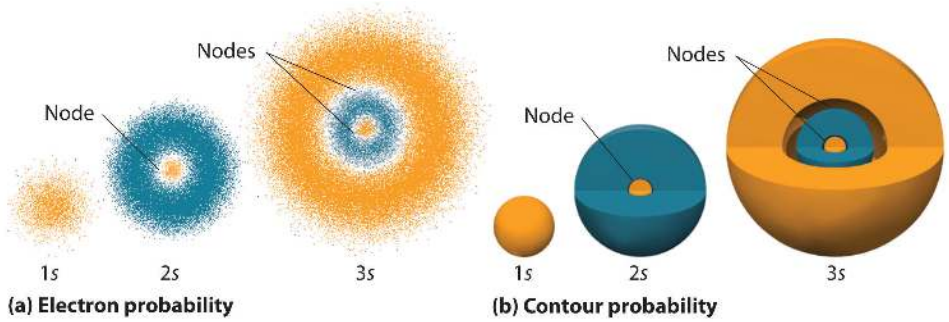


Figure 10.5.1: Depictions of atomic orbitals in hydrogen. (via LibreTexts)

When looking at the orbital depictions in Figure 10.5.1(*a*), areas where the dots are more densely shaded represent areas where finding an electron is more probable. By drawing the probabilities as contours in Figure 10.5.1(*b*), it is easy to see that the orbitals

become progressively larger as one moves to higher energy levels (shown by the numbers 1, 2, and 3). This is similar to how each orbit in the Bohr model becomes progressively farther from the nucleus. In other words, the larger the orbital, the greater the relative energy of the electron.



Figure **10.5.2**: Top to bottom: An *s* orbital, the three possible *p* orbitals, the five possible *d* orbitals, the seven possible *f* orbitals. (haade, CC BY-SA 3.0, via Wikimedia Commons)

As one moves to higher energy levels, each energy level has a greater number of possible orbitals. In addition to a greater number of possible orbitals, the types of available orbitals become increasingly complex (see Figure 10.5.2). The increasing complexity of these orbitals is analogous to the increased vibrational frequency of a guitar string as the string is stretched (see Figure 10.5.3). The vibrational frequency increases as the number of nodes (or standing waves) increases. As described in Section 10.1, energy and frequency are directly proportional. This means that as the frequency increases, so does the energy. In the case of orbitals, as the complexity of an orbital increases, so does its energy. Therefore, electrons in the same energy level that reside in orbitals that are more complex have a greater energy than those that reside in orbitals that are less complex. In other words, the greater the complexity of the orbitals, the greater the relative energy of the electron.

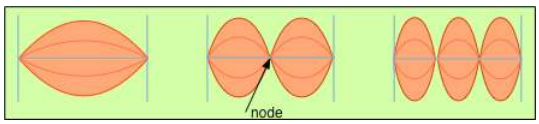


Figure **10.5.3**: Possible vibrational modes of a vibrating guitar string. The greater the number of nodes, the greater the vibrational frequency (via LibreTexts)

What does all of this mean? In the Bohr model, electrons had to jump from one orbit to another when they they absorbed or released energy on an instantaneous basis; they could not exist anywhere between the orbits. The modern model, called the **quantum mechanical model**, overcomes this issue.

In the quantum mechanical model, electrons transition from orbitals of lesser energy to ones with greater energy (either larger and/or more complex orbitals) when they absorb energy, while they transition from orbitals of greater energy to ones with lesser energy (either smaller and/or less complex orbitals) when they release energy. Since orbitals are defined based on probabilities of finding an electron in spaces that overlap each other in three-dimensional space rather than very specific, non-overlapping orbits in the Bohr model, the jumps between quantized energy levels and/or orbitals are much more easily explained.

✅ Example 10.5.1:

The energy of an electron in an atom depends on the ____ and ____ of the orbital in which it resides.

A. size
B. wave
C. shape
D. particle
E. nucleus

Solution

The energy of an electron in an atom depends on the **(A) size** and **(C) shape** of the orbital in which it resides.

✏️ Exercise 10.5.1

A. An electron in an orbital has the **most** energy if the **size** of the orbital is

1. larger
2. smaller

3. medium-sized
4. orbital size has nothing to do with the energy of the electron.

B. An electron in an orbital has the **most** energy if the **shape** of the orbital is

1. complex.
2. simple.
3. orbital shape has nothing to do with the energy of the electron.

Answer A

1. larger

Answer B

1. complex

Summary

- Quantum mechanics involves the study of material at the atomic level. This field deals with probabilities, since we cannot definitely locate a particle.
- Orbitals are mathematically derived regions of space with different probabilities of having an electron.

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10.6: Valence Electrons

Learning Objectives

- Determine the number of valence electrons for an element based on its place in the periodic table.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. Valence electrons are the electrons that reside in the outermost energy level of an atom and are, therefore, the most accessible for the formation of chemical bonds.

The number of valence electrons in one atom of each element is easily determined based on its position in the periodic table. For the main group elements (groups designated with a Roman numeral followed by the letter A), the number of valence electrons is equal to the Roman numeral. It would also be equal to the last digit of the group numbers that are written in their standard form. For example, silicon is in Group IVA (Group 14), so each atom would have four valence electrons. Chlorine is in Group VIIA (Group 17), so it would have seven valence electrons. Calcium would have two valence electrons, since it is in Group IIA (Group 2).

Helium is the only exception for the main group elements. The first energy level holds a maximum of two valence electrons. Since helium atoms only have two electrons and the outermost energy level *is* the first energy level, there can only be two valence electrons. When the outermost energy level is any other level beyond the first energy level, it may contain up to eight valence electrons. In other words, the maximum number of valence electrons that an atom may have is eight.

Transition elements (groups designated with a Roman numeral followed by the letter B) usually have two valence electrons, though we will not be working with valence electrons for the transition elements in this text.

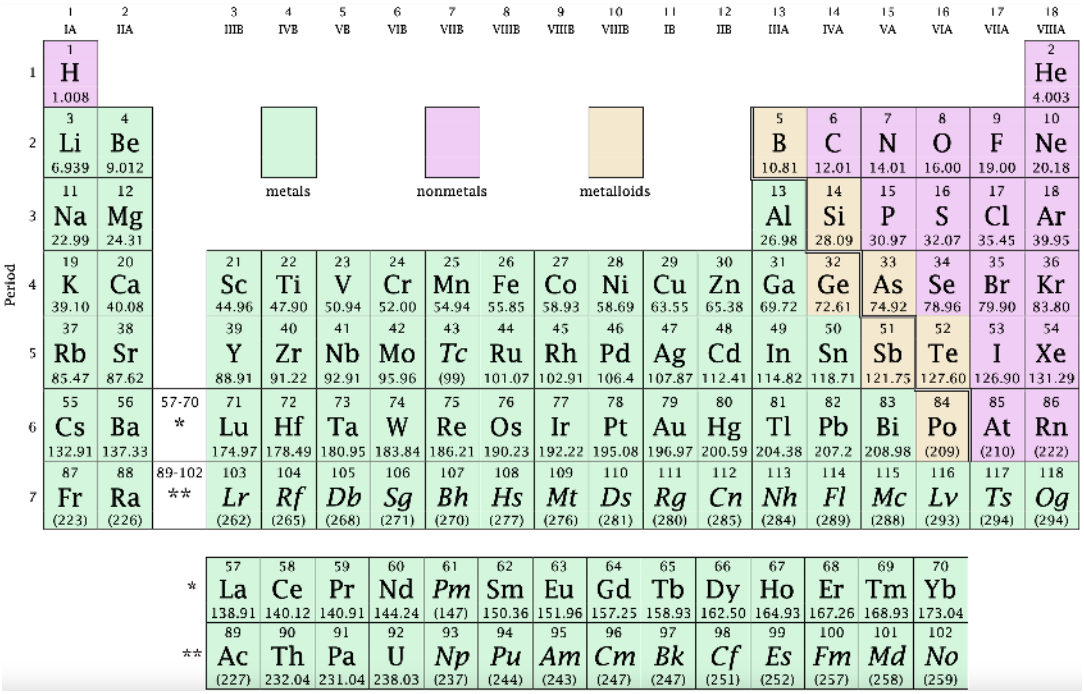


Figure 10.6.1: The modern version of the periodic table. An interactive periodic table may be found here.

✅ Example 10.6.1: Number of Valence Electrons

How many valence electrons are in one atom of each element?

A. sulfur
B. helium

C. potassium
D. aluminum

Solution

A. Sulfur (S) is located in Group VIA (Group 16), so it has 6 valence electrons.
B. Helium (He) is located in Group VIIIA (Group 18). However, one atom only has two electrons, so it could never have more than 2 valence electrons. As noted above, helium is the only exception for the main group elements.
C. Potassium (K) is located in Group IA (Group 1), so it has 1 valence electron.
D. Aluminum (Al) is located in Group IIIA (Group 13), so it has 3 valence electrons.

Exercise 10.6.1

Tell which group of elements (alkali metals, alkaline earth metals, halogens, noble gases) has atoms with the specified number of valence electrons?

A. 1 valence electron
B. 2 valence electron
C. 7 valence electrons
D. 8 valence electrons

Answer A

alkali metals

Answer B

alkaline earth metals

Answer C

halogens

Answer D

noble gases

Summary

- The number of valence electrons in an atom may be predicted based on its position in the periodic table.

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10.7: Periodic Trends

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.

Atomic Size

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 mentioned in Section 10.5, orbitals get larger with more energy. The energy level with the lowest energy, designated as $n = 1$, also has the smallest orbital and its electrons are therefore closest to the nucleus. As the value of the energy level (n) increases, the orbitals become larger. In addition, a higher energy level has more orbitals than a lower energy level and hence can hold more electrons. This is in agreement with predictions made by Bohr that the maximum number of electrons allowed in a particular energy level, n, would be $2n^2$. As we build atoms with more and more electrons, those electrons need to move to higher energy levels as lower ones are filled. For main group elements, the highest energy level containing its valence electrons corresponds to the row that the atom is in on the periodic table. Thus, magnesium's two valence electrons are in $n = 3$, and iodine's seven valence electrons are in $n = 5$.

We can use these concepts of energy levels and orbital size to describe two trends seen experimentally. The first is that as you go down a column of the periodic table, the atomic radii increase. This is because as we move down a column, the valence electrons occupy higher energy levels. The higher energy level will have larger orbitals, so the valence electrons are physically farther away from the nucleus. This trend can be summarized as follows:

$$\textbf{as} \downarrow \textbf{PT, atomic radius} \uparrow$$

where PT stands for periodic table.

The second trend is observed when going across a row on the periodic table, from left to right. Even though the valence electrons are going into the same energy level, the number of protons—and hence the positive charge of the nucleus—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. Again, we can summarize this trend as follows:

$$\textbf{as} \rightarrow \textbf{PT, atomic radius} \downarrow$$

Combining these two trends together, we see that the smallest atoms will be in the upper right corner of the periodic table and the largest atoms will be in the lower left corner. Figure **10.7.1** shows spheres representing the atoms of the main group elements from the periodic table to scale, showing the two trends for the atomic radius.

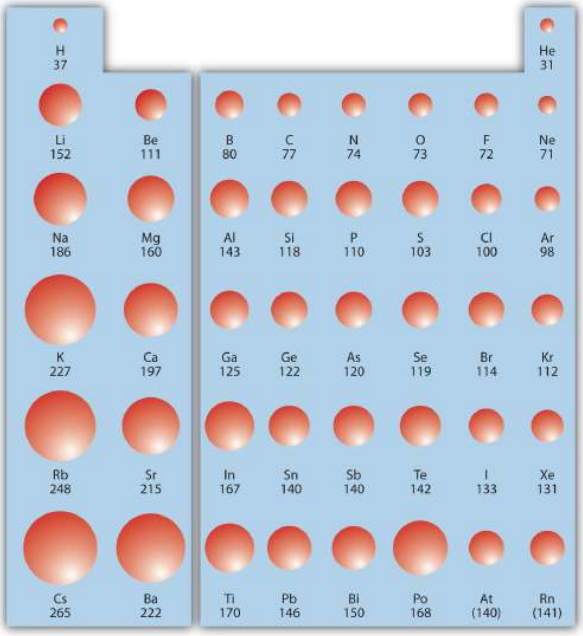


Figure **10.7.1**: Atomic radii trends on the periodic table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

✅ Example 10.7.1: Atomic Radii

Referring only to a periodic table and not to Figure **10.7.1**, which atom is larger in each pair?

A. Si or S
B. S or Te

Solution

A. 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.
B. S is above Te on the periodic table; Te is larger because as you go down the column, the atoms get larger.

✏️ Exercise 10.7.1: Atomic Radii

Referring only to a periodic table and not to Figure **10.7.1**, which atom is smaller, Ca or Br?

Answer

Br

Ionization Energy

Another property of atoms that exhibits periodic trends is ionization energy. **Ionization energy (IE)** is the amount of energy required to remove an electron from an atom in the gas phase:

$$\mathrm{A}(g) \rightarrow \mathrm{A}^{+}(g) + e^{-} \tag{10.7.1}$$

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). If the atom contains more electrons, the ionization energy for the removal of each additional electron can also be determined. For the purposes of this text, we will only consider the ionization energy for the removal of the first electron, known as the first ionization energy or IE_1.

The amount of energy that is required for IE_1 will depend on how strongly the electron is attracted to the nucleus. Electrons that are close to the nucleus will be held more tightly than electrons located farther away from the nucleus. Thus, as you go down the periodic table, it becomes easier to remove an electron from an atom (i.e., IE_1 decreases) because the valence electron is farther away from the nucleus.

$$\textbf{as} \downarrow \textbf{PT, IE} \downarrow$$

Electrons will also be held more tightly by a greater positive charge. This is reflected in the trend observed as you go across the periodic table. The greater positive charge of the nucleus draws the electrons closer in, and it takes more energy to remove an electron. As a result, IE_1 increases.

$$\textbf{as} \rightarrow \textbf{PT, IE} \uparrow$$

Combining these two trends together for the main group elements, we see that the greatest first ionization energy occurs with the smallest atoms in the upper right corner of the periodic table while the lowest first ionization energy is found in the largest atoms in the lower left corner. Figure **10.7.2** shows values of IE_1 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.

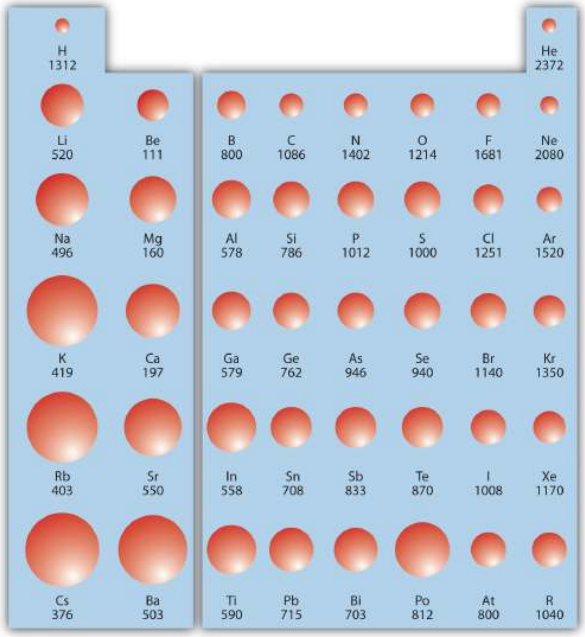


Figure **10.7.2**: Ionization energy on the periodic table. Values are in kJ/mol. Note that the larger atoms (based on the size of the spheres) tend to have the smallest IE values.

✅ Example 10.7.2: Ionization Energies

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

A. Ca or Sr
B. P or Cl

Solution

A. Because Sr is below Ca on the periodic table, it is easier to remove an electron from it; thus, Ca has the higher IE_1.
B. Because Cl is farther to the right, it has a more positive charge in its nucleus and it will be harder to remove an electron from it, so its IE_1 is larger than that of P.

Exercise 10.7.2: Ionization Energies

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

Answer

C

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. Thus, atoms with low first ionization energies will be more likely to lose their valence electrons and have chemical properties of metallic elements. They also have a high oxidation potential—they are easily oxidized (lose electrons) and are strong reducing agents. Metals also form basic oxides; the more basic the oxide, the higher the metallic character. In contrast, nonmetals have high first ionization energies making it very difficult for them to form cations. Instead, they will tend to gain electrons to fill up their energy level and form anions.

This metallic character is reflected in the regions of the periodic table described previously. Small atoms with greater first ionization energies are in the upper right corner of the periodic table. This area was designated as the nonmetals. Large atoms with lower first ionization energies are in the lower left corner of the periodic table in the region designated as the metals. The metallic character of the elements will therefore increase as we move down and to the left on the periodic table.

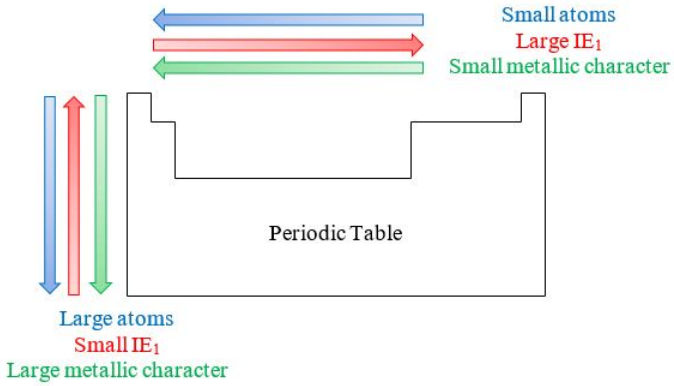


Figure **10.7.3**: Summary of periodic trends involving atomic size, first ionization energy, and metallic character. Arrows show how each property **increases** along rows and columns.

Summary

- Certain properties—notably atomic radius, ionization energies, and metallic character—can be qualitatively understood by the positions of the elements on the periodic table. The major trends are summarized in the Figure **10.7.3** above.
- There are two factors that help in the prediction of the trends in the periodic table: energy level of the valence electrons and the number of protons in the nucleus.

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10.9: Exercises

10.2: The Electromagnetic Spectrum

1. The laser on a Blu-ray player has a wavelength of 405 nm. In what region of the electromagnetic spectrum is this radiation?

Answer

Ultra-violet (UV)

2. Choose the correct word for the following statement. Blue light has a (longer or shorter) wavelength than red light.

Answer

shorter

3. Choose the correct word for the following statement. Yellow light has a (higher or lower) frequency than blue light.

Answer

lower

4. Choose the correct word for the following statement. Green light has a (larger or smaller) energy than red light.

Answer

larger

5. If "light A" has a longer wavelength than "light B", then "light A" has _____ "light B".
   a. a lower frequency than
   b. a higher frequency than
   c. the same frequency as

Answer

a lower frequency than

6. If "light C" has a shorter wavelength than "light D", then "light C" has _____ "light D".
   a. a larger energy than
   b. a smaller energy than
   c. the same energy as

Answer

a larger energy than

7. If "light E" has a higher frequency than "light F", then "light E" has _____ "light F".
   a. a longer wavelength than
   b. a shorter wavelength than
   c. the same wavelength as

Answer

a shorter wavelength than

8. If "light G" has a higher frequency than "light H", then "light G" has _____ "light H".
   a. a larger energy than
   b. a smaller energy than
   c. the same energy as

Answer

a larger energy than

9. If "light J" has larger energy than "light K", then "light J" has _____ "light K".
   a. a shorter wavelength than
   b. a longer wavelength than
   c. the same wavelength as

Answer

a shorter wavelength than

10. Which of the following statements is true?
    a. The frequency of green light is higher than the frequency of blue light and the wavelength of green light is longer than the wavelength of blue light.
    b. The frequency of green light is higher than the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.
    c. The frequency of green light is lower than the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.
    d. The frequency of green light is lower than the frequency of blue light and the wavelength of green light is longer than the wavelength of blue light.
    e. The frequency of green light is the same as the frequency of blue light and the wavelength of green light is shorter than the wavelength of blue light.

Answer

Statement d is true.

11. As the wavelength of electromagnetic radiation increases:
    a. its energy increases.
    b. its frequency increases.
    c. its speed increases.
    d. more than one of the above statements is true.
    e. none of the above statements is true.

Answer

e. none of the above statements is true.

10.3: Emission Spectra

12. The light produced by a red neon sign is due to the emission of light by excited neon atoms. Qualitatively describe the spectrum produced by passing light from a neon lamp through a prism.

Answer

The spectrum consists of colored lines, at least one of which (probably the brightest) is red.

13. The spectra of hydrogen and of magnesium are shown in Section 10.3. What causes the lines in these spectra? Why are the colors of the lines different? Suggest a reason for the observation that the spectrum of magnesium is more complicated than the spectrum of hydrogen.

Answer

The lines are caused by electrons dropping to lower energy levels. The colors are different because the electrons have different energies possible in the hydrogen and magnesium. One reason why the spectrum of magnesium could be more complicated is because it has more electrons than hydrogen.

10.4: The Bohr Model

14. According to the Bohr model, electrons in an atom can only have certain, allowable energies. As a result, we say that the energies of these electrons are _____.

Answer

quantized

15. What does it mean to say that the energy of the electrons in an atom is quantized?

Answer

Quantized energy means that the electrons can possess only certain discrete energy values; values between those quantized values are not permitted.

16. How are the Bohr model and the Rutherford model of the atom similar? How are they different?

Answer

Both have a small, heavy nucleus with the electrons moving around it. The Rutherford model has the electrons moving in orbits, much like a miniature solar system. However, the electrons will lose energy as they travel and move closer and closer to the nucleus. The amount of energy will change as they get closer to the nucleus so they will not emit discrete lines. The Bohr model also has the electrons moving in orbits. However, they are confined to specific energy values and only absorb or emit energy as they jump between levels, creating discrete line spectra.

10.6: Valence Electrons

17. How many valence electrons are in one atom of the following elements?

a. nitrogen
b. calcium
c. xenon
d. boron

Answer

a. 5
b. 2
c. 8
d. 3

18. How many valence electrons are in one atom of elements in the following groups?

a. group VIA or 16
b. group IA or 1
c. group VIIA or 17
d. group IVA or 14

Answer

a. 6
b. 1
c. 7
d. 4

10.7: Periodic Trends

19. Describe the atomic size trend for the rows in the Periodic Table.

Answer

Atoms get smaller as one moves from left to right across a row on the periodic table due to an increase in the number of protons.

20. Describe the atomic size trend for the columns on the Periodic Table.

Answer

Atoms get larger as one moves down a column on the periodic table because the valence electrons are in higher energy levels which are farther from the nucleus.

21. Which of the following would be smaller?

a. In or Ga
b. K or Cs
c. Te or Po

Answer

a. Ga
b. K
c. Te

22. Explain in your own words why Iodine is larger than Bromine.

Answer

Iodine is larger than bromine because it is farther down on the periodic table. Thus, its valence electrons are in a higher energy level that is farther from the nucleus.

23. Which of the following would have the largest atomic radii?

a. Si
b. C
c. Sn
d. Pb

Answer

Pb

24. Arrange the following in order of increasing atomic radii: Tl, B, Ga, Al, In.

Answer

B < Al < Ga < In < Tl

25. Arrange the following in order of increasing atomic radii: Ge, Sn, C.

Answer

C < Ge < Sn

26. Which of the following would be larger?

a. Rb or Sn
b. Ca or As

Answer

a. Rb
b. Ca

27. Place the following in order of increasing atomic radii: Mg, Cl, S, Na.

Answer

Cl < S < Mg < Na

28. Which of the following would have the largest atomic radii?

a. Sr
b. Sn
c. Rb
d. In

Answer

Rb

29. Which of the following would have the smallest atomic radii?

a. K
b. Kr
c. Ga
d. Ge

Answer

Kr

30. Place the following elements in order of increasing atomic radii: In, Ba, Sr, Sb, Xe.

Answer

Xe < Sb < In < Sr < Ba

31. Place the following elements in order of decreasing atomic radii: Ge, O, Sr, Cs, Ba.

Answer

Cs > Ba > Sr > Ge > O

32. Which of the following would have the largest ionization energy?

a. Na
b. Al
c. H
d. He

Answer

He

33. Which of the following would have the smallest ionization energy?

a. K
b. P
c. S
d. Ca

Answer

K

34. Place the following elements in order of increasing ionization energy: Na, O, Ne, K.

Answer

K < Na < O < Ne

35. Place the following elements in order of decreasing ionization energy: N, Si, P, Mg, He.

Answer

He > N > P > Si > Mg

36. Based on their positions in the periodic table, predict which has the smallest atomic radius: Mg, Sr, Si, Cl, I.

Answer

Cl

37. Based on their positions in the periodic table, predict which has the largest atomic radius: Li, Rb, N, F, I.

Answer

Rb

38. Based on their positions in the periodic table, predict which has the largest first ionization energy: Mg, Ba, B, O, Te.

Answer

O

39. Based on their positions in the periodic table, predict which has the smallest first ionization energy: Li, Cs, N, F, I.

Answer

Cs

40. Based on their positions in the periodic table, rank the following atoms in order of increasing metallic character: F, Li, N, Rb

Answer

F < N < Li < Rb

41. Based on their positions in the periodic table, rank the following atoms in order of increasing metallic character: Mg, O, S, Si

Answer

O < S < Si < Mg

42. Based on their positions in the periodic table, list the following atoms in order of decreasing metallic character: Mg, Ca, Rb, Cs.

Answer

Cs > Rb > Ca > Mg

43. Based on their positions in the periodic table, list the following atoms in order of decreasing metallic character: Sr, Ca, Si, Cl.

Answer

Sr > Ca > Si > Cl

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CHAPTER OVERVIEW

11: Chemical Bonding

How do atoms make compounds? Typically they join together in such a way that they lose their identities as elements and adopt a new identity as a compound. These connections are called *chemical bonds*. But how do atoms join together? Ultimately, it all comes down to electrons. Before we discuss how electrons interact, we need to introduce a tool to simply illustrate electrons in an atom.

11.1: Valence Electrons and the Periodic Table
11.2: Representing Valence Electrons with Dots
11.3: Ionic Bonds
11.4: Covalent Bonds
11.5: Multiple Covalent Bonds
11.6: Writing Lewis Structures for Molecular Compounds
11.7: Predicting the Shapes of Molecules
11.8: Templates Page for Lewis Diagrams
11.9: For Future Use
11.10: Exercises

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Thumbnail Chapter 11: A covalent bond in a hydrogen molecule. (Jacek FH via Wikimedia Commons)

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11.1: Valence Electrons and the Periodic Table

Learning Objectives

- Explain the relationship between the chemical behavior of families in the periodic table and their valence electrons.
- Identify elements that will have the most similar properties to a given element.

The chemical properties of elements are determined primarily by the number and distribution of valence electrons. Since the groups on the periodic table were originally organized based on their chemical behavior, it should be no surprise that the individual members of each group have similar electron configurations of their valence electrons.

Section 4.9 described how Dmitri Mendeleev arranged the original periodic table so that elements with the most similar properties were placed in the same group. For example, the alkali metals (Group IA) are quite soft and may easily be cut apart using a dull knife. They also reacted in a similar manner with water to produce hydrogen gas. When they combined with oxygen, they were found to do so in a 2:1 ratio of atoms. Once quantum mechanics was developed in the 1920s, it was found that the alkali metals all had one valence electron, thereby explaining their similar reactions and the 2:1 atom ratio with oxygen.

The halogens (Group VIIA) were all observed to be colorful, reactive elements that combined with oxygen in a 7:2 atom ratio and with the alkali metals in a 1:1 atom ratio. As a gas or vapor, the halogens all had a pungent odor. After the development of quantum mechanics, it was shown that the halogens all had seven valence electrons, supporting their original placement into the same group on Mendeleev's periodic table.



Figure **11.1.1**: Periodic table by Dmitri Mendeleev, 1871. (NikNaks, Public domain, via Wikimedia Commons)

The Modern Periodic Table and Chemical Bonds

As described in Section 10.6, the modern periodic table is arranged based on an atom's valence electrons. But what does this tell us about how they form chemical bonds with each other? Why do sodium atoms and chlorine atoms combine in a 1:1 ratio, while sodium and oxygen atoms combine in a 2:1 ratio? Much of this was already described in Chapter 5 in our study of chemical nomenclature. At that time, it wasn't really explained why sodium atoms always form ions with a 1+ charge, why chlorine atoms form ions with a 1– charge, or why oxygen atoms form ions with a 2– charge. By the time you complete this chapter, you will have a much better idea.

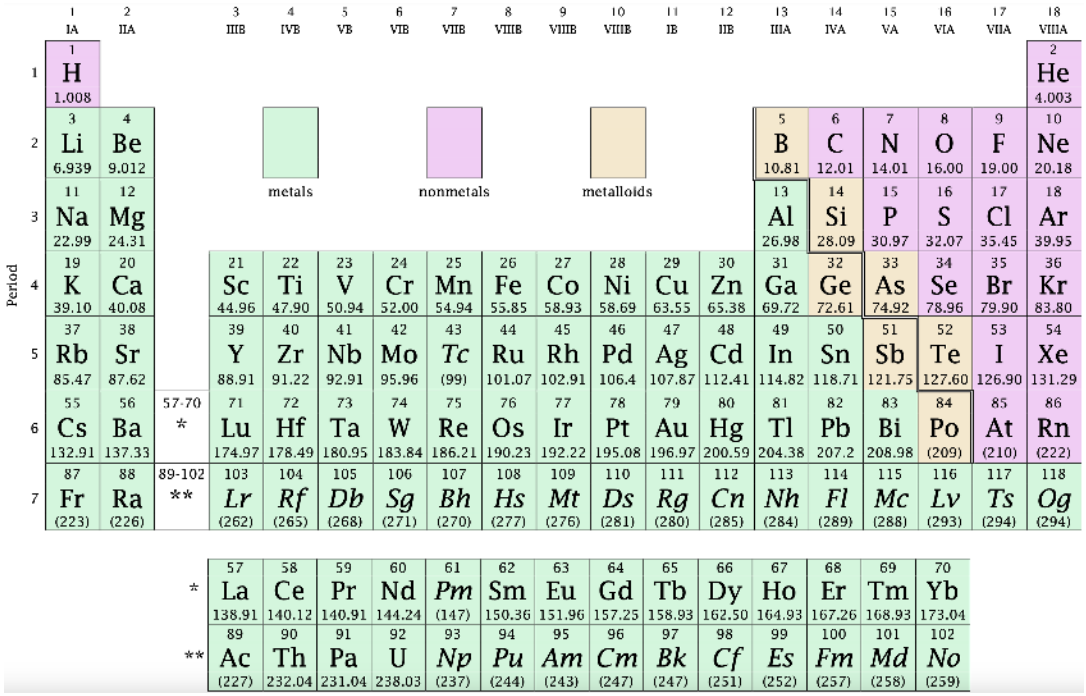


Figure **11.1.2**: The modern periodic table. An interactive periodic table may be found here.

Summary

- All of the elements in the same group on the periodic table have the same number of valence electrons and similar chemical properties.

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11.2: Representing Valence Electrons with Dots

Learning Objective

- Draw a Lewis electron dot diagram for an atom.

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 method for representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram, Lewis dot diagram, Lewis diagram, or Lewis structure) is a representation of the valence electrons of an atom that uses dots that surround the chemical symbol of the element. The number of dots equals the number of valence electrons in the atom. Dots are arranged to the right, left, above, and below the symbol, with no more than two dots on any one side. For example, the Lewis dot diagram for hydrogen is simply

$$\mathrm{H}\cdot$$

The side on which the dot is placed is not important, so the Lewis diagram for hydrogen may be drawn as:

$$\mathrm{H}\cdot \quad \text{or} \quad \dot{\mathrm{H}} \quad \text{or} \quad \cdot\mathrm{H} \quad \text{or} \quad \underset{\cdot}{\mathrm{H}}$$

As electrons are added around the chemical symbol, the general convention for dot placement is that one dot is placed on each side before doubling up, though helium is a notable exception. The Lewis diagram for helium, with its two valence electrons, is:

$$\mathrm{He}:$$

Again, the side on which the dots are placed is not important, so the Lewis diagram for helium may be drawn as:

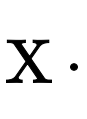
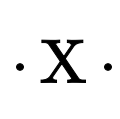
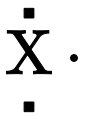
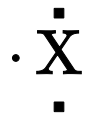
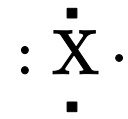
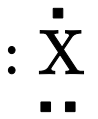
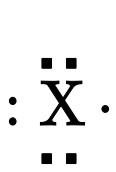
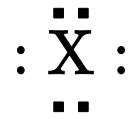
$$\mathrm{He}: \quad \text{or} \quad \ddot{\mathrm{H}}\mathrm{e} \quad \text{or} \quad :\mathrm{He} \quad \text{or} \quad \mathrm{He}\underset{\cdot}{\cdot}$$

Placing the two dots together on the same side represents two electrons that are paired up. It turns out that the first energy level of an atom holds a maximum of two electrons. Since helium atoms only have two electrons and the outermost energy level *is* the first energy level, the two electrons have no other option but to be paired up. When the outermost energy level is any other level beyond the first energy level, it may contain up to eight valence electrons. In other words, the maximum number of valence electrons that an atom may have is eight, which is why we represent the Lewis diagrams of the elements with four sides that may show up to two dots on each side.

Since the number of valence electrons for a main group element is equal to the Roman numeral, the generic Lewis diagrams for each group are:

Table 11.2.1: Generic Lewis Diagrams Based on Position in the Periodic

Table

Group	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Valence Electrons	1	2	3	4	5	6	7	8*
Lewis Diagram								

*Helium is an exception and has just two valence electrons (see above).

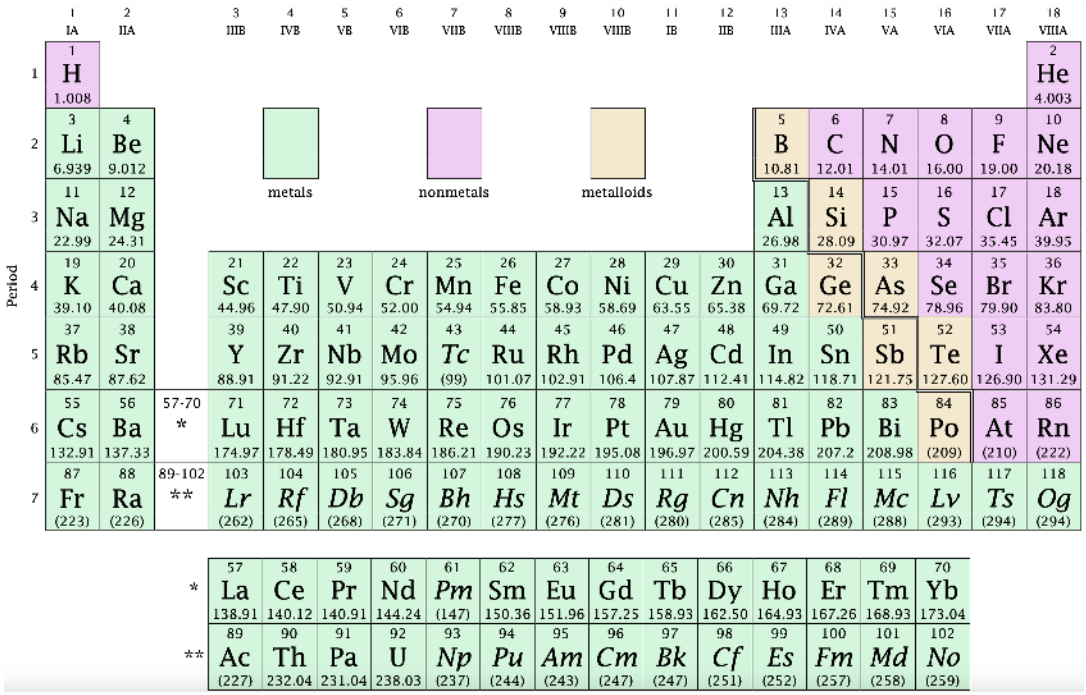


Figure **11.2.1**: Use this periodic table for the examples and exercises below. An interactive periodic table may be found here.

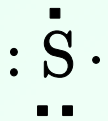
✅ Example 11.2.1: Lewis Dot Diagrams

Write the Lewis diagram for each element.

A. sulfur
B. helium
C. potassium
D. aluminum

Solution

A. Sulfur (S) is located in Group VIA, so it has 6 valence electrons.



B. Helium (He) atoms have two electrons and is the only exception for the main group elements in terms of valence electrons.



C. Potassium (K) is located in Group IA, so it has 1 valence electron.



D. Aluminum (Al) is located in Group IIIA (Group 13), so it has 3 valence electrons.

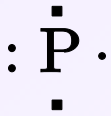


✏️ Exercise 11.2.1

Write the Lewis diagram for each element.

A. phosphorus
B. argon
C. silicon
D. magnesium

Answer A



Answer B



Answer C



Answer D



Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.

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11.3: Ionic Bonds

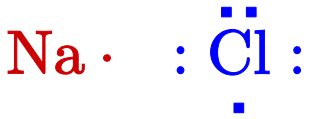
Learning Objectives

- State the octet rule.
- Define ionic bond.
- Draw Lewis structures for ionic compounds.

Back in Section 4.8, we discussed the formation of ions. It wasn't emphasized at that time, but it turns out that most monoatomic ions have eight valence electrons. Anions form when atoms gain enough electrons to wind up with a total of eight electrons in the valence shell. Anions carry a negative charge due to the gain of negatively charged electrons. Cations form when atoms lose all of the electrons from their original valence shell, revealing a new valence shell that contains eight valence electrons that lies just below the original valence shell. Cations carry a positive charge due to the loss of negatively charged electrons. For whatever reason, having eight valence electrons is a particularly energetically stable arrangement of electrons.

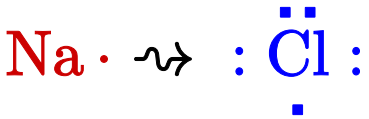
The **octet rule** explains the favorable trend of atoms having eight electrons in their valence shell, as is the case for neutral atoms of all of the noble gases, except for helium. The first energy level holds a maximum of two electrons, so atoms of helium may only have a maximum of two valence electrons. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

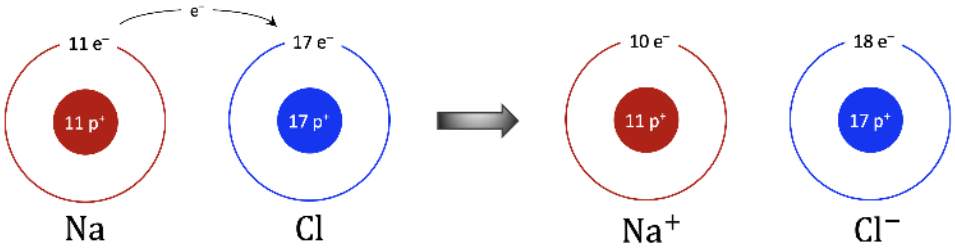
Let's consider a sodium atom in the presence of a chlorine atom. The two atoms have these Lewis diagrams:



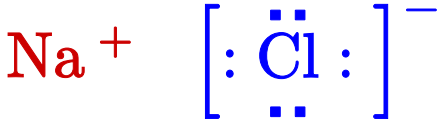
For the sodium atom to obtain an octet, it must lose a valence electron to give it the same number of valence electrons as a neutral neon atom. For the chlorine atom to obtain an octet, it must gain a valence electron to give it the same number of valence electrons as a neutral argon atom. Remember that the identity of an element is based on the number of protons in the nucleus, not on the number of electrons in the electron cloud.

An electron transfers from the sodium atom to the chlorine atom:





The sodium atom now has one less electron than it has protons, making the sodium ion have a positive charge, Na^+, while the chlorine atom has one more electron than it has protons, making the chloride ion have a negative charge, Cl^-. Since opposite charges attract, the Na^+ and Cl^- ions come together to make sodium chloride with the Lewis structure:



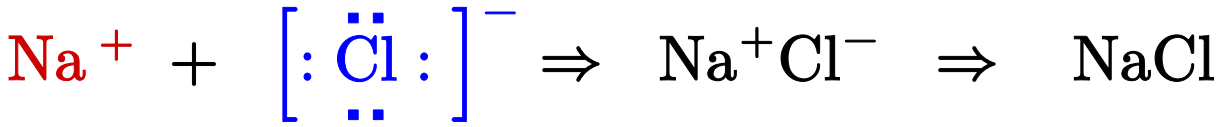
The attraction between oppositely charged ions is called an **ionic bond,** which is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by the transfer of electrons from one atom to another, typically from a metal to a nonmetal. In addition, since sodium and chlorine now have complete octets they are energetically stable.

Ionic Bond

A chemical bond that forms between oppositely charged ions and is caused by the transfer of electrons from one atom to another, typically from a metal to a nonmetal.

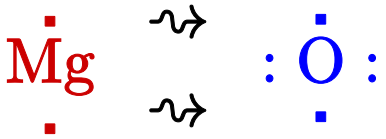
When writing Lewis structure for ionic compounds, the convention is to place square brackets, [ ], around an ion to separate dots from the charge on the ion, which is still written as a superscript. This is why the square brackets are placed around the chloride ion, Cl^- (which is surrounded by dots), while they are not placed around the sodium ion, Na^+ (which is not surrounded by dots).

From here, it is easy to see why the chemical formula of sodium chloride is NaCl. As per the convention for writing formulas for ionic compounds, charges are left off of the chemical formula:

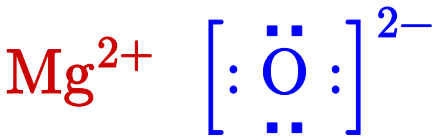


In electron transfer, the number of electrons lost must equal the number of electrons gained, as shown in the formation of NaCl.

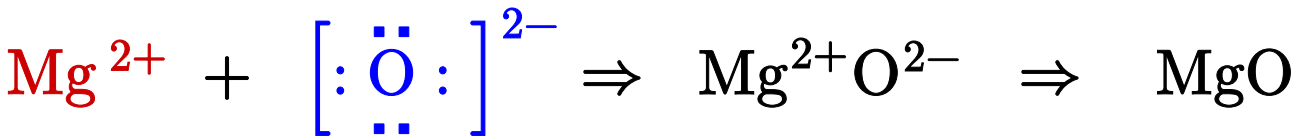
A similar process occurs between magnesium atoms and oxygen atoms. For the magnesium atom to obtain an octet, it must lose two valence electrons to give it the same number of valence electrons as a neutral neon atom. For the oxygen atom to obtain an octet, it must gain two valence electrons to give it the same number of valence electrons as a neutral argon atom. Once again, remember that the identity of an element is based on the number of protons in the nucleus, not on the number of electrons in the electron cloud.



This provides both magnesium and oxygen with complete octets. The magnesium atom now has two less electrons than it has protons, making the magnesium ion have a 2+ charge, while the oxygen atom has two more electrons than it has protons, making the oxide ion have a 2– charge. Since opposite charges attract, the Mg^{2+} and O^{2-} ions come together to make magnesium oxide with the Lewis structure:



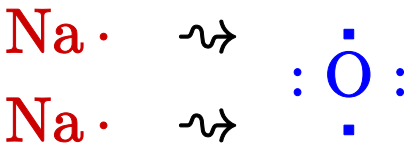
From here, it is easy to see why the formula of magnesium oxide is MgO:



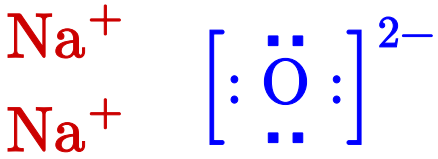
What happens when sodium atoms get together with oxygen atoms? An oxygen atom needs two electrons to complete its valence shell, while a sodium atom is only able to supply one electron:



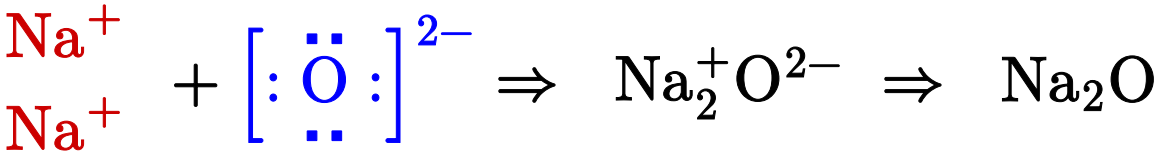
This does not provide an oxygen atom with a complete octet. Therefore, a second sodium atom is needed to provide a second electron for the oxygen atom:



The result is the formation of two sodium ions, Na^+, and one oxide ion, O^{2-}, that come together to make sodium oxide with the Lewis structure:



The resulting structure reflects the fact that the number of electrons lost by the sodium atoms must equal the number of electrons gained by the oxygen atom. It also shows that the compound is charge-balanced and why the ratio of sodium ions to oxide ions is 2:1 in sodium oxide:



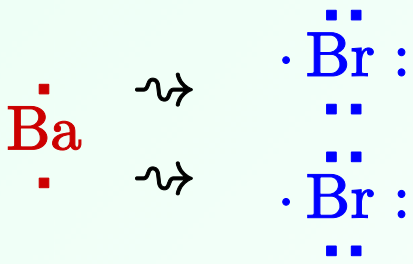
It is an expectation that the number of electrons lost by metal atoms must equal the number of electrons gained by nonmetal atoms, according to the law of conservation of mass. This will always be reflected in the Lewis structure of ionic compounds. It also explains why ionic compounds must be charge-balanced and why they have the ratios of cations to anions that they do.

✅ Example 11.3.1: Barium Bromide

Write the Lewis structure for barium bromide. Use arrows to illustrate the transfer of electrons to form barium bromide from barium atoms and bromine atoms.

Solution

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



Lewis structure: Ba^{2+} $\left[:\ddot{Br}:\right]^-$ $\left[:\ddot{Br}:\right]^-$

$Ba^{2+} + \left[:\ddot{Br}:\right]^- \left[:\ddot{Br}:\right]^- \Rightarrow Ba^{2+}Br_2^- \Rightarrow BaBr_2$

The oppositely charged ions attract each other to make $BaBr_2$.

Exercise 11.3.1

Write the Lewis structure for:

A. lithium fluoride
B. aluminum oxide

Answer A

$Li^+ \left[:\ddot{F}:\right]^-$

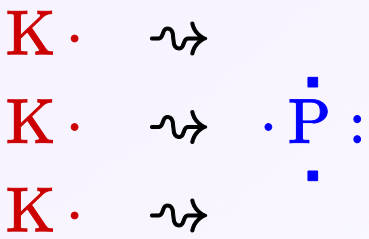
Answer B

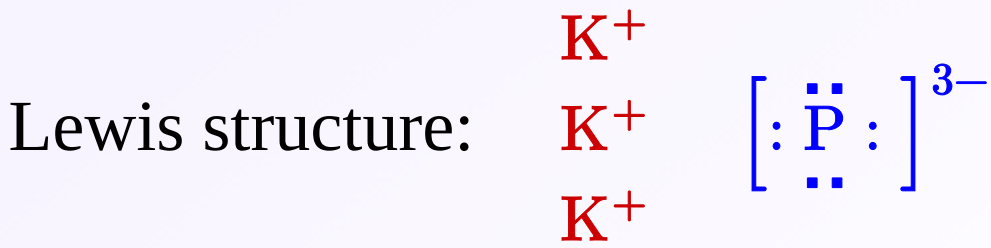
Al^{3+} Al^{3+} $\left[:\ddot{O}:\right]^{2-}$ $\left[:\ddot{O}:\right]^{2-}$ $\left[:\ddot{O}:\right]^{2-}$

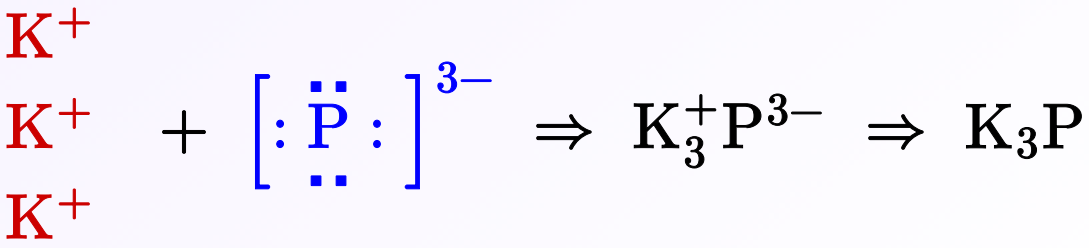
Exercise 11.3.2

Write the Lewis structure for potassium phosphide. Use arrows to illustrate the transfer of electrons to form potassium phosphide from potassium atoms and phosphorus atoms.

Answer







Summary

- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.

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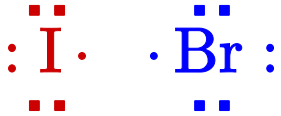
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11.4: Covalent Bonds

Learning Objectives

- Define covalent bond.
- Illustrate covalent bond formation using Lewis diagrams.

Consider a compound that forms between the elements iodine and bromine, iodine monobromide, IBr. Iodine and bromine both have 7 valence electrons and need one more electron to complete their valence shell:



If either iodine or bromine were to given up valence electrons to form a cation, they would have to give up all seven valence electrons to reveal the next energy level below as the new valence shell. Giving up seven valence electrons is energetically impossible, as it would imply the formation of an ion with a 7+ charge. Since opposite charges attract, the removal of each successive electron becomes increasingly difficult due to the increasing attraction between the nucleus and electrons in the electron cloud. It is much easier for each atom to gain one electron than it is to lose seven.

As the case illustrates above, it isn't easy for an atom to lose a whole bunch of electrons. Likewise, it isn't easy for an atom to gain a whole bunch of electrons. When working with an element like magnesium, we noticed that magnesium lost its two valence electrons to form magnesium ions, Mg^{2+}, rather than gaining six electrons to fill the valence shell to form ions with a 6– charge. Since like charges repel, the addition of each successive electron becomes increasingly difficult due to the increasing repulsions between electrons in the electron cloud. It is much easier for each atom to lose two electrons than it is to gain six.

If this is the case, how is a compound such as iodine monobromide, IBr, possible? A second type of chemical bond, called a **covalent bond**, involves the *sharing* of electrons to complete a valence shell rather than transferring electrons. A covalent bond typically forms between nonmetals and/or metalloids, since they have a valence shell that usually contains four or more valence electrons.

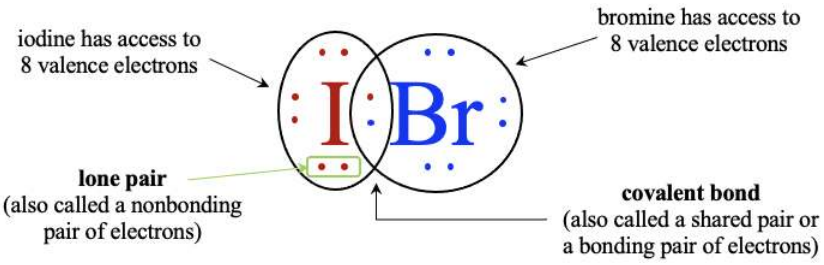
Covalent Bond

A chemical bond that forms between nonmetals and/or metalloids that is the result of sharing their valence electrons.

Let's illustrate how a covalent bond forms between iodine and bromine, with the understanding that each atom only needs one more electron to complete an octet in the valence shell. The iodine and bromine atoms each share one electron with the other.



Through sharing, the iodine atom now has access to eight valence electrons, as does the bromine atom. The portion where the circles overlap represent a shared pair of electrons, otherwise known as a covalent bond. Electrons that are not a part of a covalent bond are called lone pairs.



When two atoms share a single pair of electrons, the bond is called a single covalent bond, or simply, a **single bond**. When writing out a Lewis structure, a dash is used to represent a shared pair of electrons in place of two dots. Therefore, the Lewis structure for iodine monobromide, IBr, is shown in Figure **11.4.1**:

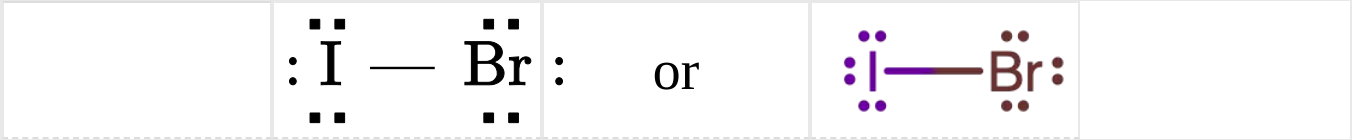
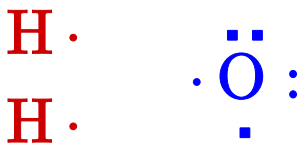


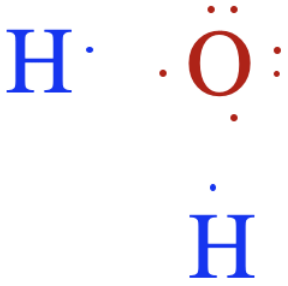
Figure **11.4.1**: The structure on the right was generated using HTML text, while the structure was generated using MarvinSketchJS. They are considered equivalent structures. Though elements generated using MarvinSketchJS have color, it is not necessary to use different colors when drawing Lewis structures.

Something might stand out when looking at the Lewis structure of IBr above – all of the electrons are paired up, whether they are bonding pairs (called covalent bonds) or nonbonding pairs (called lone pairs). In fact, this is usual standard for valence electrons in molecules – they are always paired, with a few exceptions that are not covered in this text. As we will learn in Section 11.7, the presence of lone pairs on the central atom(s) of a molecule contributes to its three-dimensional shape.

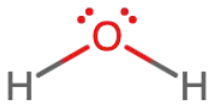
Let's look next at the structure of a water molecule. The chemical formula for water is H_2O. Each hydrogen atom has 1 valence electron. Because the first energy level holds a maximum of two valence electrons, it only needs one more electron to complete its valence shell. This sometimes called the **duet rule**. An oxygen atom has 6 valence electrons and needs two additional electrons to complete its valence shell:



Let's illustrate how covalent bonds form between hydrogen and oxygen:



Notice that the oxygen atom makes two single bonds, one to each of the two hydrogen atoms. This makes sense, since the oxygen atom needed two electrons to complete its valence shell. The Lewis structure for water is



There is nothing special about the orientation of the water molecule shown above. In fact, any of the orientations shown in Figure **11.4.2** may be drawn, as well as an infinite number of additional orientations in three-dimensional space, since molecules are in constant motion.

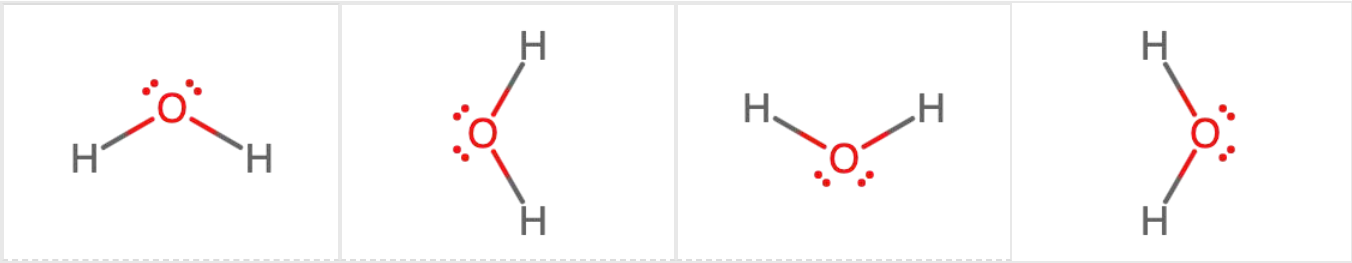


Figure **11.4.2**: Four different orientations of the Lewis structure for H_2O.

Figure **11.4.3** shows the bent three-dimensional shape of a water molecule that is due to the presence of lone pairs on the oxygen atom. The shapes of molecules are discussed in further detail in Section 11.7.

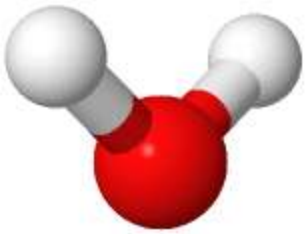


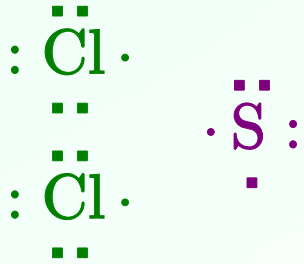
Figure **11.4.3**: A 3-dimensional model of a water molecule, H_2O.

✅ Example 11.4.1

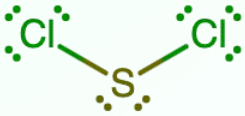
Write the Lewis structure for SCl_2.

Solution

Each chlorine atom has 7 valence electrons and needs one more electron to complete its valence shell. The sulfur atom has 6 valence electrons and needs two additional electrons to complete its valence shell:



The sulfur atom makes two single bonds, one to each of the two chlorine atoms. Once again, the orientation in three-dimensional space doesn't matter.



Exercise 11.4.1

Write the Lewis structure for HF.

Answer



Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis structures may be drawn to illustrate covalent bond formation.

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11.5: Multiple Covalent Bonds

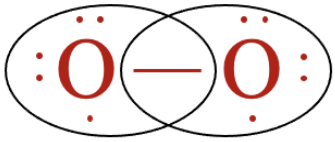
Learning Objectives

- Illustrate covalent bond formation with Lewis electron dot diagrams for simple molecules with multiple bonds.

Some molecules are unable to satisfy the octet rule by making single covalent bonds between atoms. Consider the a molecule of oxygen, O_2. Each oxygen atom has 6 valence electrons and needs two more electrons to complete its valence shell:



If the two oxygen atoms were to form a single bond with each other, they each have access to only seven valence electrons:



Recalling that a dash represents a covalent bond, or two electrons, each oxygen atom is still one electron short of satisfying the octet rule. Since there are only 12 valence electrons with which to work,

$$2 \text{ O atoms} \times 6 \text{ valence } e^- \text{ per atom} = 12 \text{ valence } e^- \text{ in the molecule}$$

each oxygen atom must achieve the octet rule by forming a double covalent bond, also called a **double bond**. A double bond is formed when two pairs of valence electrons are shared between the same two atoms. Since each dash represents a pair of electrons, we can see that the Lewis structure contains a total of 12 valence electrons and that each oxygen atom as access to eight valence electrons.

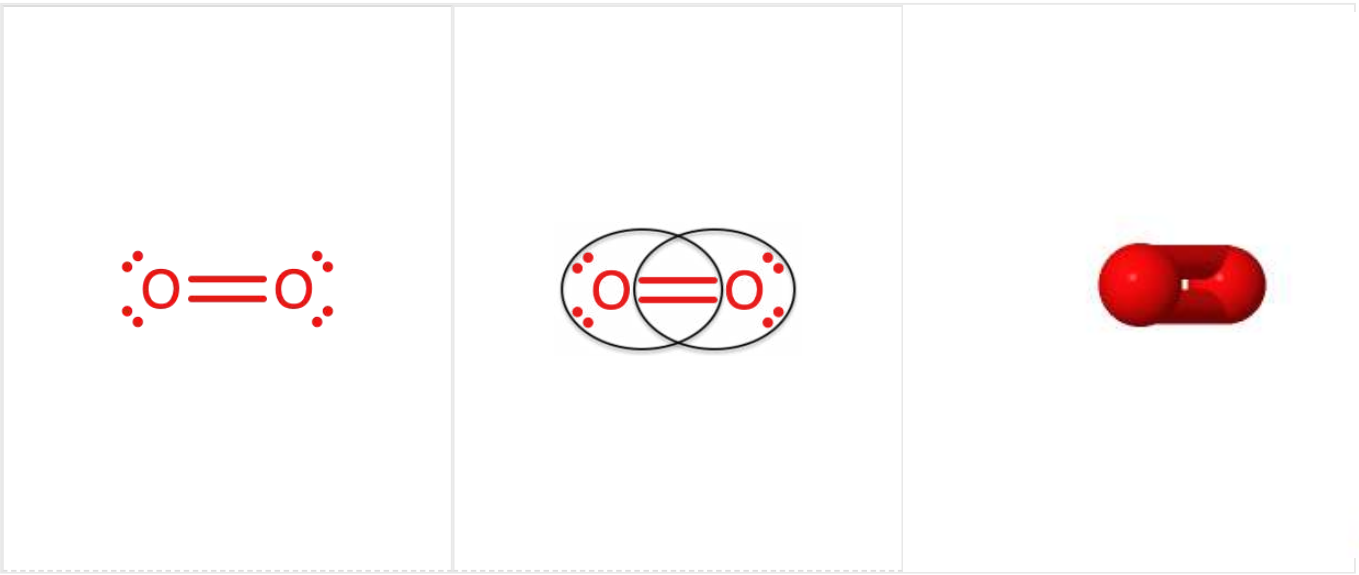


Figure **11.5.1** (from left to right): The Lewis structure of O_2. Each oxygen atom has access to 8 valence electrons by forming a double bond with the other oxygen atom. The three-dimensional rendering of an O_2 molecule.

Finally, let's consider a molecule of nitrogen, N_2. Each nitrogen atom has 5 valence electrons and needs three more electrons to complete its valence shell:



As one might suspect, since there are only 10 valence electrons with which to work,

2 N atoms × 5 valence e^- per atom = 10 valence e^- in the molecule

each nitrogen atom must achieve the octet rule by forming a triple covalent bond, or simply, a **triple bond**. A triple bond is formed when three pairs of valence electrons are shared between the same two atoms. Since each dash represents a pair of electrons, we can see that the Lewis structure contains a total of 10 valence electrons and that each nitrogen atom as access to eight valence electrons.



Figure **11.5.2** (from left to right): The Lewis structure of N_2. Each nitrogen atom has access to 8 valence electrons by forming a triple bond with the other nitrogen atom. The three-dimensional rendering of an N_2 molecule.

Summary

- Double bonds or triple bonds between atoms may be necessary to attain octets in some molecules.

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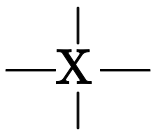
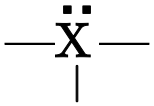
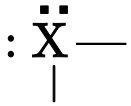
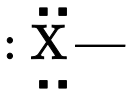
11.6: Writing Lewis Structures for Molecular Compounds

Learning Objectives

- Draw Lewis structures for molecular compounds.

By now, you have probably noticed a pattern in covalent bond formation. The normal number of covalent bonds for atoms in compounds is equal to the number of electrons required by an atom to complete its valence shell, as shown in Table 11.6.1. Recall that only nonmetals and/or metalloids typically form covalent bonds with each other.

Table 11.6.1: Normal Number of Covalent Bonds Formed

	H	Group IVA	Group VA	Group VIA	Group VIIA
Number of valence electrons:	1	4	5	6	7
Normal number of covalent bonds:	1	4	3	2	1
Lewis Structure (if all bonds are single bonds):					

When atoms form the normal number of covalent bonds with other elements, they may do so in any manner that sums to equal the normal number of bonds. For example, carbon (C) is located in Group IVA. This means it has four valence electrons and normally makes four covalent bonds. As a result, when carbon bonds with other elements, all of the following bond combinations are possible:

- 4 single bonds
- 2 single bonds and 1 double bond
- 1 single bond and 1 triple bond
- 2 double bonds

Further study in chemistry would reveal that the number of bonds an atom makes may differ when looking at polyatomic ions or in instances where the octet rule may be exceeded. Your instructor may choose to supplement this text with additional resources that examine polyatomic ions or expanded octets.

Here are some guidelines for constructing Lewis dot structures for more complex molecules than have been examined thus far:

The Basics of Constructing Lewis Dot Structures

1. Calculate the sum of the valence electrons in the molecule.

2. Construct a skeleton structure for the molecule, connecting atoms with a dash/line. Remember that each dash/line counts as two electrons.
   - The central atom(s) will usually be the element that makes the greatest number of bonds and is surrounded by atoms that make a lesser number of bonds.

- Bond angles around a central atom are rarely 180° (linear). The bond angle will only be linear when a central atom that is bonded to two other atoms has zero lone pairs (performed in the final check).
- Arrange the atoms in the skeleton structure so each atom makes the normal number of covalent bonds. If necessary, use double bonds or triple bonds to achieve the normal number of covalent bonds for each atom.
- Treat it like a puzzle that only fits together one way. In many cases, it is helpful to think of symmetry. A molecule with a few atoms is more likely to be build symmetrically around a central atom rather than in a straight line.

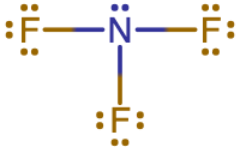
3. Add electrons as pairs to make an octet around each atom in the molecule, starting with the surrounding atoms and working inward to the central atom(s).
   - Remember that hydrogen follows the duet rule. Its valence shell holds up to two electrons.
   - Remember that an atom has access to all of the electrons in the bonds made by that atom.

4. Final check:
   - The total number of electrons shown in the molecule must equal the number of valence electrons calculated in the first step.
   - The octet rule is followed by every atom in the molecule (Note: hydrogen follows the duet rule).
   - Every atom in the molecule makes the expected number of bonds.
   - Change the bond angle to 180° (linear) for any central atom that has zero lone pairs and is also bonded to two other atoms.

✅ Example 11.6.1

Write the Lewis dot structure for NF_3.

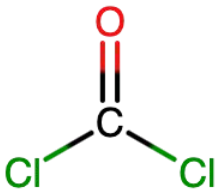
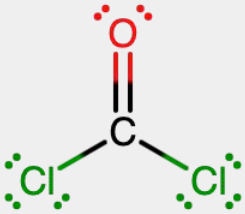
Solution

Steps for Writing Lewis Structures	
1. Calculate the sum of the valence electrons in the molecule.	1 N atom = 1 × 5 = 5 valence e^- 3 F atoms = 3 × 7 = 21 valence e^- sum of valence e^- = 26 valence e^-
2. Construct a skeleton structure for the molecule.	 N is the central atom since it makes the most bonds. An N atom makes 3 bonds, while an F atom makes 1 bond.
3. Add electrons as pairs to make an octet around each atom in the molecule.	
4. Final check.	There are 26 electrons shown in the molecule. This matches the 26 valence electrons calculated in the first step. The N atom and all 3 F atoms are following the octet rule. N is making 3 bonds, as expected. Each F atom is making 1 bond, as expected.

✅ Example 11.6.2

Write the Lewis dot structure for $COCl_2$.

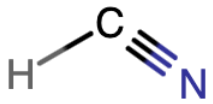
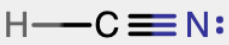
Solution

Steps for Writing Lewis Structures	
1. Calculate the sum of the valence electrons in the molecule.	1 C atom = 1 × 4 = 4 valence e^- 1 O atom = 1 × 6 = 6 valence e^- 2 Cl atoms = 2 × 7 = 14 valence e^- sum of valence e^- = 24 valence e^-
2. Construct a skeleton structure for the molecule.	 C is the central atom since it makes the most bonds (4 bonds). An O atom makes 2 bonds, while a Cl atom makes 1 bond.
3. Add electrons as pairs to make an octet around each atom in the molecule.	
4. Final check.	There are 24 electrons shown in the molecule. This matches the 24 valence electrons calculated in the first step. The C atom, O, and both Cl atoms are following the octet rule. C is making 4 bonds and O is making 2 bonds, as expected. Each Cl atom is making 1 bond, as expected.

✅ Example 11.6.3

Write the Lewis dot structure for HCN.

Solution

Steps for Writing Lewis Structures	
1. Calculate the sum of the valence electrons in the molecule.	1 H atom = 1 × 1 = 1 valence e^- 1 C atom = 1 × 4 = 4 valence e^- 1 N atom = 1 × 5 = 5 valence e^- sum of valence e^- = 10 valence e^-
2. Construct a skeleton structure for the molecule.	 C is the central atom since it makes the most bonds (4 bonds). An N atom makes 3 bonds, while an H atom makes 1 bond.
3. Add electrons as pairs to make an octet around each atom in the molecule.	
4. Final check.	There are 10 electrons shown in the molecule. This matches the 10 valence electrons calculated in the first step. H is following the duet rule. Both C and N are following the octet rule. H is making 1 bond, C is making 4 bonds, and N is making 3 bonds. These are all the expected number of bonds. Also, the molecule is linear, since the central C atom is only bonded to two other atoms and there are no lone pairs on the C.

Exercise 11.6.1

Write Lewis dot structures for each.

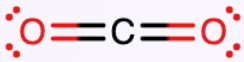
A. SI_2
B. CO_2
C. CH_2O

D. $C_2H_3F_3$

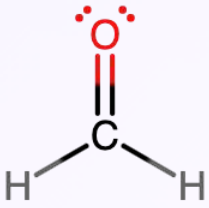
Answer A



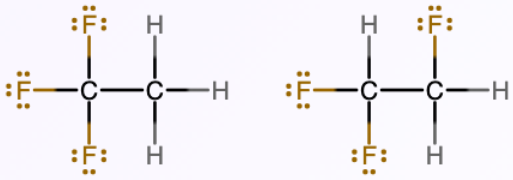
Answer B



Answer C



Answer D



Either of these two Lewis structures are acceptable. They are **isomers** of each other. Isomers are substances that have the same chemical formula, but a different molecular structure.

Notice that the structure on the left has 3 F atoms attached to one C atom and 3 H atoms attached to the other C atom. The structure on the right has 2 F atoms and 1 H atom attached to one carbon atom and 1 F atom and 2 H atoms attached to the other C atom.

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11.7: Predicting the Shapes of Molecules

Learning Objective

- Determine the shape of simple molecules.

All molecules have a three-dimensional shape. There is an abundance of experimental evidence to that effect – from their physical properties to their chemical reactivity. Small molecules – molecules with a single central atom – have shapes that are easily predicted. The model used to predict molecular shapes is called the **valence shell electron pair repulsion (VSEPR) model**. The VSEPR model is based on the fact that electron pairs, being composed of negatively charged particles, repel each other to get as far away from one another as possible in three-dimensional space.

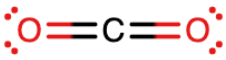
The VSEPR model makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**:

1. any type of bond – single, double, or triple.
2. lone electron pairs.

This text will not belabor how molecular geometries are determined. However, the VSPER model comes in handy when explaining why a water molecule is bent, while a carbon dioxide molecule is linear, even though they both have three atoms. When applying the VSEPR model to simple molecules, it all boils down to the number of electron groups around the central atom. Remember that a double or triple bond counts as only *one* electron group.

Linear molecules are simple. Any molecule with only two atoms is linear. A three-atom molecule whose central atom contains no lone pairs will orient the surrounding atoms as far apart from each other as possible in three-dimensional space – 180°, or linear, is as far apart apart as they can go. Carbon dioxide, CO_2, is such an example. The central carbon atom satisfies the octet rule by making two double bonds; it has no lone pairs. Therefore, the oxygen atoms lie on opposite sides of the molecule, or at 180° to each other around the carbon atom.



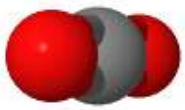


Figure **11.7.1** (left to right): The Lewis structure of CO_2. A three-dimensional rendering of CO_2.

Great! So, if water molecules have three atoms, just like carbon dioxide, then why are water molecules bent, while carbon dioxide molecules

are linear? To understand this, it may be helpful to go back to the basics of the VSEPR model. Once again, the VSEPR model is based on the fact that electron pairs repel each other to get as far away from one another as possible in three-dimensional space.

Water molecules are surrounded by four electron groups, just like molecules of methane, CH_4:



In the case of water, the four electron groups that surround the central oxygen atom includes two single bonds and two lone pairs of electrons. In the case of methane, the four electron groups that surround the central carbon atom includes four single bonds and zero lone pairs of electrons. In three-dimensional space, the farthest apart one can place the four surrounding hydrogen atoms is at 109.5° angles to the central carbon atom. If one drew imaginary lines from hydrogen atom to hydrogen atom in a methane molecule, the shape of a tetrahedron is obtained. Consequently, the geometry of four electron groups around a central atom is called a **tetrahedral** geometry.

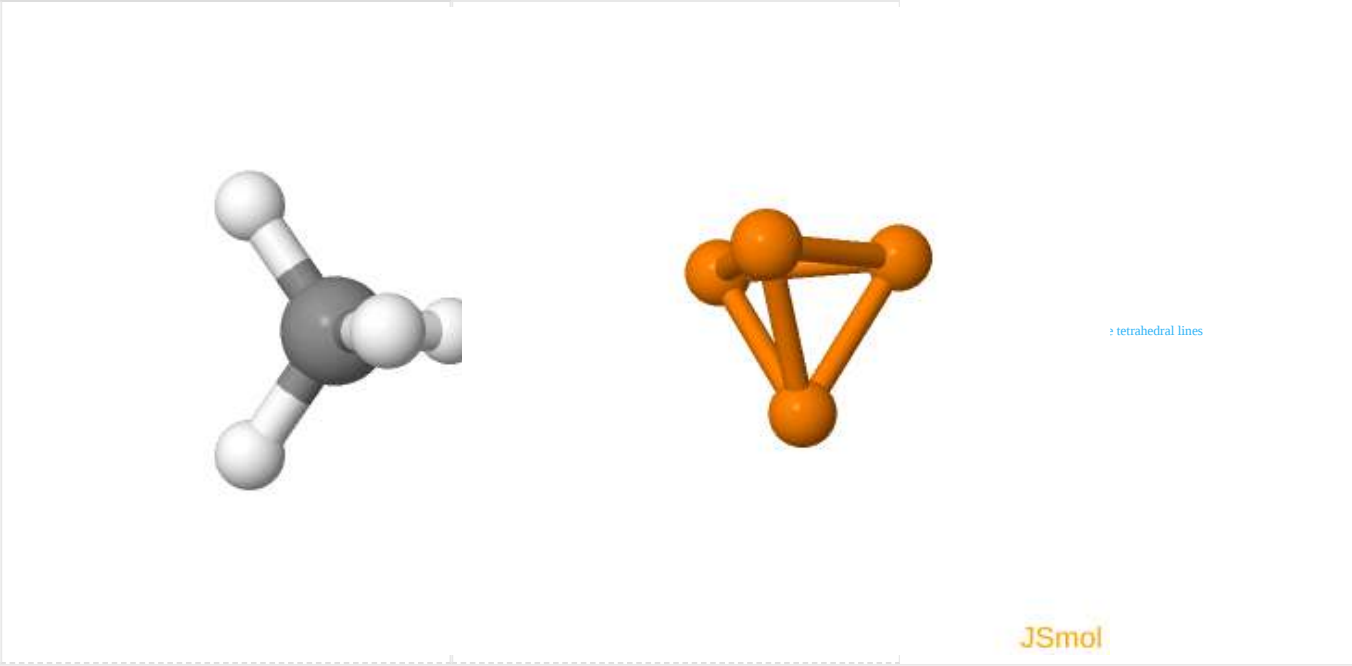
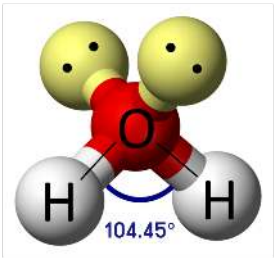


Figure **11.7.2** (left to right): A three-dimensional rendering of CH_4, a tetrahedron, and the three-dimensional structure of CH_4 drawn inside a tetrahedron.

The oxygen atom in a water molecule is still surrounded by four electron groups. It's just that two of those groups are lone pairs and those lone

pairs still have repulsions and take up space in three-dimensions. However, when we look at the molecule, we "see" the atoms and not the electrons. As a result, what we see is a water molecule that has a bent shape, rather than linear.



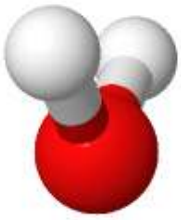


Figure **11.7.3** (left to right): The Lewis structure of H_2O, including lone pairs, superimposed over its three-dimensional rendering. A three-dimensional rendering of H_2O.

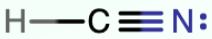
For the large part, this text will provide three-dimensional renderings of molecules that have four or more atoms. It would be helpful to be able to predict whether three atom molecules are linear or bent. Because of the important role that water plays in the chemistry of so many things around us, you should definitely be aware that water is a bent molecule and not linear.

✅ Example 11.7.1

Draw the Lewis structure of HCN that reflects its correct three-dimensional geometry (bent or linear).

Solution

HCN is a linear molecule, since it has only two electron groups and there are no lone pairs on the central carbon atom.



✏️ Exercise 11.7.1

Draw the Lewis structure of SCl_2 that reflects its correct three-dimensional geometry (bent or linear).

Answer

SCl_2 is a bent molecule, since the central sulfur atom has two lone pairs in addition to the two covalent bonds it makes to the two chlorine atoms.



Summary

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

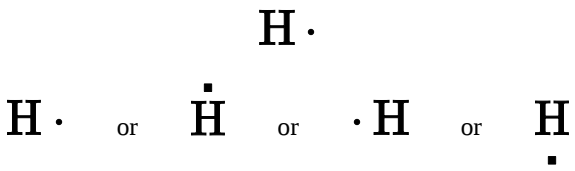
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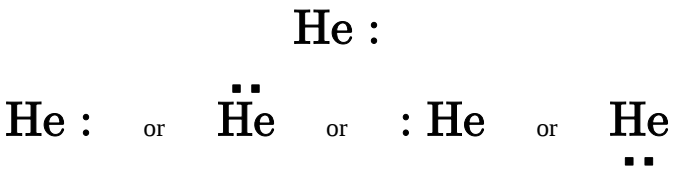
11.8: Templates Page for Lewis Diagrams

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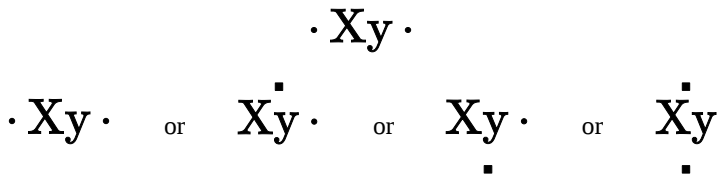
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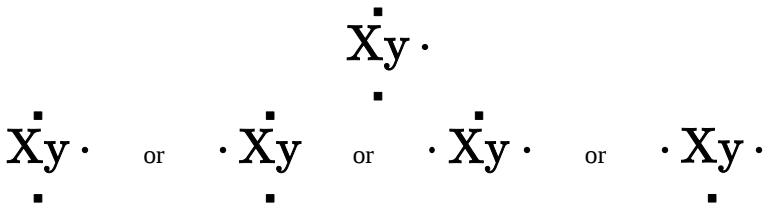
Two paired dots:



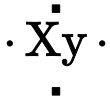
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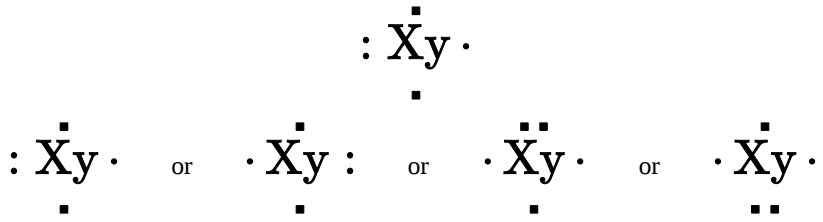
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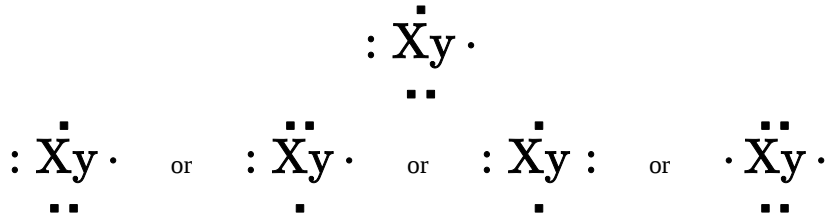
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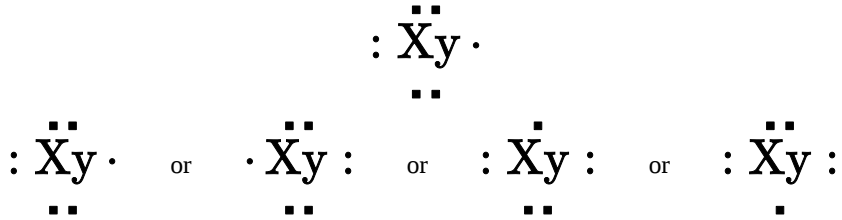
Five dots:



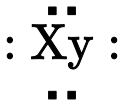
Six dots:



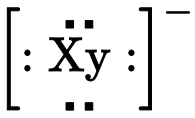
Seven dots:



Eight dots:



Monoatomic anion:

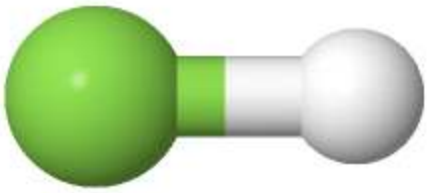


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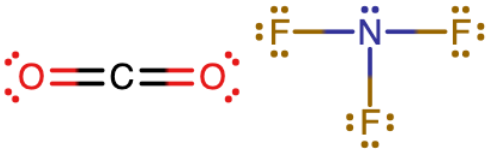
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11.10: Exercises

11.2: Representing Valence Electrons with Dots

1. Is it necessary for the first dot around an atomic symbol to go on a particular side of the atomic symbol?

Answer

No. The first dot can go on any side of the atomic symbol.

2. What column of the periodic table has Lewis electron dot diagrams with two electrons?

Answer

IIA or 2

3. What column of the periodic table has Lewis electron dot diagrams that have six electrons in them?

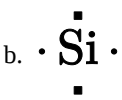
Answer

VIA or 16

4. Draw the Lewis electron dot diagram for each element.
   a. strontium
   b. silicon

Answer

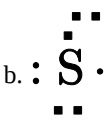
a. 

b. 

5. Draw the Lewis electron dot diagram for each element.
   a. krypton
   b. sulfur

Answer

a. 

b. 

6. Draw the Lewis electron dot diagram for each element.
   a. potassium
   b. phosphorus

Answer

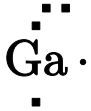
a. 

b. 

7. Draw the Lewis electron dot diagram for each element.
   a. bromine
   b. gallium

Answer

a. 

b. 

11.3: Ionic Bonds

8. How many electrons does a Ba atom have to lose to have a complete octet in its valence shell?

Answer

2

9. How many electrons does a Pb atom have to lose to have a complete octet in its valence shell?

Answer

4

10. How many electrons does an Se atom have to gain to have a complete octet in its valence shell?

Answer

2

11. How many electrons does an N atom have to gain to have a complete octet in its valence shell?

Answer

3

12. Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: P, I, Mg, Cl, In, Cs, O, Pb, Co?

Answer

P, I, Cl, and O would form anions because they are nonmetals. Mg, In, Cs, Pb, and Co would form cations because they are metals.

13. Which of the following atoms would be expected to form negative ions in binary ionic compounds and which would be expected to form positive ions: Br, Ca, Na, N, F, Al, Sn, S, Cd?

Answer

Br, N, F, and S would form anions because they are nonmetals. Ca, Na, Al, Sn, and Cd would form cations because they are metals.

14. Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:

a. P
b. Mg
c. Al
d. O
e. Cl
f. Cs

Answer

a. P^{3-}
b. Mg^{2+}
c. Al^{3+}
d. O^{2-}
e. Cl^{-}
f. Cs^{+}

15. Predict the charge on the monatomic ions formed from the following atoms in binary ionic compounds:

a. I
b. Sr
c. K
d. N
e. S
f. Li

Answer

a. I^{-}
b. Sr^{2+}
c. K^{+}
d. N^{3-}
e. S^{2-}
f. Li^{+}

16. Name two ions with the same Lewis electron dot diagram as the Cl^{-} ion.

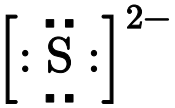
Answer

S^{2-} and P^{3-}

17. Draw the Lewis electron dot diagram for each ion.

a. Mg^{2+}
b. S^{2-}

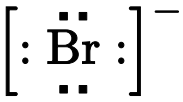
Answer

a. Mg^{2+} (no electrons)
b. 

18. Draw the Lewis electron dot diagram for each ion.

a. H^{+}
b. Br^{-}

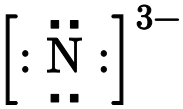
Answer

a. H^{+} (no electrons)
b. 

19. Draw the Lewis electron dot diagram for each ion.

a. Al^{3+}
b. N^{3-}

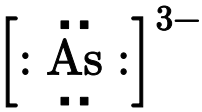
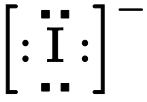
Answer

a. Al^{3+} (no electrons)
b. 

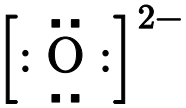
20. Write the Lewis symbols for each of the following ions:

a. As^{3-}
b. I^{-}
c. Be^{2+}
d. O^{2-}
e. Ga^{3+}
f. Li^{+}
g. N^{3-}

Answer

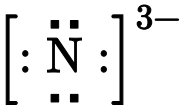
a. 
b. 

c. Be^{2+} (no electrons)

d. 

e. Ga^{3+} (no electrons)

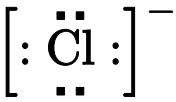
f. Li^{+} (no electrons)

g. 

21. Many monatomic ions are found in seawater, including the ions formed from the following list of elements. Write the Lewis symbols for the monatomic ions formed from the following elements:

a. Cl
b. Na
c. Mg
d. Ca
e. K
f. Br
g. Sr
h. F

Answer

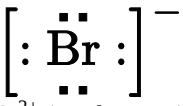
a. 

b. Na^{+} (no electrons)

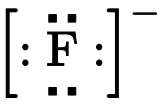
c. Mg^{2+} (no electrons)

d. Ca^{2+} (no electrons)

e. K^{+} (no electrons)

f. 

g. Sr^{2+} (no electrons)

h. 

11.4: Covalent Bonds

22. How many electrons will be in the valence shell of H atoms when it makes a covalent bond?

Answer

2

23. How many electrons will be in the valence shell of non-H atoms when they make covalent bonds?

Answer

8

24. What is the Lewis electron dot diagram of I_2?

Answer



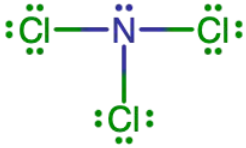
25. What is the Lewis electron dot diagram of H_2S?

Answer



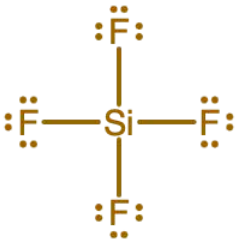
26. What is the Lewis electron dot diagram of NCl_3?

Answer



27. What is the Lewis electron dot diagram of SiF_4?

Answer

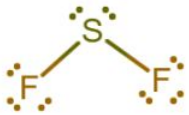


11.6: Writing Lewis Structures for Molecular Compounds

28. Draw the Lewis structure for each substance.

a. SF_2
b. C_2Cl_2 (assume two central atoms)

Answer

a. 

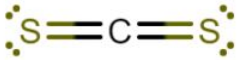
b. 

29. Draw the Lewis structure for each substance.

a. PI_3
b. CS_2

Answer

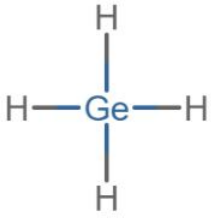
a. 

b. 

30. Draw the Lewis structure for each substance.

a. GeH_4
b. ClF

Answer

a. 

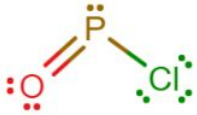
b. 

31. Draw the Lewis structure for each substance.

a. AsF_3
b. POCl

Answer

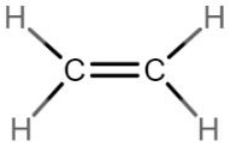
a. 

b. 

32. Draw the Lewis structure for each substance. Double or triple bonds may be needed.
    a. SiO_2
    b. C_2H_4 (assume two central atoms)

Answer

a. 

b. 

33. Write the Lewis structure for the diatomic molecule P_2, an unstable form of phosphorus found in high-temperature phosphorus vapor.

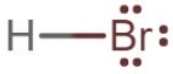
Answer

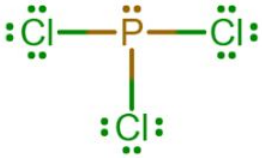


34. Write Lewis structures for the following:
    a. H_2
    b. HBr
    c. PCl_3
    d. HNNH
    e. H_2CNH

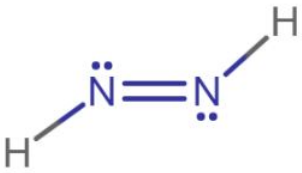
Answer

a. 

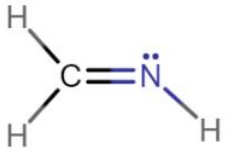
b. 



c.



d.



e.

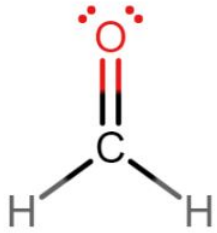
35. Write Lewis structures for the following:

a. O_2
b. H_2CO
c. ClNO
d. $SiCl_4$
e. ClCN

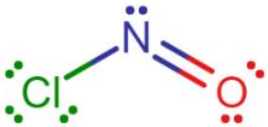
Answer



a.

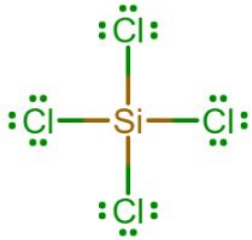


b.

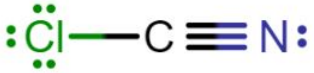


c.

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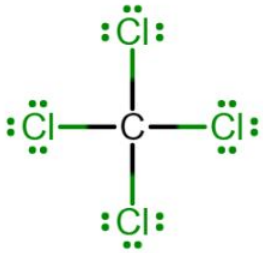
d.



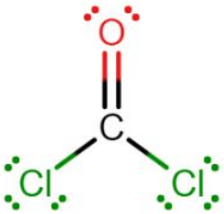
e.

36. Carbon tetrachloride was formerly used in fire extinguishers for electrical fires. It is no longer used for this purpose because of the formation of the toxic gas phosgene, Cl_2CO. Write the Lewis structures for carbon tetrachloride and phosgene.

Answer

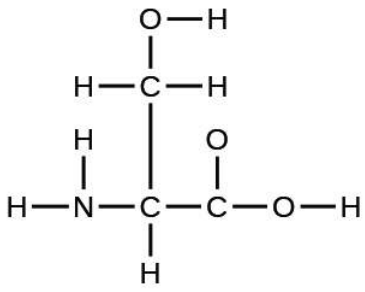


CCl_4:

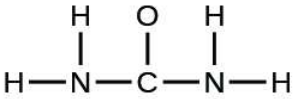


Cl_2CO:

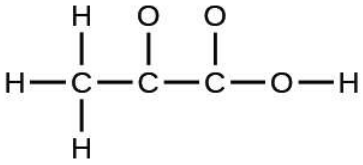
37. The arrangement of atoms in several biologically important molecules is given here. Complete the Lewis structures of these molecules by adding multiple bonds and lone pairs. Do not add any more atoms.



a. the amino acid serine:

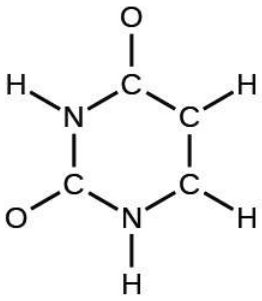


b. urea:

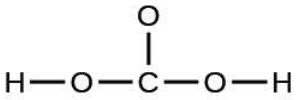


c. pyruvic acid:

d. uracil:

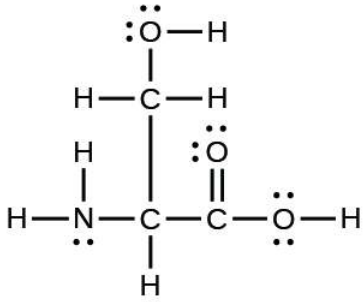


e. carbonic acid:

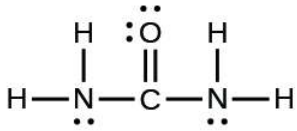


Answer

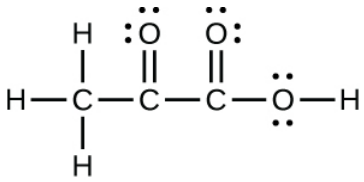
a.



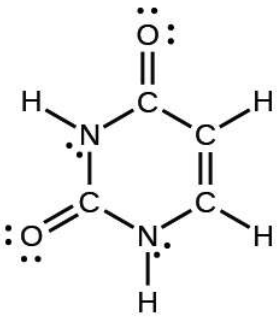
b.



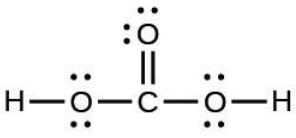
c.



d.



e.



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11.7: Predicting the Shapes of Molecules

38. What is the basic premise behind VSEPR?

Answer

Electrons groups (bonds and lone pair electrons) want to be as far from each other as possible due to the negative charges in the groups repelling each other.

39. What is the difference between the electron group geometry and the molecular geometry?

Answer

Electron group geometry tells where all electron groups are located. Molecular geometry tells where the atoms are located.

40. Identify whether the following molecules are linear or bent.

a. H_2S
b. HCN
c. CS_2

Answer

a. bent
b. linear
c. linear

41. Identify whether the following molecules are linear or bent around the central atoms.

a. C_2Cl_2 (both C atoms are central atoms and are bonded to each other)
b. N_2H_2 (both N atoms are central atoms and are bonded to each other)

Answer

a. linear
b. bent

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CHAPTER OVERVIEW

12: Organic Chemistry

Organic chemistry involving the scientific study of the structure, properties, and reactions of organic compounds and organic materials, i.e., matter in its various forms that contain carbon atoms. Study of structure includes many physical and chemical methods to determine the chemical composition and the chemical constitution of organic compounds and materials. Study of properties includes both physical properties and chemical properties, and uses similar methods as well as methods to evaluate chemical reactivity, with the aim to understand the behavior of the organic matter.

12.1: Organic Chemistry
12.2: Hydrocarbons
12.3: Alkanes
12.4: Branched Alkanes
12.5: Alkenes and Alkynes
12.6: Oxygen-Containing Organic Compounds
12.7: Alcohols, Aldehydes, Carboxylic Acids, and Ketones
12.8: Esters
12.9: Ethers
12.10: For Future Use
12.11: Exercises

Thumbnail Chapter 12: Jsmol rendering of propane (ChemLancer).

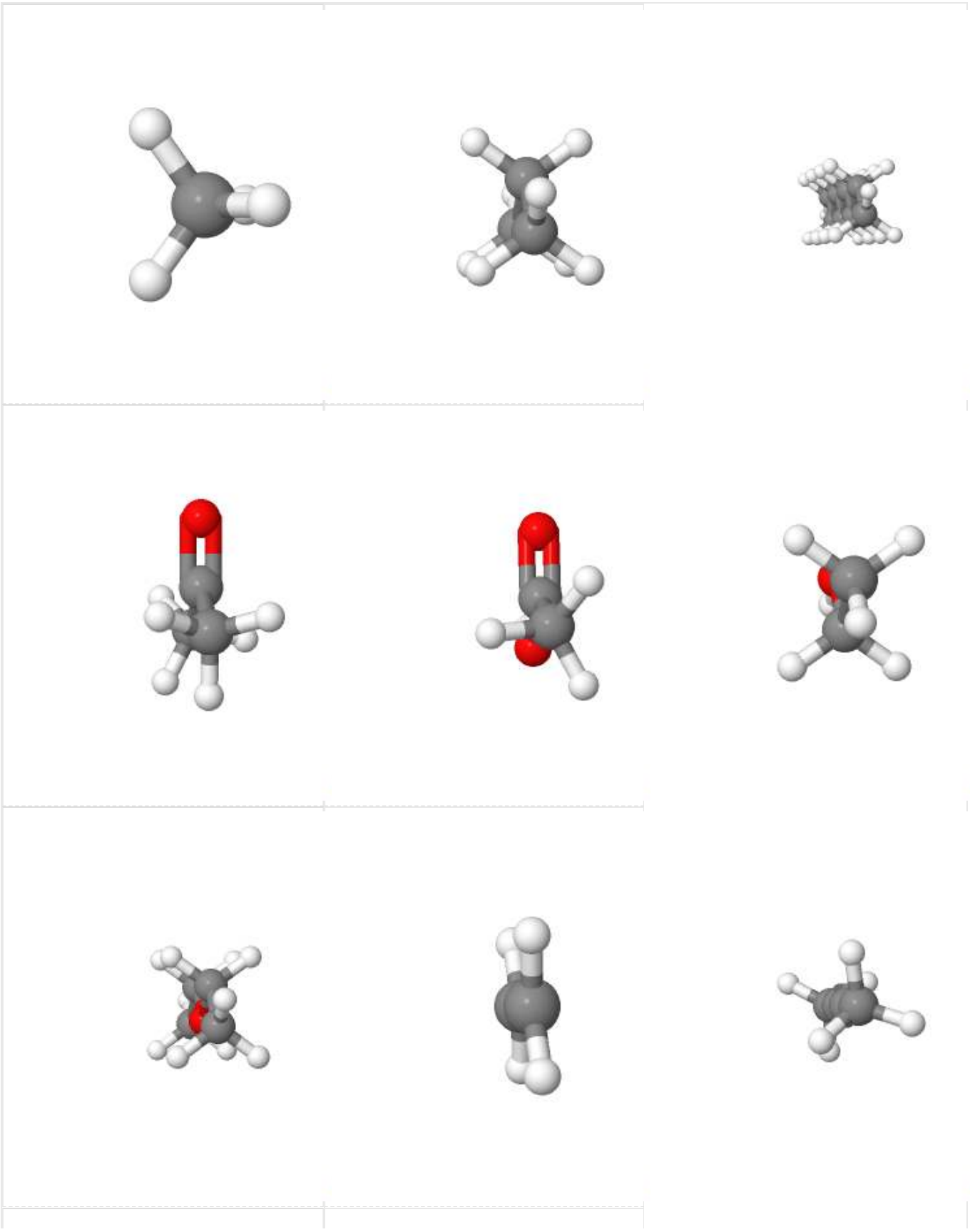
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12.1: Organic Chemistry

How many carbon-containing molecules are there? The current estimate is that there are around 20 million different known organic compounds. Why the uncertainty? Every day, scientists are coming up with new compounds. Some of these materials are of interest for a research project, while others are destined to be developed for a commercial market. As soon as we think we know how many organic compounds exist, more are

discovered and our number quickly becomes out of date. The molecule gallery in Figure 12.1.1 below shows just 12 of the 20,000,000 or so organic compounds. So, what exactly is an organic compound and why are there so many different possibilities?



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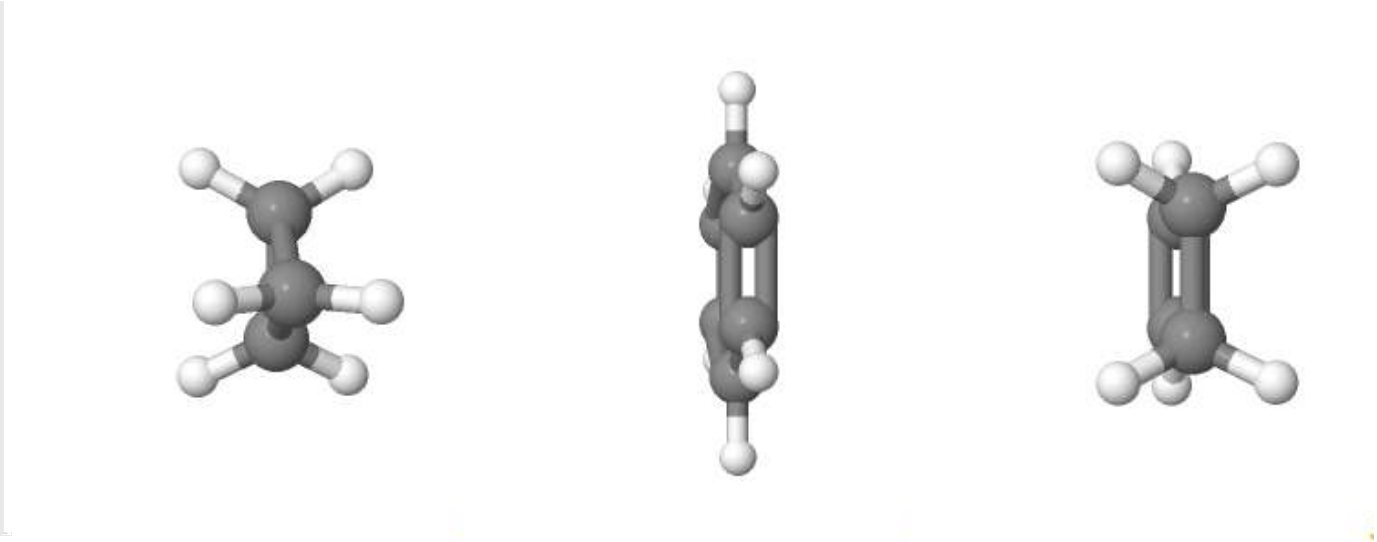


Figure 12.1.1: Row 1 (top row): methane, ethane, octane. Row 2: acetone, acetic acid, ethanol. Row 3: diethyl ether, ethylene, 2-butyne. Row 4: cyclopropane, benzene, cyclobutene. Carbon atoms are shown in gray, hydrogen in white, and oxygen in red.

Organic Chemistry

At one time in history, it was thought that only living things were capable of synthesizing the carbon-containing compounds present in cells. For that reason, the term organic was applied to those compounds. Eventually it was shown that carbon-containing compounds could be synthesized from inorganic substances, but the term "organic" has remained. Currently, **organic compounds** are defined as covalently bonded compounds containing carbon, excluding carbonates and oxides. By this definition, compounds such as carbon dioxide, CO_2, and sodium carbonate, Na_2CO_3, are considered to be inorganic compounds. **Organic chemistry** is the study of all organic compounds.

Organic chemistry is a very vast and complex subject. There are millions of known organic compounds – far more than the number of inorganic compounds. The reason lies within the uniqueness of carbon's structure and bonding capabilities. Carbon has four valence electrons, and therefore makes four separate covalent bonds in compounds. Carbon has the ability to bond to itself repeatedly, making long chains of carbon atoms, as well as ringed structures. In addition, the bonds may be single, double, or triple covalent bonds.

Carbon readily makes covalent bonds with other elements. While Figure 12.1.1 showed compounds comprised of only carbon, hydrogen, and oxygen, Figure 12.1.2 shows organic compounds that include nitrogen, sulfur, and chlorine. Other organic compounds commonly include the remaining halogens or phosphorus.

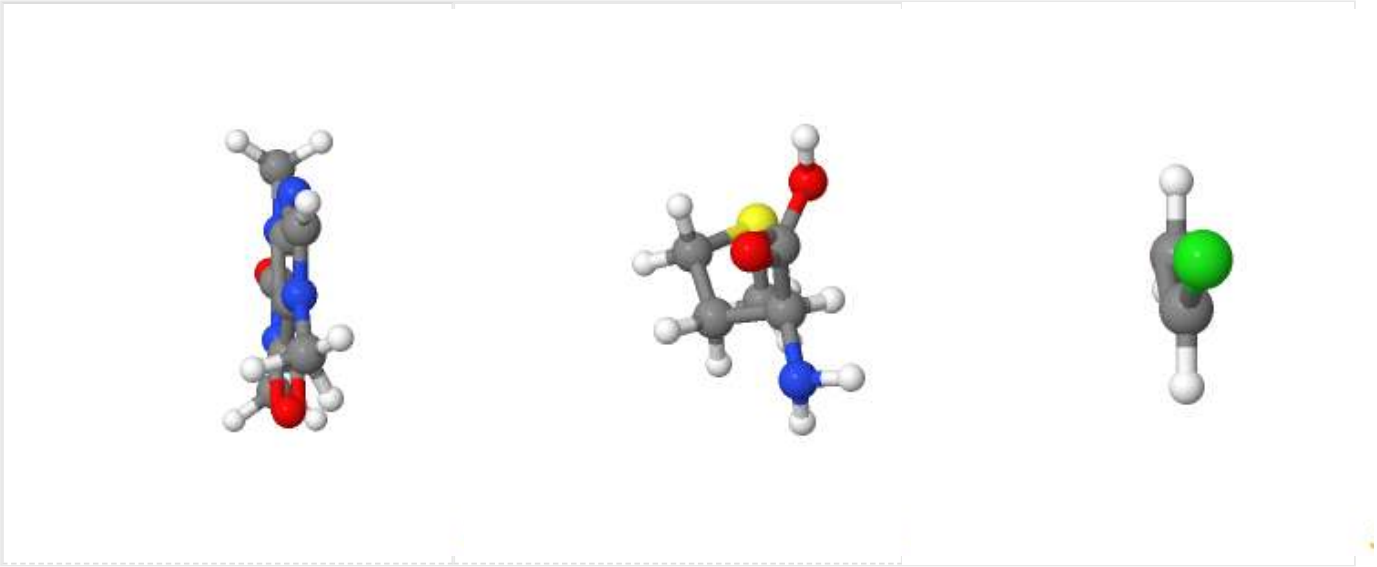


Figure 12.1.2: Caffeine (left), methionine (center, an essential amino acid), and vinyl chloride (right, a monomer of PVC plastic). Nitrogen atoms are shown in blue, sulfur in yellow, and chlorine in green.

The related field of biochemistry overlaps to some extent with organic chemistry. **Biochemistry** is the study of the chemistry of living systems. Many biochemical compounds are considered to be organic chemicals. Caffeine and methionine shown above are biochemical materials in terms of their use in the body, but organic chemicals in terms of their structure and chemical reactivity.

Summary

- Organic compounds are defined as covalently bonded compounds containing carbon, excluding carbonates and oxides. (By this definition, compounds such as carbon dioxide, CO_2, and sodium carbonate, Na_2CO_3, are considered to be inorganic.)
- Organic chemistry is the study of all organic compounds.
- Biochemistry is the study of the chemistry of living systems.

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12.2: Hydrocarbons

Learning Objectives

- Understand the difference between saturated and unsaturated hydrocarbons.

Hydrocarbons are compounds that contain only carbon and hydrogen. Even though they are composed of only two types of atoms, there is a wide variety of hydrocarbons because they may consist of varying lengths of chains, branched chains, and rings of carbon atoms, or combinations of these structures. In addition, hydrocarbons may differ in the types of carbon-carbon bonds present in their molecules.

Many hydrocarbons are found in plants, animals, and their fossils; other hydrocarbons have been prepared in the laboratory. We use hydrocarbons every day, mainly as fuels, such as natural gas, acetylene, propane, butane, and the principal components of gasoline, diesel fuel, and heating oil. The familiar plastics polyethylene, polypropylene, and polystyrene are also hydrocarbons. We can distinguish several types of hydrocarbons by differences in the bonding between carbon atoms.

Classes of Hydrocarbons

You have probably heard the terms saturated and unsaturated (as in monounsaturated or polyunsaturated) applied to different types of fats. It turns out these terms are derived from vocabulary used in organic chemistry. Hydrocarbons may also be classified as saturated or unsaturated. Their classification is based upon the number of bonds between carbon atoms.

- Saturated hydrocarbons have only single bonds between carbon atoms, so the carbon atoms are bonded to as many hydrogen atoms as possible. In other words, they are saturated with hydrogen atoms.

- Unsaturated hydrocarbons have at least one double or triple bond between carbon atoms, so the carbon atoms are not bonded to as many hydrogen atoms as possible. In other words, they are unsaturated with hydrogen atoms.

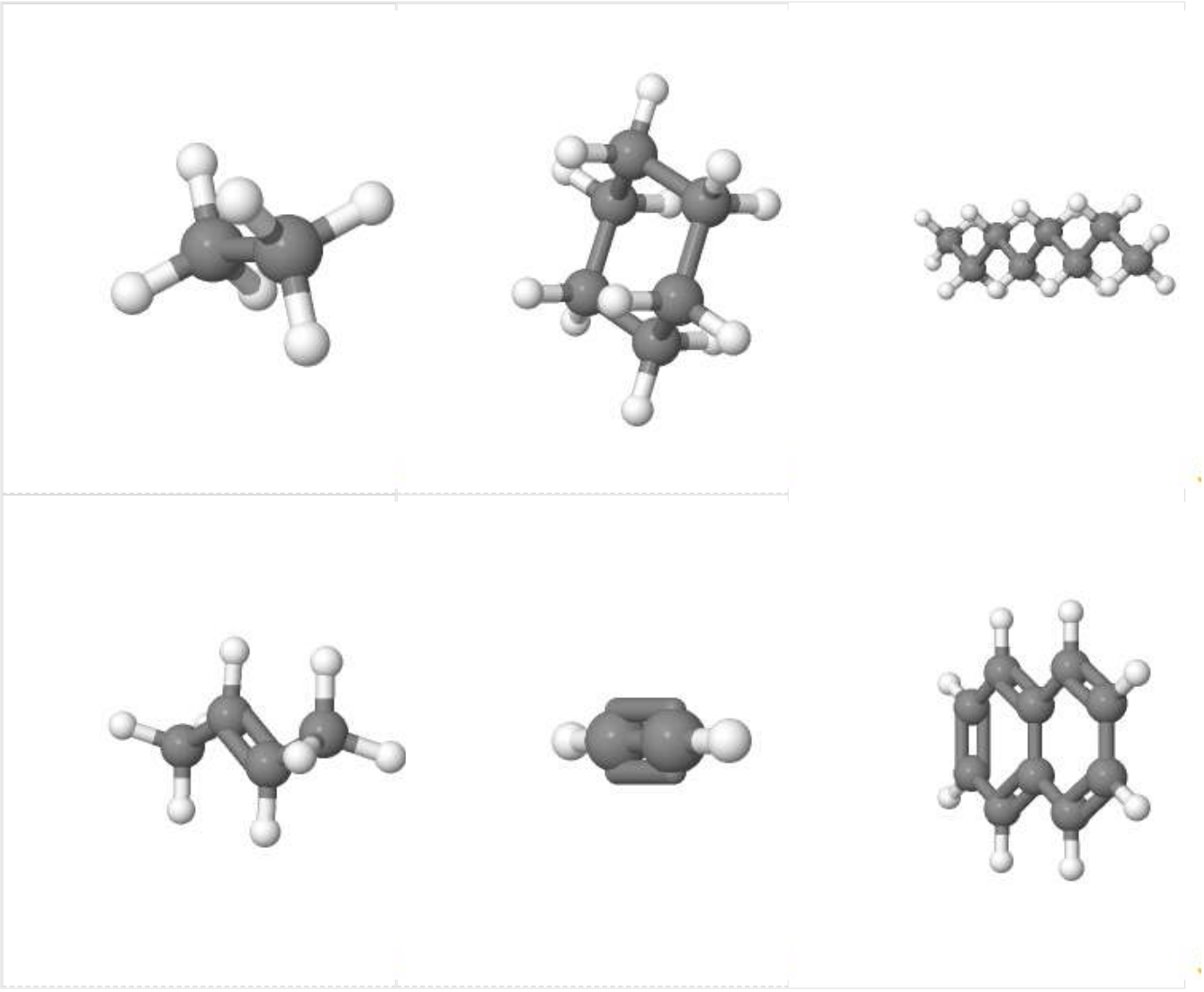


Figure 12.2.1: Row 1: saturated hydrocarbons – ethane, cyclohexane, and octane. Row 2: unsaturated hydrocarbons – trans-2-butene, acetylene, and naphthalene.

In the case of fats, a saturated fat is one in which all of the carbon-to-carbon bonds are single bonds – it is saturated with hydrogen atoms. A monounsaturated fat contains one carbon-to-carbon double bond, while a polyunsaturated fat contains many carbon-to-carbon double bonds. Though a hydrocarbon may also be considered unsaturated if it contains a carbon-to-carbon triple bond, fats will only contain carbon-to-carbon single bonds or double bonds.

Sources of Hydrocarbons

The main source of hydrocarbons is fossil fuels – coal, petroleum, and natural gas. Fossil fuels formed over hundreds of millions of years, as dead organisms were covered with sediments and placed under great pressure. Giant ferns in ancient swamps eventually turned into coal deposits. Dead organisms in ancient seas gradually formed deposits of petroleum and natural gas. Fossil fuels are usually retrieved from depths of hundreds (coal) to thousands (petroleum and natural gas) of feet.



Figure 12.2.2: First aid to the injured in an anthracite coal mine. Postcard copyright 1914 by J.W. Evans. (Janet Lindenmuth via Flickr)

Summary

- Hydrocarbons are compounds containing only carbon and hydrogen. If all carbon-to-carbon bonds are single bonds, the hydrocarbon is saturated. If the compound contains carbon-to-carbon double or triple bonds, the hydrocarbon is unsaturated.
- The primary source of hydrocarbons is fossil fuels.

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12.3: Alkanes

Learning Objectives

- Recognize straight-chain alkanes up to ten carbons in length.
- Convert between molecular formulas, structural formulas, condensed structural formulas, and line formulas for straight-chain alkanes.

An **alkane** is a hydrocarbon in which there are only single covalent bonds. The simplest alkane is methane, with the molecular formula CH_4. The carbon is the central atom and makes four single bonds to hydrogen atoms. Organic compounds are often represented by a **structural formula**. The easiest way to think of a structural formula is to think of it as a Lewis structure minus the lone pairs. The structural formula shows the layout of all of the atoms and bonds in a molecule.

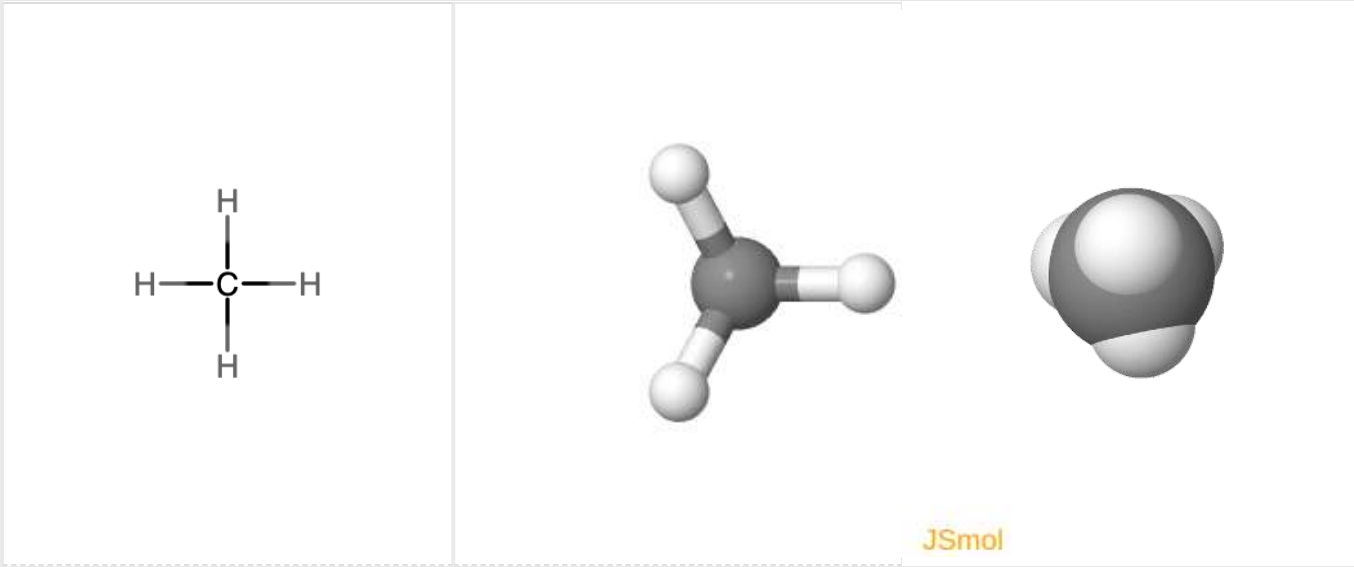
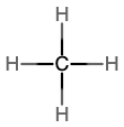
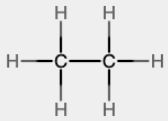
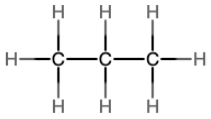
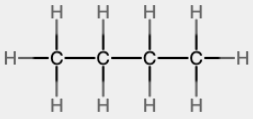
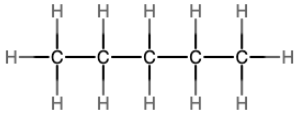
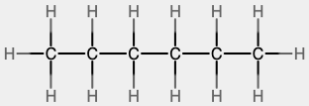
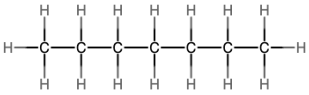
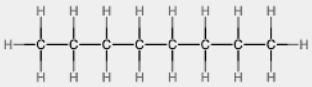
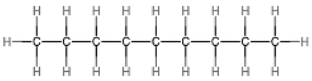
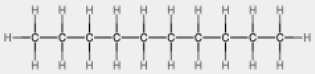


Figure 12.3.1: Methane is the simplest hydrocarbon and is shown using a structural formula, a ball-and-stick model, and a space-filling model.

The next simplest alkane is called ethane, C_2H_6, and consists of two carbon atoms with a single covalent bond between them. Each carbon is then able to bond to three hydrogen atoms. The alkane series progresses from there (see Table 12.3.1), increasing the length of the carbon chain by one carbon at a time. These alkanes are called straight-chain alkanes because the carbon atoms are connected in one continuous chain with no branches. Naming and writing structural and molecular formulas for the straight-chain alkanes is straightforward. The name of each alkane consists of a prefix that specifies the number of carbon atoms and has a suffix of *-ane*.

Notice that with the addition of each carbon atom, two additional hydrogen atoms added to the molecule. This pattern may be reflected by writing a general formula for the alkanes as C_nH_{2n+2} where n represents the number of carbon atoms in the chain. This table also shows that the boiling points of the alkanes steadily increase as the length of the carbon chain increases. This is due to an increase in the strength of the intermolecular attractive forces and is a general feature of other organic molecules as well.

Table 12.3.1: First Ten Members of the Alkane Series

Name	Molecular Formula	Structural Formula	Boiling Point (°C)
methane	CH_4		-161
ethane	C_2H_6		-89
propane	C_3H_8		-42
butane	C_4H_{10}		0.5
pentane	C_5H_{12}		36
hexane	C_6H_{14}		69
heptane	C_7H_{16}		99
octane	C_8H_{18}		126
nonane	C_9H_{20}		151
decane	$C_{10}H_{22}$		174

Simplifying Structural Formulas

The structural formulas in Table 12.3.1 take up more space and show an increasing amount repetition as the length of the chain of carbon atoms grows. Chemists often use condensed structural formulas to abbreviate the structures of organic molecules. A **condensed structural formula** has the appearance of a structure from which most or all of the bond symbols have been removed. Covalent bonds are understood to exist between each carbon atom and their associated hydrogen atoms, as well as between carbon atoms.

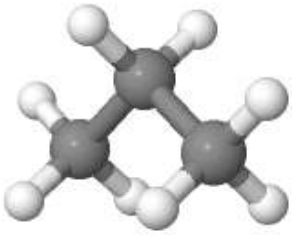
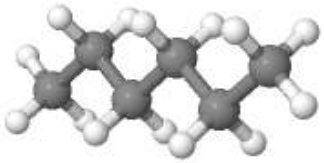
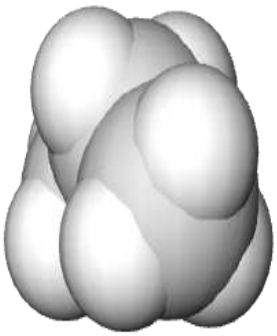
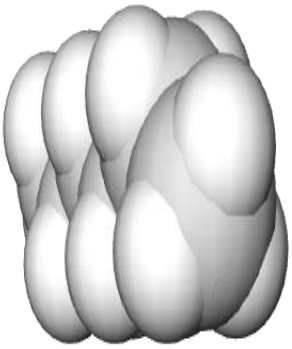
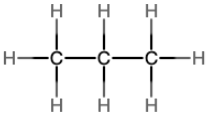
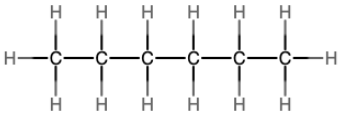
Line formulas are also used to abbreviate the structures of organic molecules. In a **line formula**, carbon atoms are represented by each end of a line or bend in a line. Hydrogen atoms are not drawn if they are attached to a carbon. Other atoms besides carbon and hydrogen are represented by their elemental symbols.

Table 12.3.2 shows the ball-and-stick models and space-filling models of propane and hexane followed by three different ways to draw each structure. Most of the organic molecules drawn in this text will use condensed structural formulas or line formulas.

Table 12.3.2: Models and Formulas of Propane and Hexane

	Propane	Hexane

https://chem.libretexts.org/@go/page/355086

	Propane	Hexane
Ball-and-Stick Model		
Space-Filling Model		
Structural Formula		
Condensed Structural Formula	$CH_3CH_2CH_3$	$CH_3CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_4CH_3$
Line Formula		

✅ Example 12.3.1:

Consider a straight-chain alkane that has 14 carbon atoms.

A. What is its molecular formula?
B. Write the condensed structural formula for this alkane.

Solution

A. The general formula for an alkane is C_nH_{2n+2}. An alkane that has 14 C atoms has 2(14) + 2 = 30 H atoms. The chemical formula is $C_{14}H_{30}$.

B. C atoms on each end are bonded to three H atoms, since they are only bonded to one other C atom. C atoms that are in between are all bonded to two other C atoms, so they only have room for two H atoms each. Therefore, the condensed structural formula is $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_{12}CH_3$.

A quick check confirms that the condensed structural formula has 14 C atoms and 30 H atoms.

Example 12.3.2: Writing Chemical Formulas

Give the molecular formula for the molecule represented with the line formula:



Solution

Each end or bend represents a carbon atom. There are 12 ends or bends, so this molecule has 12 C atoms. Alkanes have a general formula of C_nH_{2n+2}, meaning there are 2(12) + 2 = 26 H atoms. This yields a molecular formula of $C_{12}H_{26}$.

✏️ Exercise 12.3.1

Consider a straight-chain alkane that has 11 carbon atoms.

A. What is its molecular formula?
B. Write the condensed structural formula for this alkane.
C. Write the line formula for this alkane.

Answer A

$C_{11}H_{24}$

Answer B

$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$ or $CH_3(CH_2)_9CH_3$

Answer C



Summary

- An alkane is a hydrocarbon in which there are only single covalent bonds.
- Alkanes have a general formula of C_nH_{2n+2}.
- Condensed structural formulas and line formulas are often used to simplify the structures of organic compounds.

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12.4: Branched Alkanes

The ball-and-stick models you have seen thus far in this text use a rendering tool for viewing molecules online called JSmol. JSmol was designed to replicate molecular model kits (see Figure 12.4.1) that have long been used in chemistry classes to represent three-dimensional structures of molecules.



Figure **12.4.1**: A molecular model kit used to build three-dimensional molecular models. [Sonia (left) and Bin im Garten (right), via Wikimedia Commons]

A common hands-on activity is to provide students with molecular model kits to build models that have a specified chemical formula and to compare the models. There are millions of examples that may be selected for comparison, but if that formula was C_6H_{14}, students would soon notice that there are five different ways to build a molecule of C_6H_{14}. Substances that have the same molecular formula, but a different arrangement of atoms are called **isomers**. The five isomers of C_6H_{14} are shown in Figure 12.4.2:

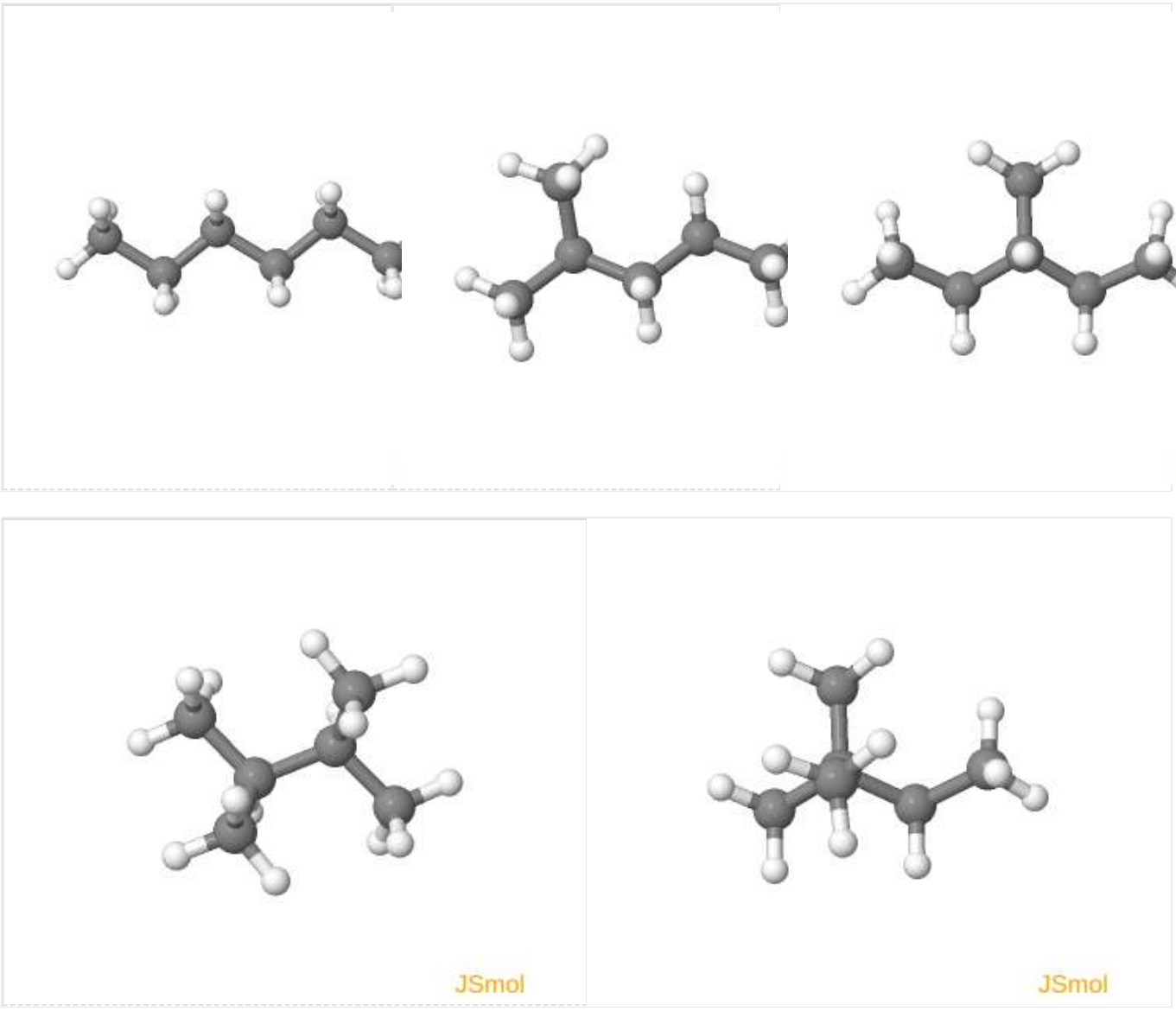


Figure **12.4.2**: The five isomers of C_6H_{14}.

Molecules rendered in JSmol may be rotated in all directions to examine molecules in different orientations. Simply "grab" onto the molecule and drag in the desired direction(s) as shown in Figure 12.4.3. Try rotating the molecules in Figure 12.4.2 above to make all 14 hydrogen atoms visible in each isomer. *Go ahead...try it!*

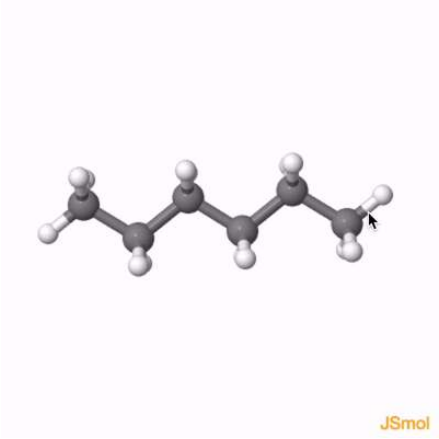


Figure **12.4.3**: Rotating a hexane molecule rendered with JSmol.

IUPAC Nomenclature

While molecules are written or drawn in two dimensions, it is important to remember that molecules are three-dimensional. Since isomers of each other are unique, they have different chemical and physical properties. The differences are sometimes small and at other times, quite significant. To distinguish isomers from one another, it is important that they be assigned a unique name. The International Union of Pure and Applied Chemistry (IUPAC) has devised a system of nomenclature that begins with the names of the alkanes and can be adjusted from there to account for more complicated structures.

Success with organic nomenclature begins with the ability to count carbon atoms. Similar to how Greek prefixes were used to designate a specific number of atoms in binary molecular compounds, IUPAC nomenclature uses prefixes based on the alkanes to designate a specific number of carbon atoms. Notice that Greek prefixes are used for specifying five or more carbon atoms.

Table 12.4.1: IUPAC Prefixes Used to Specify the Number of Carbon Atoms in Organic Compounds

Number of Carbon Atoms	IUPAC Prefix	Number of Carbon Atoms	IUPAC Prefix
1	meth-	6	hex-
2	eth-	7	hept-
3	prop-	8	oct-
4	but-	9	non-
5	pent-	10	dec-

Branched Alkanes

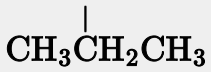
The isomers of C_6H_{14} shown in Figure 12.4.2 above are shown again in Table 12.4.2 below along with their condensed structural formulas, line formulas, and IUPAC names. While we already learned how to name hexane and other straight-chained alkanes in the previous section, the remaining isomers of C_6H_{14} are branched. How are the names for these branched molecules determined? Here are some basic guidelines:

☞ IUPAC Nomenclature of Branched Alkanes

1. Identify and name the **parent chain**.
   - The **parent chain** is the longest consecutive chain of carbon atoms in the molecule.

2. Identify and name all branches off the parent chain as **alkyl groups**.

- An **alkyl group** is an alkane that has had one H atom removed.
- If the H atom is removed from the end carbon of a straight-chain alkane, the alkyl group is named by changing the *–ane* suffix of the alkane to *–yl*.
- The name(s) of the alkyl group(s) precedes the name of the parent chain.
- Common alkyl groups:

methyl	$CH_3—$
ethyl	$CH_3CH_2—$
propyl	$CH_3CH_2CH_2—$
isopropyl	
butyl	$CH_3CH_2CH_2CH_2—$

3. Locate and assign a number to the position of each branch off the parent chain.

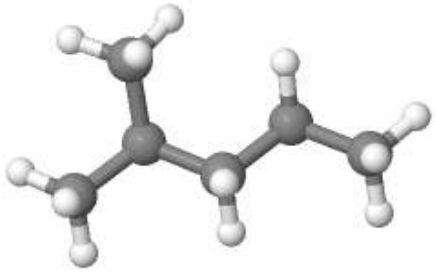
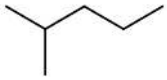
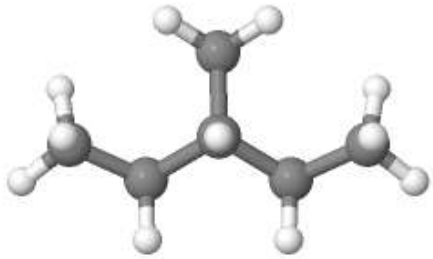
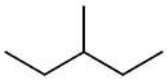
- For a single branch off the parent chain:
  - Count in the direction that gives the lowest number.
  - Place the number in front of the name of the alkyl group, separating the number and name with a hyphen (Example: 3-methyl).
- For multiple branches off the parent chain:
  - Count in the direction that gives the lowest number for any branch off the parent chain.
  - For branches of the same type:
    - Indicate the number of that type by using the Greek prefixes *di–*, *tri–*, *tetra–*, etc. in front of the name of the alkyl group.
    - Write the number of the position of each branch in ascending order, separating the numbers by commas.
    - Example: 2,3,3-trimethylheptane
  - For branches of different types:
    - The alkyl groups are named in alphabetical order, with a number indicating the position of each alkyl group off the parent chain.
    - Example: 3-ethyl-2-methylhexane

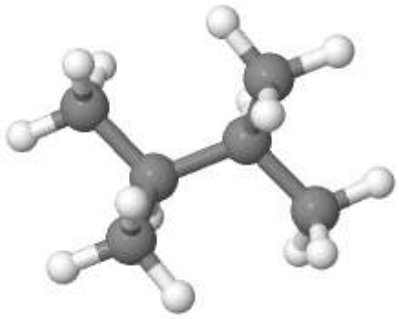
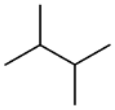
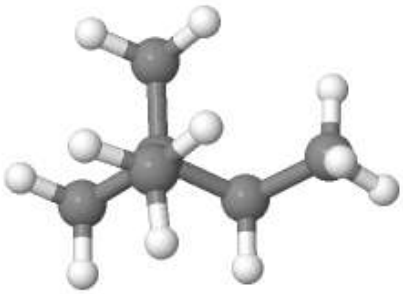
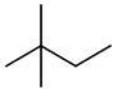
4. Additional Notes:

- The name of an alkane has no spaces.
- Numbers are separated by commas. Numbers and letters are separated by hyphens.
- Whether branched or unbranched, alkanes always have a general formula of C_nH_{2n+2}.

Table 12.4.2: The Five Isomers of C_6H_{14}.

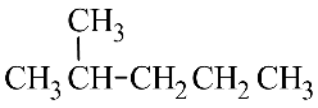
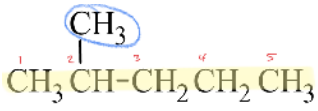
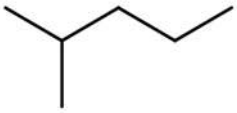
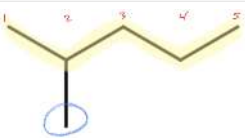
Three-Dimensional Structure	Condensed Structural Formula	Line Formula	IUPAC Name

Three-Dimensional Structure	Condensed Structural Formula	Line Formula	IUPAC Name
 JSmol	$CH_2\,CH_3$		hexane
 JSmol	$I_2\,CH_3$		2-methylpentane
 JSmol	$I_2\,CH_3$		3-methylpentane

Three-Dimensional Structure	Condensed Structural Formula	Line Formula	IUPAC Name
	CH_3		2,3-dimethylbutane
	CH_3		2,2-dimethylbutane

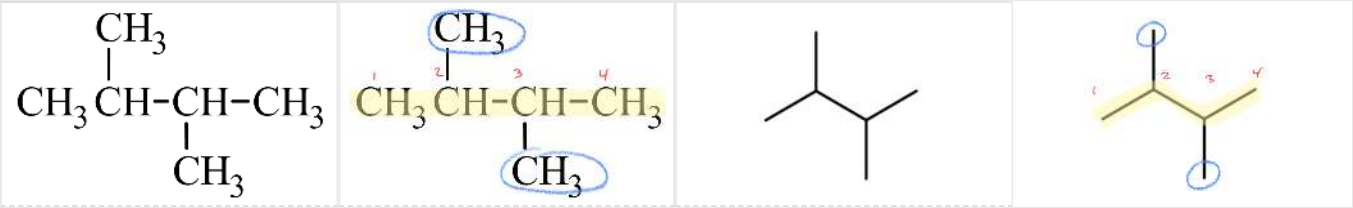
Let's look at how the name 2-methylpentane is derived. Following the guidelines for the IUPAC Nomenclature of Branched Alkanes listed above:

1. The parent chain has five carbon atoms, or *pentane*.
2. There is a one-carbon alkyl group that is branched off of the parent chain. If a one-carbon alkane is called methane, then a one-carbon alkyl group is called a *methyl group*. There are no spaces in the name, so *methyl* is written in front of *pentane*, giving us *methylpentane*.
3. The *methyl group* is branched off of the second carbon atom in the parent chain, so *2* is written in front of *methylpentane*. Letters and numbers are separated with a hyphen, making it *2-methylpentane*.

Next, lets look at how the name 2,3-dimethylbutane is derived. Following the guidelines for the IUPAC Nomenclature of Branched Alkanes listed above:

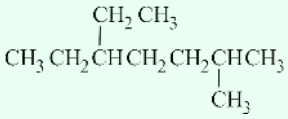
1. The parent chain has four carbon atoms, or *butane*.
2. There are two *methyl* groups branched off of the parent chain, which means that *dimethyl* is written in front of *butane*, making it *dimethylbutane*.
3. One *methyl group* is branched off the second carbon atom of the parent chain and one is branched off the third carbon atom, so *2,3-* is written in front of *dimethylbutane*, making it *2,3-dimethylbutane*.



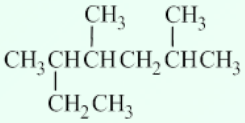
✅ Example 12.4.1

Write the IUPAC name for each molecule.

A.

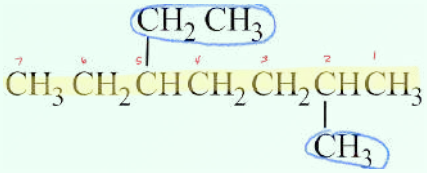


B.



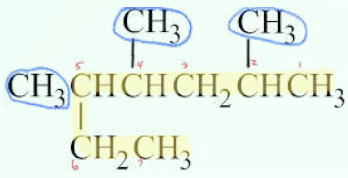
Solution

A.



1. The parent chain has seven carbon atoms, yielding a name of *heptane*.
2. There is one *ethyl* group and one *methyl* group branched off of the parent chain.
3. The parent chain is numbered from *right to left*, since this gives the lowest number for any branch off of the parent chain.
4. The *methyl* group is branched off the 2nd carbon on the parent chain (making it *2-methyl*) and the *ethyl* group is branched off the 5th carbon atom on the parent chain (making it *5-ethyl*).
5. Alphabetically, *ethyl* comes before *methyl*, so *5-ethyl* comes first in the name, making it *5-ethyl-2-methylheptane*.

B.

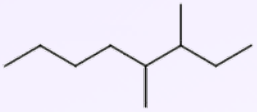


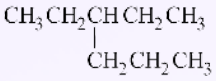
1. The parent chain has seven carbon atoms, which is longer than the six consecutive carbon atoms obtained by tracing straight across the molecule. Therefore, the parent chain is named *heptane*.
2. There are three *methyl* groups branched off of the parent chain, which means that *trimethyl* is written in front of *heptane*, yielding *trimethylheptane*.
3. The parent chain is numbered from *right to left*, since this gives the lowest number for any branch off of the parent chain.

4. The *methyl groups* are branched off the second, fourth, and fifth carbon atoms of the parent chain, so *2,4,5-* is written in front of *trimethylheptane*, making it *2,4,5-trimethylheptane*.

Exercise 12.4.1

Write the IUPAC name for each molecule.

A. 

B. 

Answer A

3,4-dimethyloctane

Answer B

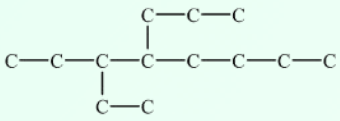
3-ethylhexane

Example 12.4.2

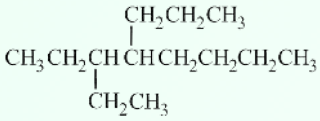
Write the condensed structural formula and line formula for each.

A. 3-ethyl-4-propyloctane
B. 3,3-dimethylhexane

Solution

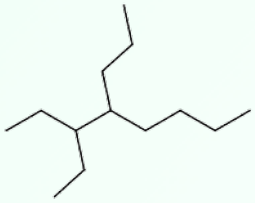
A. 

It's often the easiest to begin by drawing out the skeleton structure. The name *octane* indicates the parent chain has eight carbon atoms. An *ethyl* group has two carbon atoms and is branched off the 3rd carbon atom on the parent chain. A *propyl* group has three carbon atoms and is branched off the 4th carbon atom on the parent chain.



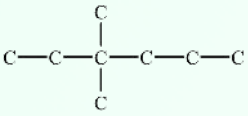
From here, fill in the H atoms. Remember that hydrogen can only make one bond, while carbon makes four. If a carbon atom is bonded to one other atom by a single bond, it can hold three hydrogen atoms. If a carbon atom makes three bonds to other atoms, it can only hold one hydrogen atom.

Once all of the hydrogen atoms have been added, you can check to see if the structure conforms to the general formula for alkanes, C_nH_{2n+2}. Since there are 13 carbon atoms, there should be 2(13) + 2 = 28 hydrogen atoms. Summing together all of the hydrogen atoms in the condensed structural formula shows the formula is indeed $C_{13}H_{28}$.

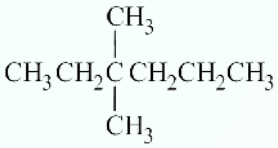


The line formula is shown above. The parent chain shows a total of eight ends or bends. There are two carbon atoms branched off the third carbon of the parent chain. There are three carbon atoms branched off the fourth carbon of the parent chain.

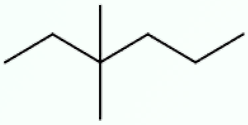
B.



Begin by drawing the skeleton structure. The name *hexane* indicates the parent chain has six carbon atoms. A *methyl* group has one carbon atom. The prefix *di–* indicates there are two *methyl* groups. Both *methyl* groups are branched off the 3rd carbon atom on the parent chain.



From here, fill in the H atoms. Each hydrogen makes one bond, while carbon makes four. Once all of the hydrogen atoms have been added, you can check to see if the structure conforms to the general formula for alkanes, C_nH_{2n+2}. Since there are 8 carbon atoms, there should be 2(8) + 2 = 18 hydrogen atoms. Summing together all of the hydrogen atoms in the condensed structural formula shows the formula is indeed C_8H_{18}.



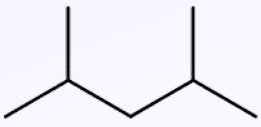
The line formula is shown above. The parent chain shows a total of six ends or bends. There are two one-carbon atom branches off the third carbon of the parent chain.

Exercise 12.4.2

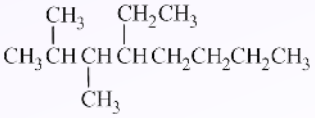
A. Draw the line structure of 2,4-dimethylpentane.

B. Write the condensed structural formula of 4-ethyl-2,3-dimethyloctane.

Answer A



Answer B



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Summary

- A structural isomer is one of multiple molecules that have the same molecular formula, but different structural formulas.
- Nomenclature rules for branched hydrocarbons are given.

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12.5: Alkenes and Alkynes



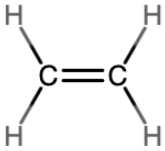
Figure **12.5.1**: Cutting with an acetylene torch. (jason.kaechler via Wikimedia Commons)

One of the most effective ways to cut metal is with an oxy-acetylene torch. Very high temperatures are obtained when acetylene burns in oxygen. Mixed 1:1 with oxygen, a temperature of over 3000°C may be achieved. The amount of energy released is high – the net heat of combustion is 1300 kJ/mol. Safety precautions need to be observed, as the gas is very explosive. The oxy-acetylene torch is one of the top tools for welding and cutting. Acetylene is also known as ethyne.

Alkenes and Alkynes

Alkenes and alkynes are two different classes of unsaturated hydrocarbons. An **alkene** has one or more carbon-carbon double bonds, while an **alkyne** has one or more carbon-carbon triple bonds. The fewest carbon atoms possible in either an alkene or alkyne is two. Whenever an additional bond is placed between two carbon atoms, it displaces two hydrogen atoms. Since alkanes have a general formula of C_nH_{2n+2}, this means that alkenes have a general formula of C_nH_{2n}, and alkynes have a general formula of C_nH_{2n-2}. Table 12.5.1 provides a side-by-side comparison of alkenes and alkynes.

Table 12.5.1: A Comparison of Alkenes and Alkynes

	Alkenes	Alkynes
Suffix	–ene	–yne
Distinguishing Feature	C-to-C double bond	C-to-C triple bond
General formula	C_nH_{2n}	C_nH_{2n-2}
Simplest Example	ethene	ethyne
Molecular Formula of Simplest	C_2H_4	C_2H_2
Ball-and-Stick Model of Simplest		
Structural Formula of Simplest		
Condensed Structural Formula of Simplest	$CH_2 = CH_2$	$CH \equiv CH$

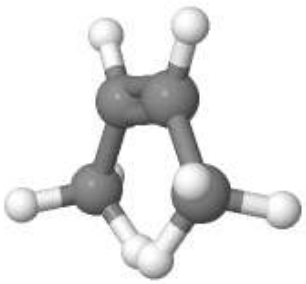
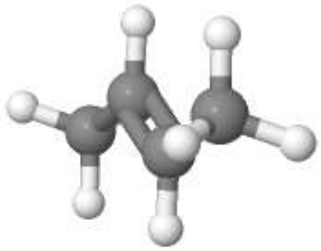
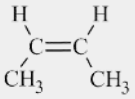
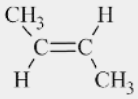
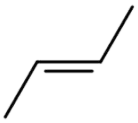
Cis-Trans Isomerism

Isomers were previously defined as substances that have the same molecular formula, but a different arrangement of atoms. Many alkenes form a second type of isomer called a **geometric isomer**. In a set of geometric isomers, the same types of atoms are attached to each other in an identical order, but the geometries of the two molecules differ.

Geometric isomers in hydrocarbons are possible when the carbon atoms involved in a carbon-to-carbon double bond are located in any position other than an end/terminal carbon atom. In other words, an alkene would need at least four carbon atoms for geometric isomers to be possible. Geometric isomers of alkenes differ in the orientation of the groups on either side of a carbon-to-carbon double bond. Unlike a single bond, where atoms are free to rotate along the axis between two atoms, a double bond prevents rotation along the axis. As a result, the atoms bonded to the double-donded carbon atoms are held in fixed positions.

The two geometric isomers of 2-butene, $\mathrm{CH_3CH = CHCH_3}$, are shown in Table 12.5.2 below. Notice that in the *cis*– isomer, the chain builds in each direction on the *same* side of the double bond, while in the *trans*– isomer, the chain builds in each direction on the *opposite* sides of the double bond.

Table 12.5.2: The Two Geometric Isomers of 2-Butene

	***cis*-2-butene**	***trans*-2-butene**
Ball-and-Stick Model		
Condensed Structural Formula		
Line formula		

IUPAC Nomenclature of Alkenes and Alkynes

The IUPAC nomenclature of alkenes and alkynes is an extension of the guidelines used to name the alkanes. While the rules used to name branched alkenes and alkynes have the potential to be complex, this text only examines the nomenclature of straight-chain alkenes and alkynes.

Here are some basic guidelines:

☞ IUPAC Nomenclature of Straight-Chain Alkenes and Alkynes

1. Count the number of carbon atoms in the parent chain and write the name of the **alkane** that has the same number of carbon atoms.
   - For an **alkene**, drop *–ane* from the alkane name and replace with the suffix *–ene*.

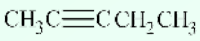
- For an **alkyne**, drop *–ane* from the alkane name and replace with the suffix *–yne*.

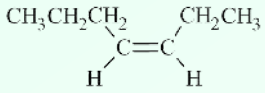
2. Locate and assign a number to the slot position of the carbon-to-carbon double or triple bond.
   - Count in the direction that gives the lowest number.
   - Place the number in front of the entire name, separating the number and name with a hyphen (Example: 2-pentyne).

3. For **alkenes**: If the double bond is in any slot position other than the first, you must tell whether the alkene is a *cis–* or *trans–* isomer.
   - If the chain builds in each direction on the same side of the double bond, it is a *cis–* isomer.
   - If the chain builds in each direction across the double bond, it is a *trans–* isomer.
   - The prefix *cis–* or *trans–* is written before the slot position (Example: *trans*-2-pentene).

✅ Example 12.5.1

Write the IUPAC name for each molecule.

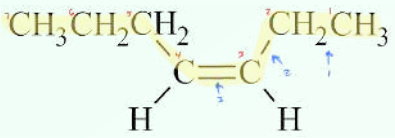
A. 

B. 

Solution

A. 

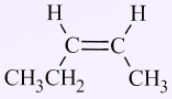
1. The parent chain has five carbon atoms, so the prefix is *pent–*. A triple bond makes this an alkyne, so the suffix is *–yne*. Putting the two together gives us *pentyne*.
2. The triple bond is located in the second slot. Therefore, this is *2-pentyne*.

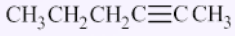
B. 

1. The parent chain has seven carbon atoms, so the prefix is *hept–*. A double bond makes this an alkene, so the suffix is *–ene*. Putting the two together gives us *heptene*.
2. The double bond is located in the third slot. Therefore, this is *3-heptene*.
3. The double bond is not in the first slot, so geometric isomers are possible. Since the chain builds on the same side of the molecule (the top side, in this case), it is a *cis–* isomer. The full name is *cis-3-heptene*.

✏️ Exercise 12.5.1

Write the IUPAC name for each molecule.

A. 

B. 

Answer A

cis-2-pentene

Answer B

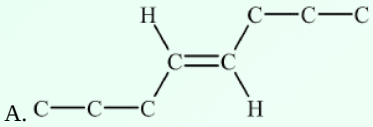
2-hexyne

✅ Example 12.5.2

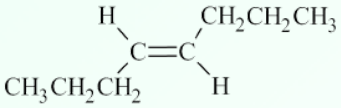
Write the condensed structural formula for each.

A. *trans*-4-octene
B. 1-hexyne

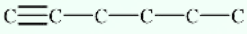
Solution

A. 

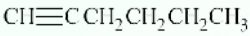
Begin by drawing the skeleton structure. The prefix *oct*– indicates the parent chain has eight carbon atoms. A suffix of *–ene* indicates this is an alkene, so it has a carbon-to-carbon double bond. The prefix *trans*– means that the molecule builds across the double bond. The number *4* indicates the double bond is in the fourth slot.



From here, fill in the H atoms. Each hydrogen makes one bond, while carbon makes four. Once all of the hydrogen atoms have been added, you can check to see if the structure conforms to the general formula for alkenes, C_nH_{2n}. Since there are 8 carbon atoms, there should be 2(8) = 16 hydrogen atoms. Summing together all of the hydrogen atoms in the condensed structural formula shows the formula is indeed C_8H_{16}.

B. 

Begin by drawing the skeleton structure. The prefix *hex*– indicates the parent chain has six carbon atoms. A suffix of *–yne* indicates this is an alkyne, so it has a carbon-to-carbon triple bond. The number *1* indicates the triple bond is in the first slot.



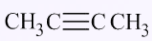
From here, fill in the H atoms. Each hydrogen makes one bond, while carbon makes four. Once all of the hydrogen atoms have been added, you can check to see if the structure conforms to the general formula for alkynes, $C_nH_{2n–2}$. Since there are 6 carbon atoms, there should be 2(6) – 2 = 10 hydrogen atoms. Summing together all of the hydrogen atoms in the condensed structural formula shows the formula is indeed C_6H_{10}.

✏️ Exercise 12.5.2

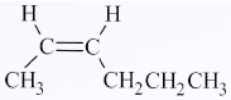
Write the condensed structural formula for each.

A. 2-butyne
B. *cis*-2-hexene

Answer A



Answer B



Summary

- An alkene is a hydrocarbon with one or more carbon-to-carbon double covalent bonds.
- An alkyne is a hydrocarbon with one or more carbon-to-carbon triple covalent bonds.
- Geometric isomers are molecules with identical bonding arrangements but different geometries.
- Nomenclature rules for straight-chain alkenes and alkynes are given.

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12.6: Oxygen-Containing Organic Compounds

Learning Objectives

- Classify oxygen-containing organic compounds based on the structure of the functional group that is present.
- Identify hydroxyl groups, carbonyl groups, and carboxyl groups within oxygen-containing organic compounds.

Oxygen-containing organic compounds may be classified based on the structural arrangement of oxygen atoms in relation to the carbon and hydrogen atoms within a molecule. Such structural arrangements often allow oxygen-containing compounds to function in a similar manner chemically and physically from molecule-to-molecule and are called **functional groups**. Figure 12.6.1 shows three common oxygen-containing functional groups found in organic molecules.

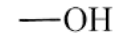
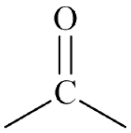
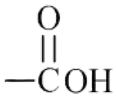
hydroxyl group	
carbonyl group	
carboxyl group	

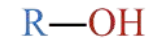
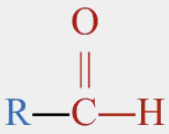
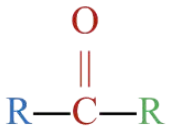
Figure **12.6.1**: Common oxygen-containing functional groups.

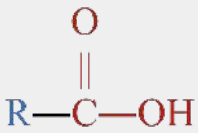
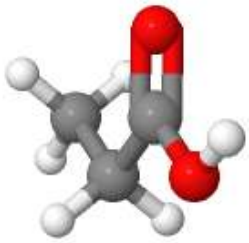
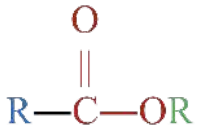
Classifying Oxygen-Containing Organic Compounds

The functional groups shown in Figure 12.6.1 are just three of many different functional groups. However, the presence and location of functional groups in organic molecules allows us to lump them together into different classes that function similarly, both chemically and physically. Table 12.6.1 shows the classification of several types of oxygen-containing organic compounds. Though the list is not complete, it provides some perspective as to the wide variety of oxygen-containing organic compounds that are possible.

Notice the presence of the functional groups shown above in the different families of compounds. The letter **R** represents an alkyl group (or more generally, a hydrocarbon group). As a reminder, molecular models rendered with JSmol may be manipulated by grabbing the molecules and rotating it to your desired orientation.

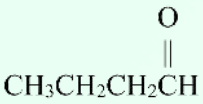
Table 12.6.1: Classification of Oxygen-Containing Organic Molecules

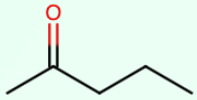
Family	Structure	Example
alcohol		
aldehyde		
ketone		

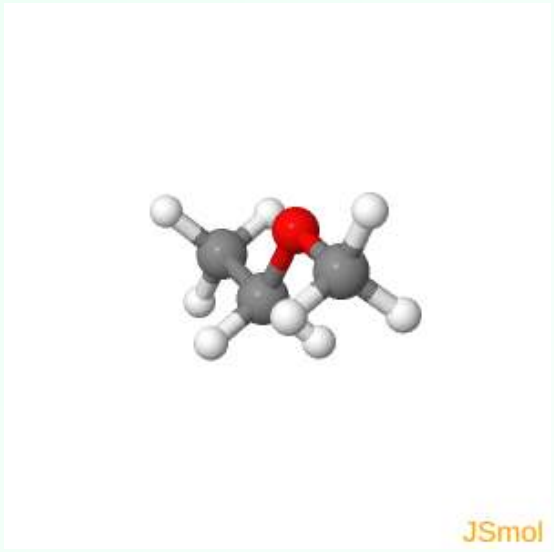
Family	Structure	Example
carboxylic acid		
ester		
ether		

✅ Example 12.6.1

Classify each molecule as an alcohol, aldehyde, ketone, carboxylic acid, ether, or ester.

A. 

B. 

C. 

D. CH_3OH

Solution

A. This molecule has an alkyl group that is bonded to a carbon atom that is double bonded to an oxygen atom and single bonded to a hydrogen atom. This gives it a general formula that fits an aldehyde,  .

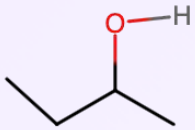
B. This molecule has an alkyl group attached to each side of a carbonyl group,  . This gives it a general formula that fits a ketone,  .

C. This molecule has an alkyl group attached to each side of an oxygen atom. This gives it a general formula that fits an ether, R—O—R.

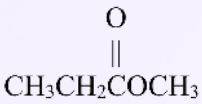
D. This molecule has an alkyl group attached to a hydroxyl group, —OH. This gives it a general formula that fits an alcohol, R—OH.

✏️ Exercise 12.6.1

Classify each molecule as an alcohol, aldehyde, ketone, carboxylic acid, ether, or ester.



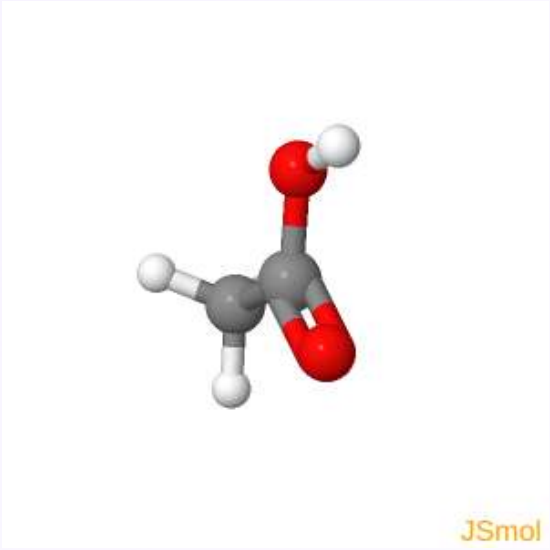
A.



B.

C. $CH_3OCH_2CH_2CH_2CH_3$

D.



Answer A

alcohol

Answer B

ester

Answer C

ether

Answer D

carboxylic acid

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12.7: Alcohols, Aldehydes, Carboxylic Acids, and Ketones

Alcohols, aldehydes, carboxylic acids, and ketones bear some relation to each other as they are easily converted from one to another through the process of oxidation or reduction. An oxidation reaction is ordinarily thought of as the loss of electrons, while a reduction reaction is ordinarily thought of as the gain of electrons. In organic chemistry, oxidation is typically observed when a molecule gains an oxygen atom (that makes it easy, huh?) and/or loses two hydrogen atoms. Reduction is the opposite process, by which a molecule loses an oxygen atom or gains two hydrogen atoms.

Oxidation of Alcohols

Alcohols may be classified according to the number of carbon atoms directly bonded to the carbon atom that is bonded to the hydroxyl group. If only one carbon atom is directly attached to that carbon, the alcohol is called a primary alcohol. If two carbon atoms are directly attached, the alcohol is called a secondary alcohol. If three carbon atoms are directly attached, the alcohol is called a tertiary alcohol. That's great, but how are alcohols, aldehydes, carboxylic acids, and ketones related?

It turns out that primary alcohols may be oxidized to aldehydes which, in turn, may be oxidized to carboxylic acids. Figure 12.7.1 shows how ethanol, a primary alcohol that has two carbon atoms, is eventually oxidized to ethanoic acid, a two-carbon carboxylic acid. Ethanol is the alcohol found in alcoholic beverages and is produced through the fermentation of sugars and starches in fruits and grains. Ethanoic acid is also called acetic acid, the acid found in vinegar. The oxidation of ethanol is the reason why a bottle of wine will begin to taste sour and take on a vinegary taste several days after it has been opened and exposed to the air.

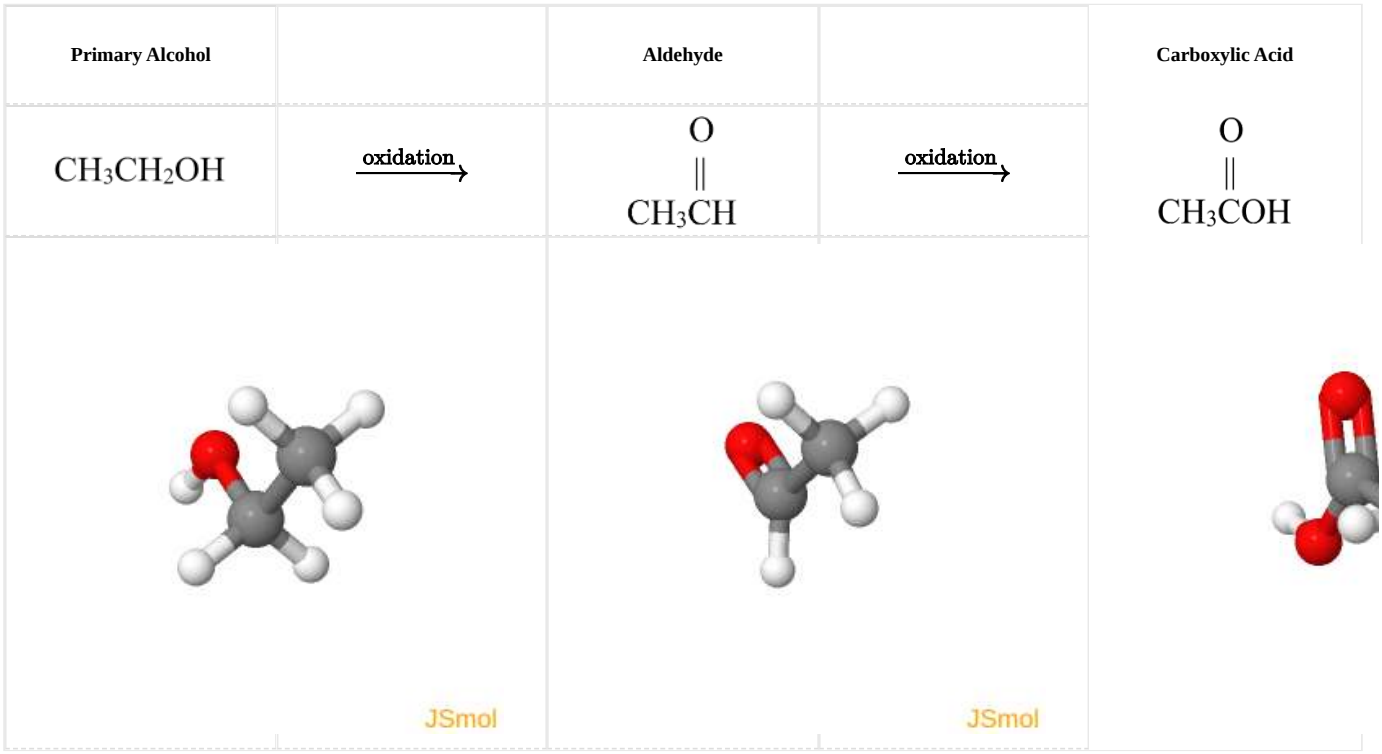
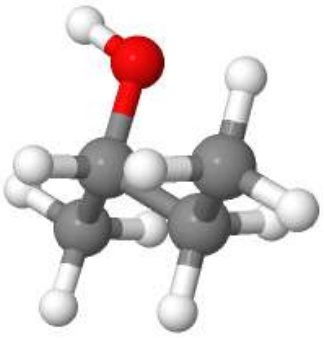


Figure **12.7.1**: The oxidation of ethanol to ethanal to ethanoic acid.

Secondary alcohols, on the other hand, may be oxidized to ketones. Figure 12.7.2 shows how 2-butanol, a secondary alcohol, is oxidized to 2-butanone, a ketone. Notice that the number of carbon atoms does not change, nor does the position to which the function group is attached. 2-butanol is a secondary alcohol because

the carbon atom to which the hydroxyl group is attached is directly bonded to two other carbon atoms.

	Secondary Alcohol		Ketone
	$CH_3CH_2CH(OH)CH_3$	oxidation →	$CH_3CH_2C(=O)CH_3$



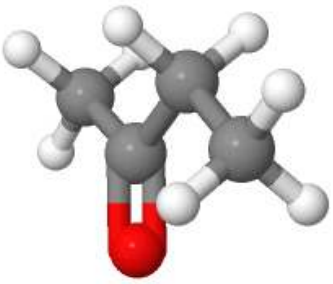
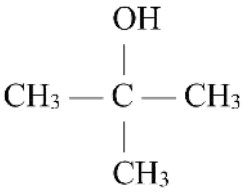


Figure **12.7.2**: The oxidation of 2-butanol to 2-butanone.

Tertiary alcohols, such as 2-methyl-2-propanol (see Figure 12.7.3), may not be oxidized in a same manner. 2-methyl-2-propanol is classified as a tertiary alcohol since the carbon atom to which the hydroxyl group is attached is directly bonded to three other carbon atoms.



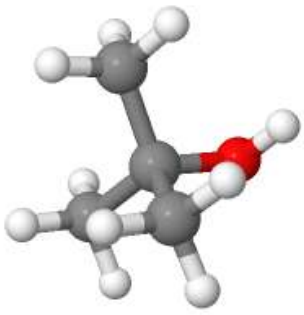


Figure **12.7.3**: Tertiary alcohols such as 2-methyl-2-propanol are not easily oxidized.

Nomenclature of Alcohols, Aldehydes, Ketones, and Carboxylic Acids

Alcohols, aldehydes, ketones, and carboxylic acids also bear resemblance to each other in terms of their nomenclature. Just like the alkenes and alkynes, the IUPAC nomenclature of alcohols, aldehydes, carboxylic acids, and ketones is an extension of the guidelines used to name the alkanes. While the rules used to name branched oxygen-containing organic molecules have the potential to be complex, this text only examines the nomenclature of straight-chain alcohols, aldehydes, carboxylic acids, and ketones.

Here are the basic guidelines:

☞ IUPAC Nomenclature of Straight-Chain Alcohols, Aldehydes, Carboxylic Acids, and Ketones

Alcohols, aldehydes, carboxylic acids, and ketones are named using the following convention:

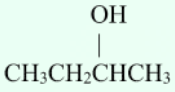
1. Count the number of carbon atoms in the parent chain and write the name of the **alkane** that has the same number of carbon atoms.
2. Drop *–e* from the alkane name.
3. If the compound is a(n):
   - **alcohol**, replace with the suffix *–ol*.
     - The position of the hydroxyl group is assigned a number by counting in the direction that gives the lowest number if the alcohol has three or more carbon atoms in the parent chain.
   - **aldehyde**, replace with the suffix *–al*.
   - **carboxylic acid**, replace with a suffix of *–oic acid* (*Note: there is a space in the name*).
   - **ketone**, replace with *–one*.
     - The position of the carbonyl group is numbered is assigned a number by counting in the direction that gives the lowest number.

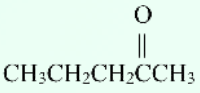
✅ Example 12.7.1

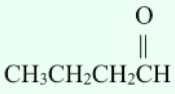
Write the IUPAC name for each molecule.

A. 

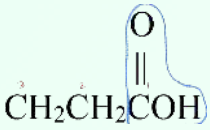
B.

C. 

D. 

E. 

Solution

A. 

1. This molecule has an alkyl group that is bonded to a carboxyl group,  . This gives it a structure that fits a carboxylic acid,  .
2. The parent chain has three carbon atoms. An alkane with three carbons is called *propane*.

3. To name the carboxylic acid, drop *–e* from *propane* and replace with the suffix *–oic acid* to give *propanoic acid*.

B. CH_3CH_2OH

1. This molecule has an alkyl group attached to a hydroxyl group, —OH. This gives it a structure that fits an alcohol, R—OH.
2. The parent chain has two carbon atoms. An alkane with two carbons is called *ethane*.
3. To name the alcohol, drop *–e* from *ethane* and replace with the suffix *–ol* to give *ethanol*.
4. There are less than three C atoms in the alcohol, so the position of the hydroxyl group is not assigned a number.

C. $CH_3CH_2CH(OH)CH_3$

1. This molecule has an alkyl group attached to a hydroxyl group, —OH. This gives it a structure that fits an alcohol, R—OH.
2. The parent chain has four carbon atoms. An alkane with four carbons is called *butane*.
3. To name the alcohol, drop *–e* from *butane* and replace with the suffix *–ol* to give *butanol*.
4. The hydroxyl group is attached to the 2nd carbon atom. The full name is *2-butanol*.

D. $CH_3CH_2CH_2C(=O)CH_3$

1. This molecule has an alkyl group attached to each side of a carbonyl group,  . This gives it a structure that fits a ketone,  .
2. The parent chain has five carbon atoms. An alkane with three carbons is called *pentane*.
3. To name the ketone, drop *–e* from *pentane* and replace with the suffix *–one* to give *pentanone*.
4. The carbonyl group is located at the 2nd carbon atom. The full name is *2-pentanone*.

E. $CH_3CH_2CH_2C(=O)H$

1. This molecule has a carbonyl group,  , located between an alkyl group and an H atom. This gives it a structure that fits an aldehyde,  .
2. The parent chain has four carbon atoms. An alkane with four carbons is called *butane*.
3. To name the aldehyde, drop *–e* from *butane* and replace with the suffix *–al* to give *butanal*.

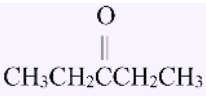
Exercise 12.7.1

Write the IUPAC name for each molecule.

A. $CH_3CH_2CH_2CH_2OH$

B. $CH_3CH_2CH_2CH_2CH_2C(=O)H$

C. $CH_3C(=O)OH$

D. 

Answer A

1-butanol

Answer B

hexanal

Answer C

ethanoic acid

Answer D

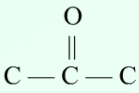
3-pentanone

✅ Example 12.7.2

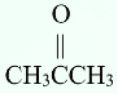
Write the condensed structural formula for each.

A. 2-propanone
B. pentanal

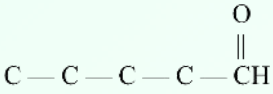
Solution

A. 

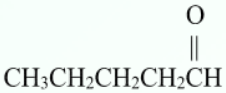
Begin by drawing the skeleton structure. The prefix *prop–* indicates the parent chain has three carbon atoms. A suffix of *–one* indicates this is a ketone, so it has a carbonyl group,  . The *2* indicates that the carbonyl group is located at the 2nd carbon atom in the parent chain.



From here, fill in the H atoms. Each hydrogen makes one bond, while carbon makes four. This structure fits a ketone,  .

B. 

Begin by drawing the skeleton structure. The prefix *pent–* indicates the parent chain has five carbon atoms. A suffix of *–al* indicates this is an aldehyde, so it has a carbonyl group,  , located at the first carbon atom.



From here, fill in the H atoms. Each hydrogen makes one bond, while carbon makes four. This structure fits an aldehyde,  .

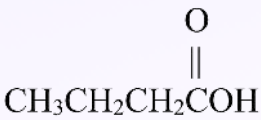
✏️ Exercise 12.7.2

Write the condensed structural formula for each.

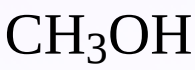
A. butanoic acid
B. methanol

https://chem.libretexts.org/@go/page/355092

Answer A



Answer B



Exercise 12.7.3

Why are numbers not assigned to the functional group in either aldehydes or carboxylic acids?

Answer

A number is not assigned to the position of the carbonyl group in an aldehyde, since it is always located at the first carbon atom. If the carbonyl group is somewhere in the middle of the molecule, the compound would be classified as a ketone. A number is not assigned to the position of the carboxyl group in a carboxylic acid, since it is always located at the first carbon atom.

Summary

- Alcohols can be classified as primary, secondary, or tertiary based on the number of carbon atoms bonded to the carbon containing the hydroxyl group.
- Primary alcohols may be oxidized to aldehydes and carboxylic acids. Secondary alcohols may be oxidized to ketones.
- Nomenclature rules for straight-chain alcohols, aldehydes, carboxylic acids, and ketones are given.

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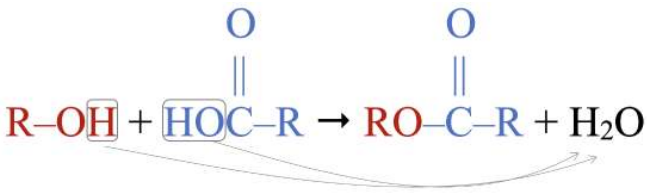
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12.8: Esters

Perfumes contain ingredients from a number of sources. Musk is obtained from animals, but the vast majority of perfume components are obtained from plants. Approximately 2,000 plant species have been used as sources for perfume materials. The needed chemicals are extracted using solvent extraction or distillation. The oils are diluted with ethanol to varying degrees, depending on the price of the finished product – the less ethanol present (meaning there is a higher percentage of the active scent ingredients), the more expensive the perfume.

Esters

An **ester** is an organic compound that is a derivative of a carboxylic acid. While esters occur naturally, they are easily synthesized by reacting an alcohol (shown in red) with a carboxylic acid (shown in blue):



Note that the structure of the carboxylic acid was flipped 180º to show the carboxyl group on the left, so it is easier to see how the portions of the ester (shown to the right of the arrow in red and blue) derived from the alcohol and the carboxylic acid fit together. The alcohol sheds an H atom, while the carboxylic acids sheds an –OH group. The H and –OH come together to form water, shown as the other product.

One of the cool things about the reaction to form an ester is that the progress of the reaction may be followed by the sense of smell, rather than visually. Esters tend to have pleasant odors and are very commonly found in plants – they are responsible for many distinct odors and flavors. For example, methyl salicylate has the odor and flavor of oil of wintergreen, while propyl ethanoate has that of a pear (see Figure 12.8.1).

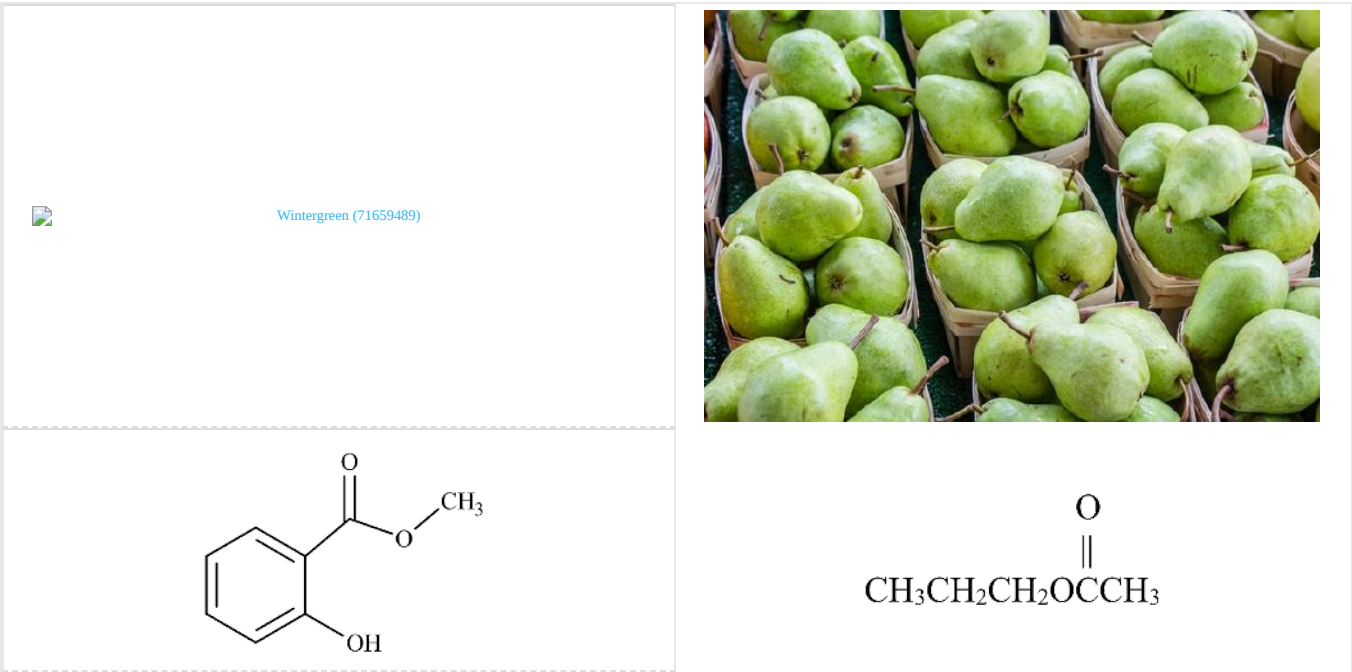


Figure **12.8.1**: Methyl salicylate is found in wintergreen (ill Szal via Wikimedia Commons) and propyl ethanoate is found in pears (Jonathon Mast via Unsplash).

On the other hand, carboxylic acids tend to have strong, foul odors. Hexanoic acid has a common name of caproic acid, which comes from the Latin word caper, which means goat. In other words, hexanoic acid has the scent of billy goats. The smell associated with blue cheese or parmesan cheese comes from butanoic acid, which has a common name of butyric acid, and was named from the smell of rancid butter.

Just like an ester may be chemically derived from an alcohol and carboxylic acid, so may its name. Here are the basic guidelines:

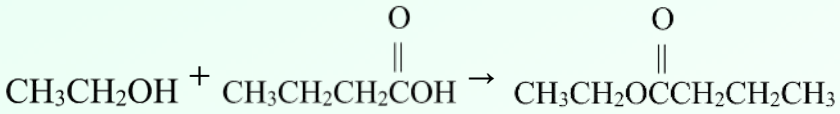
☞ IUPAC Nomenclature of Esters

Esters are named using the following convention:

1. The portion of the ester derived from the **alcohol** is named as an **alkyl group**.
   - Note: the number 1 is dropped from the name of an alkyl group.
   - Example: butyl group (the alkyl group derived from 1-butanol)
2. The portion derived from the **carboxylic acid** named by
   - Dropping *–ic acid* from the acid name and replacing with a suffix of *–ate*.
   - Example: propyl butanoate (derived from 1-propanol and butanoic acid)

✅ Example 12.8.1

Write the IUPAC name for each molecule shown in the reaction below:

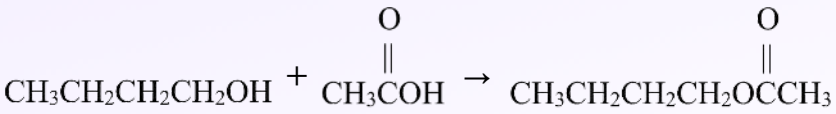


Solution

- The first molecule is an alcohol,  , that has two carbon atoms. This gives it the name *ethanol*.
- The second molecule is a carboxylic acid,  , that has four carbon atoms. This gives it the name *butanoic acid*.
- The third molecule is an ester,  .
  - The portion derived from the alcohol has two carbon atoms and is named as an alkyl group, or *ethyl*.
  - The portion derived from the carboxylic acid has four carbon atoms. It is named by dropping *–ic acid* from *butanoic acid* and replacing with a suffix of *–ate* to give *butanoate*.
- Ultimately, this allows us to arrive at a name of *ethyl butanoate* for the ester.

✏️ Exercise 12.8.1

Write the IUPAC name for each molecule shown in the reaction below, from left to right:



Answer

1-butanol, ethanoic acid, butyl ethanoate

Summary

- An ester is an organic compound that is a derivative of a carboxylic acid, in which the hydrogen atom of the hydroxyl group has been replaced with an alkyl group.
- Nomenclature rules for straight-chain esters are given.

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12.9: Ethers

Race car drivers are always looking for that "edge" that will make their car (legally) faster than their competitors' cars. One way to get more speed is to burn the gasoline in the car engine more efficiently. Methyl t-butyl ether (MTBE) has been used for this purpose, but is being discontinued due to concerns about the contamination of drinking water by leaking fuel tanks that contain this compound.

Ethers

An **ether** is an organic compound in which two alkyl groups are bonded to the same atom of oxygen. An ether is represented by the general formula $\ce{R-O-R}$. The R groups may be the same or they may be different. Ethers often go by their common name rather than their IUPAC name. Here are the basic guidelines:

☞ Nomenclature of Ethers

Ethers are named using the following convention:

1. Name each **alkyl group** attached to the oxygen in alphabetical order, followed by the word *ether* (Example: butyl ethyl ether).
2. If the two attached alkyl groups are identical, use the prefix *di–* for the alkyl group (Example: dimethyl ether).

✅ Example 12.9.1

A. Assign a name to the molecule below:

$$\ce{CH3OCH2CH3}$$

B. Write the condensed structural formula of diethyl ether.

Solution

A. The alkyl group to the left of the oxygen has one C atom, which makes it a *methyl* group. There alkyl group to the right of the oxygen has two C atoms, which makes it an *ethyl* group. Alphabetically, *ethyl* comes before *methyl*, giving it a name of *ethyl methyl ether*.

B. The name *diethyl ether* means there are two *ethyl* groups attached to the O atom. An ethyl group has two C atoms. Since an O atom only makes two bonds, the condensed structural formula is:

$$\ce{CH3CH2OCH2CH3}$$

✏️ Exercise 12.9.1

A. Assign a name to the molecule below:

$$\ce{CH3OCH3}$$

B. Write the condensed structural formula of butyl methyl ether.

Answer A

dimethyl ether

Answer B

$$CH_3CH_2CH_2CH_2OCH_3$$

Summary

- An ether is an organic compound in which two hydrocarbon groups are bonded to the same atom of oxygen.
- Common naming rules for ethers are given.

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12.10: For Future Use

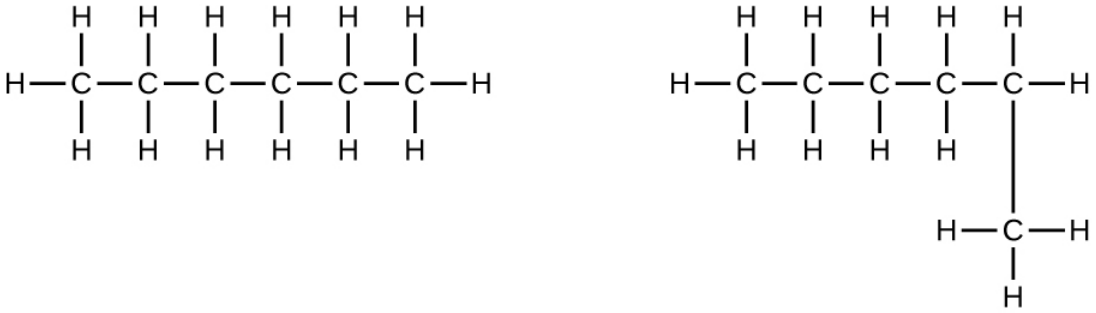
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12.11: Exercises

12.4: Branched Alkanes

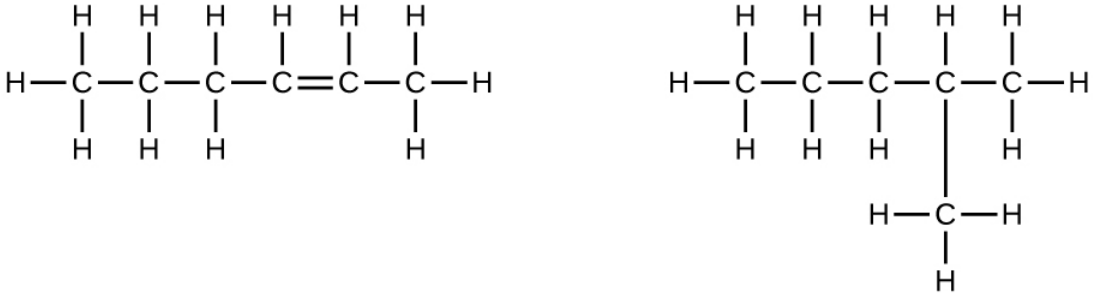
1. Explain why these two molecules are not isomers:



Answer

They are the same compound because each is a saturated hydrocarbon containing an unbranched chain of six carbon atoms.

2. Explain why these two molecules are not isomers:

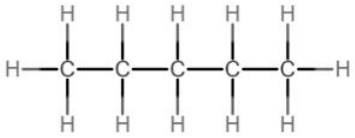
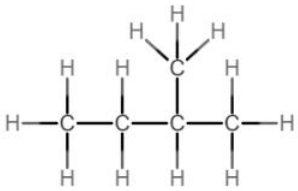
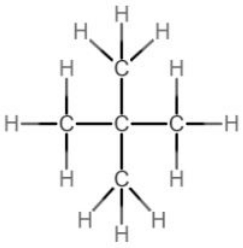


Answer

The molecule on the left has a molecular formula of C_6H_{12} while the one on the right has a molecular formula of C_6H_{14}. Since they have different numbers of hydrogens, they cannot be isomers.

3. Write the chemical formula and three different Lewis structures for a five-carbon alkane.

Answer

C_5H_{12};  ;  ; 

4. What are the IUPAC names of the alkanes drawn in Exercise 3?

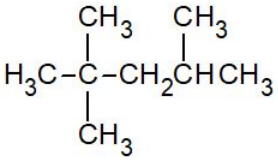
-

Answer

pentane; 2-methylbutane; 2,2-dimethylpropane

5. In the gasoline industry, what is called *isooctane* is actually 2,2,4-trimethylpentane. Draw the condensed structure of isooctane.

Answer



6. Isooctane (see Exercise 5) is an isomer of what straight-chain alkane?

Answer

octane

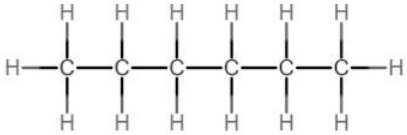
7. Name the other three alkane isomers that contain a five-carbon chain with three methyl substituents.

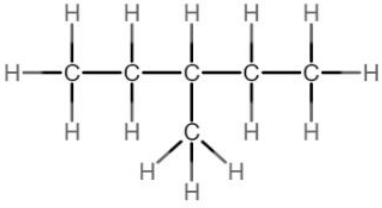
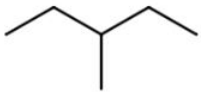
Answer

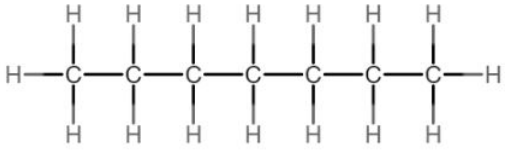
2,2,3-trimethylpentane; 2,3,4-trimethylpentane; 2,3,3-trimethylpentane

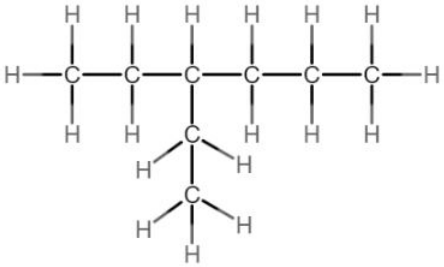
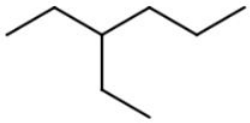
8. Draw the Lewis structure and line structure for each of the following hydrocarbons:
   a. hexane
   b. 3-methylpentane
   c. heptane
   d. 3-ethylhexane

Answer

a.  ; 

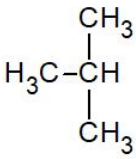
b.  ; 

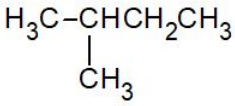
c.  ; 

d.  ; 

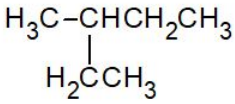
9. Give the complete IUPAC name for each of the following compounds:

a. $CH_3CH_2CH_2CH_3$

b. 

c. 

d. $CH_3CH_2CH_3$

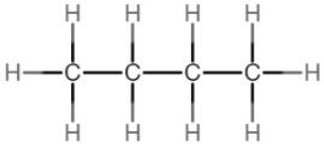
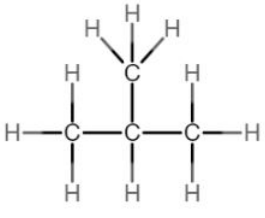
e. 

Answer

a. butane
b. 2-methylpropane
c. 2-methylbutane
d. propane
e. 3-methylpentane

10. Butane is used as a fuel in disposable lighters. Write the Lewis structure for each isomer of butane.

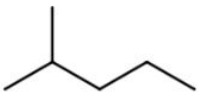
Answer

 ; 

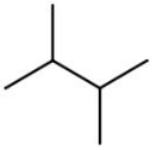
11. Draw line structures of and name the five structural isomers of hexane.

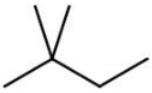
Answer

hexane, 

2-methylpentane, 

3-methylpentane, 

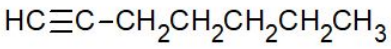
2,3-dimethylbutane, 

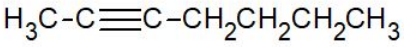
2,2-dimethylbutane, 

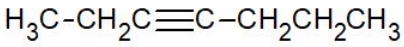
12.5: Alkenes and Alkynes

12. Draw and name all possible normal (that is, straight-chain) isomers of heptyne.

Answer

1-heptyne, 

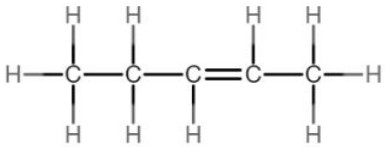
2-heptyne, 

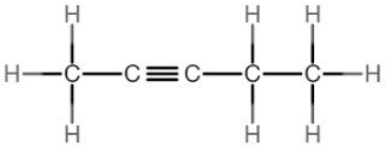
3-heptyne, 

13. Write the chemical formula and a Lewis structure of the following, each of which contains five carbon atoms:

a. an alkene
b. an alkyne

Answer

a. C_5H_{10}; 

b. C_5H_8; 

14. Explain why unbranched alkenes can form geometric isomers while unbranched alkanes cannot.

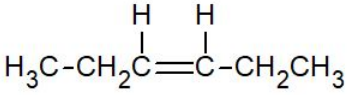
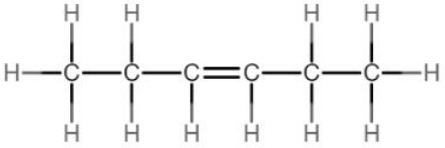
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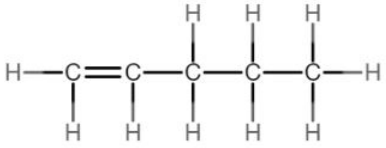
It is very difficult to rotate round a carbon-carbon double bond in an alkene so the location of the groups coming off of it is fixed. There is free rotation about a carbon-carbon single bond in an alkane so the groups coming off of the bond occupy all positions.

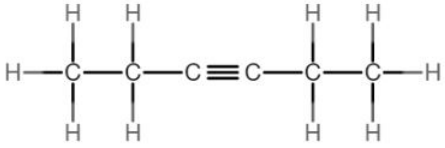
15. Write the chemical formula, condensed formula, and Lewis structure for each of the following hydrocarbons:

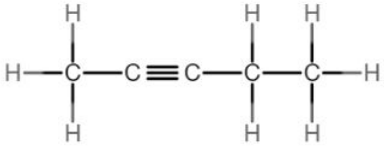
a. *cis*-3-hexene
b. 1-pentene
c. 3-hexyne
d. 2-pentyne
e. *trans*-3-heptene
f. 1-hexene
g. 2-heptyne

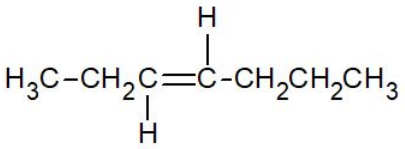
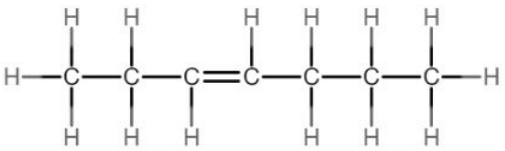
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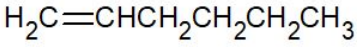
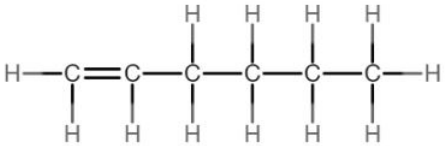
a. C_6H_{12};  ; 

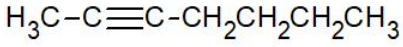
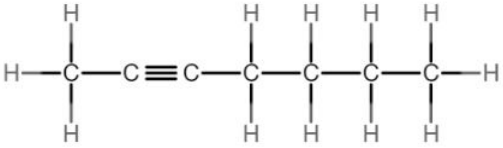
b. C_5H_{10}; $H_2C{=}CHCH_2CH_2CH_3$; 

c. C_6H_{10}; $H_3C\text{-}CH_2C{\equiv}C\text{-}CH_2CH_3$; 

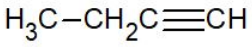
d. C_5H_8; $H_3C\text{-}C{\equiv}C\text{-}CH_2CH_3$; 

e. C_7H_{14};  ; 

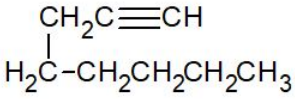
f. C_6H_{12};  ; 

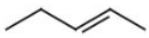
g. C_7H_{12};  ; 

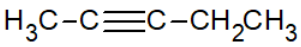
16. Give the complete IUPAC name for each of the following compounds:

a. 

b. $CH_3CH_2CH_2CH{=}CH_2$

c. 

d. 

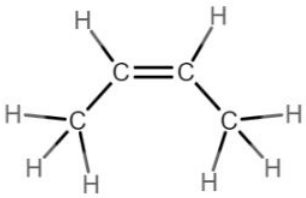
e. 

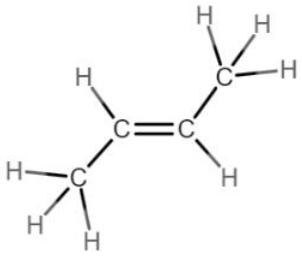
Answer

a. 1-butyne
b. 1-pentene
c. 1-octyne
d. *trans*-2-pentene
e. 2-pentyne

17. Write Lewis structures for the *cis*–*trans* isomers of $CH_3CH{=}CHCH_3$.

Answer

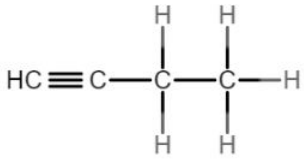
cis: 



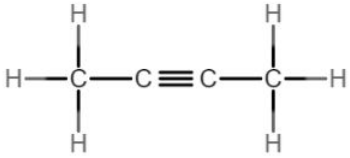
trans:

18. Write Lewis structures and IUPAC names for the alkyne isomers of C_4H_6.

Answer



1-butyne:



2-butyne:

12.6: Oxygen-Containing Organic Compounds

19. Write the condensed structures of both isomers with the formula C_2H_6O. Label the functional group of each isomer.

Answer

CH_3CH_2OH, alcohol group; CH_3OCH_3, ether

20. Write the condensed structures of all isomers with the formula $C_2H_6O_2$. Label the functional group (or groups) of each isomer.

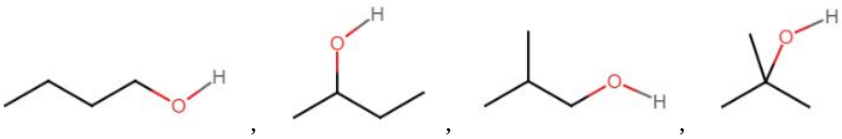
Answer

$HOCH_2CH_2OH$, two alcohol groups; CH_3OCH_2OH, ether and alcohol groups

12.7: Alcohols, Aldehydes, Carboxylic Acids, and Ketones

21. Draw the line structures of all the possible isomers of butanol. Include branched isomers.

Answer



22. Why do the compounds hexane, hexanol, and hexene have such similar names?

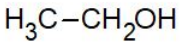
Answer

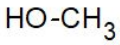
They all have six carbon chains.

23. Write condensed formulas and provide IUPAC names for the following compounds:

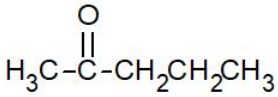
a. ethyl alcohol (in beverages)
b. methyl alcohol (used as a solvent, for example, in shellac)

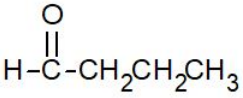
Answer

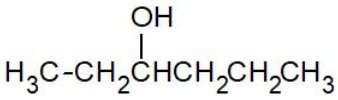

ethanol;

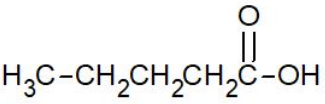

methanol;

24. Give the complete IUPAC name for each of the following compounds:

a.


b.


c.


d.


Answer

a. 2-pentanone
b. butanal
c. 3-hexanol
d. pentanoic acid

25. Explain why it is not possible to prepare a ketone that contains only two carbon atoms.

Answer

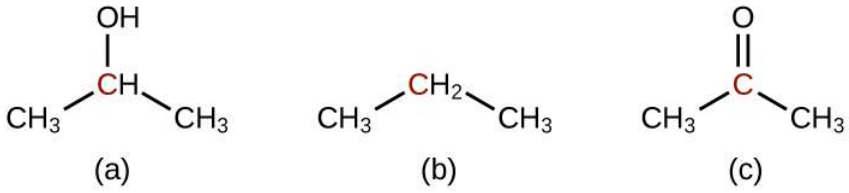
A ketone contains a group bonded to two additional carbon atoms; thus, a minimum of three carbon atoms are needed.

26. Fatty acids are carboxylic acids that have long hydrocarbon chains attached to a carboxylate group. How does a saturated fatty acid differ from an unsaturated fatty acid? How are they similar?

Answer

Since they are both carboxylic acids, they each contain the –COOH functional group and its characteristics. The difference is the hydrocarbon chain in a saturated fatty acid contains no double or triple bonds, whereas the hydrocarbon chain in an unsaturated fatty acid contains one or more multiple bonds.

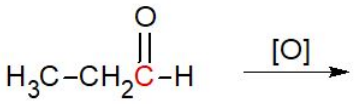
27. Order the following molecules from least to most oxidized, based on the marked carbon atom:



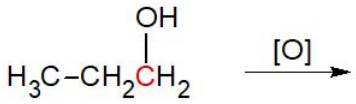
Answer

Most oxidized is c; intermediate is a; least oxidized is b.

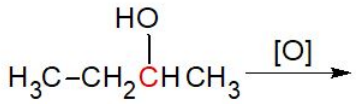
28. Predict the products of oxidizing the molecules shown in this problem. In each case, identify the product that will result from the minimal increase in oxidation state for the highlighted carbon atom:



a.

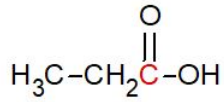


b.

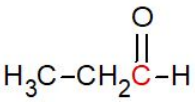


c.

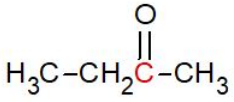
Answer



a.

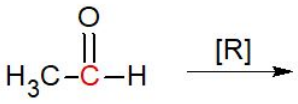


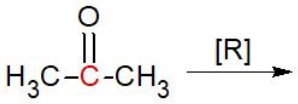
b.

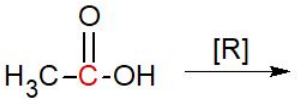


c.

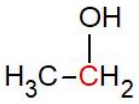
29. Predict the products of reducing the following molecules. In each case, identify the product that will result from the minimal decrease in oxidation state for the highlighted carbon atom:

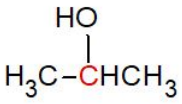
a. 

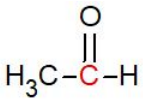
b. 

c. 

Answer

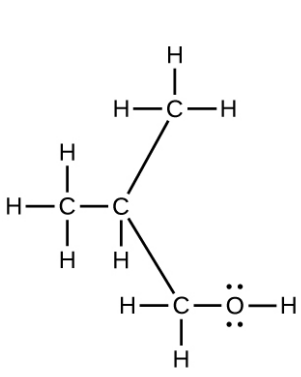
a. 

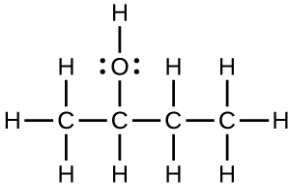
b. 

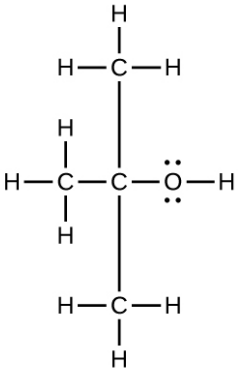
c. 

30. Alcohols A, B, and C all have the composition $C_4H_{10}O$. Molecules of alcohol A contain a branched carbon chain and can be oxidized to an aldehyde; molecules of alcohol B contain a linear carbon chain and can be oxidized to a ketone; and molecules of alcohol C can be oxidized to neither an aldehyde nor a ketone. Write the Lewis structures of these molecules.

Answer


Compound A

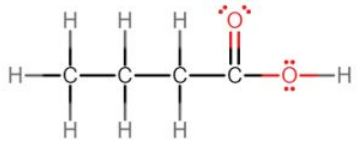

Compound B

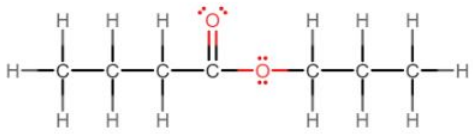

Compound C

12.8: Esters

31. The foul odor of rancid butter is caused by butanoic acid, $CH_3CH_2CH_2CO_2H$, commonly known as butyric acid.
    a. Draw the Lewis structure of butyric acid.
    b. The esters formed from butyric acid are pleasant-smelling compounds found in fruits and used in perfumes. Draw the Lewis structure for the ester formed from the reaction of butyric acid with 1-propanol.

Answer

a. 

b. 

32. Write balanced equations for each of the following reactions using condensed formulas:
    a. ethanol reacts with propanoic acid
    b. 1-butanol reacts with acetic acid (ethanoic acid)

Answer

a. 

b. 

12.9: Ethers

33. Draw the condensed structure of diethyl ether, once used as an anesthetic.

Answer

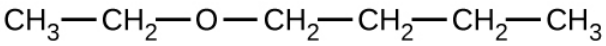
$CH_3CH_2OCH_2CH_3$

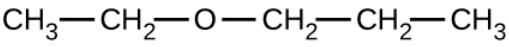
34. Draw the condensed formulas for each of the following compounds:
    a. dipropyl ether
    b. ethyl methyl ether

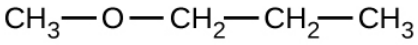
Answer

a. $CH_3CH_2CH_2OCH_2CH_2CH_3$
b. $CH_3CH_2OCH_3$

35. Give the common name for each of the following compounds:

a. 

b. 

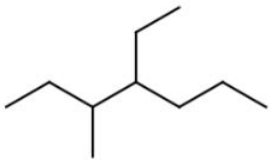
c. 

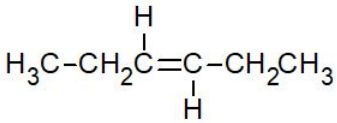
Answer

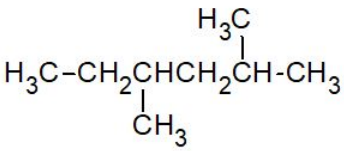
a. butyl ethyl ether
b. ethyl propyl ether
c. methyl propyl ether

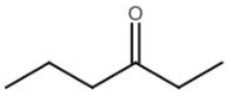
Additional Exercises

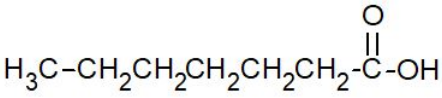
36. Give the IUPAC name for each of the following molecules.

a. 

b. 

c. 

d. 

e. 

Answer

a. 4-ethyl-3-methylheptane
b. *trans*-3-hexene
c. 2,4-dimethylhexane
d. 3-hexanone

e. heptanoic acid

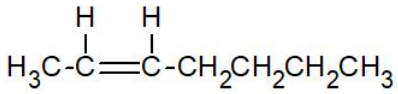
37. Write a condensed structural formula for each of the following molecules.

a. propene
b. 1-butanol
c. ethyl propyl ether
d. *cis*-2-heptene
e. 2,2,3-trimethylhexane
f. methanal (commonly known as formaldehyde)

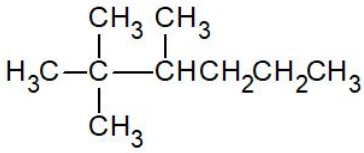
Answer

a. $CH_2=CHCH_3$
b. $CH_3CH_2CH_2CH_2OH$
c. $CH_3CH_2OCH_2CH_2CH_3$

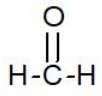
d.



e.



f.

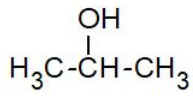


38. Write a condensed structural formula for each of the following molecules.

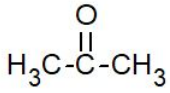
a. 2-propanol
b. acetone (2-propanone)
c. dimethyl ether
d. acetic acid (ethanoic acid)
e. 1-hexene

Answer

a.

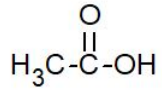


b.



c. CH_3OCH_3

d.



e. $CH_2=CHCH_2CH_2CH_2CH_3$

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CHAPTER OVERVIEW

13: States of Matter

In Chapter 9, we discussed the properties of gases. Here, we consider some properties of liquids and solids. As a review, the Table below lists some general properties of the three phases of matter.

Phase	Shape	Density	Compressibility
Gas	fills entire container	low	high
Liquid	fills a container from bottom to top	high	low
Solid	rigid	high	low

13.1: Phase Properties
13.2: Evaporation and Condensation
13.3: Melting, Freezing, Sublimation, and Deposition
13.4: Energetics of Phase Changes
13.5: Electronegativity and Polarity
13.6: Polarity and Properties
13.7: Intermolecular Forces
13.8: For Future Use
13.9: Exercises

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Thumbnail Chapter 13: Ice water. (hamasakihaus via Giphy)

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13.1: Phase Properties

Learning Objectives

- Describe the solid and liquid phases.

Solids and liquids are collectively called *condensed phases* because their particles are in virtual contact. The two states share little else, however.

Solids

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



Figure **13.1.1**: A crystalline arrangement of a quartz crystal cluster. (JJ Harrison via Wikipedia Commons)

Ionic solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a crystal. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure 13.1.1. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

Liquids

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



Figure **13.1.2**: The formation of a spherical droplet of liquid water minimizes the surface area, a natural result of surface tension in liquids. (José Manuel Suárez via Wikimedia Commons)

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either.

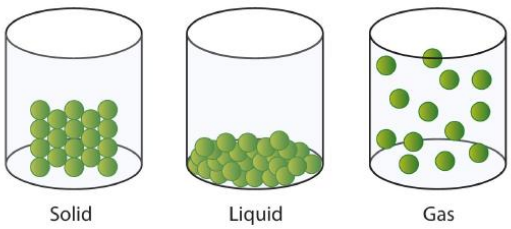


Figure **13.1.3**: A representation of the solid, liquid, and gaseous states. A solid has definite volume and shape. A liquid has a definite volume, but no definite shape. A gas has neither a definite volume nor shape.

The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, usually by a factor of 1000 or more. Figure 13.1.3 shows the differences among solids, liquids, and gases at the molecular level, while Table 13.1.1 lists the different characteristics of these states.

Table 13.1.1: Characteristics of the Three States of Matter

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

✅ Example 13.1.1

What state or states of matter does each statement describe?

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

Solution

A. This statement describes either the liquid state or the solid state.
B. This statement describes either the liquid state or the gas state.
C. This statement describes the liquid state.

✏️ Exercise 13.1.1

What state or states of matter does each statement describe?

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

Answer A

solid

Answer B

gas

Answer C

liquid or gas

Looking Closer: Water, the Most Important Liquid

Earth is the only known body in our solar system that has liquid water existing freely on its surface; life on Earth would not be possible without the presence of liquid water.

Water has several properties that make it a unique substance among substances. It is an excellent solvent; it dissolves many other substances and allows those substances to react when in solution. In fact, water is sometimes called the *universal solvent* because of this ability. Water has unusually high melting and boiling points (0°C and 100°C, respectively) for such a small molecule. The boiling points for similar-sized molecules, such as methane (BP = −162°C) and ammonia (BP = −33°C), are more than 100° lower. Though a liquid at normal temperatures, water molecules experience a relatively strong intermolecular interaction that allows them to maintain the liquid phase at a higher temperature than substances that have a similar molecular mass as water yet have weaker intermolecular forces.

Unlike most substances, the solid form of water is less dense than its liquid form, which allows ice to float on water. In colder weather, lakes and rivers freeze from the top, allowing animals and plants to continue to live underneath. Water also requires an unusually large amount of energy to change temperature. While 100 J of energy will change the temperature of 1 g of Fe by 230°C, this same amount of energy will change the temperature of 1 g of H_2O by only 100°C. Thus, water changes its temperature slowly as heat is added or removed. This has a major impact on weather, as storm systems like hurricanes can be impacted by the amount of heat that ocean water can store.

Water's influence on the world around us is affected by these properties. Isn't it fascinating that such a small molecule can have such a big impact?

Summary

- Solids and liquids are phases that have their own unique properties.

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13.2: Evaporation and Condensation

Learning Objectives

- Explain the processes of vaporization, evaporation, and condensation.

Outside the abandoned barracks shown in Figure 13.2.1 below are devices known as a "swamp coolers". This piece of equipment traces its origin back to the ancient Egyptians who hung wet blankets across the doors of their homes. As the warm air passed through the blankets, water would evaporate and cool the air. The royalty went one step further and had servants fan wet cloths over jugs of water to get more evaporation and cooling.



Figure **13.2.1**: Evaporative coolers (aka "swamp coolers") at the abandoned WWII Army Air Corps base in Wendover, Utah. (arbyreed via Flickr)

The origin of the term "swamp cooler" is not known – they certainly don't work in a swamp. The best conditions for cooling with a swamp cooler include a high temperature (over 80°F) and a low humidity (preferably less than 30%). These coolers work well in desert areas, but don't provide any cooling in the humid areas of the country.

Evaporation

A puddle of water left undisturbed eventually disappears. The liquid molecules escape into the gas phase, becoming water vapor. **Vaporization** is the process by which a liquid is converted to a gas. **Evaporation** is the conversion of a liquid to its vapor below the boiling temperature of the liquid. If the water is instead kept in a closed container, the water vapor molecules do not have a chance to escape into the surroundings and so the water level does not change. As some water molecules become vapor, an equal number of water vapor molecules condense back into the liquid state. **Condensation** is the change of state from a gas to a liquid. The process where the rate of evaporation of a substance is the same as the rate of condensation in a closed container is called a **dynamic equilibrium**.

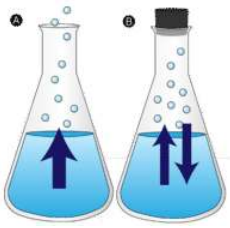


Figure **13.2.2**: The water in an open contain will eventually disappear as it evaporates (A). In a closed container, the rate of evaporation equals the rate of condensation (B).

A given liquid sample will have molecules with a wide range of kinetic energies. In order for a liquid molecule to escape into the gaseous state, the molecule must have enough kinetic energy to overcome the intermolecular attractive forces in the liquid. Liquid molecules that have this certain threshold kinetic energy escape the surface and become vapor. As a result, the liquid molecules that remain now have lower kinetic energy. As evaporation occurs, the temperature of the remaining liquid decreases. You have observed the effects of evaporative cooling. On a hot day, the water molecules in your perspiration absorb body heat and evaporate from the surface of your skin. The evaporating process leaves the remaining perspiration cooler, which in turn absorbs more heat from your body.

A given liquid will evaporate more quickly when it is heated. This is because the heating process results in a greater fraction of the liquid's molecules having the necessary kinetic energy to escape the surface of the liquid. The figure below shows the kinetic energy distribution of liquid molecules at two temperatures. The numbers of molecules that have the required kinetic energy to evaporate are shown in the shaded area under the curve at the right. The higher temperature liquid (T_2) has more molecules that are capable of escaping into the vapor phase than the lower temperature liquid (T_1).

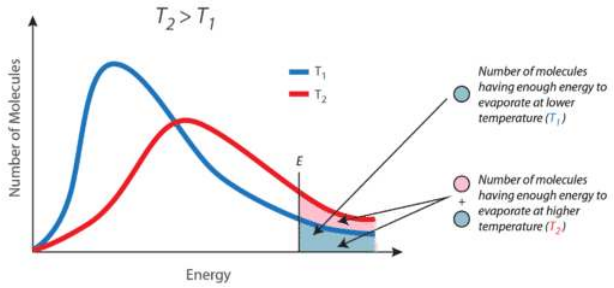


Figure **13.2.3**: Kinetic energy distribution curves for a liquid at two temperatures T_1 and T_2. The shaded area represents the molecules with enough kinetic energy to escape the liquid and become vapor.

Boiling

As a liquid is heated, the average kinetic energy of its particles increases. The rate of evaporation increases as more and more molecules are able to escape the liquid's surface into the vapor phase. Eventually a point is reached when the molecules all throughout the liquid have enough kinetic energy to **vaporize**. At this point the liquid begins to boil. The **boiling point** is the temperature at which the vapor pressure of a liquid is equal to the external pressure. The **vapor pressure** is the pressure exerted by the molecules that escape from the surface of the liquid and increases with an increase in temperature. The figure below illustrates the boiling of liquid.

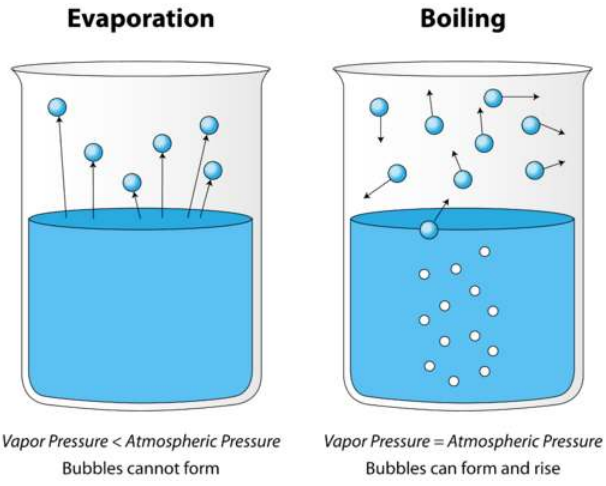


Figure **13.2.4**: Comparison between evaporation and boiling.

In the picture on the left, the liquid is below its boiling point, yet some of the liquid evaporates. On the right, the temperature has been increased until bubbles begin to form in the body of the liquid. When the vapor pressure inside the bubble is equal to the external atmospheric pressure, the bubbles rise to the surface of the liquid and burst. The temperature at which this process occurs is the boiling point of the liquid.

The normal boiling point is the temperature at which the vapor pressure of the liquid is equal to standard pressure. Because atmospheric pressure can change based on location, the boiling point of a liquid changes with the external pressure. The normal boiling point is a constant because it is defined relative to the standard atmospheric pressure of 760 mm Hg (or 1 atm or 101.325 kPa).

At 29,029 feet (8848 m), Mount Everest in the Himalayan range on the border between China and Nepal is the highest point on the earth. Its altitude presents many practical problems to climbers. The oxygen content of the air is much lower than at sea level, making it necessary to bring oxygen tanks along (although a few climbers have reached the peak without oxygen). One other problem is that of boiling water for cooking food. Although water boils at 100°C at sea level, the boiling point on top of Mount Everest is only about 70°C. This difference makes it very difficult to get a decent cup of tea (which definitely frustrated some of the British climbers).

Is it possible to boil water without heating? Check out the videos below to see how!



Video **13.2.1**: Boiling water by reducing the pressure.



Video **13.2.2**: Boiling water using ice.

✅ Example 13.2.1

What are two ways to cause water to boil?

Solution

Since the boiling point is defined as the temperature at which the vapor pressure of a liquid is equal to the external pressure, it is is possible to boil a liquid either by increasing its temperature or by reducing the pressure above the liquid.

✏️ Exercise 13.2.1

A. What is the name given to the process of a liquid being converted to a gas when the temperature of the liquid is below its boiling point?

B. What is the name of the phase change given to a gas or vapor being converted to a liquid?

Answer A

evaporation

Answer B

condensation

Summary

- The boiling point is the temperature at which the vapor pressure of a liquid is equal to the external pressure.
- As the altitude increases, the boiling point decreases.
- Evaporation is the conversion of a liquid to its vapor below the boiling temperature of the liquid.
- Condensation is the change of state from a gas to a liquid.
- As the temperature increases, the rate of evaporation increases.

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13.3: Melting, Freezing, Sublimation, and Deposition

Learning Objectives

- Explain the processes of melting, freezing, sublimation, and deposition.

In the winter, many people find the snow and ice beautiful; they enjoy getting out to ski or ice-skate. When the snow melts, however, the roads get very sloppy and messy. Some people look forward to spring, when the ice and snow are gone and the weather is warmer. All of these events and factors are dependent on the melting point of a solid and the freezing point of a liquid.

Melting and Freezing

Solids are similar to liquids in that both are condensed states, with particles that are far closer together than those of a gas. However, while liquids are fluid, solids are not. The particles of most solids are packed tightly together in an orderly arrangement. The motion of individual atoms, ions, or molecules in a solid is restricted to **vibrational motion** about a fixed point. Solids are almost completely **incompressible** and are the densest of the three states of matter.

As a solid is heated, its particles vibrate more rapidly as it absorbs kinetic energy. Eventually, the organization of the particles within the solid structure begins to break down and the solid starts to melt. The **melting point** is the temperature at which a solid changes into a liquid. At its melting point, the disruptive vibrations of the particles of the solid overcome the attractive forces operating within the solid.

As with boiling points, the melting point of a solid is dependent on the strength of those attractive forces. Sodium chloride, NaCl, is an ionic compound that consists of a multitude of strong ionic bonds. Sodium chloride melts at 801°C. Ice (solid H_2O) is a molecular compound in which molecules that are held together by an attractive force called a hydrogen bond. Though hydrogen bonds are the strongest of the intermolecular attractive forces, the strength of hydrogen bonds is much less than that of ionic bonds. The melting point of ice is 0°C.



Figure **13.3.1**: Melting ice cubes illustrate the process of this phase transition. (Public Domain; Moussa).

The melting point of a solid is the same temperature as the **freezing point** of its liquid. They are simply approached from opposite directions. When a substance is at its melting point or freezing point, the solid and liquid states of that substance are present in equilibrium. For water, this equilibrium occurs at 0°C.

$$\mathrm{H_2O}\ (s) \rightleftharpoons \mathrm{H_2O}\ (l)$$

We tend to think of solids as those materials that are solid at room temperature. However, all materials have melting points of some sort. Substances that are gases at room temperature eventually become solids at extremely low temperatures, as do liquids. Table 13.3.1 shows the melting points of several common materials.

Table 13.3.1: Melting Points of Common Materials

Materials	Melting Point (°C)
Hydrogen	–259
Oxygen	–219
Diethyl Ether	–116
Ethanol	–114
Water	0
Silver	961
Gold	1063
Iron	1538

Sublimation and Deposition

There is also a phase change where a solid goes directly to a gas, a process called **sublimation**:

$$\text{solid} \rightarrow \text{gas}$$

We encounter sublimation in several ways. You may already be familiar with dry ice, which is simply solid carbon dioxide (CO_2) (see Video 13.3.1). At −78.5°C (−109°F), solid carbon dioxide sublimes, changing directly from the solid phase to the gas phase:

$$CO_2\,(s) \xrightarrow{-78.5^\circ C} CO_2\,(g)$$

Solid carbon dioxide is called dry ice because it does not pass through the liquid phase. Instead, it goes directly to the gas phase. (Carbon dioxide *can* exist as liquid but only under high pressure.) Dry ice has many practical uses, including the long-term preservation of medical samples.

Even at temperatures below 0°C, solid H_2O will slowly sublime. For example, a thin layer of snow or frost on the ground may slowly disappear as the solid H_2O sublimes, even though the outside temperature may be below the freezing point of water. Similarly, ice cubes in a freezer usually get smaller over time. Although frozen, the solid water slowly sublimes, redepositing on the colder cooling elements of the freezer, which necessitates periodic defrosting (frost-free freezers minimize this redeposition). Lowering the temperature in a freezer will reduce the need to defrost as often.

Under similar circumstances, water will also sublime from frozen foods (e.g., meats or vegetables), giving them an unattractive, mottled appearance called freezer burn. It is not really a “burn,” and the food has not necessarily gone bad, although it looks unappetizing. Freezer burn can be minimized by lowering a freezer’s temperature and by wrapping foods tightly so water does not have any space to sublime into.



Video **13.3.1**: Solid CO_2, also known as dry ice, goes directly from a solid phase to a gaseous phase in a process called sublimation.

The reverse process of sublimation is called **deposition**, a process in which a gas goes directly to a solid, bypassing the liquid phase. Deposition is involved in the formation of frost (see Video 13.3.2). Deposition may be observed in a freezer in a humid environment, such as a freezer kept in a garage during the summer or on cold metal surfaces and single pane windows in the middle of winter, particularly if they are exposed to relatively high humidity.



Video **13.3.2**: Timelapse of frost forming on rose (Rosa) hips, in which water vapor goes directly from a gaseous phase to a solid phase in a process called deposition.

Exercise 13.3.1

A. How is the freezing point of a substance related to its melting point?

B. What is the name given to the phase change when a solid is converted directly to a gas?

Answer A

The freezing point and melting point occur at the same temperature for a substance, but are approached from opposite directions. Freezing occurs when a substance changes from a liquid to a solid, while melting occurs when a substance changes from a solid to a liquid.

Answer B

Sublimation.

Summary

- There is an energy change associated with any phase change.
- Sublimation is the change of state from a solid to a gas, without passing through the liquid state.
- Deposition is the change of state from a gas to a solid.
- Carbon dioxide is an example of a material that easily undergoes sublimation.
- The melting point is the temperature at which a solid changes into a liquid.
- Intermolecular attractive forces have a strong influence on melting point.

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13.4: Energetics of Phase Changes

Let's assume we have a block of ice, H_2O (*s*), at –10°C and 1 atm of pressure, and we begin heating it. We would monitor the temperature and continue heating until the ice first turns into liquid water, H_2O (*l*), and with enough heat, eventually to steam, H_2O (*g*). If we were to make a plot of temperature vs. time, what would the graph look like? Is the plot linear? What would we observe at the melting point of water, 0°C, or at its boiling point at 100°C. Would the temperature shoot up? Would it level off?

Heating Curves

As heat is steadily added to the ice block, the water molecules begin to vibrate faster and faster as they absorb kinetic energy. Eventually, when the ice has warmed to 0°C, the added energy will start to overcome the attractive intermolecular forces that hold the water molecules in place while in its **solid** form. As the ice melts, its temperature does not rise. All of the energy that is being put into the ice goes into the melting process and not into any increase in temperature. During the melting process, the two states – solid and **liquid** – are in equilibrium with one another. If the system was isolated at that point and no energy was allowed to enter or leave, the ice-water mixture at 0°C would remain. Temperature is always constant during a change of state.

After the ice has completely melted, continued heating of the water will now increase the kinetic energy of the liquid molecules and the temperature will, once again, rise. Assuming an atmospheric pressure of 1 atm is maintained, the temperature will rise steadily until it reaches 100°C. At this point, the added energy from the heat will cause the liquid to begin to vaporize. As with the previous phase change, the temperature will remain at 100°C as the water molecules overcome the remaining intermolecular forces that hold them together in the liquid phase as they transition to the **gaseous** phase. Once all the liquid has completely boiled away, continued heating of the steam will allow the temperature to resume increasing, this time at temperatures above 100°C.

The experiment described above can be summarized in a graph called a heating curve, as shown in Figure 13.4.1:

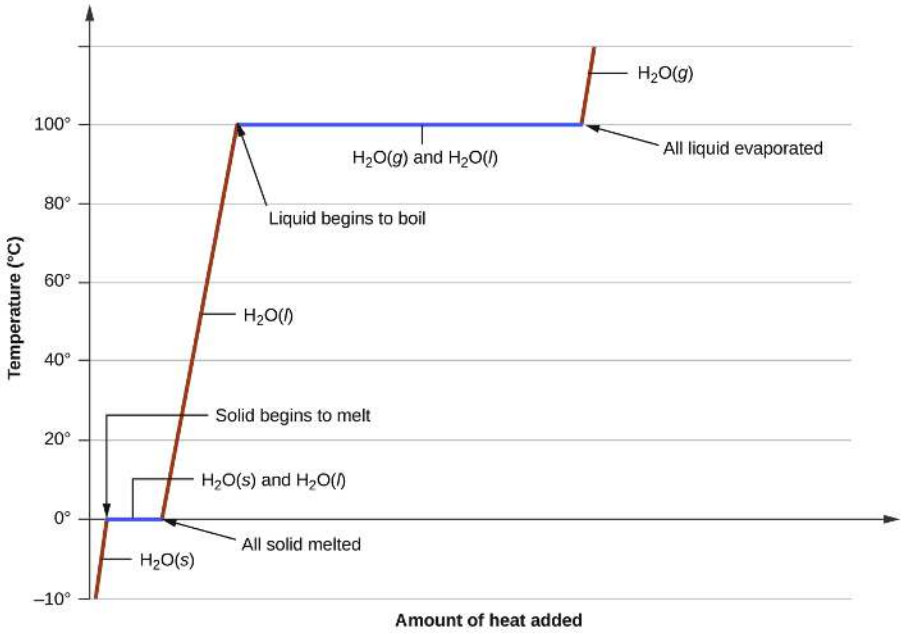


Figure 13.4.1: The heating curve for water at 1 atm. Temperature plateaus in the curve are due to phase transitions.

Phase Change Energetics

The previous section described the phase transitions that took place heating water, causing it to change from a solid to a gas. The addition of heat energy to a system from its surroundings is an **endothermic** process. In other words, ice absorbs energy from the surroundings in order to transition into liquid. Likewise, liquid water absorbs energy from the surroundings in order to transition

into a gas. This means that phase transitions in the opposite direction are **exothermic** – they release energy to the surroundings. Converting steam to liquid water is exothermic, as is converting liquid water to ice.

The energetics of phase transitions are summarized in Figure 13.4.2 below. The gaseous phase of a substance is the one with the highest energy, while the solid phase of a substance is the one with the lowest energy. In order to move to a higher energy phase, energy must be absorbed from the surroundings (endothermic). In order to move to a lower energy phase, energy must be released to the surroundings (exothermic).

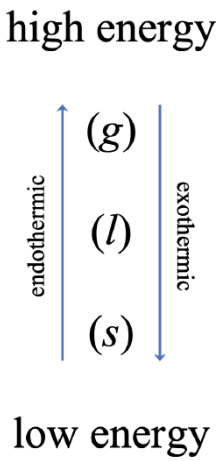


Figure **13.4.2**: The energetics of phase transitions.

Applications of Phase Changes

Many plants are damaged when they experience frost or a hard freeze. In northern climates, a late spring frost can damage young plants or the buds or blossoms on fruit trees. In southern climates, where temperatures rarely go below 0°C (32°F), a frost may be particularly damaging to the fruits and trees.

It likely seems ironic, but one of the ways that growers protect plants and trees from damage due to frost or a hard freeze is by spraying them with water (see Figure 13.4.3) while the temperatures approach and go below the freezing point. Ice begins to accumulate on the trees. How can this be a good thing? As we can see from the graphic above, the freezing is an *exothermic* process – it releases energy to the surroundings. In other words, heat energy is released to the surrounding buds, flowers, foliage, or fruit as long as water continues to freeze. Furthermore, as a layer of ice forms, it acts as an insulation from the freezing temperatures. Cool, huh?



Figure **13.4.3** (left to right): Frost protection by spraying the buds and blossoms of apple trees with water. (Haneburger via Wikimedia Commons); Freeze protection by spraying citrus trees with water (weatherguyd via Flickr)

An application of phase change energetics that you have likely experienced is sweating. When humans sweat, sweat droplets and/or a thin layer layer of sweat covers the skin. As the water in the sweat evaporates, the faster molecules are able to overcome the intermolecular attractions that hold them together, leaving behind the slower moving molecules. We observe slower moving molecules as being cooler. Additionally, from the phase change energetics graphic in Figure 13.4.2 above, we see that the transition from a liquid into a gas is an *endothermic* process. In other words, heat is absorbed from the surroundings – in this case, the surroundings are our skin and bodies. The effect is to cool us off. This is also why pigs and other animals roll in mud or water on a warm and sunny day. As the water evaporates, it serves to keep them cool.

Figure 13.4.4 shows how a central air conditioner works. Outside of the house sits the central air unit. If you've ever stood outside and felt the air coming off an air conditioner, whether it is a central air unit or a window unit, you will notice that the air feels warm or even hot. That's because there is a hose or pipe that carries a substance in the gaseous phase from the inside of the house to the outside where it goes through the process of condensation in the compressor. Referring to the graphic in Figure 13.4.2 once again, we can see that the the transition from a gas to a liquid is *exothermic* – heat is released to the surroundings.

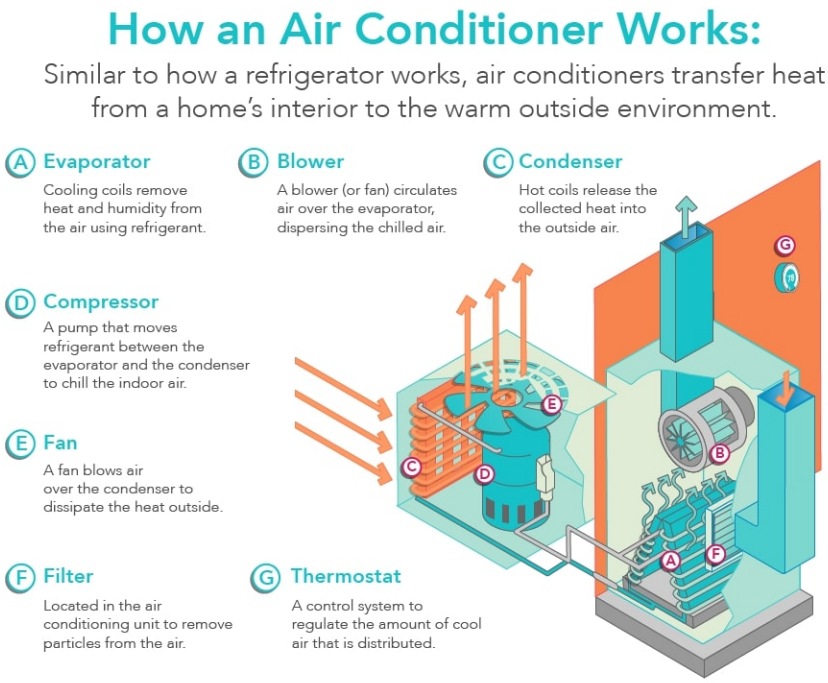


Figure **13.4.4**: Central air conditioning condenser and coil diagram. (via U.S. Department of Energy)

From here, the liquid is sent back into the house where it goes through the process of evaporation in the evaporator. The process of evaporation, the transition from a liquid to a gas, is *endothermic* – heat is absorbed from the surroundings. In this case, the surroundings are the inside of your house. A fan blows air from the inside of the house over the coils in the evaporator to make the process more efficient.

Refrigerators and freezers work in a similar manner to air conditioners. The compressor sits outside the insulated portion of the refrigerator or freezer, which is where the condensation takes place and why you feel warm air coming out from underneath or behind the unit. In most refrigerators and freezers, the evaporator coils remain hidden in the walls or shelves within the insulated portion of the unit. This is where the evaporation takes place and why it feels cool inside the unit.

✅ Example 13.4.1

Which phase changes are *endothermic*?

Solution

Referring to Figure 13.4.2, any phase change to a state of higher energy is *endothermic*, i.e. it absorbs energy from the surroundings. The phase changes include:

- melting (solid to liquid)
- boiling/evaporation (liquid to gas)
- sublimation (solid to gas)

Exercise 13.4.1

What happens to the temperature during a phase change when a substance is being heated? Explain why.

Answer

The temperature remains constant, because the energy is being used to overcome intermolecular forces that:

- hold the particles in a solid in fixed positions when a substance is melting.
- hold the particles in a liquid in close contact when a substance is evaporating or boiling.

Summary

- The temperature of a substance remains constant while a phase change occurs.
- Endothermic processes go from low energy to high energy: solid to liquid to gas.
- Exothermic processes go from high energy to low energy: gas to liquid to solid.
- Energetics of phase transitions have multiple real-world applications in botany and cooling our homes and food.

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13.5: Electronegativity and Polarity

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.

We learned in Section 11.3 that when a metal and a nonmetal come together, the metal donates its valence electrons to the nonmetal, forming ions which are held together by an ionic bond. However, when two nonmetals come together, they will share electrons with each other to form covalent bonds as we learned in Section 11.4. When describing a covalent bond, the implication was that the two electrons in the bond were shared equally between the two nuclei involved. While this is sometimes true, it seems unlikely that *all* nonmetal nuclei have identical attractiveness for the electrons. How can we tell if there is a difference and what would a difference do to the bond?

Electronegativity

The ability of an atom in a molecule to attract electrons is called **electronegativity**. Scientists have developed an electronegativity scale, as shown in Figure 13.5.1, to reflect the relative abilities of different atoms to attract the electrons. Electronegativity is a dimensionless number; the greater the electronegativity value, the greater the attraction for shared electrons.

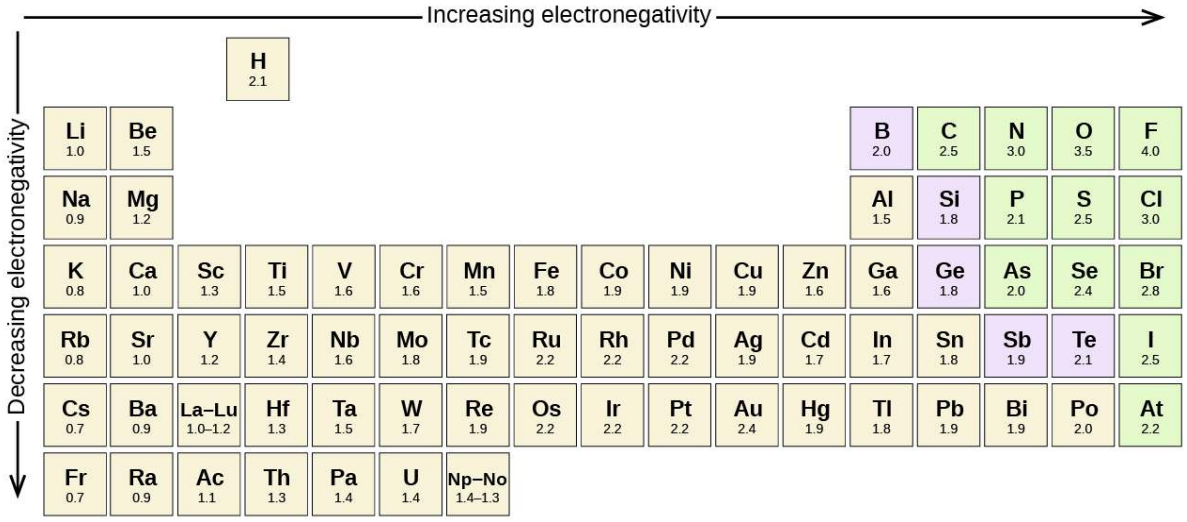


Figure **13.5.1**: Electronegativities of the elements. Electronegativities are used to determine the polarity of covalent bonds. An interactive version of this table may be found here.

The type of bond formed by two atoms can be determined by calculating the *difference* in electronegativity values of the two atoms involved in the bond, as summarized in Table 13.5.1:

Table 13.5.1: Bond Polarity

Electronegativity Difference	Bond Type
<0.4	nonpolar covalent
0.4 – 1.9	polar covalent
≥2.0	likely ionic

If the difference between the electronegativities of the two atoms is small, neither atom will take the shared electrons completely away from the other atom, and the bond will be covalent. This is the case when two nonmetals are bonded to each other. If the difference between the electronegativities is large, the less electronegative atom loses the bonding electrons to the more electronegative 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.

It should also be noted that the cutoffs for the difference in electronegativity are somewhat arbitrary. Even if the difference in electronegativity was only 0.2, there were still be a hint of polar character. Electrons would not be shared entirely equally. Likewise, if the difference in electronegativity was 1.9, the bond may be defined as polar covalent, but it would also have some ionic characteristics. There wouldn't be anything "magic" that happens once that difference in

electronegativity reaches 2.0. We may say it is likely ionic, but it would still have some covalent characteristics. The greater the difference becomes, the greater the ionic character.

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 2.0 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, sometimes simply called a **nonpolar bond**, the distribution of electrical charge is balanced between the two atoms.

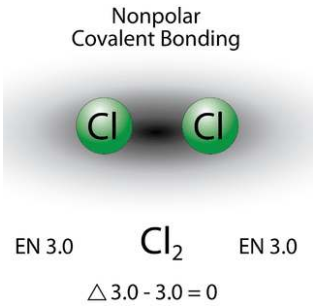


Figure **13.5.2**: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal. The Δ symbol stands for "difference."

The two chlorine atoms in Figure 13.5.2 share the pair of electrons in the single covalent bond equally, and the electron density surrounding the Cl_2 molecule is symmetrical. This is the case for any bond in which the electronegativity difference is very small (<0.4). An example would be a bond between chlorine and bromine (ΔEN = 3.0 – 2.8 = 0.2). Any shifts in electron density are too small to be noticed and the electron cloud around the nuclei is essentially symmetrical. Thus the bond is considered to be a nonpolar bond.

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.4 to 1.9 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons, so the electrons are shared unequally. In a polar covalent bond, sometimes simply called a **polar bond**, the distribution of electrons around the molecule is no longer symmetrical.



Figure **13.5.3**: In the polar bond of HF, electrons are shared unequally. There is a lower density (blue) around the hydrogen atom (EN=2.1) and a higher density (red) around the fluorine atom (EN=4.0).

One way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta, δ, as shown in Figure 13.5.4 below. The atom with the greater electronegativity, fluorine (EN=4.0) acquires a partial negative charge (δ–), since it has a higher density of negatively charged electrons surrounding it, while the atom with the lesser electronegativity, hydrogen (EN=2.1), acquires a partial positive charge (δ+), since it has a lower density of negatively charged electrons surrounding it. The delta symbol is used to indicate that the quantity of charge is less than one.

A crossed arrow, called a dipole moment vector, may also be used to indicate the direction of greater electron density. The arrow points toward fluorine, since it has a greater electronegativity than hydrogen.



Figure **13.5.4** (left to right): δ used to indicate partial charge; a crossed arrow used to indicate polarity.

✅ Example 13.5.1: Bond Polarity

What is the polarity of each bond?

A. N—Cl
B. O—H

Solution

Using Figure 13.5.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

A. For the N—Cl bond, the difference in the electronegativities is 3.0 − 3.0 = 0.0. Thus we predict that this bond will be nonpolar.
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 polar.

✏️ Exercise 13.5.1

What is the polarity of each bond?

A. Rb—F
B. P—Cl

Answer A

likely ionic

Answer B

polar

Molecular Polarity

In addition to determining whether a specific bond is polar or nonpolar, we can also consider how all of the bonds in a molecule combine to create polar and nonpolar molecules. To determine if a molecule is polar or nonpolar, it is generally 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 and/or bonded to atoms that have different electronegativities. This distinction 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 should know it's three-dimensional structure and/or have a strong grasp of Lewis structures and VSEPR theory. Assuming that you do, you can usually look at the structure of each one and decide if a molecule is polar or nonpolar, *whether or not you know the individual atom's electronegativity*. This is because the bonds between dissimilar elements are usually 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 partially positive (δ+), while the other end is partially negative (δ–), as shown in Figure 13.5.5 below. A two-atom molecule that consists of a polar covalent bond, such as HF (shown in Figure 13.5.4 above), 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**. Using this logic, hydrogen fluoride may also be called a dipole.

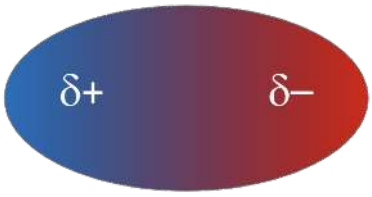


Figure **13.5.5**: A polar molecule, also called a dipole, is a molecule that has a partially positive charge on one end (blue) and a partially negative charge on the other end (red).

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. Figure 13.5.6 below shows a comparison between carbon dioxide and water. Carbon dioxide, CO_2, is a linear molecule. The oxygen atoms (EN=3.5) are more electronegative than the carbon atom (EN=2.5), so there are two individual arrows pointing outward from the central carbon atom to the oxygen atoms on each end. However, since the arrows are of equal "strength" (the electronegativity difference is the same) and are oriented in opposite directions, they cancel each other out and the overall molecular polarity of CO_2 is zero. This will always be the case for molecules in which all sides around the central atom are identical (symmetric molecules).

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the hydrogen atoms (EN=2.1) toward the central oxygen atom (EN=3.5). Because of the shape, the arrows do not cancel each other out; there is a net upward direction of the arrows and the water molecule is polar. In the figure below, the net dipole is shown by the arrow for the overall polarity and points upward. The inability of the arrows to cancel out is what causes

molecules that are asymmetrical about the central atom to be polar.

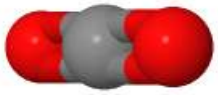
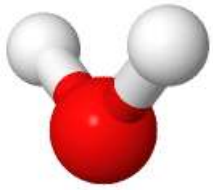
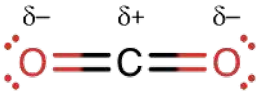
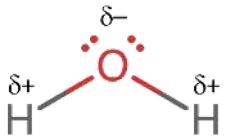
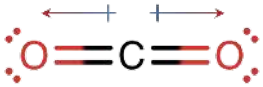
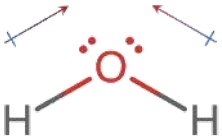
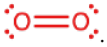
3D Model		
Partial Charges		
Dipoles		
Overall Polarity	arrows cancel	
Molecular Polarity	nonpolar molecule	polar molecule

Figure **13.5.6**: The molecular geometry of a molecule affects its polarity. CO_2 is a nonpolar molecule, while H_2O is a polar molecule.

It should be noted that if the all of the bonds in a molecule are nonpolar, then the molecule will be nonpolar. Consider a molecule of O_2,  . The double bond between the two oxygen atoms is a nonpolar bond, since the electronegativity difference between the two atoms is zero. Since the only bond in the molecule is nonpolar, this means that a molecule of O_2 is nonpolar as well. There are no arrows that can be drawn toward a more electronegative atom.

☞ Steps to Identify Molecular Polarity

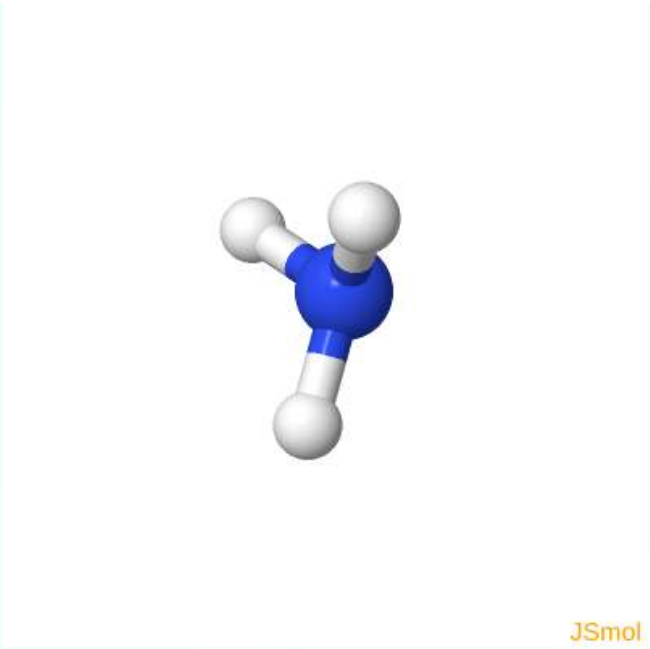
To summarize, a polar molecule must:

1. contain at least one polar bond.
2. have a molecular structure such that there is a net direction of the arrows. The strength and directions of the arrows do not cancel.

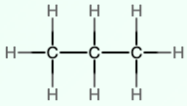
✅ Example 13.5.2: Molecular Polarity

Label each molecule as polar or nonpolar. Manipulate molecules rendered in JSmol, if necessary, to get a better idea of their three dimensional geometry.

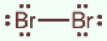
A. Ammonia:



B. Propane, C_3H_8:

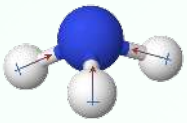


C. Bromine, Br_2:



Solution

A. Ammonia is polar. When properly rotated, one can see that the nitrogen atom (EN=3.0) sits on a different plane than the hydrogen atoms (EN=2.1). All of the arrows point in the direction of the nitrogen atom, giving it a net direction (the strength and direction of the arrows do not cancel).



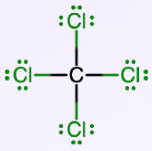
B. Propane is nonpolar. This is a symmetric molecule. All of the arrows would point inward from the hydrogen atoms (EN=2.1) toward the carbon atoms (EN=2.5). The net result is that the arrows cancel each other out. In fact, this is a property of hydrocarbons – they are nonpolar molecules.

C. Bromine is nonpolar. The molecule has only one Br—Br bond in which the EN=2.8 for each Br atom, so each Br atom exerts a pull on the electrons by exactly the same amount.

Exercise 13.5.2

Label each molecule as polar or nonpolar.

A. CCl_4:



B. HCl: 

Answer A

nonpolar

Answer B

polar

Summary

- Differences in electronegativity between two atoms can be used to determine if their bond is nonpolar, polar, or ionic.
- Nonpolar covalent bonds have an equal distribution of electron density between the two nuclei.
- Polar covalent bonds have an unequal distribution of electron density with the more electronegative atom having greater electron density.
- Nonpolar molecules are formed when bond dipoles cancel each other out. This happens when the molecule is symmetrical about the central atom.
- Polar molecules are formed when bond dipoles do not cancel each other out. This occurs when the molecule is asymmetrical and contains at least one polar bond.

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13.6: Polarity and Properties

We saw in the previous section that molecules can be classified as polar or nonpolar, depending on the types of bonds present in the molecule and its overall molecular geometry. The major difference between the two types of molecules was the presence of partially positive and partially negative regions in the polar molecule. These charged regions will cause polar molecules to have very different properties than nonpolar molecules.

Properties of Polar Molecules

Polar molecules tend to align themselves 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 13.6.1). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted.

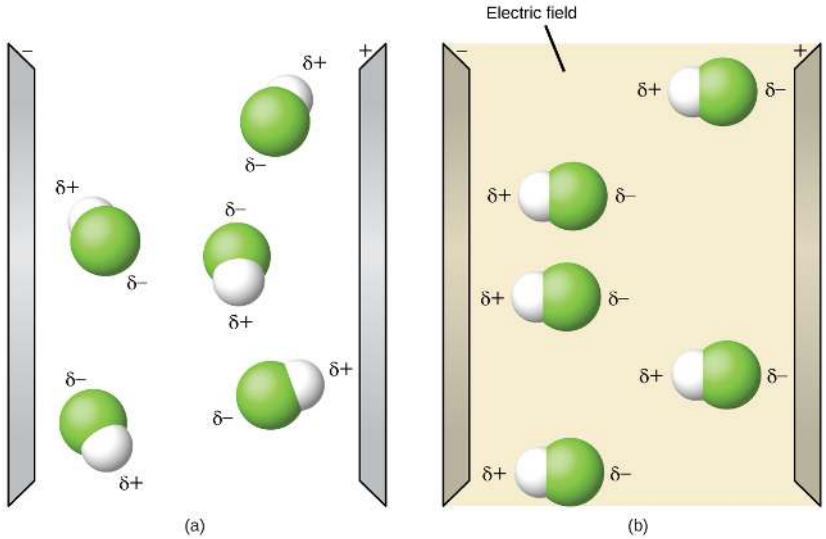


Figure 13.6.1: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA)

In the same way that polar molecules are attracted to the positive and negative plates in an electric field, they can be attracted to the positive and negative regions of neighboring polar molecules. Because of this, 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

It should be noted that while molecules can be described as *polar*, this is often a relative term, with one molecule simply being *more polar* or *less polar* than another.

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 substances comprised of nonpolar molecules tend to dissolve other substances comprised of nonpolar molecules. Substances comprised of polar molecules tend to dissolve other substances comprised of polar molecules. In addition, substances comprised of highly polar molecules, like water, often dissolve ionic compounds. Again, this can be loosely attributed to the presence or absence of charged regions in the molecule. Polar molecules will be attracted to the charged regions in other polar molecules, causing them to interact with each other and form a solution. In contrast, when polar and nonpolar molecules are mixed together, the lack of charged regions in the nonpolar molecules will cause the polar molecules to only be attracted to themselves, excluding the nonpolar molecules and creating two distinct regions instead of a solution. When two nonpolar molecules are mixed together, neither requires a charged region for attraction and they are thus able to mix freely and form a solution.

Some substances comprised of nonpolar molecules will dissolve in water, but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, a nonpolar molecule, dissolves to a small degree in water – it is this oxygen that the fish take in through their gills. The reason we can enjoy carbonated sodas is also due to a nonpolar compound that dissolves in water. 7-Up and all 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 concept of "like dissolves like" is summarized in Table 13.6.1:

Table 13.6.1: Summary of Solubilities

Polarity of Solute	Polarity of Solvent	Dominant Intermolecular Force	Is Solution Formed?
polar	polar	dipole-dipole force and/or hydrogen bond	yes
nonpolar	nonpolar	dispersion force	yes
polar	nonpolar	not applicable	no
nonpolar	polar	not applicable	no
ionic	polar	ion-dipole	usually
ionic	nonpolar	not applicable	no

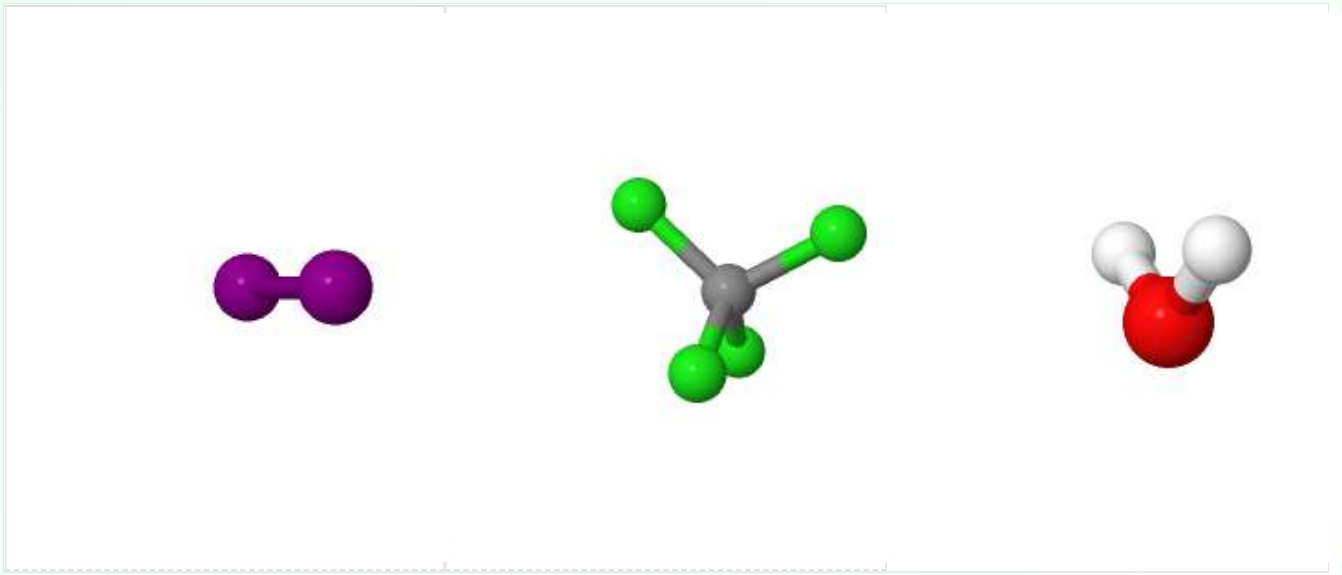
A more precise explanation for what makes a solute soluble in some solvents, but not others, lies in intermolecular interactions. The intermolecular interactions include dispersion forces, dipole-dipole interactions, and hydrogen bonding (as described in the next section). From experimental studies, it has been determined that if molecules of a solute experience similar intermolecular forces to the solvent, the solute will likely dissolve in that solvent. So, NaCl, an ionic compound, dissolves in water since the ions are able to interact with the partially charged regions of polar water molecules, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water.



Figure 13.6.2: Polar molecules of water (clear liquid) and nonpolar molecules of oil (yellow liquid) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

✅ Example 13.6.1: Polar and Nonpolar Solvents

Would I_2 be more soluble in CCl_4 or H_2O? Explain your answer. If necessary, manipulate their molecular models below.

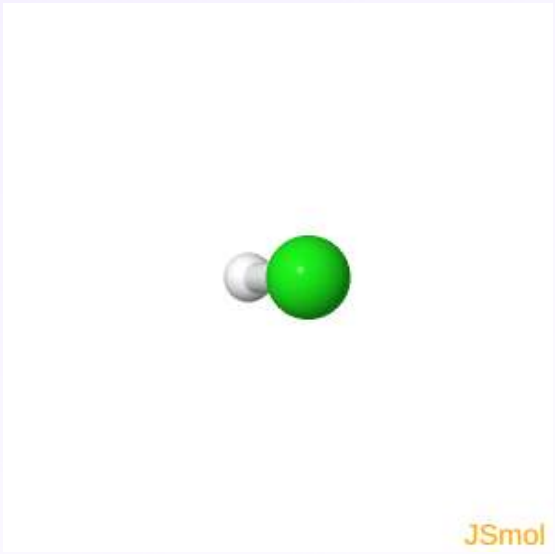


Solution

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 13.6.1

Would HCl (see the molecular model below) be more soluble in CCl_4 or H_2O? Explain your answer.



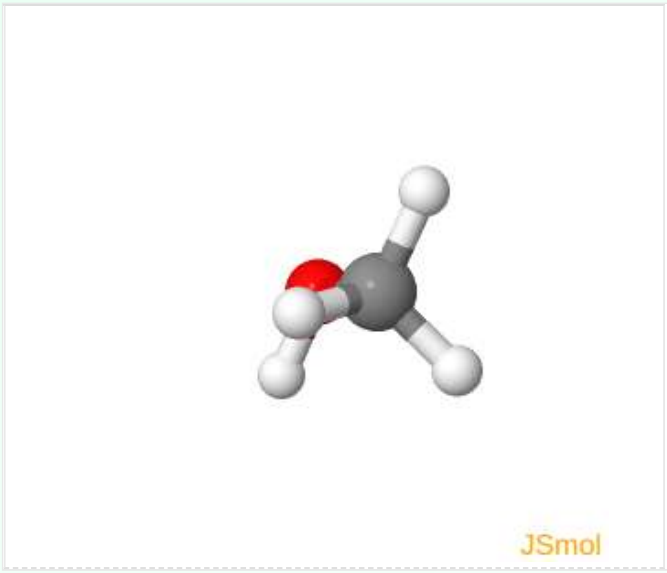
Answer

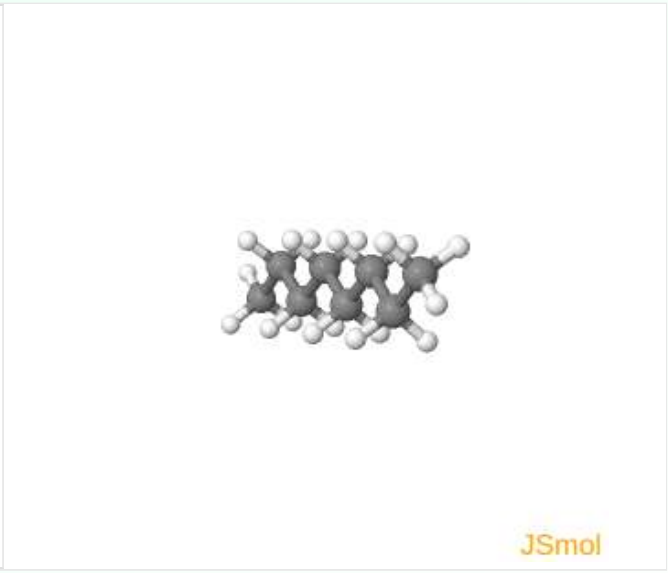
H_2O, because HCl and H_2O are both polar molecules. CCl_4 is a nonpolar molecule.

✅ Example 13.6.2

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

A. methanol (CH_3OH), shown below
B. sodium sulfate (Na_2SO_4)
C. octane (C_8H_{18}), shown below





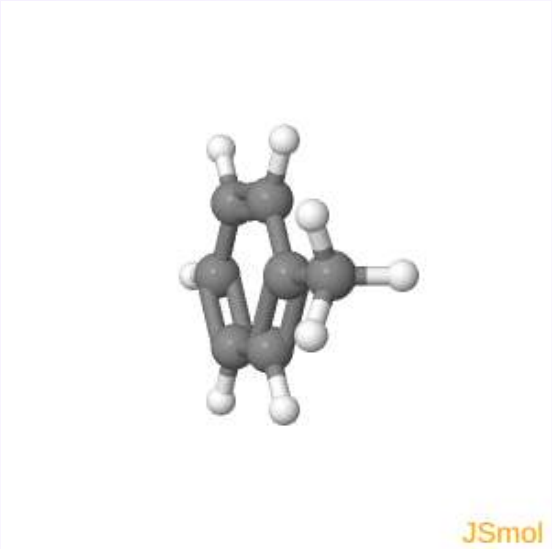
Solution

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

A. Because of the shape of the molecule and the polar —OH grouping in methanol, we expect its molecules to be polar and for it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
B. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
C. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

✏️ Exercise 13.6.2

Toluene, $C_6H_5CH_3$, (see the molecular model below) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?



A. water (H_2O)
B. sodium sulfate (Na_2SO_4)
C. octane (C_8H_{18})

Answer

Water (H_2O) is polar, sodium sulfate (Na_2SO_4) is ionic, and octane (C_8H_{18}) is nonpolar. Therefore, octane is the only substance that should dissolve in toluene.

Summary

- Polar molecules tend to have higher melting points and boiling points than nonpolar molecules.
- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.

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13.7: Intermolecular Forces

Learning Objectives

- To describe the intermolecular forces in molecules.

The properties of liquids are intermediate between those of gases and solids, but are more similar to solids. In contrast to ***intra*molecular forces**, such as the covalent bonds that hold atoms together in molecules and polyatomic ions, ***inter*molecular forces** hold molecules together in a liquid or solid. Intermolecular forces are generally much weaker than covalent bonds. Given the large difference in the strengths of intramolecular and intermolecular forces, changes between the solid, liquid, and gaseous states almost invariably occur for molecular substances *without* breaking covalent bonds.

Intermolecular forces determine bulk properties, such as the melting points of solids and the boiling points of liquids. Liquids boil when the molecules have enough thermal energy to overcome the attractive intermolecular forces that hold them together, thereby forming bubbles of vapor within the liquid. Similarly, solids melt when the molecules acquire enough thermal energy to overcome the intermolecular forces that lock them into place in the solid.

In this section, we explicitly consider three kinds of intermolecular interactions, the first two of which are often described collectively as van der Waals forces. There are two additional types of electrostatic interactions: the ion–ion interactions that are responsible for ionic bonding with which you are already familiar, and the ion–dipole interactions that occur when ionic substances dissolve in a polar substance such as water which was introduced in the previous section and will be discussed more in the next chapter.

Dipole-Dipole Forces

As previously described, polar molecules have one end that is partially positive (δ+) and another end that is partially negative (δ–). Because molecules in a liquid move freely and continuously, molecules experience both attractive and repulsive forces while interacting with each other. Polar molecules tend to align themselves so that the positive end of one dipole is near the negative end of a different dipole and vice versa, as shown in Figure 13.7.1. These predominant *attractive* intermolecular forces between polar molecules are called **dipole–dipole forces**.

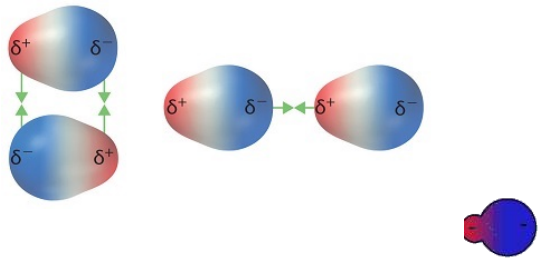


Figure **13.7.1**: Dipole-dipole forces involve molecular orientations in which the positive end of one dipole (δ^+) is near the negative end of another (δ^-) of a different dipole, causing an attraction between the two molecules. (Animation adapted from Wisc-Online, licensed by CC BY-NC-SA)

The attraction between partially positive and partially negative regions of a polar molecule that makes up dipole-dipole forces is the same type of attraction that occurs between cations and anions in an ionic compound. However, because each end of a dipole possesses only a fraction of the charge of an electron, dipole–dipole forces are substantially weaker than the forces between two ions, each of which has a charge of at least ±1, or between a dipole and an ion, in which one of the species has at least a full positive or negative charge. In addition, the attractive interaction between dipoles falls off much more rapidly with increasing distance than do the ion–ion interactions. Thus a substance such as HCl, which is partially held together by dipole–dipole interactions, is a gas at room temperature and 1 atm pressure. Conversely, NaCl, which is held together by interionic interactions, is a high-melting-point solid. Within a series of compounds of similar molar mass, the strength of the intermolecular interactions increases as the polarity of the molecules increases. This increase in the strength of the intermolecular interaction is reflected in an increase in melting point or boiling point, as shown in Table 13.7.1.

Table 13.7.1: Relationships Between the Polarity and Boiling Point for Organic Compounds of Similar Molar Mass

Compound	Molar Mass (g/mol)	*Dipole Moment (D)	Boiling Point (K)
C_3H_6 (cyclopropane)	42	0	240
CH_3OCH_3 (dimethyl ether)	46	1.30	248
CH_3CN (acetonitrile)	41	3.9	355

*The dipole moment is a measure of molecular polarity. The larger the numeric value, the greater the polarity of the molecule.

London Dispersion Forces

Thus far, we have considered only interactions between polar molecules. Other factors must be considered to explain why many nonpolar molecules, such as bromine, benzene, and hexane, are liquids at room temperature and why others, such as iodine and naphthalene, are solids. Even the noble gases can be liquefied or solidified at low temperatures, high pressures, or both. What kind of attractive forces can exist between nonpolar molecules or atoms? This question was answered by Fritz London (1900–1954), a German physicist who later worked in the United States. In 1930, London proposed that temporary fluctuations in the electron distributions within atoms and nonpolar molecules could result in the formation of short-lived instantaneous dipole moments, which produce attractive forces called **London dispersion forces**, or simply **London forces** or **dispersion forces**, between otherwise nonpolar substances.

Consider a pair of adjacent He atoms, for example. On average, the two electrons in each He atom are uniformly distributed around the nucleus. Because the electrons are in constant motion, however, their distribution in one atom is likely to be asymmetrical at any given instant, resulting in an instantaneous dipole moment (see image on left in Figure 13.7.2 below). The instantaneous dipole moment on one atom can interact with the electrons in an adjacent atom, pulling them toward the positive end of the instantaneous dipole or repelling them from the negative end (see image on right in Figure 13.7.2 below). The net effect is that the first atom causes the temporary formation of a dipole, called an induced dipole, in the second. Interactions between these temporary dipoles cause atoms to be attracted to one another. These attractive interactions are weak and fall off rapidly with increasing distance.

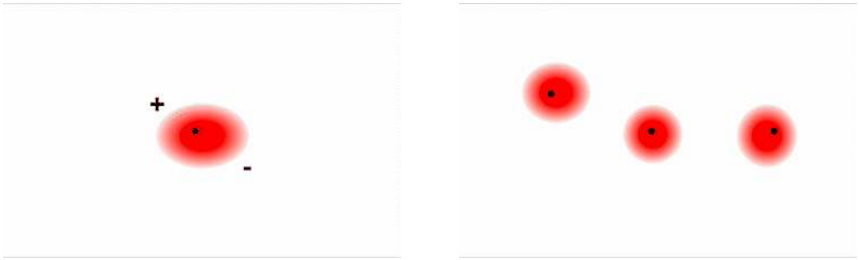


Figure 13.7.2: Instantaneous dipole moments. The formation of an instantaneous dipole moment on a He atom results in the formation of an induced dipole on an adjacent atom or molecule. The resulting temporary attractions are called dispersion forces. (Animation adapted from Wisc-Online, licensed by CC BY-NC-SA)

Dispersion forces between nonpolar molecules can produce intermolecular attractions just as they produce interatomic attractions in monatomic substances like He. This effect tends to become more pronounced as atomic and molecular masses increase (Table 13.7.2). For example, Xe boils at −108.1°C, whereas He boils at −269°C. The reason for this trend is that the strength of dispersion forces is related to the ease with which the electron distribution in a given atom can become temporarily asymmetrical. In small atoms such as He, its two electrons are held close to the nucleus in a very small volume, and electron-electron repulsions are strong enough to prevent significant asymmetry in their distribution. In larger atoms such as Xe, there are many more electrons and energy shells. As a result, it is relatively easy to temporarily deform the electron distribution to generate an instantaneous or induced dipole, since there is a greater probability of a temporary, uneven distribution of electrons.

Table 13.7.2: Normal Melting and Boiling Points of Some Elements and Nonpolar Compounds

Substance	Molar Mass (g/mol)	Melting Point (°C)	Boiling Point (°C)
Ar	40	−189.4	−185.9
Xe	131	−111.8	−108.1
N_2	28	−210	−195.8
O_2	32	−218.8	−183.0
F_2	38	−219.7	−188.1
I_2	254	113.7	184.4
CH_4	16	−182.5	−161.5
CCl_4	154	−22.2	76.7

Thus, dispersion forces are responsible for the general trend toward higher boiling points with increased molecular mass and greater surface area in a homologous series of compounds, such as the alkanes in Figure 13.7.3(a) below. The strengths of dispersion forces also depend significantly on molecular shape because shape determines how much of one molecule can interact with its neighboring molecules at any given time. For example, Figure 13.7.3(b) shows 2,2-dimethylpropane and pentane, both of which have the empirical formula C_5H_{12}. 2,2-dimethylpropane is almost spherical, with a small surface area for intermolecular interactions, whereas pentane has an extended conformation that enables it to come into close contact with other pentane molecules. As a result, the boiling point of 2,2-dimethylpropane (9.5°C) is more than 25°C lower than the boiling point of pentane (36.1°C).

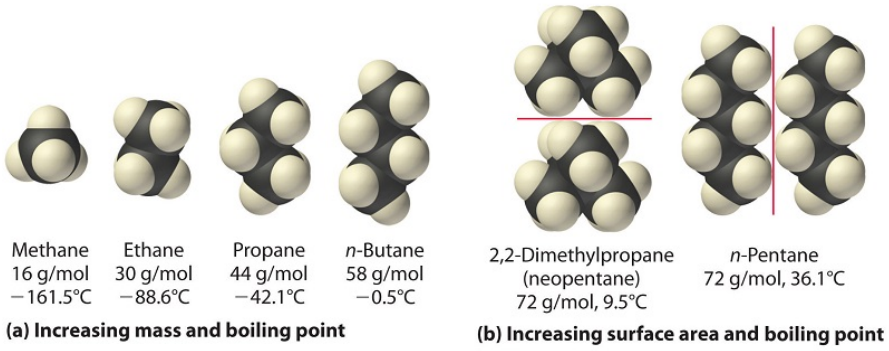


Figure 13.7.3: Mass and surface area affect the strength of dispersion forces. (a) In this series of four simple alkanes, larger molecules have stronger disperson forces between them than smaller molecules and have higher boiling points. (b) Linear pentane molecules have a larger surface area and stronger intermolecular forces than spherical neopentane molecules. As a result, 2,2-dimethylpropane is a gas at room temperature, whereas pentane is a volatile liquid.

Finally, it should be noted that *all molecules, whether polar or nonpolar*, are attracted to one another by dispersion forces *in addition* to any other attractive forces that may be present. In general, however, dipole–dipole forces in small polar molecules are significantly stronger than dispersion forces, so the dipole–dipole forces predominate.

Hydrogen Bonds

Let's apply what we have learned to the boiling points of the covalent hydrides of elements in Groups 14-17, as shown in Figure 13.7.4 below. Methane and the other hydrides of Group 14 elements are symmetrical molecules and are therefore nonpolar. Thus, the only attractive forces between molecules will be dispersion forces. These dispersion forces are expected to become stronger as the molar mass of the compound increases. The increasing strength of the dispersion forces will cause the boiling point of the compounds to increase, which is what is observed.

In Groups 15-17, lone pairs are present on the central atom, creating asymmetry in the molecules. The molecules are therefore polar to varying degrees and will contain dipole-dipole forces in addition to the dispersion forces. The presence of the stronger dipole-dipole force causes the boiling points of molecules in Groups 15-17 to be greater than the boiling point of the molecules in Group 14 in the same period.

We can still see that the boiling point increases with molar mass due to increases in the strength of the dispersion forces as we move from period 3 to period 5.

The trends break down for the hydrides of the lightest members of groups 15–17 which have boiling points that are more than 100°C greater than predicted on the basis of their molar masses. The effect is most dramatic for water: if we extend the straight line connecting the points for H_2Te and H_2Se to the line for period 2, we obtain an estimated boiling point of −70°C for water! Imagine the implications for life on Earth if water boiled at −70°C rather than 100°C. Obviously, there must be some other attractive force present in NH_3, HF, and H_2O to account for the higher boiling points in these molecules. This attractive force is known as a **hydrogen bond**.

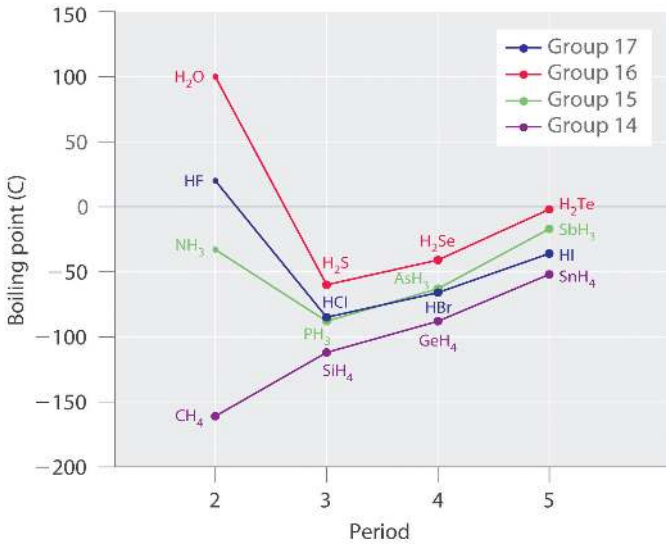


Figure **13.7.4**: The effects of hydrogen bonding on boiling points. These plots of the boiling points of the covalent hydrides of the elements of groups 14–17 show that the boiling points of the lightest members of each series for which hydrogen bonding is possible (NH_3, HF, and H_2O) are anomalously high for compounds with such low molecular masses.

Hydrogen bonds are an unusually strong version of dipole–dipole forces in which hydrogen atoms are bonded to highly electronegative atoms such as N, O, and F. In addition, the N, O, or F will typically have lone pair electrons on the atom in the Lewis structure. The large difference in electronegativity results in a large partial positive charge on hydrogen and a correspondingly large partial negative charge on the N, O, or F atom which will be concentrated on the lone pair electrons. Consequently, H—N, H—O, and H—F bonds will have very large bond dipoles, allowing the H atoms to interact strongly with the lone pairs of N, O, or F atoms on neighboring molecules. In addition, because the atoms involved are so small, these molecules can also approach one another more closely than most other dipoles. The combination of large bond dipoles and short intermolecular distances results in *very* strong dipole–dipole interactions called **hydrogen bonds**, as shown for ice in Figure 13.7.5.

Because it is such a strong intermolecular attraction, a hydrogen bond is usually indicated by a dotted line between the hydrogen atom attached to N, O, or F and the atom that has the lone pair of electrons. One thing that you may notice is that the hydrogen bond in the ice in Figure 13.7.5 is drawn to where the lone pair electrons are found on the oxygen atom. This is because the large partial negative charge on the oxygen atom (or on a N or F atom) is concentrated in the lone pair electrons. Thus, the hydrogen bond attraction will be specifically between the lone pair electrons on the N, O, or F atom and the H of a neighboring molecule. Because each water molecule contains two hydrogen atoms and two lone pairs, it can make up to four hydrogen bonds with adjacent water molecules.

In order to maximize the hydrogen bonding when fixed in position as a solid, the molecules in ice adopt a tetrahedral arrangement. Each oxygen atom is surrounded by a distorted tetrahedron of hydrogen atoms that form bridges to the oxygen atoms of adjacent water molecules. The resulting open, cage-like structure of ice means that the solid is actually slightly less dense than the liquid, which explains why ice floats on water, rather than sinks. The structure of liquid water is very similar, but in the liquid, the hydrogen bonds are continually broken and

formed because of rapid molecular motion so that the tetrahedral arrangement is not maintained.

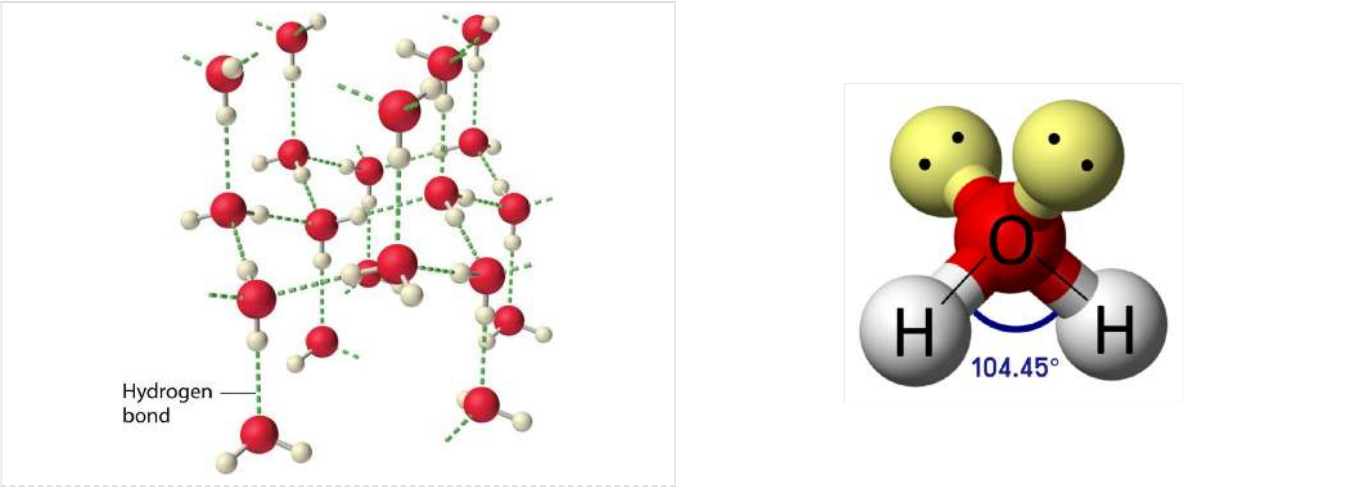
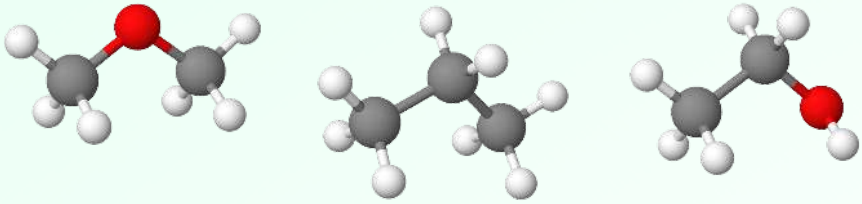


Figure **13.7.5** (left to right): The hydrogen-bonded structure of ice. The Lewis structure of H_2O, including lone pairs, superimposed over its three-dimensional rendering.

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

✅ Example 13.7.1

Which substance(s) can form a hydrogen bond to another molecule of itself? If the substance cannot form a hydrogen bond to another molecule of itself, which intermolecular force is the predominant intermolecular force for the substance?



Solution

Considering the structures from left to right:

1. This molecule cannot form hydrogen bonds to another molecule of itself since there are no H atoms directly bonded to N, O, or F. However, the molecule is polar, meaning that dipole-dipole forces are present. While all molecules, polar or nonpolar, have dispersion forces, the dipole-dipole forces are predominant.

2. This molecule cannot form hydrogen bonds to another molecule of itself since there are no H atoms directly bonded to N, O, or F. The molecule is nonpolar, meaning that the only intermolecular forces present are dispersion forces. This means that dispersion forces are also the predominant intermolecular force.

3. This molecule can form hydrogen bonds to another molecule of itself since there is an H atom directly bonded to O in the hydroxyl group (—OH).

✅ Example 13.7.2

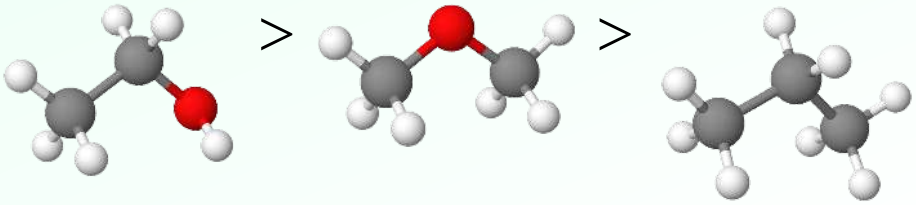
Arrange the substances shown in Example 13.7.1 above in order of *decreasing* boiling point.

Solution

Considering the structures in Example 13.7.1 from left to right, the condensed structural formulas and molar masses are:

- CH_3OCH_3 (46 g/mol), $CH_3CH_2CH_3$ (44 g/mol), and CH_3CH_2OH (46 g/mol).

Since they all have about the same molar mass, their boiling points should decrease in the order of the strongest to weakest predominant intermolecular force. Therefore, their arrangement in order of decreasing boiling point is:



✏️ Exercise 13.7.1

Which intermolecular forces are present in each substance? If there is more than one, identify the predominant intermolecular force in each substance.

A. HF
B. H_2
C. HBr

Answer A

Dispersion forces, dipole-dipole forces, hydrogen bonds are all present. Hydrogen bonds are the predominant intermolecular force.

Answer B

Dispersion forces are the only intermolecular forces present. Therefore, they are also the predominant intermolecular force.

Answer C

Dispersion forces and dipole-dipole forces are present. Dipole-dipole forces are the predominant intermolecular force.

✏️ Exercise 13.7.2

Arrange the noble gases (He, Ne, Ar, Kr, and Xe) in order of *increasing* boiling point.

Answer

He < Ne < Ar < Kr < Xe (This is in the order of increasing molar mass, since the only intermolecular forces present for each are dispersion forces.)

Summary

Click "Next" to begin a short review of this section.

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13.9: Exercises

13.1: Properties of Liquids and Solids

1. In terms of their bulk properties, how do liquids and solids differ? How are they similar?

Answer

Liquids and solids are similar in that they are matter composed of atoms, ions, or molecules. They are incompressible and have similar densities that are both much larger than those of gases. They are different in that liquids have no fixed shape, and solids are rigid.

2. Explain why liquids assume the shape of any container into which they are poured, whereas solids are rigid and retain their shape.

Answer

Liquids have enough energy to move about relative to each other while remaining in contact. Solids have less energy and remain in a fixed position.

3. The density of liquid NH_3 is 0.64 g/mL; the density of gaseous NH_3 at STP is 0.0007 g/mL. Explain the difference between the densities of these two phases.

Answer

A liquid is a condensed phase where the particles are in constant contact with each other. Gases have enough energy for the particles to separate from each other and spread out. This means that there are fewer gas particles in a given volume than liquid particles.

13.2: Evaporation and Condensation

4. Why does spilled gasoline evaporate more rapidly on a hot day than on a cold day?

Answer

As the temperature increases, the average kinetic energy of the molecules of gasoline increases and so a greater fraction of molecules have sufficient energy to escape from the liquid than at lower temperatures.

5. How does the boiling of a liquid differ from its evaporation?

Answer

When evaporating, only a few molecules at the surface of the liquid with sufficient energy are able to escape to the gas phase. When boiling, enough molecules in the interior of the liquid have sufficient energy to create small bubbles of gas that are able to move to the surface and escape.

13.3: Melting, Freezing, and Sublimation

6. Another way to minimize freezer burn is to wrap food tightly before freezing. Why would this minimize freezer burn?

Answer

Freezer burn is due to the sublimation of solid water to gaseous water. In order to form the gas, there needs to be space for the gas molecules to move about in. Minimizing this space by tightly wrapping the food will minimize the amount of freezer burn possible.

7. If a severe storm results in the loss of electricity, it may be necessary to use a clothesline to dry laundry. In many parts of the country in the dead of winter, the clothes will quickly freeze when they are hung on the line. If it does not snow, will they dry anyway? Explain your answer.

Answer

Yes, ice will sublime, although it may take it several days. Ice has a small vapor pressure, and some ice molecules form gas and escape from the ice crystals. As time passes, more and more solid converts to gas until eventually the clothes are dry.

13.4: Energetics of Phase Changes

8. Ethyl chloride (boiling point, 13°C) is used as a local anesthetic. When the liquid is sprayed on the skin, it cools the skin enough to freeze and numb it. Explain the cooling effect of liquid ethyl chloride.

Answer

The thermal energy (heat) needed to evaporate the liquid is removed from the skin.

9. Heat is added to boiling water. Explain why the temperature of the boiling water does not change. What does change?

Answer

The heat is absorbed by the water to help the molecules overcome the intermolecular attractive forces in the liquid to allow the gas to form. The water remains at 100°C until all of the water has boiled off. Only the volume of the liquid water will change as it is boiling. Once all has been converted to the gas phase, the temperature of the gas will rise.

10. Heat is added to ice at 0°C. Explain why the temperature of the ice does not change. What does change?

Answer

The heat is absorbed by the ice, providing the energy required to partially overcome intermolecular attractive forces in the solid and causing a phase transition to liquid water. The solution remains at 0 °C until all the ice is melted. Only the amount of water existing as ice changes until the ice disappears. Then the temperature of the water can rise.

13.5: Electronegativity and Polarity

11. Explain the difference between a nonpolar covalent bond, a polar covalent bond, and an ionic bond.

Answer

A nonpolar bond is one where the electrons are shared relatively equally between the two atoms involved in the bond. This will be the case when their electronegativities are very similar (less than 0.4). A polar bond is one where the electrons are shared, but the sharing is not equal causing the more electronegative atoms to have a partial negative charge and the less electronegative atom to have a partial positive charge. This will be the case when the electronegativities differ by 0.4 to about 1.9. Ionic bonds are formed when one atom is able to take the electron totally away from another, forming ions. This will be the case when metals and nonmetals are combined where the difference in electronegativities become 2.0 or more.

12. Give an example of a nonpolar covalent bond. How do you know it is nonpolar?

Answer

C—C; The two atoms in the bond are identical so they will share the electrons equally.

13. Give an example of a polar covalent bond. How do you know it is polar?

Answer

C—F; The two atoms in the bond have different electronegativities. They are both nonmetals so the difference is likely to be less than 2.0. Fluorine is more electronegative than carbon so the fluorine will have the partial negative charge.

14. From its position in the periodic table, determine which atom in each pair is more electronegative:

a. Br or Cl
b. N or O
c. S or O
d. P or S
e. Si or N
f. Ba or P
g. N or K

Answer

a. Cl
b. O
c. O
d. S
e. N
f. P
g. N

15. From its position in the periodic table, determine which atom in each pair is more electronegative:

a. N or P
b. N or Ge
c. S or F
d. Cl or S
e. H or C
f. Se or P
g. C or Si

Answer

a. N
b. N
c. F
d. Cl
e. C
f. Se

g. C

16. Identify the more polar bond in each of the following pairs of bonds:

a. H—F or H—Cl
b. N—O or C—O
c. S—H or O—H
d. P—Cl or S—Cl
e. C—H or N—H
f. S—O or P—O
g. C—N or N—N

Answer

a. H—F
b. C—O
c. O—H
d. P—Cl
e. N—H
f. P—O
g. C—N

17. Which atoms can bond to sulfur so as to produce a positive partial charge on the sulfur atom?

Answer

N, O, F, and Cl

18. Which is the most polar bond?

a. C—C
b. C—H
c. N—H
d. O—H
e. Se—H

Answer

d. O—H

19. Which bond do you expect to be more polar: an O—H bond or an N—H bond? Why?

Answer

O—H because there is a larger difference in electronegativities

20. Which bond do you expect to be more polar: an O—F bond or an S—O bond? Why?

Answer

S—O because there is a larger difference in electronegativities

21. How do you know which side of a polar bond has the partial negative charge? Identify the negatively charged side of each polar bond.
    a. H—Cl
    b. H—S

Answer

The side that has the more electronegative atom has the partial negative charge.

a. Cl
b. S

22. How do you know which side of a polar bond has the partial positive charge? Identify the positively charged side of each polar bond.
    a. H—Cl
    b. N—F

Answer

The side that has the less electronegative atom has the partial positive charge.

a. H
b. N

23. Label the bond between the given atoms as nonpolar covalent, polar covalent, or ionic.
    a. H and C
    b. C and F
    c. K and F

Answer

a. nonpolar covalent
b. polar covalent
c. ionic

24. Label the bond between the given atoms as nonpolar covalent, polar covalent, or ionic.
    a. S and Cl
    b. P and O
    c. Cs and O

Answer

a. polar covalent
b. polar covalent
c. ionic

25. From their positions in the periodic table, arrange the atoms in each of the following series in order of increasing electronegativity:
    a. C, F, H, O, N

b. H, Br, I, F, Cl
c. S, P, F, O, H
d. Al, P, O, N, Na

Answer

a. H < C < N < O < F
b. H < I < Br < Cl < F
c. H < P < S < O < F
d. Na < Al < P < N < O

26. Explain how a molecule that contains polar bonds can be nonpolar.

Answer

As long as the polar bonds are opposite of each other and can "cancel" each other out, the molecule can be nonpolar.

27. Identify the molecules with a dipole moment:

a. CF_4
b. Cl_2CCBr_2 (the two C atoms are central)
c. CH_3Cl
d. H_2CO

Answer

Cl_2CCBr_2, CH_3Cl, H_2CO

28. Which of the following molecules have dipole moments?

a. CS_2
b. SeS_2
c. CCl_2F_2
d. PCl_3
e. ClNO (N is the central atom)

Answer

SeS_2, CCl_2F_2, PCl_3, ClNO

13.7: Intermolecular Forces

29. Define the following and give an example of each:

a. dispersion force
b. dipole-dipole attraction
c. hydrogen bond

Answer

A dispersion is the attraction between a partial positive and a partial negative region on neighboring molecules that is caused by the random movement of electrons. This would be the attractive force in hydrocarbons like butane. A dipole-dipole attraction is also the attraction between a partial positive and a partial negative region on neighboring polar molecules. However, in this case it is due to the presence of a permanent dipole in the molecules and is therefore typically stronger than a dispersion force. An

example of this would be the attractive force in acetone. A hydrogen bond is a special type of dipole-dipole attraction with the highly polar bond by formed by H—N, H—O, or H—F. This is the seen in water.

30. All other things being equal, rank the intermolecular forces in order of increasing strength.

Answer

dispersion force < dipole-dipole attraction < hydrogen bond

31. Which subatomic particles (protons, neutrons, electrons) are most responsible for intermolecular forces? Explain your answer.

Answer

electrons; intermolecular forces are caused by partial charges on atoms due to the unequal distribution of electrons in bonds whether that be due to permanent dipoles in polar molecules or temporary dipoles in nonpolar molecules

32. Can a molecule experience more than one intermolecular force at the same time? Why or why not?

Answer

Yes. All molecules have dispersion forces. In addition, polar molecules can have dipole-dipole attractions and sometimes hydrogen bonding.

33. Of the properties boiling point, structure of the solid phase, and molar mass, which are influenced by hydrogen bonding? Explain your answer.

Answer

Boiling point and the structure of the solid phase are influenced by hydrogen bonding. Both of these are determined by how molecules interact with each other, which can be influenced by hydrogen bonding. Molar mass is simply the mass of the atoms in a molecule and is not affected by how they interact with each other.

34. What is the evidence that all neutral atoms and molecules exert attractive forces on each other?

Answer

All atoms and molecules will condense into a liquid or solid in which the attractive forces exceed the kinetic energy of the molecules, at sufficiently low temperature.

35. What is the relationship between the intermolecular forces in a solid and its melting temperature?

Answer

As the strength of the intermolecular forces increases, the melting point (or freezing point) will also increase.

36. The molecular mass of butanol, C_4H_9OH, is 74.14; that of ethylene glycol, $CH_2(OH)CH_2OH$, is 62.08, yet their boiling points are 117.2 °C and 174 °C, respectively. Explain the reason for the difference.

Answer

Ethylene glycol has more -OH groups than butanol so it is able to form more hydrogen bonding interactions with its neighbors, increasing the boiling point.

37. On the basis of intermolecular attractions, explain the differences in the boiling points of *n*–butane (−1 °C) and chloroethane (12 °C), which have similar molar masses.

Answer

Only rather small dispersion forces from C-H bonds are available to hold *n*-butane in the liquid state. Chloroethane, however, has rather large dipole interactions because of the Cl-C bond; the interaction is therefore stronger, leading to a higher boiling point.

38. On the basis of dipole moments and/or hydrogen bonding, explain in a qualitative way the differences in the boiling points of acetone (56.2 °C) and 1-propanol (97.4 °C), which have similar molar masses.

Answer

While both are polar molecules, 1-propanol is an alcohol so it has an -OH group in the molecule. This -OH group can participate in hydrogen bonding which will increase the strength of the intermolecular forces in the molecules and increase the boiling point.

39. The melting point of $H_2O(s)$ is 0 °C. Would you expect the melting point of $H_2S(s)$ to be −85 °C, 0 °C, or 185 °C? Explain your answer.

Answer

−85 °C. Water has stronger hydrogen bonds so it melts at a higher temperature.

40. Silane (SiH_4), phosphine (PH_3), and hydrogen sulfide (H_2S) melt at −185 °C, −133 °C, and −85 °C, respectively. What does this suggest about the polar character and intermolecular attractions of the three compounds?

Answer

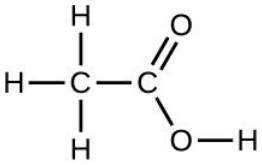
All three compounds have similar molar masses, so the difference in melting points is due to just the polar character. Silane with the lowest melting point will have the least polar character and the weakest intermolecular attractions of the three. Hydrogen sulfide with the highest melting point will have the most polar character and the greatest intermolecular attractions of the three.

41. Explain why a hydrogen bond between two water molecules is weaker than a hydrogen bond between two hydrogen fluoride molecules.

Answer

The hydrogen bond between two hydrogen fluoride molecules is stronger than that between two water molecules because the electronegativity of F is greater than that of O. Consequently, the partial negative charge on F is greater than that on O. The hydrogen bond between the partially positive H and the larger partially negative F will be stronger than that formed between H and O.

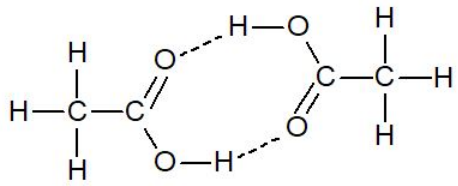
42. Under certain conditions, molecules of acetic acid, CH_3COOH, form “dimers,” pairs of acetic acid molecules held together by strong intermolecular attractions:



Draw a dimer of acetic acid, showing how two CH_3COOH molecules are held together, and stating the type of IMF that is responsible.

Answer

The IMF responsible is hydrogen bonding.



43. Arrange each of the following sets of compounds in order of increasing boiling point temperature:

a. HCl, H_2O, SiH_4
b. F_2, Cl_2, Br_2
c. CH_4, C_2H_6, C_3H_8
d. O_2, NO, N_2

Answer

a. $SiH_4 < HCl < H_2O$
b. $F_2 < Cl_2 < Br_2$
c. $CH_4 < C_2H_6 < C_3H_8$
d. $N_2 < O_2 < NO$

44. Identify the all of the intermolecular forces present in the following solids. Which is the strongest for each?

a. CH_3CH_2OH
b. $CH_3CH_2CH_3$
c. CH_3CH_2Cl

Answer

a. dispersion, dipole, hydrogen bonding; hydrogen bonding will be strongest
b. dispersion only (and thus the strongest)
c. dispersion and dipole; dipole will be strongest

45. The types of intermolecular forces in a substance are identical whether it is a solid, a liquid, or a gas. Why then does a substance change phase from a gas to liquid or to a solid?

Answer

The difference in phases is based on how much energy the molecules have relative to the strength of the intermolecular forces. If the energy of the molecules is greater than the strength of the intermolecular forces, the molecules will separate and form a gas. If the energy of the molecules is less than the strength of the intermolecular forces, the molecules will be held together in the liquid or solid phase. The temperature where this occurs will depend on the strength of the intermolecular forces present.

46. Why do the boiling points of the noble gases increase in the order He < Ne < Ar < Kr < Xe?

Answer

The London forces typically increase as the number of electrons increase.

47. Neon and HF have approximately the same molecular masses. Explain why the boiling points of Neon and HF differ.

Answer

Neon is a nonpolar atom held together by only disperson forces. HF is a polar molecule that can exhibit hydrogen bonding between molecules.

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CHAPTER OVERVIEW

14: Solutions

Solutions play a very important role in many biological, laboratory, and industrial applications of chemistry. Of particular importance are solutions involving substances dissolved in water, or aqueous solutions. Quantitative measurements of solutions are another key component of this unit. Solutions can involve all physical states – gases dissolved in gases (the air around us), solids dissolved in solids (metal alloys), liquids dissolved in solids (amalgams – liquid mercury dissolved in another metal such as silver, tin or copper). This chapter will almost exclusively be concerned with aqueous solutions – substances dissolved in water.

14.1: Cameroon Tragedy
14.2: Solution Terminology
14.3: Aqueous Solutions of Solids
14.4: Solutions of Gases in Water
14.5: Solution Concentration Measurements
14.6: Molarity
14.7: Solution Dilution
14.8: Colligative Properties of Solutions
14.9: For Future Use
14.10: Exercises

Template:HideTOC

Thumbnail Chapter 14: Colored solutions in round bottom flasks. (Raghav Bhasin via Unsplash)

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14.1: Cameroon Tragedy

Lake Nyos is a deep crater lake in the Northwest region of Cameroon, high on the flank of an inactive volcano in the Oku volcanic plain along the Cameroon line of volcanic activity. A volcanic dam impounds the lake waters. A pocket of magma lies beneath the lake and leaks carbon dioxide, CO_2, into the water, changing it into carbonic acid. Nyos is one of only three known exploding lakes to be saturated with carbon dioxide in this way.

In 1986, more than 1700 people in Cameroon were killed when a cloud of gas, almost certainly carbon dioxide, bubbled from Lake Nyos (see Figure 14.1.1). It is believed that the lake underwent a turnover due to gradual heating from below the lake, and the warmer, less-dense water saturated with carbon dioxide reached the surface. Consequently, tremendous quantities of dissolved CO_2 were released, and the colorless gas, which is denser than air, flowed down the valley below the lake and suffocated humans and animals living in the valley.





Figure **14.1.1** (left to right): Lake Nyos, Cameroon in 2013 (jbdodane via Flickr); vent used to outgas CO_2 from Lake Nyos (Bill Evans via Wikimedia Commons).

Following the Lake Nyos tragedy, scientists investigated other African lakes to see if a similar phenomenon could happen elsewhere. Lake Kivu in Democratic Republic of Congo, 2,000 times larger than Lake Nyos, was also found to be supersaturated, and geologists found evidence for out-gassing events around the lake every one thousand years or so.

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14.2: Solution Terminology

Learning Objectives

- Learn terminology involving solutions.
- Qualitatively explain the amount of solute dissolved in a solvent.

A **solution** is a homogenous mixture. The component of a solution that does the dissolving is called the **solvent**, while the component that gets dissolved is called the **solute**. Almost always, the solvent is the component that is present in the greatest proportion by mass, moles, or volume.

When one substance dissolves into another substance, we say it is **soluble** in that substance. If a substance does not dissolve into another substance, we say it is **insoluble** in that substance. When two liquids can dissolve into each other in any proportion, we say the liquids are **miscible** with each other. If they cannot dissolve into each other, we say they are **immiscible** with each other.



Figure **14.2.1**: Table salt, NaCl, is soluble in water. Salt is the solute. Water is the solvent. (CC-BY-SA 3.0; Chris 73).

Solutions exist for every possible phase combination of solute and solvent. Salt water, for example, is a solution of solid NaCl in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent. A solution in which the overall phase is a gas is called a **gaseous solution**. A solution in which the overall phase is a liquid is called a **liquid solution**. A solution in which the overall phase is a solid is called a **solid solution**. Table 14.2.1 lists some common types of solutions, with examples of each.

Table 14.2.1: Types of Solutions

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

As we have just learned, a solution need not involve water. However, when water is the solvent, it is called an **aqueous solution**. When the solvent is any substance other than water, it is called a **nonaqueous solution**.

In many cases, the relative amounts of solute and solvent may be referenced using the terms dilute and concentrated. A **dilute solution** is one that contains a relatively small amount of solute dissolved in a solvent, while a **concentrated solution** is one that contains a relatively large amount of solute dissolved in a solvent. There is no set cut-off point that distinguishes a dilute solution from a concentrated solution and what constitutes dilute or concentrated may vary with the solute and/or solvent.

Solubility and Saturation

Table salt, NaCl, readily dissolves in water. In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is specified as the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table 14.2.2 lists the solubilities of some simple ionic compounds at 25°C. These solubilities vary widely. NaCl can dissolve up to 36.1 g per 100 g of H_2O at 25°C, while AgCl can dissolve only 0.00019 g per 100 g of H_2O at 25°C.

Table 14.2.2: Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H_2O at 25°C)
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent at a specified temperature, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solvent at a specified temperature, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated at 25°C, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O at 25°C is also saturated, but rather concentrated. Figure 14.2.2 shows the distinction between an unsaturated solution and a saturated solution of NaCl.

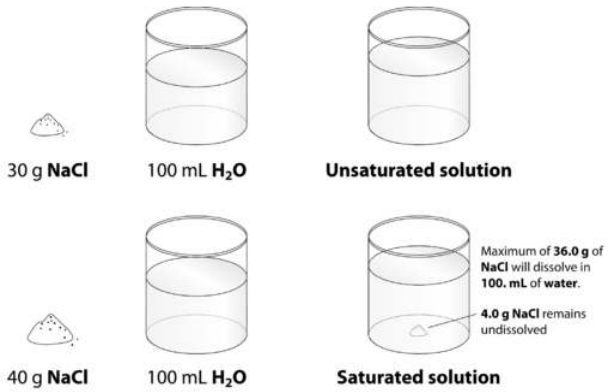


Figure 14.2.2: 30.0 g of NaCl completely dissolves in 100 mL H_2O. When 40.0 g NaCl is added, 36.0 g dissolves and 4.0 g remains undissolved. The end result is a saturated solution.

How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

This said, it is sometimes possible for a solution to have more solute dissolved than a saturated solution at a given temperature. Such a solution is called a **supersaturated** solution. A supersaturated solution may be formed by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, then allowing the solution to slowly cool. Supersaturated solutions are not stable. Given the opportunity (such as dropping a crystal of solute in the solution), excess solute will precipitate from a supersaturated solution.

Video 14.2.1 and Video 14.2.2 show some interesting applications of supersaturated solutions. Video 14.2.1 shows supersaturated sodium acetate solution being used in a reusable hand warmer, while Video 14.2.2 shows how to make rock candy from a supersaturated solution of sucrose.



Video **14.2.1**: A supersaturated sodium acetate solution may be used to make reusable hand warmers.



Video **14.2.2**: A supersaturated sucrose solution may be used to make rock candy.

✅ Example 14.2.1

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

✅ Example 14.2.2

Indicate whether gasoline and ethanol, CH_3CH_2OH, are miscible or immiscible with water.

Solution

Gasoline is a mixture of hydrocarbons. Hydrocarbons are nonpolar, while water is polar. Therefore, gasoline and water are immiscible with each other. Ethanol, CH_3CH_2OH, has a polar —OH grouping that makes ethanol a polar compound and one in which hydrogen bonds exist. Therefore, ethanol and water are miscible with each other.

Exercise 14.2.1

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

Answer

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

Exercise 14.2.2

Which solution is an example of an aqueous solution and which is an example of a nonaqueous solution?

A. Sugar dissolved in water.
B. Naphthalene dissolved in ethanol.

Answer

Sugar dissolved in water is an example of an aqueous solution.

Naphthalene dissolved in ethanol is an example of a nonaqueous solution.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Solvents and solutes can be any phase of matter combined together.
- A saturated solution has the maximum amount of solute possible at a given temperature. An unsaturated solution will have less solute than a saturated solution and, in certain cases, a supersaturated solution can be made which has more solute.

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14.3: Aqueous Solutions of Solids

Learning Objectives

- Define electrolytes and nonelectrolytes.
- Explain why solutions form.
- Discuss the idea of water as the "universal solvent."
- Explain how water molecules attract ionic solids when they dissolve in water.
- Use solubility curves to determine that saturation of a solution.

We have learned that solutions can be formed in a variety of combinations using solids, liquids, and gases. But how, exactly, do solutions form? Why is it that oil and water will not form a solution, and yet vinegar and water will? Why could we dissolve table salt in water, but not in vegetable oil? The reasons why solutions will form will be explored in this section, along with a discussion of why water is used most frequently to dissolve substances of various types.

Electrolyte Solutions: Dissolved Ionic Solids

When some substances are dissolved in water, they undergo a physical or chemical change that yields ions in solution. These substances constitute an important class of compounds called **electrolytes**. Substances that do not yield ions when dissolved are called **nonelectrolytes**. If the physical or chemical process that generates the ions is close to 100% efficient (all of the dissolved compound yields ions), then the substance is known as a strong electrolyte (good conductor of electricity). If only a relatively small fraction of the dissolved substance yields ions, the substance is a weak electrolyte (weak conductor of electricity).

Substances may be identified as strong electrolytes, weak electrolytes, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain freely mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 14.3.1).

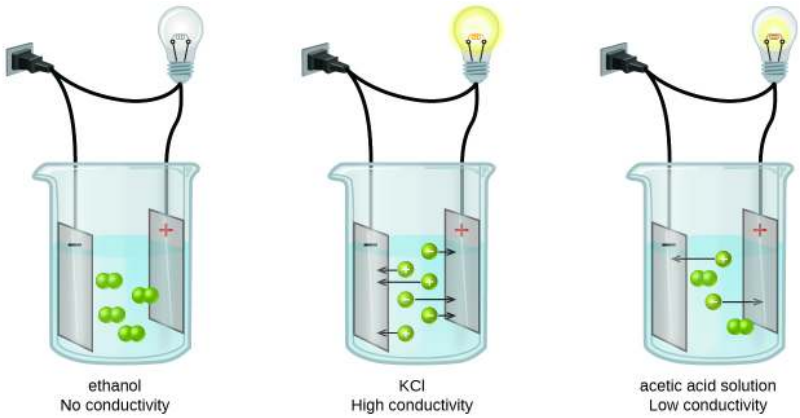


Figure **14.3.1**: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

Water and other polar molecules are attracted to ions, as shown in Figure 14.3.2, with the negative end of the molecule (oxygen in the case of water) being attracted to cations and the positive end of the molecule (hydrogen in the case of water) being attracted to anions. The electrostatic attraction between an ion and a molecule with a dipole is called an ion–dipole attraction. These attractions

play an important role in the dissolution of ionic compounds in water.

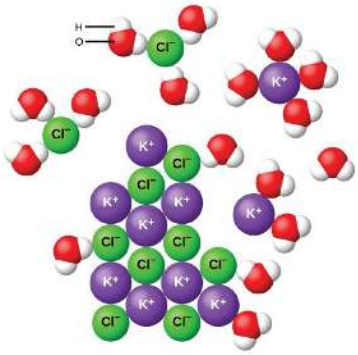


Figure 14.3.2: As potassium chloride (KCl) dissolves in water, the ions are hydrated. The polar water molecules are attracted by the charges on the K^+ and Cl^- ions. Water molecules in front of and behind the ions are not shown.

When ionic compounds dissolve in water, the ions in the solid separate and disperse uniformly throughout the solution because water molecules surround and solvate the ions, reducing the strong electrostatic forces between them. This process represents a physical change known as dissociation. Under most conditions, many ionic compounds will dissociate nearly completely when dissolved, and so they are classified as strong electrolytes.

✅ Example 14.3.1

Which compound(s) will dissolve in water to yield ions? (Reminder: Table 7.7.1 and Table 7.7.2 show solubility rules for ionic compounds.)

A. LiF
B. P_2F_5
C. CH_3CH_2OH
D. $MgSO_4$

Solution

LiF and $MgSO_4$ will separate into ions when dissolved in water, since they are both ionic compounds. P_2F_5 and CH_3CH_2OH are both molecular compounds and will remain as molecules in a solution.

✏️ Exercise 14.3.1

Which compound(s) will dissolve in water to yield ions? (Reminder: Table 7.7.1 and Table 7.7.2 show solubility rules for ionic compounds.)

A. $C_6H_{12}O_6$ (glucose)
B. CCl_4
C. $CaCl_2$
D. $Cr(NO_3)_3$

Answer

$CaCl_2$ & $Cr(NO_3)_3$

Ionic compounds are not the only substances that dissolve in water. We previously used the saying "like dissolves like" to describe the fact that polar molecules tend to dissolve in polar solvents and nonpolar molecules tend to dissolve in nonpolar solvents. Because water is a polar molecule, it will tend to dissolve other polar substances. The negative end of the water will be attracted to

the positive region of the solute and the positive end of the water will be attracted to the negative region of the solute, creating new dipole–dipole forces between the water and the solute. The interaction will be the most significant when the solute also has hydrogen bonding capabilities with the water. This is why alcohols and sugars which contain hydroxyl (—OH) groups are so soluble in water. Water is often considered to be the "universal solvent" due to the wide variety of ionic, polar, and even some nonpolar compounds that it is able to dissolve.

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per 100 g of solvent. The solubility of sodium chloride in water is 36.0 g per 100 g water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. We will consider the solubility of materials in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (Figure 14.3.4).

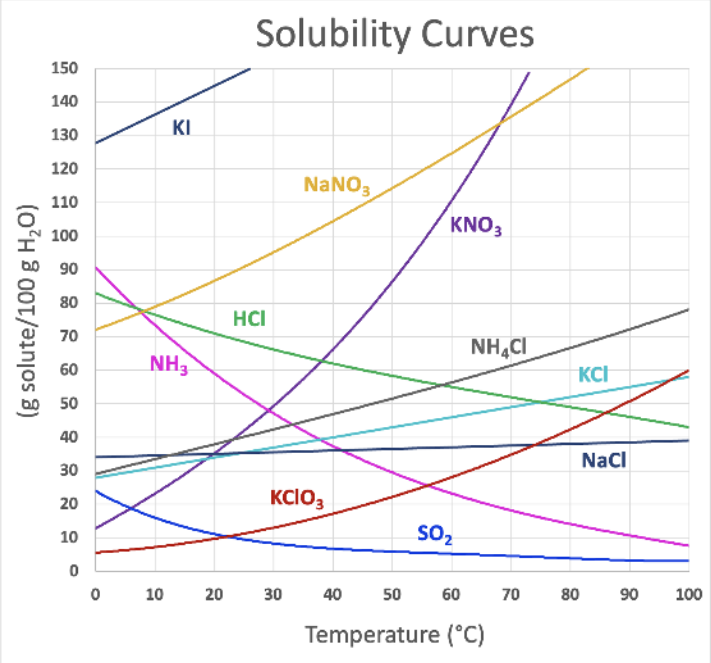


Figure 14.3.4: Solubility curves for several compounds. (adapted from LibreTexts)

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances – HCl, NH_3, and SO_2 – have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that 80 g of KNO_3 is added to 100 g of water at 30°C. According to the solubility curve, approximately 49 g of KNO_3 will dissolve at 30°C. This means that the solution will be saturated since 49 g is less than 80 g. We can also determine that there will be 80 g – 49 g = 31 g of undissolved

KNO_3 remaining at the bottom of the container. Now suppose that this saturated solution is heated to 60°C. According to the curve, the solubility of KNO_3 at 60°C is about 110 g per 100 g of water. Now the solution is unsaturated since it contains only the original 80 g of dissolved solute. Now suppose the solution is cooled all the way down to 0°C. The solubility at 0°C is about 14 g per 100 g of water, meaning that 80 g – 14 g = 66 g of the KNO_3 will recrystallize.

✅ Example 14.3.2

67 g of each of the following solutes were dissolved in 100 g H_2O at 40°C. Indicate whether the solution is unsaturated, saturated, or supersaturated.

A. KNO_3
B. $NaNO_3$
C. NH_4Cl

Solution

The graph in Figure 14.3.4 indicates that at 40°C, the solubility of

A. KNO_3 is 67 g per 100 g H_2O, so the solution would be saturated.
B. $NaNO_3$ is 104 g per 100 g H_2O, so the solution would be unsaturated.
C. NH_4Cl is 48 g per 100 g H_2O, so the solution would be supersaturated.

✏️ Exercise 14.3.2

How many grams of the indicated solute would dissolve in 100 g H_2O at the given temperature.

A. SO_2 at 10°C
B. $KClO_3$ at 50°C
C. KI at 20°C

Answer A

16 g SO_2

Answer B

22 g $KClO_3$

Answer C

146 g KI

✏️ Exercise 14.3.3

Indicate whether each solution is unsaturated, saturated, or supersaturated. Explain your answer.

A. 15 g NH_3 dissolved in 50 g H_2O at 50°C
B. 50 g NaCl dissolved in 200 g H_2O at 80°C

Answer A

Since the solubility of NH_3 is 30 g per 100 g H_2O at 50°C, half the mass of water (50 g) would dissolve up to 15 g NH_3. Therefore, the solution is saturated.

Answer B

Since the solubility of NaCl is 38 g NH_3 per 100 g H_2O at 80°C, twice the mass of water (200 g) would dissolve up to 76 g NaCl. Therefore, the solution is unsaturated.

Summary

- Solubility is the specific amount of solute that can dissolve in a given amount of solvent.
- Ionic compounds dissolve in polar solvents, especially water. This occurs when the positive cation from the ionic solid is attracted to the negative end of the water molecule (oxygen) and the negative anion of the ionic solid is attracted to the positive end of the water molecule (hydrogen).
- Water is considered the universal solvent since it can dissolve both ionic and polar solutes, as well as some nonpolar solutes (in very limited amounts).
- Saturated and unsaturated solutions can be determined using solubility curves.
- The solubility of a solid in water increases with an increase in temperature. The solubility of a gas in water decreases with an increase in temperature.

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14.4: Solutions of Gases in Water

Learning Objectives

- Explain how temperature and pressure affect the solubility of gases.

In the previous section, we saw that ionic solids will dissolve in water because of ion–dipole forces that allow the water molecules to separate the ions and disperse them through the solution. More generally, solutions form when the solute and solvent are able to create intermolecular attractions that are of similar strength to the types of attractions that were present in the pure solute and solvent. This is another manifestation of "like dissolves like". The chemical structures of the solute and solvent dictate the types of forces possible and, consequently, are important factors in determining solubility.

For example, under similar conditions, the water solubility of oxygen is approximately three times greater than that of helium, even though both are nonpolar species. However, the water solubility of oxygen is still 100 times less than the solubility of the polar trichloromethane, $CHCl_3$. Considering the role of the solvent's chemical structure, note that the solubility of oxygen in the nonpolar liquid hydrocarbon hexane, C_6H_{14}, is approximately 20 times greater than it is in water.

Other factors also affect the solubility of a given substance in a given solvent. Temperature is one such factor, with gas solubility typically decreasing as temperature increases (Figure 14.4.1). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.



Figure **14.4.1**: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

When the temperature of a river, lake, or stream is raised abnormally high, usually due to the discharge of hot water from some industrial process, the solubility of oxygen in the water is decreased. Decreased levels of dissolved oxygen often have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills (Figure 14.4.2).



Figure **14.4.2**: (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (Credit a: modification of work by Liz West; credit b: modification of work by U.S. Fish and Wildlife Service.)

The solubility of a gaseous solute is also affected by the partial pressure of the gas to which the solution is exposed. Gas solubility increases as the pressure of the gas increases. Essentially the greater pressure "pushes" more gaseous solute into the solution. This relationship between solubility and pressure is expressed as Henry's law. Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at this pressure. When the beverage container is opened, a familiar hiss is heard as the carbon dioxide gas pressure is released, and some of the dissolved carbon dioxide is typically seen leaving solution in the form of small bubbles (Figure 14.4.3). At this point, the beverage is *supersaturated* with carbon dioxide and, with time, the dissolved carbon dioxide concentration will decrease to its equilibrium value and the beverage will become "flat."

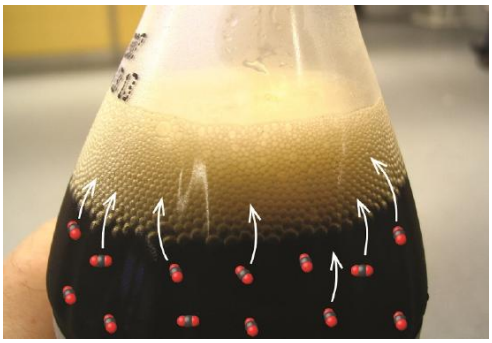


Figure **14.4.3**: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (Credit: modification of work by Derrick Coetzee.)

Case Study: Decompression Sickness ("The Bends")

Decompression sickness (DCS), or "the bends," is an effect of the increased pressure of the air inhaled by scuba divers when swimming underwater at considerable depths. In addition to the pressure exerted by the atmosphere, divers are subjected to additional pressure due to the water above them, experiencing an increase of approximately 1 atm for each 10 m of depth. Therefore, the air inhaled by a diver while submerged contains gases at the corresponding higher ambient pressure, and the concentrations of the gases dissolved in the diver's blood are proportionally higher per Henry's law.

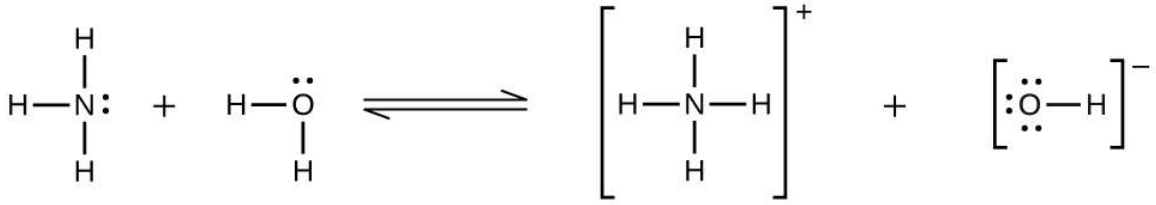
As the diver ascends to the surface of the water, the ambient pressure decreases and the dissolved gases becomes less soluble. If the ascent is too rapid, the gases escaping from the diver's blood may form bubbles that can cause a variety of symptoms ranging from rashes and joint pain to paralysis and death. To avoid DCS, divers must ascend from depths at relatively slow speeds (10 or 20 m/min) or otherwise make several decompression stops, pausing for several minutes at given depths during the ascent. When these preventative measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers (Figure 14.4.4).



(a) (b)

Figure **14.4.4**: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.

Deviations from Henry's law are observed when a chemical reaction takes place between the gaseous solute and the solvent, removing the gaseous solute from the solution. Thus, for example, the solubility of ammonia in water does not increase as rapidly with increasing pressure as predicted by the law because ammonia, being a base, reacts to some extent with water to form ammonium ions and hydroxide ions.



Gases can form supersaturated solutions. If a solution of a gas in a liquid is prepared either at low temperature or under pressure (or both), then as the solution warms or as the gas pressure is reduced, the solution may become supersaturated.

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14.5: Solution Concentration Measurements

Learning Objectives

- Express the amount of solute in a solution in various concentration units.

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms with meanings that depend on various factors. **Concentration** is the measure of how much of a given substance is mixed with another substance. For example, vinegar is 5% acetic acid in water. A solution of 10% acetic acid would be more concentrated than a vinegar solution.



Figure **14.5.1** (left to right): Solutions of a red dye in water from the most dilute to the most concentrated. (adapted from gelokaliseer deur Naudefj via Wikimedia Commons).

Mass Percentage

There are several ways of expressing the concentration of a solution by using a percentage. The mass percentage, $\%\left(\frac{m}{m}\right)$, is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\left(\frac{m}{m}\right) = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \tag{14.5.1}$$

where

$$\text{mass of solution} = \text{mass of solute} + \text{mass of solvent} \tag{14.5.2}$$

As long as the masses of the solute and solution can be measured, determining the mass percentage is easy. It is important to remember that all masses must be expressed in the same units to properly calculate the concentration.

Example 14.5.1

Calculate the mass percentage of a solution prepared by dissolving 50.0 g $C_6H_{12}O_6$ in 200.0 g H_2O.

Solution

The mass of the solution (solute + solvent) = 50.0 g + 200.0 g = 250.0 g

These masses of the solvent and solution may then be substituted into Equation 14.5.1:

$$\%\left(\frac{m}{m}\right) = \frac{50.0\ g}{250.0\ g} \times 100\% = \boxed{20.0\%\left(\frac{m}{m}\right)\ C_6H_{12}O_6}$$

Exercise 14.5.1

A 2.00×10^2 g solution of Epsom salts ($MgSO_4$) contains 15.8 g of $MgSO_4$. What is the mass percentage of the solution?

Answer

$7.90\%\left(\frac{m}{m}\right)\ MgSO_4$

Using Mass Percentage in Calculations

What if we wanted to prepare a specified amount of solution with a desired mass percentage? Using the mass percentage as a conversion is useful in this type of problem. The mass percentage may be expressed as a conversion factor in the form:

$$\frac{\text{g solute}}{\text{100 g solution}} \text{ or } \frac{\text{100 g solution}}{\text{g solute}}$$

For example, a solution that is $36\%\left(\frac{m}{m}\right)$ HCl means that there are 36 g HCl for every 100 g of solution and may be expressed as a conversion factor as:

$$\frac{\text{36 g HCl}}{\text{100 g solution}} \text{ or } \frac{\text{100 g solution}}{\text{36 g HCl}}$$

Therefore, if we wanted to know the mass of HCl required to make 250 g of a solution that is $36\%\left(\frac{m}{m}\right)$ HCl, a solution map may be generated, in a manner similar to when they were first introduced in this text:

$$\boxed{\text{g solution}} \xrightarrow[\text{100 g solution}]{\text{36 g HCl}} \boxed{\text{g HCl}}$$

✅ Example 14.5.2

Describe the preparation of 3000.0 g of $5.00\%\left(\frac{m}{m}\right)$ NaCl.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 3000.0 g solution; $5.00\%\left(\frac{m}{m}\right)$ NaCl Find: mass of solute and solvent required, so solution prep can be described
List other known quantities.	$\frac{5.0\text{ g NaCl}}{100\text{ g solution}}$
Prepare concept map using the proper conversion factor(s).	g solution $\xrightarrow{\frac{5.0\text{ g solute}}{100\text{ g solution}}}$ g NaCl
Calculate the answer.	$3000.0\ \cancel{\text{g solution}} \times \frac{5.00\text{ g NaCl}}{100\ \cancel{\text{g solution}}} = \boxed{150\text{ g NaCl}}$ Mass solvent (H_2O) required = 3000.0 g solution – 150 g NaCl solute = 2850 g H_2O
Describe preparation.	Dissolve 150 g NaCl into 2850 g H_2O.
Think about your result.	This makes sense, 150 g and 2850 g sum to 3000 g. In addition, 150 g is 5.00% of 3000.0 g.

Volume Percentage

Liquid volumes over a wide range of magnitudes are conveniently measured using common and relatively inexpensive laboratory equipment. The concentration of a solution formed by dissolving a liquid solute in a liquid solvent is therefore often expressed as a volume percentage, $\%\left(\frac{v}{v}\right)$:

$$\%\left(\frac{v}{v}\right) = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{14.5.3}$$

Just like the mass percentage, all volumes must be expressed in the same units to properly calculate the concentration. One important thing to note with a volume percentage is that volumes are not always additive, due to particle size and intermolecular forces. How could this be? Consider an example where 200 mL of sand were mixed with 200 mL of marbles. The combined volume would be less than 400 mL because the sand would fill the spaces between the marbles. How much less? It would depend on the size of the marbles and sand particles. While we couldn't predict the final volume, we could state with confidence it is less than 400 mL. Therefore, in a volume percentage, only the volume of the solute and volume of the solution are important.

A common usage is in the measurement ethanol content in alcoholic beverages. A beer that contains 5.8% ABV (alcohol by volume) means that it contains $5.8\%\left(\frac{v}{v}\right)$ ethanol. For liquor, proof is defined as twice the ABV. Therefore, a whiskey labeled as 80

proof, is also 40% ABV, or $40\%\left(\frac{v}{v}\right)$ ethanol.

As with the mass percentage, a volume percentage may also be expressed as a conversion factor in the form:

$$\frac{\text{mL solute}}{\text{100 mL solution}} \text{ or } \frac{\text{100 mL solution}}{\text{mL solute}}$$

Exercise 14.5.2

What is the volume of isopropyl alcohol found in a 1-pint (473 mL) bottle of rubbing alcohol that is $70.0\%\left(\frac{v}{v}\right)$ isopropyl alcohol?

Answer

331 mL isopropyl alcohol

Mass-Volume Concentrations

"Mixed" percentage units, derived from the mass of solute and the volume of solution, are popular for certain biochemical and medical applications. A mass-volume percentage is a ratio of a solute's mass to the solution's volume expressed as a percentage. In a **mass-volume percentage**, grams are used for solute mass and mL are used for the solution volume. For example, physiological saline solution, used to prepare intravenous fluids, has a concentration of $0.9\%\left(\frac{m}{v}\right)$ $\ce{NaCl}$, indicating that the composition is 0.9 g NaCl per 100 mL solution.

The concentration of glucose in blood (commonly referred to as "blood sugar") is also typically expressed in terms of a **mass-volume ratio**. Though not expressed explicitly as a percentage, blood glucose concentration is usually given in milligrams of glucose per deciliter (mg/dL) (1 dL = 100 mL) of blood (Figure 14.5.2).



Figure 14.5.2: "Mixed" mass-volume units are commonly encountered in medical settings. (a) The NaCl concentration of physiological saline is $0.9\%\left(\frac{m}{v}\right)$ $\ce{NaCl}$. (b) This device measures glucose levels in a sample of blood. The normal range for glucose concentration in blood (fasting) is around 70–100 mg/dL. (credit a: modification of work by "The National Guard"/Flickr; credit b: modification of work by Biswarup Ganguly).

Parts per Million and Parts per Billion

Very low solute concentrations are often expressed using appropriately small units such as parts per million (ppm) or parts per billion (ppb). Like percentage ("part per hundred") units, ppm and ppb may be defined in terms of masses, volumes, or mixed mass-volume units. There are also ppm and ppb units defined with respect to numbers of atoms and molecules.

The mass-based definitions of ppm and ppb are given here:

$$\text{ppm} = \frac{\text{mass solute}}{\text{mass solution}} \times 1,000,000 \tag{14.5.4}$$

$$\text{ppb} = \frac{\text{mass solute}}{\text{mass solution}} \times 1,000,000,000 \tag{14.5.5}$$

Both ppm and ppb are convenient units for reporting the concentrations of pollutants and other trace contaminants in water. Concentrations of these contaminants are typically very low in treated and natural waters, and their levels cannot exceed relatively low concentration thresholds without causing adverse effects on health and wildlife. For example, the EPA has identified the

maximum safe level of fluoride ion in tap water to be 4 ppm. Inline water filters are designed to reduce the concentration of fluoride and several other trace-level contaminants in tap water (Figure 14.5.3).



(a) (b)

Figure **14.5.3**: (a) In some areas, trace-level concentrations of contaminants can render unfiltered tap water unsafe for drinking and cooking. (b) Inline water filters reduce the concentration of solutes in tap water. (credit a: modification of work by Jenn Durfey; credit b: modification of work by "vastateparkstaff"/Wikimedia commons).

Exercise 14.5.3

Match the measurement with the quantity it is likely measuring:

A. cholesterol levels in the blood
B. fluoride content of toothpaste
C. ethanol content of wine
D. mercury content in drinking water

Choices:

- 180 mg/dL
- 0.8 ppb
- $0.15\%\left(\frac{m}{v}\right)$
- 12% ABV

Answer A

180 mg/dL is a typical measure of cholesterol levels in the blood.

Answer B

$0.15\%\left(\frac{m}{v}\right)$ represents the fluoride content of a typical toothpaste.

Answer C

12% ABV is the typical ethanol content of wine.

Answer D

0.8 ppb would likely be a typical mercury content in drinking water.

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14.6: Molarity

Learning Objectives

- Use molarity as a conversion factor in calculations.

Of all the quantitative measures of solution concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{Liters of solution}} \tag{14.6.1}$$

This means that molarity may be expressed in units of mol/L, though the abbreviation M (must be capitalized) is used more often to abbreviate molarity. Equation 14.6.1 may be written more simply as:

$$\text{M} = \frac{\text{mol}}{\text{L}} \tag{14.6.2}$$

Perhaps you have used a solution of 0.50 M HCl in lab. This concentration is read as "0.50 molar HCl" and means that there is 0.50 mol HCl dissolved for every one liter of solution. This concentration may also be written as 0.50 mol/L HCl. It is important to remember that “mol” in this expression refers to moles of *solute* and that “L” refers to liters of *solution*. In some instances, the molarity of a solution is instead called its **molar concentration**.

Units of molarity are useful for discussing chemical reactions in which a solute is a product or reactant. Molar masses may then be used as conversion factors to convert amounts in moles to amounts in grams. Sometimes chemists use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[Ag^+]$ refers to the molarity of the silver ion in solution.

Kool-Aid

It turns out that making a solution of known molarity is similar to the preparation of Kool-Aid (see Figure 14.6.1 below) in that the total amount of *solution* is specified rather than the amount of *solvent*. Note that the front side of the Kool-Aid package reads, "Makes 2 Quarts," while the back side instructs one to add all of the other ingredients first before adding water to make a total of two quarts of Kool-Aid. If instead, two quarts of water had been added to the other ingredients, more than two quarts of Kool-Aid would be prepared, and the drink would taste watered down. The procedure for making Kool-Aid and its resemblance to preparing solutions of known molarity lead many students to affectionately call it the "Kool-Aid Rule".



Figure **14.6.1**: The front and back sides of a package of unsweetened Kool-Aid.

Important Notes About Molarity

- The capital letter M is used for *molarity*, not a lower case m which is for a different concentration unit called *molality*.
- Molarity is calculated from the volume of *solution*, **not** the volume of *solvent*! The solute contributes to the total volume of the solution.
- The *molarity* of a solution is the same as its *molar concentration*.
- Square brackets are often used to represent molar concentration. Example: $[NH_3]$ = 0.50 M.

Before getting into more involved examples, let's look at the preparation of 1.00 L of a 1.00 M $NaHCO_3$ solution. A 1.00 M $NaHCO_3$ solution contains 1 mol $NaHCO_3$, or 84.0 g $NaHCO_3$ for every 1 L of solution (if necessary, review molar mass calculations). To make 1.00 L of this solution, a device that measures volume precisely will be needed. One option is to use a large graduated cylinder capable of measuring 1000 mL. However, there is a specific piece of equipment designed for this purpose that measures one volume very precisely called a *volumetric flask*. Volumetric flasks come in various sizes, so a 1000-mL volumetric flask would be selected to prepare 1.00 L of solution.

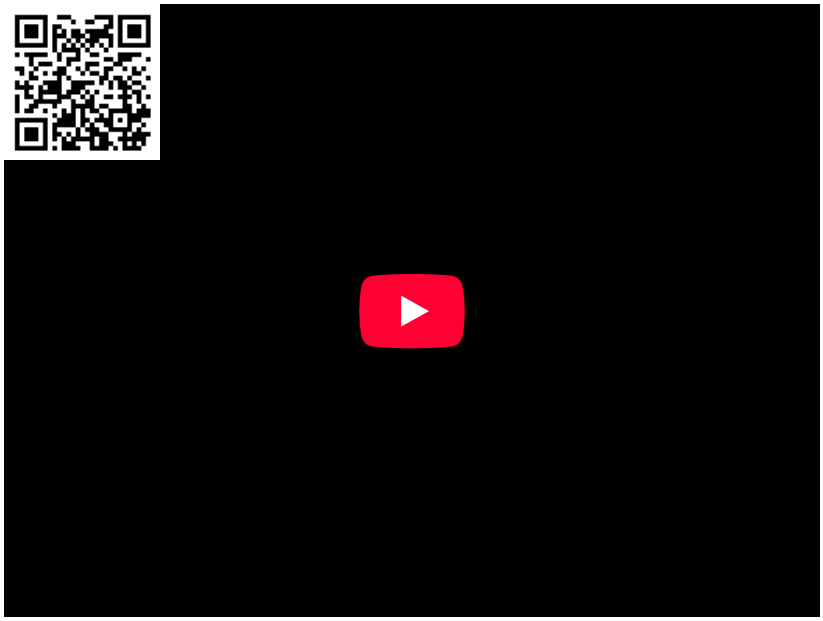


Figure **14.6.2**: An assortment of volumetric flasks. (Nuno Nogueira via Wikimedia Commons)

Next, 84.0 g $NaHCO_3$ would be weighed on a scale and added to the volumetric flask. Since solid residue could remain on the measuring device, it should be rinsed with water and added to the volumetric flask. Water is then added to the flask until it is ½ to ¾ full. The contents of the flask are then mixed thoroughly until all of the solid

dissolves. Recall that volumes are not always additive due to particle size and intermolecular forces, so make sure that the solute completely dissolves before adding water to the 1000-mL etched mark on the neck of the flask. Finally, water would be added to the 1000-mL mark on the neck and the solution would be thoroughly mixed.

Video 14.6.1 below shows how to make 1.00 L of a 1.00 M NaCl solution. Can you tell an important step that was skipped near the end of the video?



Video **14.6.1**: Making a 1.00 M NaCl solution. One important step was skipped at the very end. What is it? (CAYERCHEM via YouTube)

You would be correct if you said they forgot to mix the solution thoroughly after the last portion of water was added to the flask. It is important that solutions are mixed thoroughly to obtain a uniform concentration of solute throughout the solution.

✅ Example 14.6.1

What is the molarity of 2.50 L of solution that contains 62.6 g $Mg(NO_3)_2$?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 62.6 g $Mg(NO_3)_2$; 2.50 L solution Find: Molarity = ? M
List known relationship(s).	1 mol $Mg(NO_3)_2$ = 148.33 g $Mg(NO_3)_2$ $$M = \frac{mol}{L}$$
Prepare a concept map and use the proper conversion factor(s).	Convert mass $Mg(NO_3)_2$ to moles: $$\boxed{\text{g } Mg(NO_3)_2} \xrightarrow{\frac{1 \text{ mol } Mg(NO_3)_2}{148.33 \text{ g } Mg(NO_3)_2}} \boxed{\text{mol } Mg(NO_3)_2}$$ Molarity is calculated by dividing moles by liters: $$M = \frac{\text{mol } Mg(NO_3)_2}{2.50 \text{ L}}$$
Calculate the answer.	Now substitute the known quantities and solve. $$62.6 \cancel{\text{ g } Mg(NO_3)_2} \times \frac{1 \text{ mol } Mg(NO_3)_2}{148.33 \cancel{\text{ g } Mg(NO_3)_2}} = 0.4220 \text{ g } Mg(NO_3)_2$$ $$M = \frac{0.4220 \text{ mol } Mg(NO_3)_2}{2.50 \text{ L}} = 0.169 \frac{\text{mol } Mg(NO_3)_2}{\text{L}} = \boxed{0.169 \text{ M } Mg(NO_3)_2}$$
Think about your result.	The molarity is 0.169 M, meaning that one liter of the solution would contain 0.169 mol $Mg(NO_3)_2$. Three significant figures are appropriate.

Exercise 14.6.1

A. Calculate the molar concentration of 0.500 L of solution in which 137 g NaCl are dissolved.
B. What is the molarity of a solution that contains 66.2 g $C_6H_{12}O_6$ dissolved in 235 mL of solution?

Answer A

4.69 M NaCl

Answer B

1.56 M $C_6H_{12}O_6$

Using Molarity as a Conversion Factor

Molarity may be used as a conversion factor to convert between moles of solute and liters of solution. As such, molarity can be useful in a variety of stoichiometry problems. For example, suppose we would like to know how many moles of NaCl are present in 108 mL of 0.887 M NaCl. Since molarity is defined as the moles of solute per liter of solution, this means that there are 0.887 mol NaCl for every 1 L of solution. However, the volume of solution is given in mL. Therefore, mL must first be converted to L or L must be converted to mL. As we recall, 1 L = 1000 mL, so 0.887 M NaCl may be expressed as a conversion factor in the form:

$$\frac{0.887 \text{ mol NaCl}}{1 \text{ L solution}} = \frac{0.887 \text{ mol NaCl}}{1000 \text{ mL solution}} \text{ or } \frac{1 \text{ L solution}}{0.887 \text{ mol NaCl}} = \frac{1000 \text{ mL solution}}{0.887 \text{ mol NaCl}}$$

The conversion that enables cancellation of units would then be used in the solution map that takes us from mL of solution to mol NaCl. With the solution map in place, the moles of NaCl are easily calculated.

$$\boxed{\text{mL solution}} \xrightarrow{\frac{0.887 \text{ mol NaCl}}{1000 \text{ mL solution}}} \boxed{\text{mol NaCl}}$$

$$108 \cancel{\text{ mL solution}} \times \frac{0.887 \text{ mol NaCl}}{1000 \cancel{\text{ mL solution}}} = 0.0958 \text{ mol NaCl}$$

Had the equation approach been used instead, the same answer would be obtained. Since:

$$\text{M} = \frac{\text{mol}}{\text{L}}$$

this equation could be rearranged to solve for moles by multiplying each side by liters of solution. The result of this rearrangement is:

$$\text{mol} = \text{M} \times \text{L} \tag{14.6.3}$$

Note that using the equations requires that volumes be expressed in liters, so one would first convert 108 mL to 0.108 L, then plug the known quantities into the equation:

$$\text{mol} = 0.887 \text{ M NaCl} \times 0.108 \text{ L} = 0.887 \frac{\text{mol NaCl}}{\cancel{\text{L}}} \times 0.108 \cancel{\text{ L}} = 0.0958 \text{ mol NaCl}$$

Whether using the conversion factor method or the equation method, it is still necessary to remember that M stands for mol/L, so the proper cancellation of units may be performed. However, once again, we can see that the same answer is obtained using either method.

What if we were looking for volume and wanted to know how many liters of 2.35 M $CuSO_4$ were needed to obtain 4.88 mol of $CuSO_4$? A molarity of 2.35 M $CuSO_4$ means that there are 2.35 mol $CuSO_4$ for every 1 L of solution. As such, 2.35 M $CuSO_4$ may be expressed as a conversion factor in the form:

$$\frac{2.35 \text{ mol CuSO}_4}{1 \text{ L solution}} \text{ or } \frac{1 \text{ L solution}}{2.35 \text{ mol CuSO}_4}$$

The conversion that enables cancellation of units would then be used in the solution map that takes us from mol $CuSO_4$ to L solution. With the solution map in place, the liters of solution are easily calculated.

$$\boxed{\text{mol CuSO}_4} \xrightarrow{\frac{1 \text{ L solution}}{2.35 \text{ mol CuSO}_4}} \boxed{\text{L solution}}$$

$$4.88 \cancel{\text{ mol CuSO}_4} \times \frac{1 \text{ L solution}}{2.35 \cancel{\text{ mol CuSO}_4}} = 2.08 \text{ L solution}$$

✅ Example 14.6.2

A chemist needs to prepare 250.0 mL of a 0.150 M solution of potassium permanganate, $KMnO_4$.

A. What mass of $KMnO_4$ does she need to make the solution?
B. Describe how this solution would be prepared.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 0.150 M $KMnO_4$; 250.0 mL solution Find: Mass $KMnO_4$ = ? g $KMnO_4$
List known relationship(s).	1 mol $KMnO_4$ = 158.04 g $KMnO_4$ 0.150 mol $KMnO_4$ for every 1 L (1000 mL) solution
Prepare a concept map and use the proper conversion factor(s).	$\boxed{\text{mL solution}} \xrightarrow{\frac{0.150 \text{ mol KMnO}_4}{1000 \text{ mL solution}}} \boxed{\text{mol KMnO}_4} \xrightarrow{\frac{158.04 \text{ g KMnO}_4}{1 \text{ mol KMnO}_4}}$
Calculate the answer.	Now substitute the known quantities and solve. $250.0 \cancel{\text{ mL solution}} \times \frac{0.150 \cancel{\text{ mol KMnO}_4}}{1000 \cancel{\text{ mL solution}}} \times \frac{158.04 \text{ g KMnO}_4}{1 \cancel{\text{ mol KMnO}_4}} = \boxed{5.93 \text{ g KMnO}_4}$
Think about your result.	When 5.93 g $KMnO_4$ are used to make 250.0 mL of solution, the molarity is 0.150 M.

B. To prepare the solution, weigh out 5.93 g $KMnO_4$ solid on a scale. Place the solid in a 250-mL volumetric flask. Add water to flask and mix to dissolve the solid. Add water until the bottom of the meniscus is on the line on the neck. Mix thoroughly.

✅ Example 14.6.3

What volume of a 0.200 M NaOH solution would contain 16.0 g NaOH?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: 0.200 M NaOH; 16.0 g NaOH Find: Volume of Solution = ? L
List known relationship(s).	1 mol NaOH = 40.00 g NaOH 0.200 mol NaOH for every 1 L solution
Prepare a concept map and use the proper conversion factor(s).	$\boxed{\text{g NaOH}} \xrightarrow{\frac{1\text{ mol NaOH}}{40.00\text{ g NaOH}}} \boxed{\text{mol NaOH}} \xrightarrow{\frac{1\text{ L solution}}{0.200\text{ mol NaOH}}} \boxed{\text{L so}}$
Calculate the answer.	Now substitute the known quantities and solve. $16.0\ \cancel{\text{g NaOH}} \times \frac{1\ \cancel{\text{mol NaOH}}}{40.00\ \cancel{\text{g NaOH}}} \times \frac{1\text{ L solution}}{0.200\ \cancel{\text{mol NaOH}}} = \boxed{2.00\text{ L solution}}$
Think about your result.	When 16.0 g NaOH are used to make 2.00 L of solution, the molarity is 0.200 M.

✏️ Exercise 14.6.3

A. What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?
B. What is the volume of 0.0444 M CH_2O solution that contains 0.0773 mol CH_2O?
C. What volume of 1.50 M HCl solution, in mL, contains 10.0 g of hydrogen chloride?

Answer A

6.12 g H_2SO_4

Answer B

1.74 L

Answer C

183 mL or 0.183L

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14.7: Solution Dilution

Learning Objectives

- Explain how concentrations can be changed in the lab.
- Understand how stock solutions are used in the laboratory.

Muriatic acid, another name for HCl (*aq*), is widely used for cleaning concrete and masonry surfaces (see Figure 14.7.1). The acid must be diluted before use to get it down to a safer strength. Commercially available at concentrations of about 18%, muriatic acid may be used to remove scales and deposits (usually composed of basic materials).

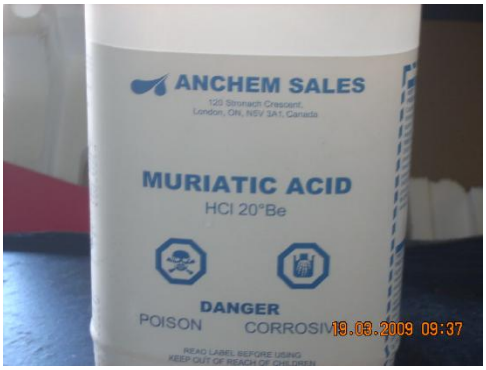


Figure **14.7.1** (left to right): A bottle of muriatic acid, another name for HCl (*aq*) (SLBB via Flickr). Air Force Staff Sgt. Marvin Martin cleans a concrete form (Cpl. Ryan R. Jackson via Wikimedia Commons).

Stock Solutions

It is often necessary to have a solution with a concentration that is very precisely known. A solution containing a precise mass of solute in a precise volume of solution is called a **stock solution** (or standard solution). A volumetric flask is usually used to prepare a stock solution. **Aliquots** (carefully measured volumes) of the stock solution can then be diluted to any desired volume. In other cases, it may be inconvenient to weigh a small mass of sample accurately enough to prepare a small volume of a dilute solution. Instead, a more concentrated solution is first prepared. An aliquot of this solution is then diluted to the desired concentration.

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. This is because the amount of solute remains the same, while the volume of the solution increases (see Figure 14.7.2).

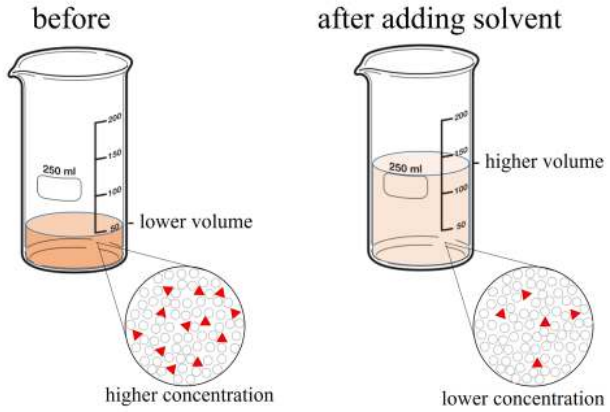


Figure 14.7.2: A solution before and after dilution. The volume increases upon dilution and the solute concentration decreases. (Theislikerice, via Wikimedia Commons)

Once again, the amount of solute remains the same, so the moles of solute before dilution (1) are equal to the moles of solute after dilution (2):

$$\text{mol}_1 = \text{mol}_2 \tag{14.7.1}$$

From the previous section, the moles of solute in a solution are equal to the molarity multiplied by the liters (mol = $M \times L$), so:

$$M_1 \times L_1 = M_2 \times L_2 \tag{14.7.2}$$

Since the two sides of the equation are equal, the volume (V) may be any unit we choose, as long as the units on each side are the same, meaning:

$$M_1 \times V_1 = M_2 \times V_2 \tag{14.7.3}$$

Likewise, the concentration, (C) may be any unit we choose, as long as the concentration unit is based on the total volume of the solution and the unit on each side is the same. This yields the most general form of the dilution equation:

$$C_1V_1 = C_2V_2 \tag{14.7.4}$$

To summarize, C stands for the concentration of the solution, V for the volume of the solution, a subscript of 1 means *before* dilution, and a subscript of 2 means *after* dilution. Preparing dilutions is a common activity in the chemistry lab and elsewhere. Once you understand this relationship, the calculations are simple.

Suppose there are 100. mL of a 2.0 M solution of HCl available. The solution is diluted by adding enough water to make 500. mL of solution. The new molarity is easily calculated by rearranging the dilution equation and solving for C_2.

$$\frac{C_1V_1}{V_2} = \frac{C_2\cancel{V_2}}{\cancel{V_2}} \quad \Rightarrow \quad C_2 = \frac{C_1V_1}{V_2} = \frac{(2.0\text{ M HCl})(100\ \cancel{\text{mL}})}{500\ \cancel{\text{mL}}} = 0.40\text{ M HCl}$$

It should make sense that if the original volume is ⅕ of the final volume, then the diluted solution should be ⅕ of its original concentration.



Video 14.7.1: How to dilute a solution. (Carolina Biological via YouTube)

✅ Example 14.7.1

What is the concentration of a solution prepared by diluting 10.00 mL of a $20.0\%\left(\frac{m}{v}\right)\ C_6H_{12}O_6$ stock solution to a final volume of 200.0 mL?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $C_1 = 20.0\%\left(\frac{m}{v}\right)\ C_6H_{12}O_6$; $V_1 = 10.00$ mL; $V_2 = 200.0$ mL Find: C_2
List known relationship(s).	Dilution equation: $C_1V_1 = C_2V_2$
Rearrange the dilution equation and solve for C_2.	$\frac{C_1V_1}{V_2} = \frac{C_2\cancel{V_2}}{\cancel{V_2}} \Rightarrow C_2 = \frac{C_1V_1}{V_2}$
Calculate the answer.	Plug the known quantities into the equation: $C_2 = \frac{C_1V_1}{V_2} = \frac{\left(20.0\%\left(\frac{m}{v}\right)\ C_6H_{12}O_6\right)\left(10.00\ \cancel{mL}\right)}{200.0\ \cancel{mL}}$ $= \boxed{1.00\%\left(\frac{m}{v}\right)\ C_6H_{12}O_6}$
Think about your result.	The volume before dilution is $\frac{1}{20}$ the final volume, so it makes sense that the concentration after dilution is $\frac{1}{20}$ its original concentration.

✅ Example 14.7.2

Nitric acid, HNO_3, is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is 16 M. How much of this stock solution, in mL, of nitric acid are needed to prepare 8.00 L of a 0.50 M solution?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: C_1 = 16 M HNO_3; C_2 = 0.50 M HNO_3; $V_2 = 8.00\ \cancel{L} \times \frac{1000\ \text{mL}}{1\ \cancel{L}} = 8000\ \text{mL}$ (V_2 converted to mL, since answer will be in mL. Volume units must be the same.) Find: V_1 (in mL)
List known relationship(s).	Dilution equation: $C_1V_1 = C_2V_2$
Rearrange the dilution equation and solve for V_1.	$\frac{\cancel{C_1}V_1}{\cancel{C_1}} = \frac{C_2V_2}{C_1} \Rightarrow V_1 = \frac{C_2V_2}{C_1}$
Calculate the answer.	Plug the known quantities into the equation: $V_1 = \frac{C_2V_2}{C_1} = \frac{(0.50\ \cancel{\text{M HNO}_3})(8000\ \text{mL})}{16\ \cancel{\text{M HNO}_3}} = \boxed{250\ \text{mL}}$
Think about your result.	The final concentration is $\frac{1}{32}$ its original concentration, so it makes sense that the volume of the stock solution used is $\frac{1}{32}$ of the final volume.

✏️ Exercise 14.7.1

A 0.885 M solution of KBr with an initial volume of 76.5 mL has more water added until its concentration is 0.500 M. What is the new volume of the solution?

Answer

135 mL

✏️ Exercise 14.7.2

A. What volume of a 6.0 ppb Hg^{2+} solution would be needed to prepare 600 mL of 0.10 ppb Hg^{2+} solution?
B. Assuming the volumes are additive, describe how the diluted solution would be prepared.

Answer A

10. mL of 6.0 ppb Hg^{2+} solution are needed.

Answer B

The solution would be prepared by thoroughly mixing 10 mL of the 6.0 ppb Hg^{2+} solution with 590 mL of water. This mixture makes 600 mL of solution.

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14.8: Colligative Properties of Solutions

Learning Objectives

- Explain what the term "colligative" means, and list the colligative properties.
- Indicate what happens to the boiling point and the freezing point of a solvent when a solute is added to it.

People who live in colder climates have seen trucks put salt on the roads when snow or ice is forecast (Figure 14.8.1), but why is this done? It turns out that the addition of a solute lowers the freezing point of a liquid, a phenomenon called freezing point depression. Freezing point depression is one the four different kinds of colligative properties of solutions. **Colligative properties** are properties of solutions that depend only on the concentration of solute particles in a solvent, but not their identity. What this means for the example above is that people in colder climates do not necessarily need salt to get the same effect on the roads – any solute will work. However, the higher the concentration of solute, the more these properties will change.



Figure **14.8.1**: Workers manually spreading salt from a salt truck in Milwaukee, WI. (Michael Pereckas via Wikimedia Commons)

The four colligative properties of solutions are:

1. freezing point depression
2. boiling point elevation
3. vapor pressure depression (not discussed in this text)
4. osmotic pressure elevation

Freezing Point Depression

The freezing point of a solution is lower than the freezing point of a pure solvent, which is why the phenomenon is called **freezing point depression**. The **freezing point** of a substance is the temperature at which the liquid changes into a solid. At a given temperature, if a substance is added to a solvent (such as water), the solute-solvent interactions prevent the solvent from transitioning into the solid phase. The solute-solvent interactions require the temperature to decrease further (slowing the particles down) in order to solidify the solution.

Salt is placed onto roads and sidewalks in the winter to prevent water from freezing at the normal temperature of 0°C and, instead, causing it to freeze at a lower temperature, sometimes as low as –10°C to –15°C. De-icing planes is another common example of freezing point depression in action. A number of solutions are used, but commonly a solution such as ethylene glycol, or a less toxic monopropylene glycol, is used to de-ice aircraft. Aircrafts are sprayed with the solution when the temperature is predicted to drop below the freezing point. Yet another example of freezing point depression is the addition of antifreeze to a car radiator. This prevents the radiator from freezing in the winter. Freezing water can cause an engine block to crack, since water expands upon freezing.

Freezing point depression is observed for any solute added to a solvent; the freezing point of the solution will always be lower than the freezing point of the pure solvent (without the solute). Thus, when anything is dissolved in water, the solution will freeze at a lower temperature than pure water.

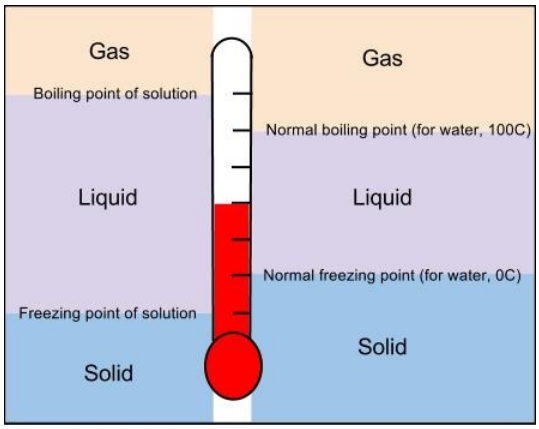


Figure **14.8.2**: Comparison of boiling and freezing points of a solution (left side) with a pure liquid (right side). (Tracy Poulsen via Flickr)

Boiling Point Elevation

Pure water boils at 100°C at 1 atm of pressure. When table salt is added to water, the resulting solution has a higher boiling point than pure water. Since the addition of a solute increases the boiling point, the colligative property is called **boiling point elevation**. In the case of salt, the boiling point increases because the ions form an attraction with the water molecules that prevents them from escaping as easily into the gaseous phase. In order for a saltwater solution to boil, the temperature must be raised above 100°C at 1 atm of pressure. Boiling point elevation is observed for any nonvolatile solute (one that does not evaporate easily) that is added to a solvent; the boiling point of a solution will always be higher than the boiling point of the pure liquid.

One such application of boiling point elevation is the addition of antifreeze to a car radiator. Interestingly enough, the term *antifreeze* is rarely heard during the summer or in warmer climates. The term *coolant* is often used instead. When it comes to vehicles, antifreeze and coolant are the same thing – they are the same chemical(s); it just depends on the season or whether a person lives in a colder climate or warmer climate. While antifreeze/coolant will prevent water from freezing when the ambient temperature is very cold, it will also prevent the water in the radiator from boiling over when the ambient temperature is very warm. Should the water in a radiator and engine block boil away, it will cause the engine to overheat causing certain parts attached directly to the engine to warp or gaskets to blow out.

Osmotic Pressure Elevation

Before the final colligative property is introduced, a new concept needs to be presented. A **semipermeable membrane** is a thin membrane that allows certain small molecules to pass through, but not others. A thin sheet of cellophane, for example, acts as a semipermeable membrane (see Figure 14.8.3(a)).

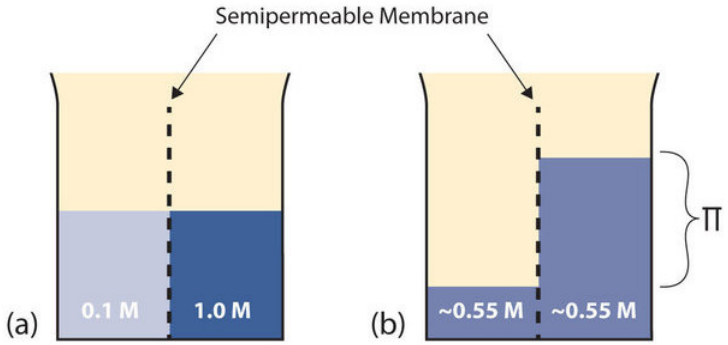


Figure **14.8.3**: Osmosis. (a) Two aqueous solutions of differing concentrations are placed on either side of a semipermeable membrane. (b) When osmosis occurs, water molecules pass through the membrane from the dilute solution to the concentrated solution until the two concentrations on either side of the membrane are the same. The pressure exerted by the different height of the solution on the right is called the osmotic pressure. (CC BY-SA-NC 3.0; anonymous)

The semipermeable membrane is used to separate two solutions having the different concentrations marked. Curiously, this situation is not stable; there is a tendency for water molecules to move from the dilute side (on the left) to the concentrated side (on the right) until the concentrations are equalized, as in Figure 14.8.3(b). This tendency is called **osmosis**. In osmosis, the solute is unable to pass through the membrane and remains on its original side of the system; only solvent molecules move through the semipermeable membrane. In the end, the two sides of the system will have different volumes. Because a column of liquid exerts a pressure, there is a pressure difference (Π) between the two sides of the system that is proportional to difference in height between the two columns. This pressure difference is called the **osmotic pressure**.

Osmotic pressure plays an important role in biological systems because cell walls are semipermeable membranes. In particular, when a person is receiving intravenous (IV) fluids, the osmotic pressure of the fluid needs to be approximately the same as blood serum to avoid any negative consequences. Figure 14.8.4 shows three red blood cells:

- A healthy red blood cell.
- A red blood cell that has been exposed to a lower solute concentration than normal blood serum (a *hypotonic* solution); the cell has plumped up as solvent moves into the cell to dilute the solutes inside.
- A red blood cell exposed to a higher solute concentration than normal blood serum (*hypertonic*); water leaves the red blood cell, so it collapses onto itself. Only when the solutions inside and outside the cell are the same (*isotonic*) will the red blood cell be able to do its job.

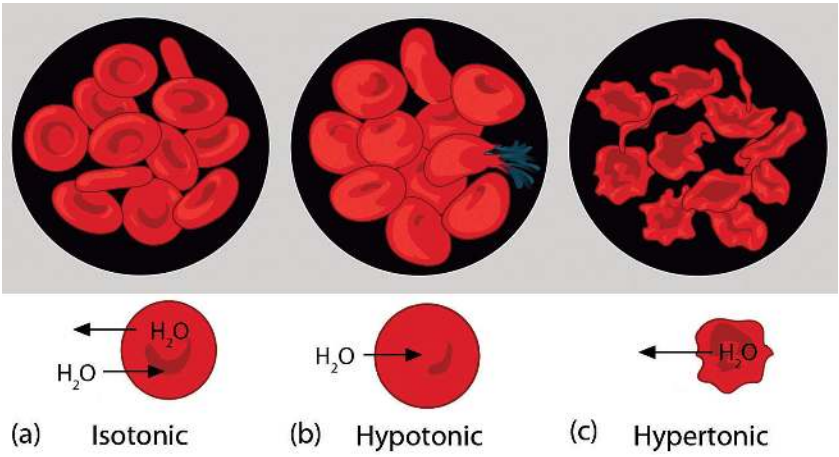


Figure **14.8.4**: Osmotic Pressure and Red Blood Cells. (a) This is what a normal red blood cell looks like. (b) When a red blood cell is exposed to a hypotonic solution, water passes through the cell membrane diluting the cell contents, potentially rupturing. (c) When a red blood cell is exposed to a hypertonic solution, water passes from the cell to the surrounding solution, diluting the hypertonic solution and collapsing the cell. Neither of these last two cases is desirable, so IV solutions must be isotonic with blood serum to prevent deleterious effects. (Public Domain; Mariana Ruiz Villareal)

Osmotic pressure is also the reason you should not drink seawater if you're stranded on a lifeboat in the ocean; seawater has a higher osmotic pressure than most of the fluids in your body. You *can* drink the water, but ingesting it will pull water out of your cells as osmosis works to dilute the seawater. Ironically, your cells will die of thirst, and you will also die. (It is okay to drink the water if you are stranded on a body of freshwater, at least from an osmotic pressure perspective.) Osmotic pressure is also thought to be important – in addition to capillary action – in getting water to the tops of tall trees.

Summary

- Colligative properties are properties that are due only to the number of particles in solution, and are not related to the chemical properties of the solute.
- Freezing points of solutions are lower than the freezing points of the pure solvents.
- Boiling points of solutions are higher than the boiling points of the pure solvents.
- Osmotic pressure is caused by concentration differences between solutions separated by a semipermeable membrane, and is an important biological consideration.

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14.10: Exercises

14.2: Solution Terminology

1. One brand of ethyl alcohol (Everclear) is 95% ethyl alcohol, with the remaining 5% being water. What is the solvent and what is the solute of this solution?

Answer

The solvent is ethyl alcohol and the solute is water.

2. Give an example of each type of solution from your own experience.
   a. A solution composed of a gas solute in a liquid solvent.
   b. A solution composed of a solid solute in a liquid solvent.
   c. A solution composed of a liquid solute in a liquid solvent.
   d. A solution composed of a solid solute in a solid solvent. (Hint: usually such solutions are made as liquids and then solidified.)

Answer

Answers will vary. Possible answers include:

a. carbonated beverages
b. sugar water
c. alcoholic beverages
d. stainless steel

3. Differentiate between the terms *unsaturated* and *dilute*.

Answer

Unsaturated means that less than the maximum amount of solute is present in solution. Dilute means a small amount of solute is present. A solution can be unsaturated and still have a large amount of solute present.

4. Differentiate between the terms *saturated* and *concentrated*.

Answer

Saturated means that the maximum amount of solute is present in solution. Concentrated means that a large amount of solute is present. A solute can be concentrated and still have less than the maximum amount of solute possible. There are also weakly soluble compounds which can be saturated and have only a small amount of solute present in solution.

14.3: Aqueous Solutions of Solids

5. Which of the following compound(s) will dissolve in water to yield ions?
   a. KOH
   b. $FeCl_3$
   c. CH_3COCH_3 (acetone)

Answer

KOH and $FeCl_3$

6. Which of the following compound(s) are electrolytes?
    a. CH_3OH
    b. Li_2SO_4
    c. $C_6H_{12}O_6$ (gluose)

Answer

Li_2SO_4

7. Using Figure 14.3.4, how many grams of the indicated solute would dissolve in 100 g H_2O at the given temperature?
    a. NH_3 at 85°C
    b. $KClO_3$ at 45°C
    c. HCl at 60°C
    d. KI at 10°C

Answer

    a. 13 g
    b. 20 g
    c. 55 g
    d. 136 g

8. Using Figure 14.3.4, indicate whether each of the following solutions is saturated, unsaturated, or supersaturated.
    a. 80 g NH_4Cl in 100 g H_2O at 90°C
    b. 120 g KNO_3 in 200 g H_2O at 50°C
    c. 95 g $NaNO_3$ in 100 g H_2O at 40°C
    d. 30 g KCl in 50 g H_2O at 60°C

Answer

    a. supersaturated
    b. unsaturated
    c. unsaturated
    d. supersaturated

9. Based on the information in Figure 14.3.4, what temperature would be needed to dissolved 10 g SO_2 in 100 g H_2O?

Answer

20°C or colder

14.5: Solution Concentration Measurements

10. The World Health Organization recommends that the maximum fluoride ion concentration in drinking water is 1.0 ppm. Assuming water has the maximum concentration, if an average person drinks 1,920 mL of water per day, how many milligrams of fluoride ion are being ingested? Assume the density of drinking water is 1.00 g/mL.

Answer

1.9 mg F^-

11. For sanitary reasons, water in pools should be chlorinated to a maximum level of 3.0 ppm. In a typical 5,000 gal pool that contains 21,200 kg of water, what mass of chlorine must be added to obtain this concentration?

Answer

64 g

12. Given its notoriety, you might think that uranium is very rare, but it is present at about 2–4 ppm of the earth's crust, which is more abundant than silver or mercury. If the earth's crust is estimated to have a mass of 8.50×10^{20} kg, what range of mass is thought to be uranium in the crust?

Answer

1.7×10^{15} - 3.4×10^{15} kg

13. Chromium is thought to be an ultratrace element, with about 8.9 ng present in a human body. If the average body mass is 75.0 kg, what is the concentration of chromium in the body in pptr (parts per trillion)?

Answer

1.2 pptr

14. What mass of nitric acid is needed to prepare 400.0 g of a 10.0% solution of HNO_3 by mass?

Answer

40.0 g HNO_3

15. What mass of a 4.00% NaOH solution by mass contains 15.0 g of NaOH?

Answer

375 g solution

16. What mass of solid NaOH is required to prepare 1.00 L of a 10.0% solution of NaOH by mass? The density of the 10.0% solution is 1.109 g/mL. (Hint: First find the total mass of the solution.)

Answer

111 g NaOH

17. What mass of HCl is contained in 45.0 mL of an aqueous HCl solution that has a density of 1.19 g/mL and contains 37.21% HCl by mass?

Answer

19.9 g HCl

18. Copper(I) iodide (CuI) is often added to table salt as a dietary source of iodine. How many moles of CuI are contained in 1.00 lb (454 g) of table salt containing 0.0100% CuI by mass?

Answer

2.38×10^{-4} mol CuI

19. A cough syrup contains 5.0% ethyl alcohol, C_2H_5OH, by mass. If the density of the solution is 0.9928 g/mL, how many grams of alcohol are in 20.0 mL of the cough syrup, the typical adult dose?

Answer

0.99 g alcohol

20. D5W is a solution used as an intravenous fluid. It is a 5.0% by mass solution of dextrose $C_6H_{12}O_6$ in water. If the density of D5W is 1.029 g/mL, how many grams of dextrose are needed to make 500.0 mL of solution?

Answer

26 g dextrose

14.6: Molarity

21. What information do we need to calculate the molarity of a sulfuric acid solution?

Answer

We need to know the number of moles of sulfuric acid dissolved in the solution and the volume of the solution.

22. What does it mean when we say that a 200 mL sample and a 400 mL sample of a solution of salt have the same molarity? In what ways are the two samples identical? In what ways are these two samples different?

Answer

The two solutions have the same ratio of moles of solute to volume of solution. They have the same concentration. They are different because they have different volumes which results in the 400 mL solution have twice as much solute as the 200 mL solution.

23. Determine the molarity for each of the following solutions:
    a. 0.444 mol of $CoCl_2$ in 0.654 L of solution
    b. 98.0 g of phosphoric acid, H_3PO_4, in 1.00 L of solution
    c. 0.2074 g of calcium hydroxide, $Ca(OH)_2$, in 40.00 mL of solution
    d. 7.0×10^{-3} mol of I_2 in 100.0 mL of solution
    e. 1.8×10^{4} mg of HCl in 0.075 L of solution

Answer

a. 0.679 M
b. 1.00 M
c. 0.06998 M
d. 0.070 M

e. 6.6 M

24. Determine the molarity of each of the following solutions:

a. 1.457 mol KCl in 1.500 L of solution
b. 0.515 g of H_2SO_4 in 1.00 L of solution
c. 20.54 g of $Al(NO_3)_3$ in 1575 mL of solution
d. 0.005653 mol of Br_2 in 10.00 mL of solution
e. 0.000889 g of glycine, $C_2H_5NO_2$, in 1.05 mL of solution

Answer

a. 0.9713 M
b. 5.25×10^{-3} M
c. 0.06122 M
d. 0.5653 M
e. 0.0113 M

25. What mass of $FeCl_2$ is present in 445 mL of 0.0812 M $FeCl_2$ solution?

Answer

4.58 g

26. What mass of SO_2 is present in 26.8 L of 1.22 M SO_2 solution?

Answer

2.09×10^3 g

27. What volume of 0.225 M $Ca(OH)_2$ solution is needed to deliver 100.0 g of $Ca(OH)_2$?

Answer

6.00 L

28. What volume of 12.0 M HCl solution is needed to obtain exactly 1.000 kg of HCl?

Answer

2.29 L

29. What is the mass of the solute in 0.500 L of 0.30 M glucose, $C_6H_{12}O_6$, used for intravenous injection?

Answer

27 g

30. What is the mass of solute in 200.0 L of a 1.556 M solution of KBr?

Answer

3.703×10^4 g

31. Calculate the number of moles and the mass of the solute in each of the following solutions:

a. 2.00 L of 18.5 M H_2SO_4, concentrated sulfuric acid
b. 100.0 mL of 3.8×10^{-5} M NaCN, the minimum lethal concentration of sodium cyanide in blood serum
c. 5.50 L of 13.3 M H_2CO, the formaldehyde used to “fix” tissue samples
d. 325 mL of 1.8×10^{-6} M $FeSO_4$, the minimum concentration of iron sulfate detectable by taste in drinking water

Answer

a. 37.0 mol; 3.63×10^3 g
b. 3.8×10^{-6} mol; 1.9×10^{-4} g
c. 73.2 mol; 2.20×10^3 g
d. 5.9×10^{-7} mol; 8.9×10^{-5} g

32. Calculate the number of moles and the mass of the solute in each of the following solutions:

a. 0.2500 L of 0.1135 M K_2CrO_4, an analytical reagent used in iron assays
b. 10.5 L of 3.716 M $(NH_4)_2SO_4$, a liquid fertilizer
c. 325 mL of 8.23×10^{-5} M KI, a source of iodine in the diet
d. 75.0 mL of 2.2×10^{-5} M H_2SO_4, a sample of acid rain

Answer

a. 0.02838 mol; 5.510 g
b. 39.0 mol; 5.16×10^3 g
c. 2.67×10^{-5} mol; 4.44×10^{-3} g
d. 1.7×10^{-6} mol; 1.6×10^{-4} g

33. What is the molarity of $KMnO_4$ in a solution of 0.0908 g of $KMnO_4$ in 0.500 L of solution?

Answer

1.15×10^{-3} M

34. What is the molarity of HCl if 35.23 mL of a solution of HCl contain 0.3366 g of HCl?

Answer

0.2621 M

35. Calculate the molarity of each of the following solutions:

a. 0.195 g of cholesterol, $C_{27}H_{46}O$, in 0.100 L of serum, the average concentration of cholesterol in human serum
b. 4.25 g of NH_3 in 0.500 L of solution, the concentration of NH_3 in household ammonia
c. 1.49 kg of isopropyl alcohol, C_3H_7OH, in 2.50 L of solution, the concentration of isopropyl alcohol in rubbing alcohol
d. 0.029 g of I_2 in 0.100 L of solution, the solubility of I_2 in water at 20 °C

Answer

a. 5.04×10^{-3} M
b. 0.499 M
c. 9.92 M
d. 1.1×10^{-3} M

36. Calculate the molarity of each of the following solutions:
    a. 293 g HCl in 666 mL of solution, a concentrated HCl solution
    b. 2.026 g $FeCl_3$ in 0.1250 L of a solution used as an unknown in general chemistry laboratories
    c. 0.001 mg Cd^{2+} in 0.100 L, the maximum permissible concentration of cadmium in drinking water
    d. 0.0079 g $C_7H_5SNO_3$ in one ounce (29.6 mL), the concentration of saccharin in a diet soft drink

Answer

12.1 M

0.09993 M

9×10^{-8} M

1.5×10^{-3} M

37. There is about 1.0 g of calcium, as Ca^{2+}, in 1.0 L of milk. What is the molarity of Ca^{2+} in milk?

Answer

0.025 M

38. A 2.00-L bottle of a solution of concentrated HCl was purchased for the general chemistry laboratory. The solution contained 868.8 g of HCl. What is the molarity of the solution?

Answer

11.9 M

39. What volume of a 0.20 M K_2SO_4 solution contains 57 g of K_2SO_4?

Answer

1.6 L

14.7: Solution Dilution

40. Explain what changes and what stays the same when 1.00 L of a solution of NaCl is diluted to 1.80 L.

Answer

The number of moles of NaCl remains the same. The volume of the solution increases.

41. What volume of a 1.00 M $Fe(NO_3)_3$ solution can be diluted to prepare 1.00 L of a solution with a concentration of 0.250 M?

Answer

0.250 L

42. If 0.1718 L of a 0.3556 M C_3H_7OH solution is diluted to a concentration of 0.1222 M, what is the volume of the resulting solution?

Answer

0.4999 L

43. If 4.12 L of a 0.850 M H_3PO_4 solution is be diluted to a volume of 10.00 L, what is the concentration the resulting solution?

Answer

0.350 M

44. What volume of a 0.33 M $C_{12}H_{22}O_{11}$ solution can be diluted to prepare 25 mL of a solution with a concentration of 0.025 M?

Answer

1.9 mL

45. What is the concentration of the NaCl solution that results when 0.150 L of a 0.556 M solution is allowed to evaporate until the volume is reduced to 0.105 L?

Answer

0.794 M

46. What is the molarity of the diluted solution when each of the following solutions is diluted to the given final volume?
    a. 1.00 L of a 0.250 M solution of $Fe(NO_3)_3$ is diluted to a final volume of 2.00 L
    b. 0.5000 L of a 0.1222 M solution of C_3H_7OH is diluted to a final volume of 1.250 L
    c. 2.35 L of a 0.350 M solution of H_3PO_4 is diluted to a final volume of 4.00 L
    d. 22.50 mL of a 0.025 M solution of $C_{12}H_{22}O_{11}$ is diluted to 100.0 mL

Answer

a. 0.125 M
b. 0.04888 M
c. 0.206 M
d. 5.6 × 10-3 M

47. What is the final concentration of the solution produced when 225.5 mL of a 0.09988 M solution of Na_2CO_3 is allowed to evaporate until the solution volume is reduced to 45.00 mL?

Answer

0.5005 M

48. An experiment in a general chemistry laboratory calls for a 2.00 M solution of HCl. How many mL of 11.9 M HCl would be required to make 250 mL of 2.00 M HCl?

Answer

42.0 mL

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CHAPTER OVERVIEW

15: Acids and Bases

Acids and bases are common substances found in many every day items, from fruit juices and soft drinks to soap. In this chapter, we will examine what the properties are of acids and bases, and learn about the chemical nature of these important compounds. You'll learn what pH is and how to calculate the pH of a solution.

15.1: Properties of Acids and Bases
15.2: Common Acids and Their Uses
15.3: Common Bases and Their Uses
15.4: Acids and Bases Defined
15.5: Strong and Weak Acids and Bases
15.6: Water as an Acid and as a Base
15.7: An Introduction to pH
15.8: pH and pOH Calculations
15.9: Reactions of Acids and Bases
15.10: For Future Use
15.11: Exercises

Template:HideTOC

Thumbnail Chapter 15: The pH scale. (Heinrich-Boll-Stiftung via Wikimedia Commons)

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15.1: Properties of Acids and Bases

Learning Objectives

- To examine properties of acids and bases.

Sour Patch Kids are a soft candy with a coating of invert sugar and sour sugar. The candy's slogan, "Sour. Sweet. Gone," refers to its sour-to-sweet taste. Sour sugar is a food ingredient that is used to impart a sour flavor, made from citric or tartaric acid and sugar. It is used to coat sour candies like Sour Patch Kids. Eating large amounts of sour sugar can cause irritation of the tongue due to the *acid*. It can also cause irreversible dental erosion.



Figure 15.1.1: Sour Patch Kids candies. Image courtesy of Evan-Amos (public domain).

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The *base* in the antacid neutralizes the excess HCl in the stomach, taking care of that unpleasant feeling.

Acids

The word **acid** is derived from the Latin word *acidus*, which means sour, so it shouldn't be surprising that acids are commonly found in foods. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid. Ascorbic acid is better known as vitamin C. Carbonated sodas contain carbonic acid. In addition, many darker carbonated beverages, such as colas, also contain phosphoric acid. Vinegar is an aqueous solution of acetic acid. Your own stomach utilizes hydrochloric acid to digest food.

Bases

The word **base** is believed to have originated in the 1700s as an evolution of a term used by alchemists. A base may be thought of a "base material" from which other substances may be derived. One such example would be the ashes obtained from burning wood. When the ashes are mixed with water, they produce an alkaline mixture from which the solids may then be filtered out. This alkaline water (or lye water) could then be combined with various acids to produce different neutral salts, or mixed with animal fats to make soap. Hence, the lye water was a base for making different substances.



Figure 15.1.2: Phenolphthalein indicator in presence of base.

Properties of Acids and Bases

Acids and bases are distinct classes of compounds because of the properties of their aqueous solutions. Table 15.1.1 provides a comparison of their mostly contrasting properties.

Table 15.1.1: A Comparison of the Properties of Acids and Bases

Acids	Bases
1. Aqueous solutions of acids are electrolytes. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.	1. Aqueous solutions of bases are electrolytes. Some bases are strong electrolytes, dissociating completely into ions in water, while other bases are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.	2. Bases often have a bitter taste and are found in foods less frequently than acids. Examples are unsweetened chocolate or the alkaloids found in coffee. Many bases, like soaps, are slippery to the touch.
3. Acids change the color of certain acid-base indicators. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.	3. Bases also change the color of certain acid–base indicators. Red litmus turns blue in the presence of a base, while phenolphthalein turns pink.
4. Acids are called **corrosives**, since they react with active metals and cause slow-healing burns and tissue damage when concentrated. Strong acids may do so even when relatively dilute. This is one of the many reasons why "taste tests" should never be used!	4. Bases are called **alkalis** or **caustics**, since they cause slow-healing burns and tissue damage when concentrated. Strong bases may do so even when relatively dilute. This is one of the many reasons why "taste tests" should never be used!
5. Acids react with bases to produce a salt and water (see Section 7.9). Example: $HNO_3\,(aq) + KOH\,(aq) \rightarrow KNO_3\,(aq) + H_2O\,(l)$	5. Acids react with bases to produce a salt and water (see Section 7.9). Example: $HNO_3\,(aq) + KOH\,(aq) \rightarrow KNO_3\,(aq) + H_2O\,(l)$
6. Acids react with active metals to yield hydrogen gas. Example: $Zn\,(s) + H_2SO_4\,(aq) \rightarrow ZnSO_4\,(aq) + H_2\,(g)$	
7. Acids react with carbonates and bicarbonates to yield carbon dioxide gas (discussed in Section 7.9). Example: $NaHCO_3\,(s) + HCl\,(aq) \rightarrow NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$	

⚠ Warning!

Tasting chemicals and touching them are NOT good lab practices and should be avoided! In other words, don't do this at home.

It should not be hard for you to name several common acids (but you might find that listing bases is a little more difficult). Section 15.2 examines some common acids and their uses, while Section 15.3 examines some common bases.

Summary

- Acids and bases are distinct classes of compounds because of the properties of their aqueous solutions.

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15.2: Common Acids and Their Uses

Learning Objectives

- Identify sources and uses of common acids.

You are likely familiar with a number of common acids and their uses. Several common acids and their uses are described below.

Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. It is a colorless, highly pungent solution of hydrogen chloride gas, HCl (g), in water. Hydrochloric acid is usually prepared by treating HCl (g) with water.

$$\mathrm{HCl}\,(g) \xrightarrow{\mathrm{H_2O}} \mathrm{HCl}\,(aq)$$

Hydrochloric acid is a strong acid, since it is completely dissociated in water. Hydrochloric acid is the preferred acid in titration for determining the amount of bases and can also be used to prepare various chloride salts. It is often sold under the common name of muriatic acid for applications like cleaning concrete and pools.

Sulfuric Acid

Sulfuric acid is a highly corrosive strong mineral acid with the molecular formula H_2SO_4. Sulfuric acid has a wide range of applications including use in domestic acidic drain cleaners, as an electrolyte in lead-acid batteries, and in various cleaning agents. It is also a central substance in the chemical industry.

Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas. Because the hydration of sulfuric acid is thermodynamically favorable (and is highly exothermic) and the affinity of it for water is sufficiently strong, sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating property, removing water from other compounds including sugar and other carbohydrates and producing carbon, heat, steam (see Video 15.2.1).

$$\mathrm{C_{12}H_{22}O_{11}}\,(s) \xrightarrow{\mathrm{H_2SO_4}} 12\,\mathrm{C}\,(s) + 11\,\mathrm{H_2O}\,(g)$$



Video 15.2.1: The dehydration of sucrose, $C_{12}H_{22}O_{11}$, with concentrated sulfuric acid, H_2SO_4. (dancingscientist via YouTube)

Nitric Acid

Nitric acid, HNO_3, is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. Nitric acid can be made by reacting nitrogen dioxide, NO_2, with water.

$$3\,\mathrm{NO_2}\,(g) + \mathrm{H_2O}\,(l) \rightarrow 2\,\mathrm{HNO_3}\,(aq) + \mathrm{NO}\,(g)$$

Nitric acid reacts with most metals, but the details depend on the concentration of the acid and the nature of the metal. Dilute nitric acid behaves as a typical acid in its reaction with most metals. For example, magnesium, manganese, and zinc liberate hydrogen gas:

$$Mg\ (s) + 2\ HNO_3\ (aq) \rightarrow Mg(NO_3)_2\ (aq) + H_2\ (g)$$

$$Mn\ (s) + 2\ HNO_3\ (aq) \rightarrow Mn(NO_3)_2\ (aq) + H_2\ (g)$$

$$Zn\ (s) + 2\ HNO_3\ (aq) \rightarrow Zn(NO_3)_2\ (aq) + H_2\ (g)$$

However, concentrated nitric acid reacts with copper metal in a much more interesting manner, yielding the reddish-brown gas nitrogen dioxide, NO_2, as one of the products.

$$Cu\ (s) + 4\ HNO_3\ (aq) \rightarrow Cu(NO_3)_2\ (aq) + 2\ NO_2\ (g) + 2\ H_2O\ (l)$$

⛓ Nitric Acid Acts Upon Copper

Nitric acid is a corrosive acid and a powerful oxidizing agent. The major hazard it poses is chemical burns, as it carries out acid hydrolysis with proteins (amides) and fats (esters) which consequently decomposes living tissue (Figure 15.2.1). Concentrated nitric acid stains human skin yellow due to its reaction with the keratin.



Figure **15.2.1**: Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).

The following excerpt is taken from the biographical memoir of Ira Remsen (1846-1927), a notable chemistry educator and founder of chemical journals, and was presented to the National Academy of Sciences at the annual meeting in 1931:

> "While reading a text-book of chemistry I came upon the statement, 'nitric acid acts upon copper'. I was getting tired of reading such absurd stuff and I determined to see what this meant. Copper was more or less familiar to me, for copper cents were then in use. I had seen a bottle marked 'nitric acid' on a table in the doctor's office where I was then 'doing time!' I did not know its peculiarities but I was getting on and likely to learn. The spirit of adventure was upon me.
>
> "Having nitric acid and copper, I had only to learn what the words 'acts upon' meant. Then the statement, 'nitric acid acts upon copper' would be something more than mere words. All was still. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession.
>
> "I put one of them on the table; opened the bottle marked 'nitric acid'; poured some of the liquid on the copper; and prepared to take an observation. But what was this wonderful thing I beheld ? The cent was already changed, and it was no small change either. A greenish blue liquid foamed and fumed over the cent and over the table. The air in the neighborhood of the

performance became colored dark red. A great colored cloud arose. This was disagreeable and suffocating. How should I stop this?

"I tried to get rid of the objectionable mess by picking it up and throwing it out of the window, which I had meanwhile opened. I learnt another fact—nitric acid not only acts on copper but it acts on fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and another fact was discovered. Nitric acid acts upon trousers.

"Taking everything into consideration, that was the most impressive experiment, and, relatively, probably the most costly I have ever performed. I tell it even now with interest. It was a revelation to me. It resulted in a desire on my part to learn more about that kind of action. Plainly the only way to learn about it was to see its results, to experiment, to work in a laboratory."

What does this reaction look like? Check out Video 15.2.2 to take a peek!



Video 15.2.2: The reaction of a copper penny with concentrated nitric acid. (Wayne Breslyn via YouTube)

Carbonic Acid

Carbonic acid is a chemical compound with the chemical formula H_2CO_3 and is also a name sometimes given to solutions of carbon dioxide in water (carbonated water), because such solutions contain small amounts of H_2CO_3 (*aq*). Carbonic acid, which is a weak acid, forms two kinds of salts: the carbonates and the bicarbonates. In geology, carbonic acid causes limestone to dissolve, producing calcium bicarbonate – which leads to many limestone features such as stalactites and stalagmites.

When carbon dioxide dissolves in water, it exists in chemical equilibrium, producing carbonic acid:

$$CO_2\ (g) + H_2O\ (l) \rightleftarrows H_2CO_3\ (aq)$$

The reaction can be pushed to favor the reactants to generate CO_2 (g) from solution, which is key to the bubbles observed in carbonated beverages (Figure 15.2.2).



Figure **15.2.2**: A glass of sparkling water. (CC BY-SA 3.0; Nevit Dilmen).

Formic Acid

Formic acid, HCOOH, is the simplest carboxylic acid and is an important intermediate in chemical synthesis and occurs naturally, most notably in some ants. The word "formic" comes from the Latin word for ant, formica, referring to its early isolation by the distillation of ant bodies. Formic acid occurs widely in nature as its conjugate base formate.

Acetic Acid

Acetic acid, $HC_2H_3O_2$, is the second simplest carboxylic acid. It is the acid found in vinegar, usually at a concentration of 5% acetic acid. Organic chemists usually write the formula of acetic acid as CH_3COOH to indicate the presence of the carboxyl group. The IUPAC name of acetic acid is ethanoic acid. In other words, acetic acid and ethanoic acid are the same substance.

The oxidation of ethanol, CH_3CH_2OH, the alcohol found in alcoholic beverages, leads to the production of acetic acid and is the reason why a bottle of wine will begin to taste sour and take on a vinegary taste several days after it has been opened and exposed to the air. Industrially, most acetic acid is produced by the carbonylation of methanol using carbon monoxide:

$$CH_3OH + CO \xrightarrow{\text{catalyst}} CH_3COOH$$

Citric Acid

Citric acid, $H_3C_6H_5O_7$, is a weak organic tricarboxylic acid that occurs naturally in citrus fruits. The citrate ion is an intermediate in the TCA cycle (Krebs cycle), a central metabolic pathway for animals, plants and bacteria. Because it is one of the stronger edible acids, the dominant use of citric acid is as a flavoring and preservative in food and beverages, especially soft drinks.



Figure **15.2.3**: Lemons, oranges, limes, and other citrus fruits possess high concentrations of citric acid (CC BY-SA 2.5; André Karwath).

Acetylsalicylic Acid

Acetylsalicylic acid (also known as aspirin) is a medication used to treat pain, fever, and inflammation. Aspirin, in the form of leaves from the willow tree, has been used for its health effects for at least 2,400 years. Aspirin is a white, crystalline, weakly acidic substance.



Figure 15.2.4: Jsmol rendering of an aspirin molecule.

Summary

Table 15.2.1: Common Acids and Their Uses

Chemical Name	Common Name	Uses
hydrochloric acid, HCl	muriatic acid (used in pools) stomach acid	Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning.
sulfuric acid, H_2SO_4		Used in car batteries and in the manufacture of fertilizers.
nitric acid, HNO_3		Used in the manufacture of fertilizers, explosives, and in extraction of gold.
acetic acid, $HC_2H_3O_2$	vinegar	Main ingredient in vinegar.
carbonic acid, H_2CO_3	responsible for the "fizz" in carbonated drinks	As an ingredient in carbonated drinks.
citric acid, $H_3C_6H_5O_7$		Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions.
acetylsalicylic acid, $HC_9H_7O_4$	aspirin	The active ingredient in aspirin.

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15.3: Common Bases and Their Uses

Learning Objectives

- Identify sources and uses of common bases.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate. Several common bases and their corresponding uses are described below.

Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula NaOH. It is a white solid ionic compound consisting of sodium cations , Na^+, and hydroxide anions, OH^-.



Figure **15.3.1**: A sample of sodium hydroxide as pellets in a watch glass. (Public Domain; Walkerma.)

Dissolution of solid sodium hydroxide in water is a highly exothermic reaction:

$$\mathrm{NaOH}\,(s) \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The resulting solution is colorless and odorless and feels slippery when it comes in contact with skin.

Potassium Hydroxide

Potassium hydroxide is an inorganic compound with the formula KOH, and is commonly called caustic potash. Along with sodium hydroxide, NaOH, this colorless solid is a prototypical strong base. It has many industrial and niche applications, most of which exploit its caustic nature and its reactivity toward acids. Its dissolution in water is strongly exothermic.

$$\mathrm{KOH}\,(s) \xrightarrow{\mathrm{H_2O}} \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

A concentrated aqueous solution is sometimes called *potash lye*.

Magnesium Hydroxide

Magnesium hydroxide is the inorganic compound with the chemical formula $Mg(OH)_2$. Magnesium hydroxide is a common component of antacids, such as milk of magnesia, as well as laxatives.



Figure **15.3.2**: Bottle of antacid tablets. (CC BY 2.5; Midnightcomm).

It is a white solid with low solubility in water. Combining a solution of many magnesium salts with basic water induces precipitation of solid $Mg(OH)_2$. However, a weak concentration of dissociated ions can be found in solution:

$$Mg(OH)_2\,(s) \rightleftharpoons Mg^{2+}\,(aq) + 2\,OH^-\,(aq)$$

Calcium Hydroxide

Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $Ca(OH)_2$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime. Calcium hydroxide is used in many applications, including food preparation. Limewater is the common name for a saturated solution of calcium hydroxide.

Calcium hydroxide is relatively insoluble in water, but is large enough that its solutions are basic according to the following reaction:

$$Ca(OH)_2\,(s) \rightleftharpoons Ca^{2+}\,(aq) + 2\,OH^-\,(aq)$$

Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula NH_3 and is a colorless gas with a characteristic pungent smell. It is the active product of “smelling salts,” and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form.

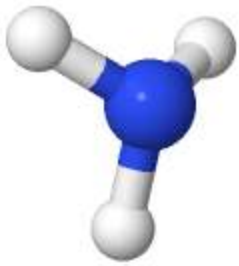


Figure 15.3.3: Jsmol rendering of the ammonia molecule.

In aqueous solution, ammonia acts as a base, acquiring hydrogen ions from H_2O to yield ammonium and hydroxide ions:

$$NH_3\,(g) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

Summary

Table 15.3.1: Common Bases and Their Uses

Chemical Name	Common Name	Uses
sodium hydroxide, NaOH	lye or caustic soda	Used in the manufacture of soaps and detergents, and as the main ingredient in oven and drain cleaners.
potassium hydroxide, KOH	lye or caustic potash	Used in the production of liquid soaps and soft soaps. Used in alkaline batteries.
magnesium hydroxide, $Mg(OH)_2$	milk of magnesia	Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater.
calcium hydroxide, $Ca(OH)_2$	slaked lime	Used in the manufacture of cement and lime water. Also added to neutralize acidic soil.
ammonia, NH_3		Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers.

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15.4: Acids and Bases Defined

Learning Objectives

- Identify an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

There are three major classifications of substances known as acids or bases. The theory developed by Svante Arrhenius in 1883, the Arrhenius definition, states that an acid produces H^+ in solution and a base produces OH^-. Later, two more sophisticated and general theories were proposed. These theories are the Brønsted-Lowry and Lewis definitions of acids and bases. This text considers the Arrhenius and Brønsted-Lowry theories.

The Arrhenius Theory of Acids and Bases

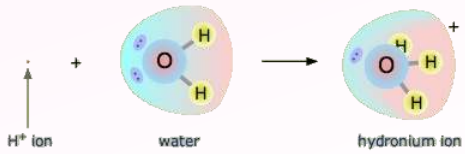
In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. According to Arrhenius, an acid is a compound that increases the concentration of hydrogen ions, H^+, when dissolved in water. Over time, this definition has slightly evolved. Today, an **Arrhenius acid** is often thought of as a compound that increases the concentration of hydronium ions, H_3O^+, when dissolved in water.

If you're thinking, "Wait a minute! I have never heard of hydronium ions," you are correct at least as far as chemical nomenclature is concerned. This is because hydronium ions are not found in compounds. Hydronium ions are only found in aqueous solutions of acids. In other words, you will never see a compound such as hydronium sulfate or hydronium chloride. So what, exactly, are hydronium ions? Read on:

The Hydronium Ion

A hydrogen ion, H^+, represents a hydrogen atom that has lost its only electron. 99% of hydrogen atoms are comprised of the isotope hydrogen-1, $^1_1\mathbf{H}$, meaning that 99% hydrogen atoms have no neutrons in their nucleus. Because of this, an H^+ ion is often thought of as a single proton, p^+. Due to their tiny size in comparison to the size of an atom, individual protons represent an extremely concentrated charge. As a result, they attract the lone pairs of neighboring water molecules.

A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $H_5O_2^+$ or $H_9O_4^+$ rather than H_3O^+. It is simpler, however, to use H_3O^+ to represent the hydronium ion.



To represent the process using chemical equations, consider HCl (*aq*). The HCl (aq) releases H^+ ions:

$$\mathbf{HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-\,(aq)} \tag{15.4.1}$$

The H^+ ions combine with H_2O to form H_3O^+ ions as shown above:

$$\mathbf{H^+\,(aq) + H_2O\,(\mathit{l}) \rightarrow H_3O^+\,(aq)} \tag{15.4.2}$$

Summing Equation 15.4.1 and Equation 15.4.2 yields an equation that shows how hydronium ions are formed when an acid such as HCl is dissolved in water:

$$\mathbf{HCl\,(aq) + H_2O\,(\mathit{l}) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{15.4.3}$$

An **Arrhenius base** is a compound that increases the concentration of hydroxide ions, OH^-, when dissolved in water. The dissociation of NaOH in water is represented by the following equation:

$$\mathbf{NaOH\ (s) \xrightarrow{H_2O} Na^+\ (aq) + OH^-\ (aq)} \qquad \mathbf{(15.4.4)}$$

In this reaction, sodium hydroxide (NaOH) dissociates into sodium ions, Na^+, and hydroxide ions, OH^-, when dissolved in water, thereby releasing OH^- ions into solution.

Arrhenius Theory of Acids and Bases

1. An **Arrhenius acid** is a compound that increases the concentration of hydronium ions, H_3O^+, when dissolved in water.
2. An **Arrhenius base** is a compound that increases the concentration of hydroxide ions, OH^-, when dissolved in water.

The Brønsted-Lowry Theory of Acids and Bases

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. As described above, an H^+ ion may be thought of as a single proton, p^+. A **Brønsted-Lowry acid** is any species that donates a proton (H^+) to another species, while a **Brønsted-Lowry base** is any species that accepts a proton from another species. In short, **a Brønsted-Lowry acid** is a **proton donor**, while **a Brønsted-Lowry base** is a **proton acceptor**.

Let's use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$NH_3\ (aq) + H_2O\ (l) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion from the water molecule. We can illustrate this as follows:

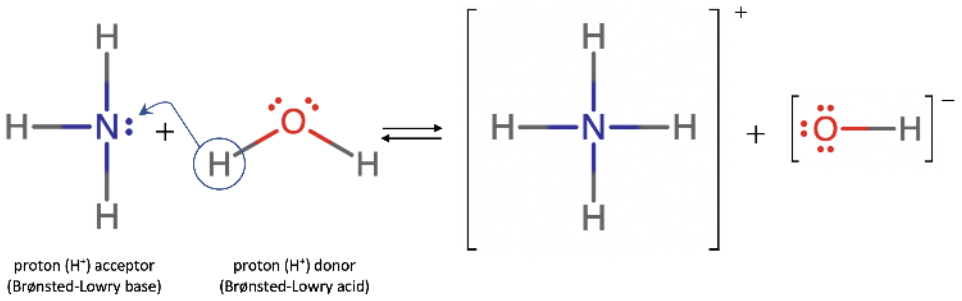


Figure **15.4.1**: A molecule of NH_3 accepts a proton (H^+) from a molecule of H_2O.

Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule – which accepts the hydrogen ion – is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense (since it produces OH^- ions when dissolved in water) and the Brønsted-Lowry sense (since it acts as a proton acceptor).

What about hydrochloric acid? Is an Arrhenius acid like HCl still an acid in the Brønsted-Lowry sense? Yes. Once again, the reaction that occurs when HCl is dissolved in H_2O looks like this:

$$HCl\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + Cl^-\ (aq)$$

This process may be depicted using Lewis dot diagrams:

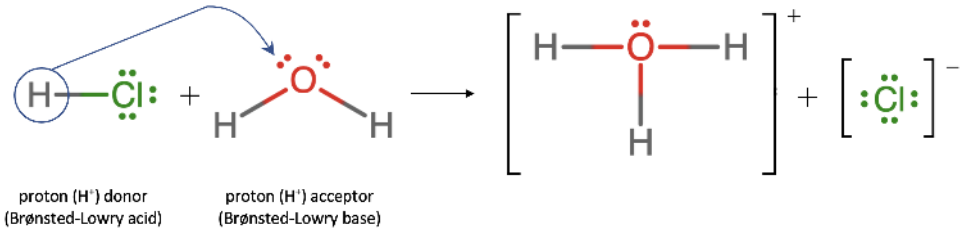


Figure **15.4.2**: A molecule of HCl donates a proton (H^+) to a molecule of H_2O.

This equation shows that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule yielding chloride ions and hydronium ions. As the proton donor, HCl acts as a Brønsted-Lowry acid; as a proton acceptor, H_2O is a Brønsted-Lowry base. In other words, HCl is both an Arrhenius acid and a Brønsted-Lowry acid. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl.

The Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base. However, the Arrhenius definition would not have labeled H_2O a base for the reaction shown in Figure 15.4.2, nor would it have labeled H_2O as an acid in Figure 15.4.1.

⚓ Brønsted-Lowry Theory of Acids and Bases

1. A **Brønsted-Lowry acid** is a proton (H^+) donor.
2. A **Brønsted-Lowry base** is a proton (H^+) acceptor.

Note: All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

✅ Example 15.4.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule, just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction may be written as follows:

$$C_6H_5NH_2\,(aq) + H_2O\,(l) \rightleftharpoons C_6H_5NH_3^+\,(aq) + OH^-\,(aq)$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

✏️ Exercise 15.4.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$H_3PO_4 + H_2O \rightleftharpoons H_3O^+ + H_2PO_4^-$$

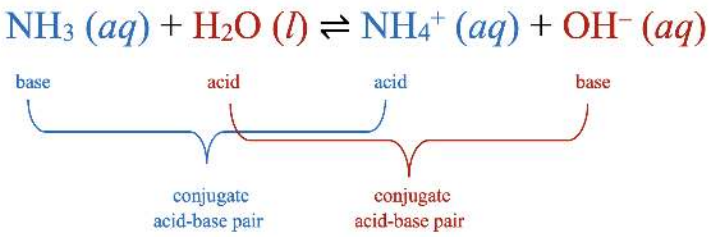
Answer

Brønsted-Lowry acid: H_3PO_4; Brønsted-Lowry base: H_2O

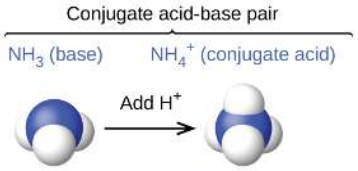
Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. Let's consider again the acid-base reaction that takes place when ammonia is dissolved in water. As previously established, NH_3 is a base since it accepts a proton from H_2O and H_2O is an acid since it donates a proton to the NH_3.

What if the reaction were to happen in reverse? In the reverse reaction, the ammonium ion, NH_4^+, donates a proton (H^+) to the hydroxide ion, OH^-, yielding a molecule of NH_3 and a molecule of H_2O. Since NH_4^+ acts as a proton donor in the reverse reaction, it would be considered an acid. Since OH^- acts as a proton acceptor in the reverse reaction, it would be considered a base.



Notice that NH_3 and NH_4^+ differ from each other by one proton (H^+). This relationship is called a **conjugate acid-base pair** or sometimes simply a **conjugate pair**. Within a conjugate acid-base pair, the species that holds the extra H^+ is considered the acid within the pair. This means that within this pair, NH_3 is the base and NH_4^+ is the acid.



By now you also probably noticed that H_2O and OH^- differ from each other by one proton (H^+). This means they are also a conjugate acid-base pair. Since H_2O holds the extra H^+ in the pair, H_2O is the acid and OH^- is the base.

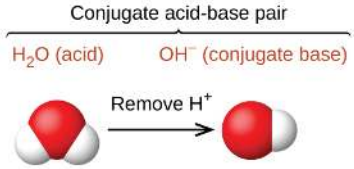


Figure 15.4.3 shows several conjugate acid-base pairs. The acid appears in the first column, while its conjugate base appears in the same row in the second column. Notice again that all conjugate acid-base pairs differ from each other by a single proton (H^+).

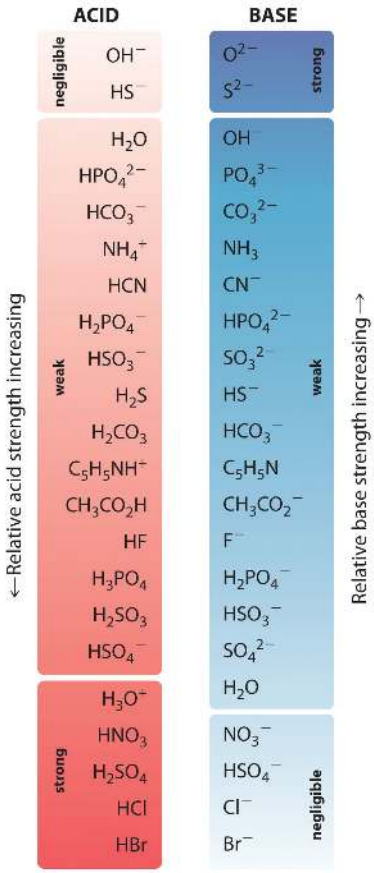


Figure **15.4.3**: Several common conjugate acid-base pairs. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base. Likewise, the conjugate acid of a strong base is a very weak acid.

✅ Example 15.4.2

Identify the conjugate acid-base pairs in this equilibrium.

$$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

Solution

In this reaction, acetic acid, $HC_2H_3O_2$, donates a proton to H_2O. This means $HC_2H_3O_2$ is the acid and H_2O is the base. In the reverse reaction, H_3O^+ is the acid since it donates a proton to the acetate ion, $C_2H_3O_2^-$. This means that $C_2H_3O_2^-$ is a base.

The two conjugate acid-base pairs:

- $HC_2H_3O_2$ (acid) and $C_2H_3O_2^-$ (base).
- H_2O (base) and H_3O^+ (acid).

Example 15.4.3

What is the conjugate acid of NO_3^-?

Solution

The conjugate acid in the pair always has one more proton (H^+) than its base. Adding an H^+ to NO_3^- yields the acid HNO_3.

Exercise 15.4.2

Identify the conjugate acid-base pairs in this reaction.

$$N_2H_4 + H_2O \rightleftharpoons N_2H_5^+ + OH^-$$

Answer

- N_2H_4 (base) and $N_2H_5^+$ (acid)
- H_2O (acid) and OH^- (base)

Exercise 15.4.3

What is the conjugate base of H_2CO_3?

Answer

HCO_3^-

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15.5: Strong and Weak Acids and Bases

Learning Objectives

- Define a strong and a weak acid and base.
- Recognize an acid or a base as strong or weak.

Electrolytes were previously described as substances that yield ions when dissolved in water, which means that aqueous solutions of electrolytes are able to conduct electricity. It should be clear that soluble ionic compounds are electrolytes. While it may seem that molecular compounds should all be nonelectrolytes, it turns out that this isn't always the case. Molecular compounds that are classified as acids or bases are electrolytes – some are strong electrolytes; most, however, are weak electrolytes.

An acid that is a strong electrolyte is called a **strong acid**, while an acid that is a weak electrolyte is a **weak acid**. A base that is a strong electrolyte is called a **strong base**, while a base that is a weak electrolyte is called a **weak base**. Classifying acids or bases as strong or weak has nothing to do with their concentration. It is possible to have a dilute solution of a strong acid or base and it is also possible to have a concentrated solution of a weak acid or base.



Figure **15.5.1**: A bottle of 100% acetic acid, also called glacial acetic acid. The term "glacial" refers to the fact that it would often solidify in many old, poorly heated chemical stockrooms. Pure acetic acid has a freezing point of 17°C (62°F). (W. Oelen via Wikimedia Commons)

Weak acids and bases may be just as dangerous as their stronger counterparts when concentrated. For example, vinegar is an aqueous solution of acetic acid, $HC_2H_3O_2$, usually with a concentration of 5%. Vinegar is used to provide flavor and preserve foods. One might think of acetic acid as being quite harmless due to its widespread presence in foods. However, acetic acid in its pure form, called glacial acetic acid (see Figure 15.5.1), can cause severe skin burns in its concentrated form. One of the authors of this section has had such an experience back in his days in college when working with glacial acetic acid as a student in lab.

Strong Acids and Strong Bases

The terms **strong acid** and **strong base** are used to indicate that these substances are strong electrolytes when dissolved in water. The hydroxides of elements in Group IA and Group IIA on the periodic table are all strong electrolytes [except for $Mg(OH)_2$ and $Be(OH)_2$ since they are mostly insoluble in water]. As soluble ionic compounds, they dissociate completely to yield the metal ion and the hydroxide ion. Consider an aqueous solution of NaOH:

$$\mathrm{NaOH}\,(aq) \xrightarrow{100\%} \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq) \tag{15.5.1}$$

NaOH is a base, due to the formation of OH^- in water and it is a strong electrolyte, since it dissociates completely in water. We say it **dissociates**, because NaOH is already an ionic compound and the ions simply dissociate, or separate, from each other when NaOH is dissolved in water. Because NaOH is a base and a strong electrolyte, NaOH is classified as a **strong base**.

When HCl is dissolved in water, it yields a solution that is strongly conducting. This is an indication that it is completely converted into ions. Since HCl is a molecular compound, we can't say it dissociates into its ions, but rather, we say that it **ionizes** completely in the presence of water. The term ionize means that ions are formed from something that was not originally ionic. To explain the existence of ions, we can show that HCl reacts with water to yield H_3O^+ (*aq*) and Cl^- (*aq*) ions:

$$\mathrm{HCl}\,(aq) + \mathrm{H_2O}\,(l) \xrightarrow{100\%} \mathrm{H_3O^+}\,(aq) + \mathrm{Cl^-}\,(aq) \tag{15.5.2}$$

Since HCl ionizes completely to form ions, it is a strong electrolyte. HCl is classified as an acid due to the formation of H_3O^+ ions in water. Consequently, HCl is classified as a **strong acid**.

Weak Acids and Weak Bases

The terms **weak acid** and **weak base** are used to indicate that these substances are weak electrolytes when dissolved in water. Most acids and bases are weak electrolytes and are also molecular compounds. As molecular compounds, this means they must ionize (react with water) to be an electrolyte, but that they do not ionize completely.

For example, when acetic acid is placed in water, approximately 1% of the molecules *ionize*. The remaining 99% of $HC_2H_3O_2$ molecules remain intact and unreacted. Since most of the molecules do not ionize, the solution is only weakly conducting. This reaction may be visualized in equation form as:

$$\mathrm{HC_2H_3O_2}\,(aq) + \mathrm{H_2O}\,(l) \xrightarrow{\sim 1\%} \mathrm{H_3O^+}(aq) + \mathrm{C_2H_3O_2^-}\,(aq) \tag{15.5.3}$$

Since this reaction does not go 100% to completion, it is more appropriate to write it as a **reversible reaction** to indicate that most of the reaction mixture remains as reactants:

$$\mathrm{HC_2H_3O_2}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{C_2H_3O_2^-}\,(aq) \tag{15.5.4}$$

Since $HC_2H_3O_2$ ionizes only partially to form ions, it is a weak electrolyte. $HC_2H_3O_2$ is classified as an acid due to the formation of H_3O^+ ions in water. Consequently, $HC_2H_3O_2$ is classified as a **weak acid**.

When NH_3 is dissolved in water, it *ionizes* to yield a solution that is weakly conducting. Therefore, we show its reaction with water as a reversible reaction to indicate that most of the reaction mixture remains as reactants and not as ions:

$$\mathrm{NH_3}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{NH_4^+}(aq) + \mathrm{OH^-}\,(aq) \tag{15.5.5}$$

NH_3 is a base, due to the formation of OH^- in water and it is a weak electrolyte, since it ionizes only partially in water. Because NH_3 is a base and a weak electrolyte, NH_3 is classified as a **weak base**.

Classification of Acids and Bases as Strong or Weak

As it turns out, there are very few strong acids and very few strong bases. Those classified as strong acids and strong bases are listed in Table 15.5.1. If an acid is not listed here, it is a weak acid. If a base is not listed here, it is a weak base.

Table 15.5.1: Strong Acids and Strong Bases

Strong Acids	Strong Bases
HCl	Group IA and Group IIA hydroxides*, *including*:
HBr	LiOH
HI	NaOH
HNO_3	KOH
H_2SO_4	RbOH
$HClO_3$	CsOH
$HClO_4$	$Ca(OH)_2$
	$Sr(OH)_2$
	$Ba(OH)_2$

*$Mg(OH)_2$ is sometimes classified as a strong base, since it ionizes completely. However, because it is relatively insoluble, little dissolves and relatively few OH^- ions are formed in solution. Therefore, is is also sometimes classified as a weak base. It will be considered a weak base in this text.

As mentioned above, Table 15.5.1 is an all-inclusive list of strong acids and a strong bases. If an acid or base does not appear in on the list in Table 15.5.1, it is classified as weak. Table 15.5.2 below shows a few common weak acids and weak bases:

Table 15.5.2: Some Common Weak Acids and Weak Bases

Weak Acids	Weak Bases
$HC_2H_3O_2$	$Mg(OH)_2$
H_3PO_4	NH_3
HF	
H_2CO_3	

✅ Example 15.5.1: Identifying Strong and Weak Acids and Bases

Identify each acid or base as strong or weak.

A. HCl
B. $Mg(OH)_2$
C. C_5H_5N

Solution

A. Because HCl is listed in Table 15.5.1, it is a strong acid.
B. Because $Mg(OH)_2$ is listed in Table 15.5.2, it is a weak base.
C. Since it is implied this must be an acid or a base, it must be a base since it does not have H out at the front of the formula. It can be assumed that C_5H_5N would act as a proton acceptor and therefore act as a base. Since it does not contain a OH^- ion, it cannot be considered a strong base; it is a weak base.

✏️ Exercise 15.5.1

Identify each acid or base as strong or weak.

A. LiOH
B. HNO_2

Answer A

strong base

Answer B

weak acid

✅ Example 15.5.2: Characterizing Base Ionization

Write the balanced chemical equation for the dissociation of $Ba(OH)_2$ and indicate whether it proceeds 100% to products or not.

Solution

This is an ionic compound of Ba^{2+} ions and OH^- ions. When an ionic compound dissolves, it separates into its constituent ions:

$$\mathrm{Ba(OH)_2}\,(aq) \xrightarrow{100\%} \mathrm{Ba^{2+}}\,(aq) + 2\,\mathrm{OH^-}\,(aq)$$

$Ba(OH)_2$ is listed in Table 15.5.1, so this reaction proceeds 100% to products.

✏️ Exercise 15.5.2

Write the balanced chemical equation for the ionization of nitrous acid (HNO_2) and indicate whether it proceeds 100% to products or not.

Answer

The reaction is as follows:

$$\mathrm{HNO_2}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{NO_2^-}\,(aq)$$

It does not proceed 100% to products because nitrous acid is a weak acid.

Key Takeaways

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.

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15.6: Water as an Acid and as a Base

Learning Objectives

- Describe the autoionization of water.
- Calculate the concentrations of H_3O^+ and OH^- in aqueous solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid, as with its reaction with NH_3:

$$\underbrace{NH_3}_{\text{base}} + \underbrace{H_2O}_{\text{acid}} \rightleftharpoons NH_4^+ + OH^- \tag{15.6.1}$$

and that H_2O can act as a base, as with its reaction with HCl:

$$\underbrace{HCl}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \longrightarrow H_3O^+ + Cl^- \tag{15.6.2}$$

Therefore, it may not surprise you to learn that within any given sample of water, some H_2O molecules act as an acid and some H_2O molecules act as a base. The chemical equation is as follows:

$$\underbrace{H_2O}_{\text{acid}} + \underbrace{H_2O}_{\text{base}} \rightleftharpoons H_3O^+ + OH^- \tag{15.6.3}$$

However, this process (shown in Figure 15.6.1), called the **autoionization of water**, occurs to a very small extent, with only 6 out of every 100,000,000 H_2O molecules participating in autoionization at 25°C.

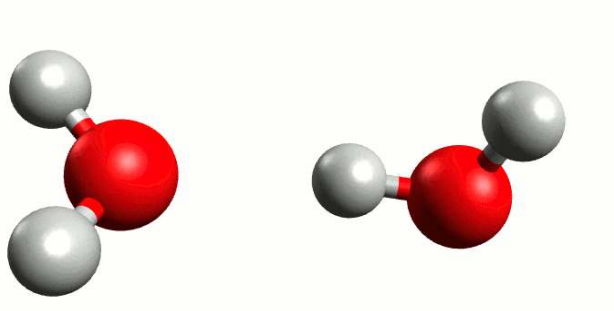


Figure **15.6.1**: The autoionization of water, resulting in hydronium ions and hydroxide ions. (Manuel Almagro Rivas via Wikimedia Commons)

This represents a concentration of 1.0×10^{-7} M for both H_3O^+ (*aq*) and OH^- (*aq*) in a sample of pure H_2O (at 25°C). As previously discussed, square brackets, [ ], around a dissolved species represents the molarity of that species. This means that for any sample of pure water:

$$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ \text{M} \tag{15.6.4}$$

The product of these two concentrations is 1.0×10^{-14}:

$$[H_3O^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14} \tag{15.6.5}$$

It turns out that for an aqueous solution at 25°C, the *product* of the two concentrations, $[H_3O^+] \times [OH^-]$, is *always* equal to 1.0×10^{-14}, regardless of whether the aqueous solution is acidic, basic, or neutral. In other words,

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14} \tag{15.6.6}$$

The product of these two concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_w:

$$K_w = [H_3O^+]\,[OH^-] = 1.0 \times 10^{-14} \tag{15.6.7}$$

This means that if $[H_3O^+]$ is known for an aqueous solution, $[OH^-]$ may be calculated, since their product is always 1.0×10^{-14}; or if $[OH^-]$ is known, that $[H_3O^+]$ may be calculated. This also implies that as one concentration goes up, the other must go down to ensure their product always equals the value of K_w.

⚓ Temperature Matters

The degree of autoionization of water, as shown in Equation 15.6.3 changes with temperature. This means that the value of K_w also changes with temperature. Therefore, Equation 15.6.4 through Equation 15.6.7 are only valid at room temperature (25°C). In the event the temperature is not known, it is assumed to be 25°C.

Since pure water is neutral and $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ M$ in pure water, then any aqueous solution in which $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}\ M$ must also be neutral (at 25°C). From this, we can conclude the following:

- Since acids produce H_3O^+ when dissolved in water, $[H_3O^+] > 1.0 \times 10^{-7}\ M$ in acidic solutions (at 25°C).
- Since bases produce OH^- when dissolved in water, $[OH^-] > 1.0 \times 10^{-7}\ M$ in basic solutions (at 25°C).

✅ Example 15.6.1

What is the $[OH^-]$ of an aqueous solution if the $[H_3O^+]$ is 1.0×10^{-4} M?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $[H_3O^+]$ =1.0 × 10^{-4} M Find: $[OH^-]$ = ? M
List known relationship(s).	$[H_3O^+]\,[OH^-] = 1.0 \times 10^{-14}$
Plan the problem.	Rearrange the equation to solve for $[OH^-]$. $[OH^-] = \frac{\cancel{[H_3O^+]}\,[OH^-]}{\cancel{[H_3O^+]}} = \frac{1.0 \times 10^{-14}}{[H_3O^+]}$
Calculate the answer.	Now substitute the known quantities into the equation and solve. $[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = \boxed{1.0 \times 10^{-10}\ \text{M}}$
Think about your result.	The $[H_3O^+]$ is high (> 1 × 10^{-7} M), so $[OH^-]$ should be low. Furthermore, a quick answer check shows: $[H_3O^+]\,[OH^-] = (1.0 \times 10^{-4})\,(1.0 \times 10^{-10}) = 1.0 \times 10^{-14}$

Exercise 15.6.1

What is $[H_3O^+]$ in solution where $[OH^-]$ = 0.00032 M?

Answer

3.1×10^{-11} M

$[H_3O^+]$ and $[OH^-]$ of Strong Acids and Bases

Finding the $[H_3O^+]$ and $[OH^-]$ of *strong* acids and *strong* bases is relatively straightforward. Finding the $[H_3O^+]$ and $[OH^-]$ of *weak* acids and *weak* bases is a bit more complicated and beyond the scope of this text.

As previously discussed, strong acids ionize completely. For *monoprotic* strong acids (strong acids with only one H),

$$[H_3O^+] = [\text{acid}] \tag{15.6.8}$$

This means that $[H_3O^+]$ = 0.10 M in a solution of 0.10 M HNO_3, a strong acid, since one H_3O^+ ion forms for each molecule of HNO_3:

$$HNO_3\,(aq) + H_2O\,(l) \xrightarrow{100\%} H_3O^+\,(aq) + NO_3^-\,(aq)$$

For H_2SO_4, the only strong acid that is *diprotic* (has two H atoms) , $[H_3O^+]$ = [acid] is a close enough approximation, though the $[H_3O^+]$ is slightly greater than [acid], since more than one H_3O^+ ion forms for each molecule of H_2SO_4. However, it is much for difficult for the second proton (H^+) to be released from H_2SO_4, which is why $[H_3O^+]$ = [acid] remains a close enough approximation.

Also previously discussed, strong bases are soluble in water and dissociate completely. For strong bases,

$$[OH^-] = [\text{base}] \times \text{number of } OH^- \text{ ions in the chemical formula} \tag{15.6.9}$$

This means that $[OH^-]$ = 0.10 M in a solution of 0.10 M NaOH, a strong base, since there is only one OH^- ion in NaOH:

$$NaOH\ (aq) \xrightarrow{100\%} Na^+\ (aq) + OH^-\ (aq)$$

and that $[OH^-]$ = 0.20 M in a solution of 0.10 M $Ba(OH)_2$, also a strong base, since there are two OH^- ions in $Ba(OH)_2$:

$$Ba(OH)_2\ (aq) \xrightarrow{100\%} Ba^{2+}\ (aq) + 2\ OH^-\ (aq)$$

✅ Example 15.6.2

What is the:

A. $[H_3O^+]$ in 0.0023 M HBr?
B. $[H_3O^+]$ in 0.0023 M HNO_3?
C. $[OH^-]$ in 0.0023 M KOH?
D. $[OH^-]$ in 0.0023 M $Sr(OH)_2$?

Solution

A. $[H_3O^+]$ = [acid] in a strong acid, so $[H_3O^+]$ = 0.0023 M
B. $[H_3O^+]$ = [acid] in a strong acid, so $[H_3O^+]$ = 0.0023 M
C. $[OH^-] = [\text{base}] \times \text{number of } OH^- \text{ ions in the chemical formula}$ for a strong base, so $[OH^-]$ = 0.0023 M, since there is only one OH^- ion in KOH.
D. $[OH^-] = [\text{base}] \times \text{number of } OH^- \text{ ions in the chemical formula}$ for a strong base, so $[OH^-]$ = 0.0046 M, since there are two OH^- ions in $Sr(OH)_2$. (0.0023 M × 2 = 0.0046 M).

✅ Example 15.6.3

What is $[H_3O^+]$ in a 0.0044 M solution of $Ca(OH)_2$?

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: $Ca(OH)_2$ = 0.0044 M Find: $[H_3O^+]$ = ? M
List known relationship(s).	$[OH^-] = [\text{base}] \times \text{number of } OH^- \text{ ions in the chemical formula}$ $[H_3O^+]\,[OH^-] = 1.0 \times 10^{-14}$
Plan the problem.	Rearrange the equation to solve for $[H_3O^+]$. $[H_3O^+] = \dfrac{[H_3O^+]\,\cancel{[OH^-]}}{\cancel{[OH^-]}} = \dfrac{1.0 \times 10^{-14}}{[OH^-]}$
Calculate.	$[OH^-] = 0.0044\ \text{M} \times 2 = 0.0088\ \text{M}$ $[H_3O^+] = \dfrac{1.0 \times 10^{-14}}{[OH^-]} = \dfrac{1.0 \times 10^{-14}}{0.0088} = \boxed{1.1 \times 10^{-12}\ \text{M}}$
Think about your result.	The $[OH^-]$ is high (> 1 × 10^{-7} M), so $[H_3O^+]$ should be low (< 1 × 10^{-7} M).

Exercise 15.6.3

Calculate the $[H_3O^+]$ and $[OH^-]$ of 6.1×10^{-3} M HI.

Answer

$[H_3O^+] = 6.1\times10^{-3}$ M; $[OH^-] = 1.6\times10^{-12}$ M

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15.7: An Introduction to pH

Learning Objectives

- Define pH and acidic, basic, and neutral pH values.
- Determine the magnitude of change in $[H_3O^+]$ for changing pH values.

Perhaps you have heard of the term *pH* used to describe the acidity or basicity of a substance or you already know that a pH = 7 is neutral – neither acidic, nor basic. What is the pH and what does it mean? If a carbonated beverage has a pH = 3, how does it compare to a solution that has a pH = 6? Is it half as acidic? Twice as acidic? Or is it something entirely different? (Hint: It is!)



Figure **15.7.1**: A typical cola has a pH ≤ 3. (rawpixel via FreeRange)

The pH Scale

We have seen that $[\mathbf{H_3O^+}]$ and $[\mathbf{OH^-}]$ values may be very small and vary widely from one solution to another. Before it was known that H^+ ions were unlikely to exist in aqueous solution, Danish scientist Søren Sørensen (1868-1939) proposed a system for representing the $[\mathbf{H^+}]$ called the pH scale. The letters **pH** stand for the "power of the hydrogen ion" and was originally defined as the negative logarithm of the hydrogen ion concentration, as shown in Equation **15.7.1** (read about Logarithmic Scales below):

$$\mathbf{pH} = -\log\left[\mathbf{H^+}\right] \tag{15.7.1}$$

Today, we know that acids produce hydronium ions in solution rather than hydrogen ions, so the definition of **pH** may be slightly reworded to be the negative logarithm of the hydronium ion concentration:

$$\mathbf{pH} = -\log\left[\mathbf{H_3O^+}\right] \tag{15.7.2}$$

In pure water, we have already established that $[\mathbf{H_3O^+}] = [\mathbf{OH^-}] = \mathbf{10^{-7}\ M}$. Substituting $[\mathbf{H_3O^+}] = \mathbf{10^{-7}\ M}$ into the equation, $\mathbf{pH} = -\log[\mathbf{H_3O^+}]$:

$$\mathbf{pH} = -\log\left[10^{-7}\right] = -(-7) = 7$$

Thus, the pH = 7 for pure water (or any neutral solution). It should also be noted that *pH values have no units*. The pH of pure water is simply 7.

Logarithmic Scales

What is a logarithm and why is it used to represent the acidity or basicity of a solution? The $[\mathbf{H_3O^+}]$ of an aqueous solution is almost always in the range of 10^{-14} M to 1 M. This is an extremely wide range that is over several orders of magnitude. A logarithmic scale condenses this range of values to numbers that are easier to use, almost always in the range of 0 to 14.

Logarithms have different "bases". Base-10 logarithms, which are based on magnitudes of 10, are used to calculate pH. Some example logarithm calculations:

$$\log\left(10^5\right) = 5$$

$$\log\left(10^{-9}\right) = -9$$

Therefore, in general:

$$\log(10^x) = x$$

Each integer value increment represents a factor of 10. In the case of the pH, we define it as being the *negative* logarithm of the hydronium ion concentration, $\mathbf{pH = -\log[H_3O^+]}$. Therefore, a solution with a pH = 3 has a $[H_3O^+]$ = 10^{-3} M, which is 10 times larger than a solution with a pH = 4, which has a $[H_3O^+]$ = 10^{-4} M, and a solution with a pH = 4 has a $[H_3O^+]$ that is 10 times larger than a solution with a pH = 5, which has a $[H_3O^+]$ = 10^{-5} M. This means that a solution with a pH = 3 has a $[H_3O^+]$ that is 100 times larger than a solution with a pH = 5. In other words, each whole number decrease in pH value represents a 10-fold increase in the $[H_3O^+]$.

To answer the question asked at the beginning of this section, "If a carbonated beverage has a pH = 3, how does it compare to a solution that has a pH = 6?" We can now say that a solution with a pH = 3 is 10×10×10, or 10^3 times the acidity of a solution with a pH = 6, in terms of the $[H_3O^+]$; *not* twice as acidic or half as acidic.



Figure **15.7.2**: Earthquake damage in Portsmouth, Dominica from magnitude 6.2 Les Saintes earthquake. (21 November, 2004). (UWI Seismic Research Centre via Flickr)

Other measurements based on a logarithmic scale include the Richter scale, used to measure earthquake intensity and sound intensity, which is measured in units called decibels (dB). In the case of the Richter scale, each whole number increment in magnitude represents a 10-fold increase in earthquake intensity. An earthquake with a magnitude of 7.2 has 10 times the intensity of an earthquake with a magnitude of 6.2, which has 10 times the intensity of an earthquake with a magnitude of 5.2. This means that an earthquake with a magnitude of 7.2 has 100 times the intensity of an earthquake with a magnitude of 5.2.

In the case of sound intensity, each decibel (dB) represents ⅒ bel (B). This means that each increment of 10 dB represents a 10-fold increase in sound intensity. A sound intensity level of 98 dB has 10 times the intensity of 88 dB, which has 10 times the intensity of 78 dB. In other words, a sound intensity level of 98 dB has 100 times the intensity of 78 dB.

As established in the previous section, a solution with a $[H_3O^+]$ greater than $\mathbf{1.0 \times 10^{-7}\ M}$ is acidic, while a solution with a $[H_3O^+]$ less than $\mathbf{1.0 \times 10^{-7}\ M}$ is basic. In terms of pH, this may be summarized as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

Figure **15.7.3** shows the pH scale in relation to several common materials (the H^+ concentration in the figure below is the same as the $[H_3O^+]$).

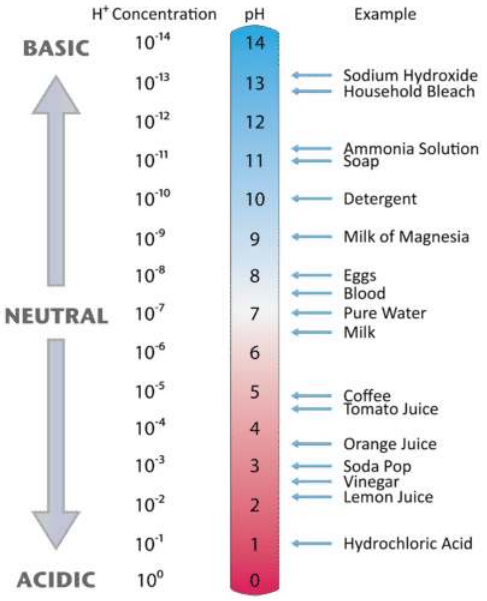


Figure **15.7.3**: The pH values for several common materials.

✅ Example 15.7.1

Identify each solution as acidic, basic, or neutral.

A. milk of magnesia, pH = 10.5
B. a salt solution, pH = 7
C. wine, pH = 3.0

Solution

A. With a pH >7, milk of magnesia is basic. [Note: Milk of magnesia is the common name for $Mg(OH)_2$.]
B. With a pH = 7, a salt solution is neutral.
C. With a pH < 7, wine is acidic.

✏️ Exercise 15.7.1

Identify each substance as acidic, basic, or neutral.

A. Human blood with a pH = 7.4
B. Household ammonia with a pH = 11.0
C. Cherries with a pH = 3.6

Answer A

basic

Answer B

basic

Answer C

acidic

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15.8: pH and pOH Calculations

Learning Objectives

- Define pH and pOH.
- Determine the pH of acidic and basic solutions.
- Determine the hydronium ion concentration and pOH from pH.

The pOH scale

In the previous section, the **pH** was defined as the negative logarithm of the hydronium ion concentration:

$$\mathbf{pH} = -\log\left[\mathbf{H_3O^+}\right] \tag{15.8.1}$$

It is likely you have only heard of the pH scale. However, there is a pH counterpart called the **pOH** (the "power of the hydroxide ion"), which is defined as the negative logarithm of the hydroxide ion concentration:

$$\mathbf{pOH} = -\log\left[\mathbf{OH^-}\right] \tag{15.8.2}$$

For aqueous solutions at 25°C, the sum of the pH and pOH is always 14.00:

$$\mathbf{pH} + \mathbf{pOH} = 14.00 \tag{15.8.3}$$

While its derivation is beyond the scope of this text, Equation 15.8.3 is a mathematical extension of the relationship between the $[\mathbf{H_3O^+}]$ and the $[\mathbf{OH^-}]$, established in Section 15.6, in which their product is 1.0×10^{-14} at 25°C:

$$[\mathbf{H_3O^+}][\mathbf{OH^-}] = 1.0 \times 10^{-14}$$

Simple pH and pOH Calculations

In pure water, we have already established that $[\mathbf{H_3O^+}] = [\mathbf{OH^-}] = 10^{-7}$ **M**. Substituting $[\mathbf{OH^-}] = 10^{-7}$ **M** into the equation, $\mathbf{pOH} = -\log\left[\mathbf{OH^-}\right]$:

$$\mathbf{pOH} = -\log\left[10^{-7}\right] = -(-7) = 7$$

Thus, both the pH = 7 and pOH = 7 for pure water. This makes sense, since the pH and pOH should sum to 14, as shown in Equation 15.8.3 above. As previously noted, temperature matters. However, if the temperature is not cited, a temperature of 25°C (room temperature) may be assumed.

Example 15.8.1

What is the pH of these solutions?

A. pOH = 5.55
B. $[\mathbf{H_3O^+}] = 10^{-11}$ M
C. $[\mathbf{OH^-}] = 10^{-8}$ M

Solution

A. From Equation 15.8.3, pH + pOH = 14.00. Therefore, pH = 14.00 – pOH = 14.00 – 5.55 = 8.45.
B. From Equation 15.8.1, $\mathbf{pH} = -\log\left[\mathbf{H_3O^+}\right] = -\log\left(10^{-11}\right) = -(-11) = 11$
C. From Equation 15.8.2, $\mathbf{pOH} = -\log\left[\mathbf{OH^-}\right] = -\log\left(10^{-8}\right) = -(-8) = 8$. From Equation 15.8.3, pH + pOH = 14.00. Therefore, pH = 14.00 – pOH = 14.00 – 8 = 6.

Exercise 15.8.1

What is the pOH of these solutions?

A. $[\mathbf{OH^-}] = 10^{-4}$ M

B. $[H_3O^+] = 10^{-12}$ M

C. pH = 9.01

Answer A

pOH = 4

Answer B

pOH = 2

Answer C

pOH = 4.99

Figure 15.8.1 summarizes the relationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH. Note that the product of $[H_3O^+]$ and $[OH^-]$ is always 1.0×10^{-14} M and that the sum of the pH and pOH is always 14.

$[H_3O^+]$	$[OH^-]$		pH	pOH
1 M	10^{-14} M		0	14
10^{-1} M	10^{-13} M		1	13
10^{-2} M	10^{-12} M		2	12
10^{-3} M	10^{-11} M	acidic	3	11
10^{-4} M	10^{-10} M		4	10
10^{-5} M	10^{-9} M		5	9
10^{-6} M	10^{-8} M		6	8
10^{-7} M	10^{-7} M	neutral	7	7
10^{-8} M	10^{-6} M		8	6
10^{-9} M	10^{-5} M		9	5
10^{-10} M	10^{-4} M		10	4
10^{-11} M	10^{-3} M	basic	11	3
10^{-12} M	10^{-2} M		12	2
10^{-13} M	10^{-1} M		13	1
10^{-14} M	1 M		14	0

Figure 15.8.1: A table summarizing the relationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH. (ChemLancer)

pH and pOH Calculations That Require a Calculator

Calculations more complex than the ones shown above likely require a calculator. It is impossible to provide instructions for every available model of calculator, but we will look at three different models. The large majority of calculators share one of the sequences used by the calculators shown in Figure 15.8.2.

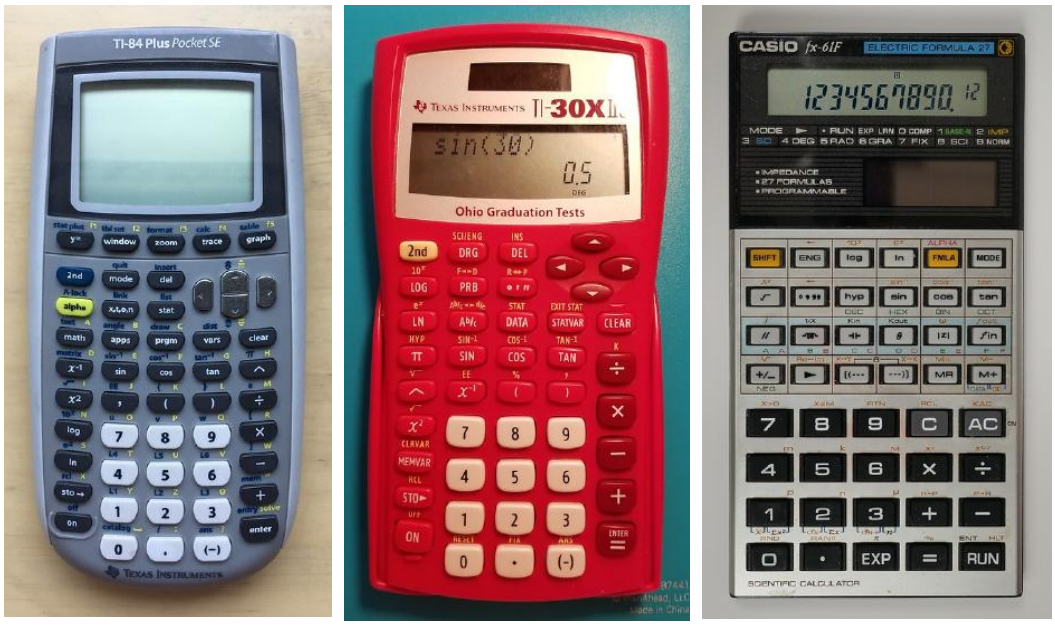


Figure 15.8.2: Three different models of calculators (left to right): TI-84 Plus (bfishadow via Flickr), TI-30XIIS (Alice H. via Flickr), Casio fx-61F (Dave Jones via Flickr)

✅ Example 15.8.2

Calculate the pH of the following solutions:

A. $[H_3O^+] = 5.6 \times 10^{-5}$ M
B. $[H_3O^+] = 5.7 \times 10^{-5}$ M
C. $[H_3O^+] = 5.7 \times 10^{-4}$ M

Solution

A. $\mathbf{pH = -\log[H_3O^+] = -\log(5.6 \times 10^{-5}) = 4.2518 = \boxed{4.25}}$
B. $\mathbf{pH = -\log[H_3O^+] = -\log(5.7 \times 10^{-5}) = 4.2441 = \boxed{4.24}}$
C. $\mathbf{pH = -\log[H_3O^+] = -\log(5.7 \times 10^{-4}) = 3.2441 = \boxed{3.24}}$

The procedure for determining significant figures in pH calculations is described below. The method for using a calculator to calculate pH for Part A follows.

Look carefully at the $[H_3O^+]$ for Parts A and B in Example 15.8.2 above. Each measurement has two significant figures and the uncertain digit is the second significant figure (last digit). Upon calculating the pH, we see in the *Solution* that the uncertain digit is the second digit after the decimal. When comparing the $[H_3O^+]$ for Parts B and C, we notice that changing the exponent on the power of 10 changes the number that is in front of the decimal in the calculated pH.

Since each measurement can only have one uncertain digit, this leads us to conclude:

- The calculated pH will have the same number of decimal places as the number of significant figures in the $[H_3O^+]$.

In other words, if the $[H_3O^+]$ has only one significant figure, then the calculated pH will only have one decimal place. A similar statement may be made about the number of significant figures with regard to pOH and the $[OH^-]$:

- The calculated pOH will have the same number of decimal places as the number of significant figures in the $[H_3O^+]$.

✍ Using a Scientific Calculator to Calculate pH

To calculate the answer to Example 15.8.2 Part A:

On a TI-83, TI-84, or TI-30XIIS:

- (–) log 5.6 2nd EE (–) 5)
- The EE function is located above the , button on a TI-83 and TI-84. It is located above the x^{-1} button on a TI-30XIIS.
- At this point your display should look like this: $-\log(5.6\text{E}-5)$ and you can press the Enter button.
- An answer of 4.2518... will be displayed and need to be rounded to the proper number of significant figures.

On a Casio fx-61F:

- 5.6 EXP 5 +/– log +/– =
- An answer of 4.2518... will be displayed and need to be rounded to the proper number of significant figures.

✅ Example 15.8.3

Calculate the pH of the following solutions:

A. 0.10 M HCl
B. 0.0044 M $Ca(OH)_2$

Solution

Steps for Problem Solving	0.10 M HCl	0.0044 M $Ca(OH)_2$
Identify the "given" information and what the problem is asking you to "find."	Given: HCl = 0.10 M Find: pH	Given: $Ca(OH)_2$ = 0.0044 M Find: pH
List known relationship(s).	HCl is a strong acid. Therefore: $[H_3O^+] = [\text{acid}]$ $pH = -\log[H_3O^+]$	$Ca(OH)_2$ is a strong base. Therefore: $[OH^-] = [\text{base}] \times$ number of OH^- ions in the chemical formula $pOH = -\log[OH^-]$ $pH + pOH = 14.00$
Plan the problem.	Calculate the pH by plugging the $[H_3O^+]$ into the equation.	Calculate the $[OH^-]$. Calculate the pOH by plugging the $[OH^-]$ into the equation. Calculate the pH by rearranging and plugging in the pOH: pH = 14.00 – pOH
Calculate the answers.	$pH = -\log[H_3O^+] = -\log(0.10) = \boxed{1.00}$	$[OH^-] = 0.0044\ M \times 2 = 0.0088\ M$ $pOH = -\log[OH^-] = -\log(0.0088) = 2.06$ $pH = 14.00 - pOH = 14.00 - 2.06 = \boxed{11.94}$
Think about your results.	HCl is a strong acid. A pH = 1.00 would be very acidic, so the answer makes sense.	$Ca(OH)_2$ is a strong base. A pH = 11.94 would be quite basic, so the answer makes sense.

Exercise 15.8.2

What is the pH of these solutions?

A. 0.10 M NaOH
B. 6.1×10^{-3} M HI
C. $[OH^-] = 2.0 \times 10^{-7}$ M

Answer A

pH = 13.00

Answer B

pH = 2.21

Answer C

pH = 7.30

Calculating Hydronium Concentration from pH

In some cases, the pH of a solution is known and the $[H_3O^+]$ or concentration of the acid solution is needed. To convert pH to $[H_3O^+]$, Equation 15.8.1 may be rearranged to solve for $[H_3O^+]$. This involves taking the antilog (or inverse log) of the negative value of pH. Yeah, sounds complicated...so we'll spare you the details and show you the end result:

$$[H_3O^+] = 10^{-pH} \tag{15.8.4}$$

A similar equation exists to retrieve the $[OH^-]$ from the pOH:

$$[OH^-] = 10^{-pOH} \tag{15.8.5}$$

✅ Example 15.8.4

Calculate the hydronium ion and hydroxide ion concentrations of a solution with a pH of 12.60.

Solution

Steps for Problem Solving	
Identify the "given" information and what the problem is asking you to "find."	Given: pH = 12.60 Find: $[H_3O^+]$ = ? M; $[OH^-]$ = ? M
List known relationship(s).	$[H_3O^+] = 10^{-pH}$ $pH + pOH = 14.00$ $[OH^-] = 10^{-pOH}$
Plan the problem.	Find $[H_3O^+]$ by plugging pH into the equation. Calculate the pOH by rearranging and plugging in the pH: pOH = 14.00 – pH Find $[OH^-]$ by plugging pOH into the equation.
Calculate the answers.	$[H_3O^+] = 10^{-pH} = 10^{-12.60} = \boxed{2.5 \times 10^{-13}\ M}$ $pOH = 14.00 - pH = 14.00 - 12.60 = 1.40$ $[OH^-] = 10^{-pOH} = 10^{-1.4} = \boxed{0.040\ M}$
Think about your results.	The $[H_3O^+]$ and $[OH^-]$ should each have two significant figures, since the pH had two decimal places. A quick check shows that $[H_3O^+]\,[OH^-] = (2.5 \times 10^{-13})(0.040) = 1.0 \times 10^{-14}$.

✍ Using a Scientific Calculator to Calculate $[H_3O^+]$

To calculate the $[H_3O^+]$ in Example 15.8.4 above:

On a TI-83, TI-84, or TI-30XIIS:

- 10 [^] [(] [(–)] 12.60 [)]
- At this point your display should look like this: $10(-12.60)$ and you can press the [Enter] button.
- An answer of 2.511886...E–13 will be displayed and will need to be rewritten in scientific notation and rounded to the proper number of significant figures (2.5×10^{-13}).

On a Casio fx-61F:

- 10 [SHIFT] [x^y] 12.60 [+/–] [=]
- An answer of 2.511886...$^{-13}$ will be displayed and is very easily misinterpreted. It needs to be properly rewritten in scientific notation rounded to the proper number of significant figures (2.5×10^{-13}).

Exercise 15.8.3

A. If moist soil has a pH of 7.84, find the $[H_3O^+]$ and $[OH^-]$ of the soil sample.
B. When lime (CaO) is added to this soil, its pH increases to 8.42. Find the new $[H_3O^+]$ and $[OH^-]$ of the soil sample.
C. When sulfur is added to this soil, its pH decreases to 6.57. What is the new $[H_3O^+]$ and $[OH^-]$ of the sample?

Answer A

$[H_3O^+] = 1.4 \times 10^{-8}$ M; $[OH^-] = 6.9 \times 10^{-7}$ M

Answer B

$[H_3O^+] = 3.8 \times 10^{-9}$ M; $[OH^-] = 2.6 \times 10^{-6}$ M

Answer C

$[H_3O^+] = 2.7 \times 10^{-7}$ M; $[OH^-] = 3.7 \times 10^{-8}$ M

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15.9: Reactions of Acids and Bases

For a review of acid-base reactions and reactions of acids with sulfides, carbonates and bicarbonates, sulfites and bisulfites, and metals, visit Section 7.9.

Contributions & Attributions

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15.11: Exercises

15.4: Acids and Bases Defined

1. Write equations showing $H_2PO_4^-$ acting as both a Brønsted-Lowry acid and as a base with water.

Answer

acid: $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$; base: $H_2PO_4^- + H_2O \rightleftharpoons H_3PO_4 + OH^-$

2. Write the equation for the chemical reaction that occurs when caffeine ($C_8H_{10}N_4O_2$) acts as a Brønsted-Lowry base in water.

Answer

$C_8H_{10}N_4O_2 + H_2O \rightleftharpoons HC_8H_{10}N_4O_2^+ + OH^-$

3. Citric acid ($C_6H_8O_7$) is the acid found in citrus fruits. It can lose a maximum of three H^+ ions in the presence of a base. Write the chemical equations for citric acid acting stepwise as a Brønsted-Lowry acid in water.

Answer

$C_6H_8O_7 + H_2O \rightleftharpoons C_6H_7O_7^- + H_3O^+$; $C_6H_7O_7^- + H_2O \rightleftharpoons C_6H_6O_7^{2-} + H_3O^+$; $C_6H_6O_7^{2-} + H_2O \rightleftharpoons C_6H_5O_7^{3-} + H_3O^+$

4. What is the conjugate acid of each of the following?
   a. OH^-
   b. H_2O
   c. HCO_3^-
   d. NH_3
   e. HSO_4^-
   f. H_4N_2

Answer

   a. H_2O
   b. H_3O^+
   c. H_2CO_3
   d. NH_4^+
   e. H_2SO_4
   f. $H_5N_2^+$

5. What is the conjugate base of each of the following?
   a. $H_2PO_4^-$
   b. HS^-
   c. $H_3O_2^+$
   d. H_4N_2
   e. H_2S
   f. CH_3OH

Answer

   a. HPO_4^{2-}

b. S^{2-}
c. H_2O_2
d. $H_3N_2^-$
e. HS^-
f. CH_3O^-

6. Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:
   a. $HNO_3 + H_2O \rightleftharpoons H_3O^+ + NO_3^-$
   b. $CN^- + H_2O \rightleftharpoons HCN + OH^-$
   c. $H_2SO_4 + Cl^- \rightleftharpoons HCl + HSO_4^-$
   d. $HSO_4^- + OH^- \rightleftharpoons SO_4^{2-} + H_2O$

Answer

a. BL acid = HNO_3; conjugate base = NO_3^-; BL base = H_2O; conjugate acid = H_3O^+
b. BL acid = H_2O; conjugate base = OH^-; BL base = CN^-; conjugate acid = HCN
c. BL acid = H_2SO_4; conjugate base = HSO_4^-; BL base = Cl^-; conjugate acid = HCl
d. BL acid = HSO_4^-; conjugate base = SO_4^{2-}; BL base = OH^-; conjugate acid = H_2O

7. Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:
   a. $NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$
   b. $HBr + H_2O \rightleftharpoons H_3O^+ + Br^-$
   c. $HS^- + H_2O \rightleftharpoons H_2S + OH^-$
   d. $H_2PO_4^- + HCl \rightleftharpoons H_3PO_4 + Cl^-$

Answer

a. BL acid = H_2O; conjugate base = OH^-; BL base = NO_2^-; conjugate acid = HNO_2
b. BL acid = HBr; conjugate base = Br^-; BL base = H_2O; conjugate acid = H_3O^+
c. BL acid = H_2O; conjugate base = OH^-; BL base = HS^-; conjugate acid = H_2S
d. BL acid = HCl; conjugate base = Cl^-; BL base = $H_2PO_4^-$; conjugate acid = H_3PO_4

15.5: Strong and Weak Acids and Bases

8. Identify each of the following acids or bases as strong or weak:
   a. $HC_2H_3O_2$
   b. $Ba(OH)_2$
   c. $HClO_4$
   d. CH_3NH_2

Answer

a. weak acid
b. strong base
c. strong acid
d. weak base

9. Identify each of the following acids or bases as strong or weak:
    a. NaOH
    b. $C_{17}H_{21}NO_4$ (cocaine)
    c. H_2SO_3
    d. HBr

Answer

    a. strong base
    b. weak base
    c. weak acid
    d. strong acid

15.6: Water as an Acid and as a Base

10. Are the concentrations of hydronium ion and hydroxide ion in a solution of an acid or a base in water directly proportional or inversely proportional? Explain.

Answer

Inversely proportional. As the concentration of one increases, the concentration of the other will decrease.

11. The hydronium ion concentration in a sample of rainwater is found to be 1.7×10^{-6} M. What is the concentration of hydroxide ions in the rainwater?

Answer

$[OH^-] = 5.9 \times 10^{-9}$ M

12. The hydroxide ion concentration in household ammonia is 3.2×10^{-3} M. What is the concentration of hydronium in the solution?

Answer

$[H_3O^+] = 3.1 \times 10^{-12}$ M

13. Calculate the hydroxide ion concentration for each of the following solutions. Is the solution acidic, basic, or neutral?
    a. $[H_3O^+] = 3.5 \times 10^{-4}$ M
    b. $[H_3O^+] = 8.1 \times 10^{-8}$ M
    c. $[H_3O^+] = 7.0 \times 10^{-3}$ M

Answer

    a. $[OH^-] = 2.9 \times 10^{-11}$ M; acidic
    b. $[OH^-] = 1.2 \times 10^{-7}$ M; basic
    c. $[OH^-] = 1.4 \times 10^{-12}$ M; acidic

14. Calculate the hydronium ion concentration for each of the following solutions. Is the solution acidic, basic, or neutral?
    a. $[OH^-] = 1.0 \times 10^{-7}$ M
    b. $[OH^-] = 9.3 \times 10^{-11}$ M

c. $[OH^-] = 2.5 \times 10^{-2}$ M

Answer

a. $[H_3O^+] = 1.0 \times 10^{-7}$ M; neutral
b. $[H_3O^+] = 1.1 \times 10^{-4}$ M; acidic
c. $[H_3O^+] = 4.0 \times 10^{-13}$ M; basic

15. Calculate the hydronium and hydroxide ion concentrations of the following solutions:

a. 0.200 M HCl
b. 0.0143 M NaOH
c. 3.0 M HNO_3
d. 0.0031 M $Ca(OH)_2$

Answer

a. $[H_3O^+] = 0.200$ M; $[OH^-] = 5.0 \times 10^{-14}$ M
b. $[H_3O^+] = 6.99 \times 10^{-13}$ M; $[OH^-] = 0.0143$ M
c. $[H_3O^+] = 3.0$ M; $[OH^-] = 3.3 \times 10^{-15}$ M
d. $[H_3O^+] = 1.6 \times 10^{-12}$ M; $[OH^-] = 0.0062$ M

16. Calculate the hydronium and hydroxide ion concentrations of each of the following solutions:

a. 0.000259 M $HClO_4$
b. 0.21 M NaOH
c. 0.000071 M $Ba(OH)_2$
d. 2.5 M KOH

Answer

a. $[H_3O^+] = 0.000259$ M; $[OH^-] = 3.86 \times 10^{-11}$ M
b. $[H_3O^+] = 4.8 \times 10^{-14}$ M; $[OH^-] = 0.21$ M
c. $[H_3O^+] = 7.1 \times 10^{-11}$ M; $[OH^-] = 0.00014$ M
d. $[H_3O^+] = 4.0 \times 10^{-15}$ M; $[OH^-] = 2.5$ M

15.7: An Introduction to pH

17. Identify each of the following solutions as acidic, basic, or neutral.

a. pH = 8.2 (baking soda)
b. pH = 2.54 (vinegar)
c. pH = 9.4 (laundry detergent)

Answer

a. basic
b. acidic
c. basic

18. Identify each of the following solutions as acidic, basic, or neutral.

a. pH = 7.44 (human tears)
b. pH = 3.4 (white wine)

c. pH = 4.85 (coffee)

Answer

a. basic
b. acidic
c. acidic

19. Normal clean rain has a pH of 5.0 to 5.5. Typical acid rain has a pH of 4.0. How much more acidic is acid rain than normal rain with a pH of 5.0?

Answer

One pH unit corresponds to a factor of 10. So typical acid rain is 10 times more acidic than normal rain.

20. When a sparkling water was first opened, it had a pH of 3.7. Tap water was found to have a pH of 6.7. Which water is more basic? How much more basic is it?

Answer

The tap water has a higher pH, so it is more basic. The tap water's pH is 3 pH units greater. This means that the tap water is 10^3 or 1000 times more basic.

15.8: pH and pOH Calculations

21. What is the pOH range for an acidic solution?

Answer

The pOH is greater than 7 for an acidic solution (typically up to 14).

22. What is the pOH range for a basic solution?

Answer

The pOH is less than 7 for a basic solution (typically down to 0).

23. Calculate the pH of each solution and classify it as acidic or basic.

a. pOH = 5.18
b. pOH = 9.26
c. pOH = 2.84

Answer

a. pH = 8.82; basic
b. pH = 4.74; acidic
c. pH = 11.16; basic

24. Calculate the pOH of each solution and classify it as acidic or basic.

a. pH = 8.31

b. pH = 5.20
c. pH = 7.43

Answer

a. pOH = 5.69; basic
b. pOH = 8.80; acidic
c. pOH = 6.57; basic

25. The concentration of commercial HCl is about 12 M. What is its pH and pOH?

Answer

pH = −1.08; pOH = 15.08

26. The concentration of concentrated H_2SO_4 is about 18 M. Assuming only one H^+ comes off the H_2SO_4 molecule, what is its pH and pOH?

Answer

pH = −1.26; pOH = 15.26

27. Calculate the pH and pOH of each of the following solutions:

a. 0.200 M HCl
b. 0.0143 M NaOH
c. 3.0 M HNO_3
d. 0.0031 M $Ca(OH)_2$

Answer

a. pH = 0.699; pOH = 13.301
b. pH = 12.155; pOH = 1.845
c. pH = −0.48; pOH = 14.48
d. pH = 11.79; pOH = 2.21

28. Calculate the pH and pOH of each of the following solutions:

a. 0.000259 M $HClO_4$
b. 0.21 M NaOH
c. 0.000071 M $Ba(OH)_2$
d. 2.5 M KOH

Answer

a. pH = 3.587; pOH = 10.413
b. pH = 13.32; pOH = 0.68
c. pH = 10.15; pOH = 3.85
d. pH = 14.40; pOH = −0.40

29. What are the hydronium and hydroxide ion concentrations in a solution whose pH is 6.52?

Answer

$[H_3O^+] = 3.0 \times 10^{-7}$ M; $[OH^-] = 3.3 \times 10^{-8}$ M

30. What are the hydronium and hydroxide ion concentrations in a solution whose pH is 9.20?

Answer

$[H_3O^+] = 6.3 \times 10^{-10}$ M; $[OH^-] = 1.6 \times 10^{-5}$ M

31. What are the hydronium and hydroxide ion concentrations in a solution whose pOH is 2.18?

Answer

$[H_3O^+] = 1.5 \times 10^{-12}$ M; $[OH^-] = 6.6 \times 10^{-3}$ M

15.9: Reactions of Acids and Bases

32. Write the balanced chemical equation between Zn metal and HCl(*aq*). The other product is $ZnCl_2$.

Answer

$Zn(s) + 2\ HCl(aq) \rightleftharpoons ZnCl_2(aq) + H_2(g)$

33. Write the neutralization reaction in which $ZnCl_2$, also found in Exercise 32, is the salt product.

Answer

$Zn(OH)_2 + 2\ HCl \rightleftharpoons ZnCl_2 + 2\ H_2O$

34. Write the balanced chemical equation between aluminum hydroxide and sulfuric acid.

Answer

$2\ Al(OH)_3 + 3\ H_2SO_4 \rightleftharpoons Al_2(SO_4)_3 + 6\ H_2O$

35. Write the balanced chemical equation between phosphoric acid and barium hydroxide.

Answer

$2\ H_3PO_4 + 3\ Ba(OH)_2 \rightleftharpoons Ba_3(PO_4)_2 + 6\ H_2O$

36. Complete and balance the following acid-base reactions:

a. HCl solution reacts with solid $Ca(OH)_2(s)$
b. a solution of $Sr(OH)_2$ is added to a solution of HNO_3

Answer

a. $2\ HCl(aq) + Ca(OH)_2(s) \rightleftharpoons CaCl_2(aq) + 2\ H_2O(l)$
b. $Sr(OH)_2(aq) + 2\ HNO_3(aq) \rightleftharpoons Sr(NO_3)_2(aq) + 2\ H_2O(l)$

37. Complete and balance the following acid-base reactions:

a. aqueous H_2SO_4 reacts with NaOH
b. $Ba(OH)_2$ reacts with HF gas

Answer

a. $H_2SO_4(aq) + 2\ NaOH(aq) \rightleftharpoons Na_2SO_4(aq) + 2\ H_2O(l)$
b. $Ba(OH)_2(aq) + 2\ HF(g) \rightleftharpoons BaF_2(aq) + 2\ H_2O(l)$

38. Gastric juice, the digestive fluid produced in the stomach, contains hydrochloric acid, HCl. Milk of Magnesia, a suspension of solid $Mg(OH)_2$ in an aqueous medium, is sometimes used to neutralize excess stomach acid. Write a complete balanced equation for the neutralization reaction.

Answer

$2\ HCl(aq) + Mg(OH)_2(s) \rightleftharpoons MgCl_2(aq) + 2\ H_2O(l)$

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CHAPTER OVERVIEW

16: Environmental Chemistry

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16.1: Chemistry is Good...and Unavoidable

Many people are intimidated by the idea of chemistry and try to avoid it. But avoiding chemistry is impossible. That is because all matter — the air we must breathe, the water we must drink, and all living organisms — is made of chemicals. People who try to avoid everything that they regard as chemical may fail to realize that chemical processes are continuously being carried out in their own bodies. These are processes that far surpass in complexity and variety those that occur in chemical manufacturing operations. So, even those people who want to do so cannot avoid chemistry; they are, themselves, complex chemical factories. The best course of action with anything that cannot be avoided and that might have an important influence on our lives is to try to understand it. Hopefully that has happened as you have used this book.

During the approximately two centuries that chemical science has been practiced on an ever-increasing scale, it has enabled the production of a wide variety of goods that are valued by humans. These include such things as pharmaceuticals that have improved health and extended life, fertilizers that have greatly increased food productivity and prevented widespread starvation, and semiconductors that have made possible computers and other electronic devices. Without the persistent efforts of chemists and the enormous productivity of the chemical industry, nothing approaching the high standard of living enjoyed in modern industrialized societies would be possible.

But there can be no denying that in years past, and even at present, chemistry has been misused in many respects, such as the release of pollutants and toxic substances and the production of non-biodegradable materials, which has resulted in harm to the environment and living things, including humans. As the chemical industry developed and grew during the early and mid-1900s, most practitioners of chemistry remained unconcerned with and largely ignorant of the potential for harm of their products and processes. Environmental chemistry was essentially unknown and certainly not practiced by most chemists. Incidents of pollution and environmental damage, which were many and severe, were commonly accepted as a cost of doing business or blamed upon the industrial or commercial sectors. The unfortunate attitude that prevailed is summarized in a quote from a standard book on industrial chemistry from 1954 (*American Chemical Industry—A History*, W. Haynes, Van Nostrand Publishers, 1954): “By sensible definition any by-product of a chemical operation for which there is no profitable use is a waste. The most convenient, least expensive way of disposing of said waste — up the chimney or down the river — is best."



Figure 16.1.1: A 2015 view of the Yellow River as it passes between Wubu and Jundu, China. The impact of chemical waste can be seen in the river itself and in the atmosphere above the cities. (N509FZ, CC BY-SA 4.0, via Wikimedia Commons)

It is now obvious that chemical science must be turned away from emphasis upon the exploitation of limited resources and the production of increasing amounts of products that ultimately end up as waste and toward the application of chemical science in ways that provide for human needs without damaging the Earth's support system and depleting its natural capital upon which all living things depend. Fortunately, the practice of chemical science and industry is moving steadily in the direction of environmental

friendliness and resource sustainability. The practice of chemistry in a manner that maximizes its benefits while eliminating, or at least greatly reducing, its adverse impacts has come to be known as **green chemistry**.

Despite their potential to cause harm, nobody is more qualified to accept responsibility for environmental damage from chemical products or processes than are chemists who have the knowledge to understand how such harmful effects came about. As the detrimental effects of chemical manufacture and use became more obvious and severe, chemists were forced, often reluctantly, to deal with them. At present, enlightened chemists and chemical engineers do not view the practice of environmentally beneficial chemistry and manufacturing as a burden, but rather as an opportunity that challenges human imagination and ingenuity.

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16.2: Chemistry and the Environment

Learning Objectives

- Identify the five spheres of the environment and how they can interact with each other.
- Identify the areas of environmental chemistry that relate to each of the five spheres.

Environmental Spheres

The **environment** consists of our surroundings, which may affect us and which, in turn, we may affect. Obviously the chemical nature and processes of matter in the environment are important. Compared to the generally well-defined processes that chemists study in the laboratory, those that occur in the environment are rather complex and must be viewed in terms of simplified models. A large part of this complexity is due to the fact that the environment consists of the five overlapping and interacting spheres — the atmosphere, the hydrosphere, the geosphere, the biosphere, and the anthrosphere — shown from the viewpoint of their chemical interactions in Figure **16.2.1**. In order to better understand the chemistry that occurs in these spheres, they are briefly described here.

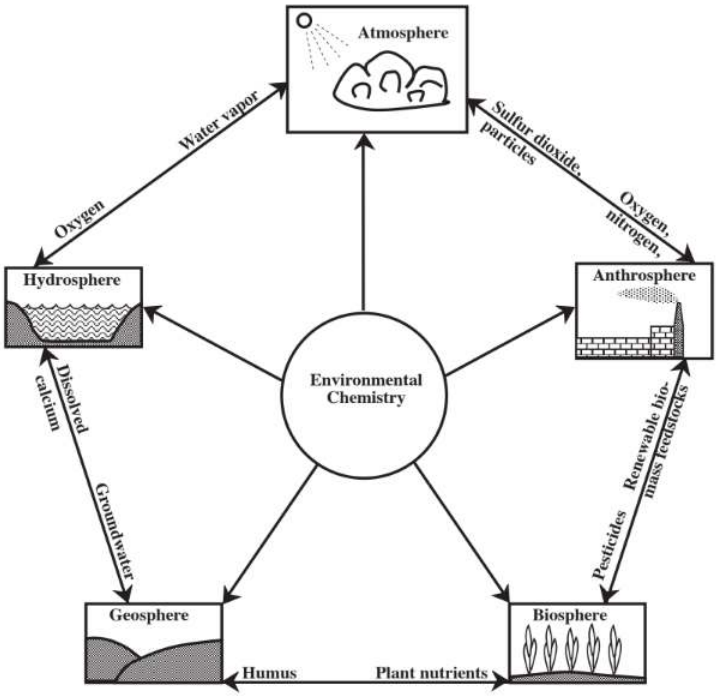


Figure **16.2.1**. Illustration of the five major spheres of the environment and some of their chemical interactions. These spheres are closely tied together, interact with each other, and exchange materials and energy. A meaningful examination of environmental sciences must include all five of these spheres, including the anthrosphere.

The **atmosphere** is a very thin layer compared to the size of Earth, with most atmospheric gases lying within a few kilometers of sea level. In addition to providing oxygen for living organisms, the atmosphere provides carbon dioxide required for plant photosynthesis, and nitrogen that organisms use to make proteins. The atmosphere serves a vital protective function in that it absorbs highly energetic ultraviolet radiation from the sun that would kill living organisms exposed to it. A particularly important part of the atmosphere in this respect is the stratospheric layer of ozone, an ultraviolet-absorbing form of elemental oxygen. In addition, the atmosphere stabilizes Earth's surface temperature by absorbing infrared radiation (heat) which the Earth would otherwise lose by radiating it out to space. The atmosphere also serves as the medium by which solar energy, that falls with greatest intensity in equatorial regions, is redistributed away from the Equator. It is the medium in which water vapor, evaporated from oceans as the first step in the hydrologic cycle, is transported over land masses to fall as rain over land.

Earth's water is contained in the **hydrosphere**. Although frequent reports of torrential rainstorms and flooded rivers produced by massive storms might give the impression that a large fraction of Earth's water is fresh water, in reality more than 97% of the water on Earth is seawater in the oceans. Of the remaining 3% that is fresh water, most is present as ice in polar ice caps and glaciers. A small fraction of the total water is present as vapor in the atmosphere. The remaining liquid fresh water is what is available for growing plants and other organisms and for industrial uses. This water may be present on the surface as lakes, reservoirs, and streams, or it may be underground as groundwater.

The solid part of earth, the **geosphere**, includes all rocks and minerals. A particularly important part of the geosphere is soil which supports plant growth, the basis of food for all living organisms. The **lithosphere** is a relatively thin solid layer extending from Earth's surface to depths of 50–100 km. The even thinner outer skin of the lithosphere known as the **crust** is composed of relatively lighter silicate-based minerals. It is the part of the geosphere that is available to interact with the other environmental spheres and that is accessible to humans.

The **biosphere** is composed of all living organisms. For the most part, these organisms live on the surface of the geosphere on soil, or just below the soil surface. The oceans and other bodies of water support high populations of organisms. Some life forms exist at considerable depths on ocean floors. In general, though, the biosphere is a very thin layer at the interface of the geosphere with the atmosphere. The biosphere is involved with the geosphere, hydrosphere, and atmosphere in *biogeochemical cycles* through which materials such as nitrogen and carbon are circulated.

Through human activities, the **anthrosphere**, that part of the environment made and operated by humans, has developed strong interactions with the other environmental spheres. Many examples of these interactions could be cited. By cultivating large areas of soil for domestic crops, humans modify the geosphere and influence the kinds of organisms in the biosphere. Humans divert water from its natural flow, use it, sometimes contaminate it, then return it to the hydrosphere. Emissions of particles to the atmosphere by human activities affect visibility and other characteristics of the atmosphere. The emission of large quantities of carbon dioxide to the atmosphere by combustion of fossil fuels are modifying the heat-absorbing characteristics of the atmosphere to the extent that global warming is almost certainly taking place. The anthrosphere perturbs various biogeochemical cycles.

Environmental Chemistry and the Five Spheres

Environmental chemistry, an important branch of chemical science, studies the sources, reactions, transport, effects, and fates of chemical species in water, soil, air, and living environments and the effects of technology on them. Figure **16.2.1**, which shows the five environmental spheres, may provide an idea of the complexity of environmental chemistry as a discipline. Enormous quantities of materials and energy are continually exchanged among the five environmental spheres. In addition to variable flows of materials, there are variations in temperature, intensity of solar radiation, mixing, and other factors, all of which strongly influence chemical conditions and behavior.

There are several highly interconnected and overlapping categories of environmental chemistry. **Atmospheric chemistry** is the branch of environmental chemistry that considers chemical phenomena in the atmosphere. Two things that make this chemistry unique are the extreme dilution of important atmospheric chemicals and the influence of photochemistry. Photochemistry occurs when molecules absorb photons of high-energy visible light or ultraviolet radiation, become energized ("excited"), and undergo reactions that lead to a variety of products, such as photochemical smog. In addition to reactions that occur in the gas phase, many important atmospheric chemical phenomena take place on the surfaces of very small solid particles suspended in the atmosphere and in droplets of liquid in the atmosphere. Although no significant atmospheric chemical reactions are mediated by organisms in the atmosphere, microorganisms play a strong role in determining species that get into the atmosphere. As examples, bacteria growing in the absence of oxygen, such as in cows' stomachs and under water in rice paddies, are the single greatest source of hydrocarbon in the atmosphere because of the large amounts of methane that they emit. The greatest source of organic sulfur compounds in the atmosphere consists of microorganisms in the oceans that emit dimethyl sulfide.

Aquatic chemistry deals with chemical phenomena and processes in water. Aquatic chemical processes are very strongly influenced by microorganisms in the water, so there is a strong connection between the hydrosphere and biosphere insofar as such processes are concerned. Aquatic chemical processes occur largely in "natural waters" consisting of water in oceans, bodies of

fresh water, streams, and underground aquifers. These are places in which the hydrosphere can interact with the geosphere, biosphere, and atmosphere and is often subjected to anthrospheric influences.

Chemical processes that occur in the geosphere involving minerals and their interactions with water, air, and living organisms are addressed by the topic of **geochemistry**. A special branch of geochemistry, soil chemistry, deals with the chemical and biochemical processes that occur in soil.

Environmental biochemistry addresses biologically mediated processes that occur in the environment. Such processes include the biodegradation of organic waste materials in soil or water and processes within biogeochemical cycles, such as denitrification, which returns chemically bound nitrogen to the atmosphere as nitrogen gas. The toxic effects of chemicals on a variety of living organisms are of utmost concern to chemists and the public as well.

Although there is not a formally recognized area of chemistry known as “anthrospheric chemistry,” most of chemical science and engineering developed to date deals with chemistry carried out in the anthrosphere. Included is industrial chemistry, which is very closely tied to the practice of green chemistry. A good way to view “anthrospheric chemistry” from a green chemistry perspective is within the context of **industrial ecology**. Industrial ecology considers industrial systems in a manner analogous to natural ecosystems. In a system of industrial ecology, various manufacturing and processing operations carry out “industrial metabolism” on materials. A successful industrial ecosystem is well balanced and diverse, with various enterprises that generate products for each other and use each other’s products and potential wastes. A well-functioning industrial ecosystem recycles materials to the maximum extent possible and produces little — ideally no — wastes. Therefore, a good industrial ecosystem is a green chemical system.

Summary

- The environment consists of five overlapping and interacting spheres: the atmosphere, the hydrosphere, the geosphere, the biosphere, and the anthrosphere.
- Environmental chemistry studies the sources, reactions, transport, effects, and fates of chemical species in water, soil, air, and living environments.
- There are categories of environmental chemistry which deal with each of the five spheres of the environment.

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16.3: The Air We Breathe

Learning Objectives

- Understand the composition of the air.
- Know the relative reactivities of the major components of air.

A Sea of Gas

We live and breathe in the **atmosphere**, a sea of gas consisting primarily of elemental O_2 and N_2. The fundamental properties of gases determine the properties of the atmosphere. Recall that gases consist of molecules or atoms (in the case of noble gases) with large amounts of space between them. The gas molecules are in constant, rapid motion which causes gases to exert pressure. The motion of gas molecules becomes more rapid with increasing temperature. The relationships among the the volume, temperature, and pressure of a gas can be calculated by the gas laws discussed in Chapter 9.

Whereas seawater in the ocean has a well-defined volume and a distinct surface, the same cannot be said for the mass of gases comprising the atmosphere. Although most of the atmosphere is within a few kilometers of Earth's surface, there is no distinct point at higher altitude where the atmosphere ends. Instead, air becomes progressively thinner with increasing altitude. This is noticeable to humans who have traveled to higher altitudes on mountains where the thinner air makes breathing more difficult. Indeed, climbers who scale the highest mountain peaks commonly carry oxygen to aid breathing.

Atmospheric Composition

What is air? At our level, it is a mixture of gases of uniform composition. On a dry basis, air is 78.1% (by volume) nitrogen, 21.0% oxygen, 0.9% argon, and 0.04% carbon dioxide. Normally, air is 1–3% water vapor by volume. Trace gases at levels below 0.002% in air include ammonia, carbon monoxide, helium, hydrogen, krypton, methane, neon, nitrogen dioxide, nitrous oxide (dinitrogen monoxide), ozone, sulfur dioxide, and xenon.

By a wide margin, oxygen and nitrogen are the most abundant gases in the atmosphere. Because of the extremely high stability and low reactivity of the N_2 molecule, the chemistry of atmospheric elemental nitrogen is singularly unexciting, although nitrogen molecules are the most common species that absorb excess energy from atmospheric chemical reactions to help stabilize those reactions. Elemental nitrogen is an important commercial gas extracted from the atmosphere by nitrogen-fixing bacteria and in the industrial synthesis of ammonia. Oxides of nitrogen actively participate in atmospheric chemical reactions.

Oxygen is a reactive species in the atmosphere. It will produce oxidation products from oxidizable gases in the atmosphere such as sulfur dioxide gas, SO_2, and pollutant hydrocarbons. Molecular O_2 does not react with these substances directly but only indirectly through the action of reactive intermediates, especially **hydroxyl radical**, HO•. A crucially important atmospheric chemical phenomenon involving oxygen is the formation of stratospheric ozone, O_3. Oxygen in the atmosphere is consumed in the burning of hydrocarbons and other carbon-containing fuels. It is also consumed when oxidizable minerals undergo chemical weathering, such as in the formation of Fe_2O_3, commonly known as rust.

$$4\ FeO + O_2 \rightarrow 2\ Fe_2O_3$$

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16.4: Air Pollution

Learning Objectives

- Define air pollution.
- Know the types and effects of major outdoor pollutants
- Know the sources of major outdoor pollutants
- Identify the different pollutant gases in automobile emissions

Air pollution refers to the introduction into the atmosphere of substances that have harmful effects on humans, other living organisms, and the environment. These substances can occur as solid particles, liquid droplets or gases. Air pollution can result from natural processes such as dust storms, forest fires, and volcanic eruptions, or from human activities such as biomass burning, vehicular emissions, mining, agriculture, and industrial processes. Improved technology and government policies have helped reduce most types of outdoor air pollution in many industrialized countries, including the United States, in recent decades. However, outdoor air quality is still a problem in less industrialized nations, especially in megacities of rapidly industrializing nations such as China and India.

Pollutants are categorized into two major types based on how they originated: namely primary and secondary pollutants. **Primary pollutants** are those released directly from the source into the air in a harmful form. The primary pollutants that account for nearly all air pollution problems are carbon monoxide (58%), volatile organic compounds (VOCs, 11%), nitrogen oxides (15%), sulfur dioxides (13%), and particulate material (3%). **Secondary pollutants** are produced through reactions between primary pollutants and normal atmospheric compounds. For example, ground-level ozone forms over urban areas through reactions powered by sunlight between primary pollutants (oxides of nitrogen) and other atmospheric gases such as VOCs.

Criteria pollutants

Under the US Clean Air Act, the Environmental Protection Agency (EPA) establishes air quality standards to protect public health and the environment. The EPA has set national air quality standards for six common air pollutants: carbon monoxide, ground-level ozone, nitrogen dioxide, sulfur dioxide, lead, and particulate matter (also known as particle pollution). The EPA calls these pollutants "criteria" air pollutants because it regulates them by developing human health-based and/or environmentally-based criteria (science-based guidelines) for setting permissible levels. The set of limits based on human health is called **primary standards**. Another set of limits intended to prevent environmental and property damage is called **secondary standards**. Of the six pollutants, particle pollution and ground-level ozone are the most widespread health threats.

1. **Carbon Monoxide (CO)** is a colorless, odorless gas emitted from combustion processes, specifically the incomplete combustion of fuel. Nationally, and particularly in urban areas, the majority of CO emissions to ambient air occurs when something is burned. The greatest sources of CO to outdoor air are cars, trucks and other vehicles or machinery that burn fossil fuels. A variety of items in your home such as unvented kerosene and gas space heaters, leaking chimneys and furnaces, and gas stoves also release CO and can affect air quality indoors.
2. **Ground-level ozone (O_3)** is a colorless gas with a slightly sweet odor that is not emitted directly into the air, but is created by the interaction of sunlight, heat, oxides of nitrogen (NO_x) and volatile organic compounds (VOCs). Ozone is likely to reach unhealthy levels on hot sunny days in urban environments. Emissions from industrial facilities and electric utilities, motor vehicle exhaust, gasoline vapors, and chemical solvents are some of the major sources of NO_x and VOCs. Ozone can be transported long distances by wind, so even rural areas can experience high ozone levels.
3. **Nitrogen dioxide (NO_2)** is one of a group of highly reactive gases known as "oxides of nitrogen," or "nitrogen oxides (NO_x)." Other nitrogen oxides include nitrous oxide (N_2O, dinitrogen monoxide), nitric oxide (NO, nitrogen monoxide), and nitric acid. NO_2 is a yellowish-brown to reddish-brown foul-smelling gas that is a major contributor to smog and acid rain. Nitrogen oxides result when atmospheric nitrogen and oxygen react at the high temperatures created by combustion engines. Most emissions in the U.S. result from combustion in vehicle engines, electrical utilities, and industrial combustion. They may also be produced by natural sources such as lightning. Nitrous oxide is generated by bacteria and its release is one of the ways in which chemically-fixed nitrogen in the biosphere is returned to the atmosphere.

4. **Sulfur dioxide (SO_2)** is one of a group of highly reactive gases known as “oxides of sulfur.” The largest sources of SO_2 emissions are from fossil fuel combustion at power plants (73%) and other industrial facilities (20%). Smaller sources of SO_2 emissions include industrial processes such as extracting metals from their ores, and the burning of high sulfur containing fuels by locomotives, large ships, and non-road equipment. They may also be directly released by volcanoes or formed by the oxidation of H_2S and dimethyl sulfide, $(CH_3)_2S$, emitted by bacteria and marine organisms, respectively.
5. **Lead (Pb)** is a metal found naturally in the environment as well as in manufactured products. The major sources of lead emissions have historically been from fuels in motor vehicles (such as cars and trucks) and industrial sources. As a result of EPA's regulatory efforts to remove lead from gasoline, emissions of lead from the transportation sector dramatically declined by 95 percent between 1980 and 1999, and levels of lead in the air decreased by 94 percent during that time period. The major sources of lead emissions today are ore and metal processing and piston-engine aircraft operating on leaded aviation gasoline. Today, the highest levels of lead in air are usually found near lead smelters.
6. **Particulate material (PM)**, sometimes known simply as “**particulates**” refers to solid particles and liquid droplets suspended in the air we breathe. Particulate pollution is made up of a variety of components, including acids (nitrates and sulfates), organic chemicals, metals, soil or dust particles, and allergens (pollen and mold spores). The size of the particles is directly linked to their potential for causing health problems. Particles that are 10 micrometers in diameter or smaller generally pass through the throat and nose and enter the lungs. EPA groups these into two types: "**inhalable coarse particles**" with diameters between 2.5 and 10 micrometers and "**fine particles**" with diameters that are 2.5 micrometers and smaller. How small is 2.5 micrometers? Think about a single hair from your head. The average human hair is about 70 micrometers in diameter – making it 30 times larger than the largest fine particle (Figure 16.4.2). Our respiratory systems are equipped to filter larger particles out of the air once it is inhaled. However, the lungs are vulnerable to both coarse particles (PM_{10}), and fine particles ($PM_{2.5}$). These can slip past the respiratory system's natural defenses and get deep into the lungs and some may even get into the bloodstream. Coarse particles come from road dust while fine particles come from combustion processes.

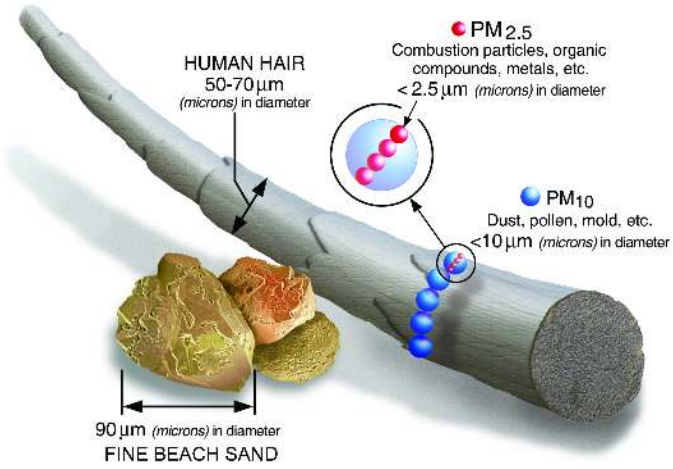


Figure 16.4.1: Graphic showing size comparison of particulate material (PM_{10} and $PM_{2.5}$) compared to fine beach sand and human hair. (Source: US EPA http://www.epa.gov/air/particlepollution/basic.html)

Volatile Organic Compounds

Volatile organic compounds (VOCs) are carbon-containing chemicals emitted as gases from natural and human-made sources. The largest natural source of VOCs is plants. They can also be emitted by bacteria in the guts of termites and ruminant animals. These compounds are generally oxidized to carbon monoxide and carbon dioxide in the atmosphere. VOCs are of great concern because they are precursors for the formation of ozone, a secondary air pollutant.

A large number of synthetic organic chemicals such as benzene, toluene, formaldehyde, vinyl chloride, chloroform, and phenols are widely used as ingredients in countless household products. Paints, paint strippers, varnishes, as well as many cleaning, disinfecting, cosmetic, degreasing, and hobby products all contain VOCs. Fuels are also made up of organic chemicals. All of these products can release organic compounds while you are using them and, to some degree, when they are stored. The “new car smell” characteristic of new cars is from a complex mix of dozens of VOCs. In addition, concentrations of many VOCs are consistently

higher indoors (up to ten times higher) than outdoors. They are often held responsible for sick building syndrome, an illness resulting from indoor pollution in which the specific cause is not identified.

Toxic pollutants

Toxic air pollutants, also known as hazardous air pollutants, are those pollutants that are known or suspected to cause cancer or other serious health effects, such as reproductive effects or birth defects, or adverse environmental effects. Examples of toxic air pollutants include **benzene**, which is found in gasoline; **perchloroethylene,** which is emitted from some dry cleaning facilities; **methylene chloride**, which is used as a solvent and paint stripper by a number of industries; and others such as dioxin, asbestos, toluene, and metals such as cadmium, mercury, chromium, and lead compounds.

Most air toxics originate from human-made sources, including vehicles, factories, refineries, and power plants, as well as indoor sources such as some building materials and cleaning solvents. Some air toxics are also released from natural sources such as volcanic eruptions and forest fires. Exposure to air toxics is mainly through breathing but some toxic air pollutants such as mercury can deposit onto soils or surface waters where they are taken up by plants and ingested by animals and are eventually magnified up through the food chain. Like humans, animals may experience health problems if exposed to sufficiently high quantities of air toxics over time.

Effects of Air Pollution on Human Health and the Environment

The World Health Organization (WHO) and other international agencies recognize air pollution as a major threat to human health. Numerous scientific studies have linked air pollution to a variety of health problems (Table 16.4.1) including aggravation of respiratory and cardiovascular diseases; decreased lung function; increased frequency and severity of respiratory symptoms such as difficulty breathing and coughing; increased susceptibility to respiratory infections; effects on the nervous system, including the brain, such as IQ loss and impacts on learning, memory, and behavior; cancer; and premature death. Immediate effects of air pollution may show up after a single exposure or repeated exposures. Other health effects may show up either years after exposure has occurred or only after long or repeated periods of exposure.

Immediate effects of air pollution include irritation of the eyes, nose, and throat, headaches, dizziness, and fatigue. Such immediate effects are usually short-term and treatable. Sometimes the treatment is simply eliminating the person's exposure to the source of the pollution, if it can be identified. Symptoms of some diseases, including asthma, hypersensitivity pneumonitis, and humidifier fever, may also show up soon after exposure to some indoor air pollutants.

Table 16.4.1: Health and environmental effects of criteria pollutants

Pollutant	Health Effects	Environmental Effects
Carbon Monoxide (CO)	Breathing elevated levels of CO reduces the amount of oxygen reaching the body's organs and tissues. For those with heart disease, this can result in chest pain and other symptoms leading to hospital admissions and emergency room (ER) visits. At very high levels, CO can cause dizziness, confusion, unconsciousness, and death.	Contributes to the formation of CO_2 and ozone, greenhouse gases that warm the atmosphere.
Ground-level Ozone (O_3)	Reduces lung function and causes respiratory symptoms, such as coughing and shortness of breath. Aggravates asthma and other lung diseases leading to increased medication use, hospital admissions, and ER visits. May also increase premature mortality from respiratory causes.	Damages vegetation by injuring leaves, reducing photosynthesis, impairing reproduction and growth and decreasing crop yields. Damage may alter ecosystem structure, reduce biodiversity, and decrease plant uptake of CO_2. Is a greenhouse gas and main ingredient of smog.
Oxides of Nitrogen (NO_x)	Aggravates respiratory diseases, particularly asthma, leading to respiratory symptoms, hospital admissions, and ER visits. Long-term exposure may contribute to asthma development and increased susceptibility to respiratory infection.	React with VOCs to form ozone. React with ammonia and other compounds to form particle pollution. Contributes to acidification and nutrient enrichment of soils and surface waters.
Sulfur Dioxide (SO_2)	Aggravates asthma and increases respiratory symptoms. Contributes to particle formation with associated health risks.	Reacts with ammonia and other compounds to form particle pollution. Contributes to acidification of soils and surface waters. Can also cause injury to vegetation and species loss in aquatic and terrestrial systems.
Lead	Damages the developing nervous system of children, resulting in lower IQs, learning deficits and behavioral problems. Contributes to cardiovascular effects such as high blood pressure and heart disease in adults.	Decreases growth and reproductive rates in plants and animals.
Particulate Matter (PM)	Causes harmful effects on the cardiovascular system, including heart attacks and strokes. Linked to harmful respiratory effects, including asthma attacks.	Causes reduced visibility (haze) in parts of US, including parks and wilderness areas. May make lakes and streams acidic, change nutrient balance in coastal waters and large river basins, damage sensitive forests and farm crops, and affect diversity of ecosystems. Stains and damages stone and other materials.

Source: www.epa.gov

The likelihood of immediate reactions to air pollutants depends on several factors. Age and preexisting medical conditions are two important influences. Some sensitive individuals appear to be at greater risk for air pollution-related health effects, for example, those with pre-existing heart and lung diseases (such as heart failure/ischemic heart disease, asthma, emphysema, and chronic bronchitis), diabetics, older adults, and children. In other cases, whether a person reacts to a pollutant depends on individual sensitivity, which varies tremendously from person to person. Some people can become sensitized to biological pollutants after repeated exposures, and it appears that some people can become sensitized to chemical pollutants as well.

In addition to posing a threat to human health, many air pollutants have damaging effects on the environment as a whole. Both nitrogen oxides and sulfur dioxide react with gases in the atmosphere and then combine with water droplets to form acid rain. As will be seen in Section 16.5, ground-level ozone is a main ingredient in "smog" that occurs in urban areas and results in lowers visibility. The haziness can be aggravated by particulate matter formed by nitrogen oxides as well. Ozone can harm sensitive vegetation and ecosystems such as forests, parks and wildlife refugees, particularly during the growing season. Sulfur dioxide causes leaf necrosis (death of leaf tissue) and chlorosis (bleaching or yellowing of green leaves). Health effects seen in humans with pollutants such as lead can also be observed in animals in nature.

Sources of Pollution

As already mentioned, air pollution is caused by the introduction of substances that have harmful effects on biological systems into the atmosphere. It can be introduced by natural processes such as dust storms, forest fires, volcanic eruptions, or biogenics (emissions from vegetation and soil bacteria). However, there is little that we can do to reduce these types of natural emissions. Instead, we will focus on pollution that is anthropogenic, the result of human activities, as this is something that we can influence.

Outdoor pollutants can come from **mobile** or **stationary** sources (Figure **16.4.2**). Mobile sources of air pollutants move from place to place while emitting pollutants. Examples of mobile sources include vehicles, aircrafts, ships, and trains. Stationary sources have a fixed location. Examples include power plant smokestacks, industrial processes, burning, residential heating, construction sites, farmlands and surface mines among others. Stationary sources can be further broken down into **point** and **nonpoint** sources. Point sources refer to larger emitters at a fixed location, such as an industrial facility. Nonpoint sources tend to be abundant and smaller in magnitude, such as residential heating.



Figure **16.4.2**: Various sources of natural and anthropogenic air pollution. A volcanic eruption is a natural source. Planes and vehicles are mobile sources. A smokestack is a stationary source. (volcano by Sakaori; plane by Air Force file photo/Airman 1st Class Chad Kellum, Public domain; vehicle by Air-pollution.JPG: Zakysant at the German Wikipediaderivative work: F. Lamiot; smokestack by Petr Štefek, all via Wikimedia Commons)

Much of the pollution from human activities is due to the burning of fossil fuels. We utilize fuel combustion to heat homes, generate electricity, and propel a variety of vehicles. **Exhaust gas** or flue gas is emitted as a result of the combustion of fuels such as natural gas, gasoline, petrol, biodiesel blends, diesel fuel, fuel oil, or coal. It is a major component of both motor vehicle (and stationary internal combustion engine) emissions. Motor vehicle emissions contribute to air pollution and are a major ingredient in the creation of smog in some large cities. A 2013 study by MIT indicates that 53,000 early deaths occur per year in the United States alone because of vehicle emissions. In addition to the products of combustion, exhaust gases can also include unreacted gasoline vapors. Many fuels such as coal and wood also directly release particulate matter into the atmosphere.

There are anthropogenic sources of pollution that do not come from combustion reactions. Some of these pollutants are by-products of industrial processes while others are the actual chemicals themselves. Many of these are volatile organic compounds (VOCs) which are carbon containing compounds used as ingredients in many household products. These compounds can vaporize and thus enter into the atmosphere. Others pollutants are compounds such as asbestos or heavy metals like mercury and lead which disperse as particulate matter and can deposit onto soils and surface waters to be taken up by plants and enter into the food chain. Emissions of the most prevalent pollutants are monitored by the government (Figure **16.4.3**).

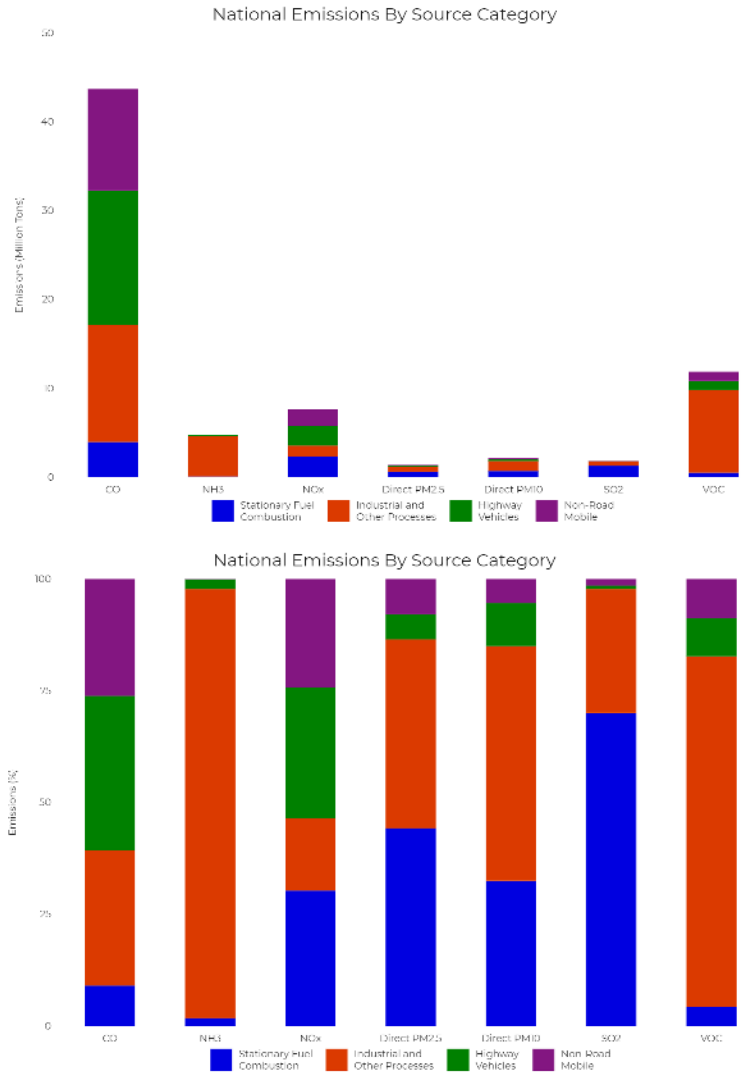


Figure **16.4.3**: National emissions for the US in 2017. The first figure shows the total amount of pollutant emitted for the year (in millions of tons) along with the source. The second figure expands each bar to show the percent of each of four anthropogenic sources contributing to the emission of that pollutant. (Source: EPA air report, https://gispub.epa.gov/air/trendsreport/2022/#sources)

Summary

- Air pollution refers to the introduction, into the atmosphere, of substances that have harmful effects on humans, other living organisms, and the environment either as solid particles, liquid droplets or gases.
- Major sources of air pollution are gas emissions from fossil-fueled vehicles (and other processes which burn fossil fuels) and their reaction products (CO, NO_x, ozone) as well as particulate matter, SO_2, and VOCs.
- Ozone is created by chemical reactions between oxides of nitrogen (NO_x) and volatile organic compounds (VOC).
- Examples of toxic air pollutants include benzene, which is found in gasoline; perchloroethylene, which is emitted from some dry cleaning facilities; methylene chloride, which is used as a solvent and paint stripper by a number of industries; and others such as dioxin, asbestos, toluene, and metals such as cadmium, mercury, chromium, and lead compounds.
- Numerous scientific studies have linked air pollution to a variety of health problems.

Contributors and Attributions

- Libretext: Introduction to Environmental Science (Zendher et al.)
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- US Environmental Protection Agency
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16.5: Photochemical Smog

Learning Objectives

- Describe how photochemical smog is formed.
- Describe how industrial smog is formed.
- Identify harmful effects of smog.

Smog is a mixture of air pollutants (sulfur dioxide, nitrogen oxides, ozone, and particulates) that often form over urban areas as a result of fossil fuel combustion. Thus, smog is a byproduct of modern industrialization. Due to industry and the number of motor vehicles, it is more of a problem in large cities that have a warm, sunny, and dry climate. The term was coined from the terms “smoke” and “fog” referring to a brownish haze that pollutes the air, greatly reducing visibility and making it difficult for some people to breathe (Figure 16.5.1 and 16.5.2). There are two main types of smog: photochemical and industrial smog. **Photochemical smog** is formed when sunlight drives chemical reactions between primary pollutants from automobiles and normal atmospheric compounds. The product is a mix of over 100 different chemicals with the most abundant being ground-level ozone. **Industrial smog** is produced primarily by the burning of fossil fuels which produces carbon dioxide (from complete combustion), carbon monoxides (from partial combustion), sulfur, and mercury. The sulfur reacts with other chemicals in the atmosphere producing several sulfur compounds including sulfur dioxide. These compounds along with particulate material make up industrial smog.



Figure 16.5.1: Smog over Almaty city, Kazakhstan. (Igors Jefimovs, via Wikimedia Commons)



Figure 16.5.2: Smog over Santiago in Chile. (https://commons.wikimedia.org/wiki/F...tiago30std.jpg via Wikimedia Commons)

Formation of Photochemical Smog

Photochemical smog occurs in dry, stagnant air masses, usually stabilized by a temperature inversion (which traps gases close to the surface of the Earth as seen in Figure 16.5.1), that are subjected to intense sunlight. The driving energy force behind smog formation is electromagnetic radiation from the sunlight with a wavelength at around 400 nm or less. This is in the ultraviolet region, just shorter than the lower limit for visible light. Energy absorbed by a molecule from this radiation can result in the formation of active species, thus initiating **photochemical reactions**.

Smog is produced in a series of chain reactions. There are two different series of chemical reactions which can take place in the formation of photochemical smog. The first involves only nitrogen oxides formed by combustion reactions.

Step 1: People begin driving in the morning, nitrogen reacts with oxygen in the extreme temperatures of the combustion engine.

$$N_2 + O_2 \rightarrow 2\,NO \tag{16.5.1}$$

Step 2: After a few hours, NO combines with O_2 in another combustion reaction.

$$2\,NO + O_2 \rightarrow 2\,NO_2 \tag{16.5.2}$$

Step 3: Nitrogen dioxide absorbs light energy ($h\nu$), reforming NO and generating a highly reactive oxygen atom. The NO can then go back to react with more O_2 in step 2 to continue cycling through the reaction or it can be used in later reactions.

$$NO_2 + h\nu \rightarrow NO + O \tag{16.5.3}$$

Step 4: In sunlight, the atomic oxygen combines with oxygen gas to form ozone.

$$O + O_2 \rightarrow O_3 \tag{16.5.4}$$

Step 5: The ozone is consumed by NO to produce nitrogen dioxide and O_2. Again, these products are able to be reused in earlier steps of the reaction to continue the cycle.

$$NO + O_3 \rightarrow NO_2 + O_2 \tag{16.5.5}$$

These reactions result in the accumulation of various nitrogen oxides and some ozone which have the health and environmental effects discussed previously in Section 16.4. However, NO and NO_2 can also enter into a second cycle of chemical reactions by reacting with volatile organic compounds (VOCs, shown as R) and oxygenated organic and inorganic compounds (RO_x) instead of ozone to form a variety of other volatile compounds.

$$NO_2 + R \rightarrow \text{oxygenated products} \tag{16.5.6}$$

$$NO + RO_x \rightarrow NO_2 + \text{other products} \tag{16.5.7}$$

Because the nitrogen oxides react with VOCs in this second series of reactions, ozone is no longer broken down in the first cycle. Instead, ozone is allowed to build up. The accumulation of ozone and volatile organic compounds, along with the energy from the sun, forms the brown, photochemical smog seen on hot, sunny days (Figure **16.5.3**).

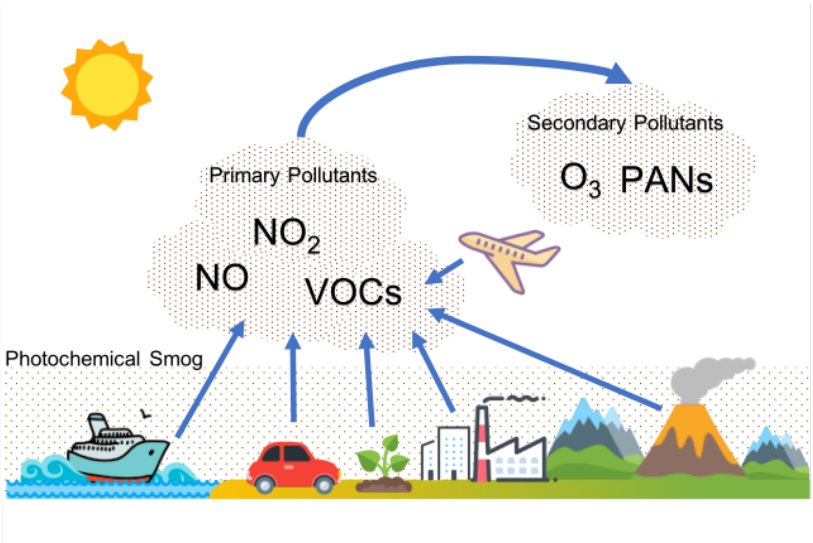
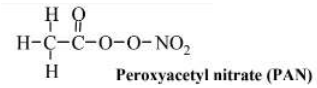


Figure **16.5.3** A photochemical smog formation diagram. (Liweichao.vivian, CC BY-SA 4.0, via Wikimedia Commons)

Because of the wide variety of VOCs present in the atmosphere, numerous noxious products are generated. As already stated, one of the main ones of these is ozone, O_3, and it is the single species most characteristic of smog. Another class of materials formed with smog consists of oxygen-rich organic compounds containing nitrogen of which peroxyacetyl nitrate, PAN,



Peroxyacetyl nitrate (PAN)

is the most common example. This compound, and ones similar to it, is a potent oxidizer and highly irritating to eyes and mucous membranes of the respiratory tract. Also associated with smog are aldehydes, which are irritants to eyes and the respiratory tract. The simplest aldehyde, and one commonly found in smoggy atmospheres, is formaldehyde:



Formaldehyde

Harmful Effects of Smog

Smog adversely affects human health and comfort as well as plants, materials, and atmospheric quality. Ozone is the smog constituent that is generally regarded as being most harmful to humans, plants, and materials, although other oxidants and some of the noxious organic materials, such as aldehydes, are harmful as well. These effects may be especially pronounced for individuals undergoing vigorous exercise because of the large amounts of air that they inhale. On smoggy days, air pollution alerts may advise against exercise and outdoor activities.

Plants are harmed by exposure to nitrogen oxides, ozone, and peroxyacetyl nitrate (PAN, see above), all species present in a smoggy atmosphere. PAN is the most harmful of these constituents, damaging younger plant leaves, especially. Some plant species, including sword-leaf lettuce, black nightshade, quickweed, and double-fortune tomato, are extremely susceptible to damage by chemicals in smog and are used as bioindicators of the presence of smog. Costs of crop and orchard damage by smog run into millions of dollars per year in areas prone to this kind of air pollution, such as southern California.

Materials that are adversely affected by smog are generally those that are attacked by oxidants. The best example of such a material is rubber, especially natural rubber, which is attacked by ozone. Indeed, the hardening and cracking of natural rubber has been used as a test for atmospheric ozone.

Visibility-reducing atmospheric aerosol particles are the most common manifestation of the harm done to atmospheric quality by smog. The smog-forming process occurs by the oxidation of organic materials in the atmosphere, and carbon-containing organic materials are the most common constituents of the aerosol particles in an atmosphere afflicted by smog. Conifer trees (pine and cypress) and citrus trees are major contributors to the organic hydrocarbons that are precursors to organic particle formation in smog.

Industrial Smog

Industrial smog or London-type smog is mainly a product of burning large amounts of high sulfur coal (Figure **16.5.4**). Clean air laws passed in 1956 have greatly reduced smog formation in the United Kingdom; however, in other parts of the world London-type smog is still very prevalent. The main constituent of London-type smog is soot; however, these smogs also contain large quantities of fly ash, sulfur dioxide, sodium chloride and calcium sulfate particles. If concentrations are high enough, sulfur dioxide can react with atmospheric hydroxide to produce sulfuric acid, which will precipitate as acid rain.

$$SO_2 + OH \rightarrow HOSO_2$$

$$HOSO_2 + O_2 \rightarrow HO_2 + SO_3$$

$$SO_3 + H_2O \rightarrow H_2SO_4$$

Concerns about the harmful effects of acid rain have led to strong pressure on industry to minimize the release of SO_2. For example, coal-burning power plants now use SO_2 “scrubbers,” which trap SO_2 by its reaction with lime (CaO) to produce calcium sulfite dihydrate ($CaSO_3 \cdot 2H_2O$) thus preventing its release into the atmosphere.



Figure **16.5.4**: An example of industrial or London smog. (Willem van de Poll, CC0, via Wikimedia Commons)

Summary

- Photochemical smog is a mixture of pollutants that are formed (mostly during the hot summer months) when nitrogen oxides and volatile organic compounds (VOCs) react with sunlight, creating a brown haze above cities.
- Smog is a serious problem in many cities and continues to harm human health and are especially harmful for senior citizens, children, and people with heart and lung conditions such as emphysema, bronchitis, and asthma.
- Industrial smog is a mixture of soot and other particulate matter, as well as sulfur dioxide gas.
- If concentrations are high enough, the sulfur dioxide in industrial smog can form sulfuric acid and acid rain.

Contributors and Attributions

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16.6: Clean Air

Learning Objectives

- Identify ways to limit particulate emissions to the atmosphere.
- Describe how various command and control measures have affected air pollution levels.
- Describe how the Clean Air Acts and the resulting regulations have affected air pollution levels and productivity.

As we have seen in the preceding sections, there are many different chemicals which are released into the atmosphere, both due to natural processes and to human activity. These chemicals can undergo reactions which create various forms of air pollution resulting in decreased visibility, adverse health affects, and damage to buildings and infrastructure. No one enjoys the effects of air pollution and so efforts must be made to deal with it. One rather obvious way of eliminating the chemicals that result from human activities is to just stop most of the activities that generate them. It was demonstrated that it is possible to curb some types of air pollution at the 2008 and 2022 Beijing Olympics and during the lockdown that occurred during the pandemic in 2020 by simply stopping the human activities that cause them. However, we do need the energy that is created by these activities in order to function efficiently. So this is not a practical, long term solution. What other options are there?

Limiting Particulate Emissions

The first widespread measures to limit air pollution were directed at control of particle emissions. Because of their ability to reduce visibility and light, atmospheric particles are the most visible form of air pollution. Commonly called **particulates** in air pollution terminology, atmospheric *aerosols* are solids or liquids less than 100 micrometers in diameter, but are most often found in a size range of 0.001 to 20 μm. One type is **dispersion aerosols** formed by grinding solids, dispersing dusts, or atomizing liquids. Common dispersion aerosols include water droplets from sea spray, solid particles of NaCl left over when water evaporates from sea spray droplets, cement dust, soil dust dispersed by wind, foundry dust, and pulverized coal. A second type of atmospheric aerosol is **condensation aerosols** produced by the condensation of gases or vapors, often formed as the result of atmospheric chemical processes. Carbon black, metal fumes, and combustion nuclei form as condensation aerosols from combustion or partial combustion reactions. Liquid particle **mists** include raindrops, fog, cloud droplets, and droplets of sulfuric acid produced when atmospheric SO_2 is oxidized. Organisms also produce an abundance of particles. For those afflicted with allergies, the most annoying such particles are plant pollen. Other particles of biological origin include viruses, bacteria, and spores of bacteria and fungi.

In the past and even now in some areas of the world, one of the more troublesome sources of atmospheric particles was **fly ash**, a byproduct residue from combustion of liquids or very finely divided coal. Often the most abundant component of fly ash is elemental carbon left over from incompletely burned fuel. Fly ash commonly includes oxides of aluminum, calcium, iron, and silicon, as well as some magnesium, sulfur, titanium, phosphorus, potassium, and sodium. With properly operating emission control devices, fly ash emissions are now well controlled. These measures have become very effective so that the “smoke” that one sees emanating from smokestacks usually consists of droplets of water formed by condensation of steam.

The simplest method of particle control from stack gas and other gases released to the atmosphere consists of **sedimentation** in which particles contained in stack gas are allowed to settle by gravity in relatively large chambers. Sedimentation is most effective for larger particles. **Inertial mechanisms** operate by spinning a gas in a round chamber such that particles impinge upon the container walls by centrifugal force. Fabric filters contained in **baghouses** act to filter particles from air or stack gas (Figure **16.6.1**). The mechanism employed provides for periodic shaking of the fabric filters to collect particles held on their walls, thus restoring gas flow through the fabric. Numerous factors including moisture levels, particle abrasion, particle size, and acidity or alkalinity of the gas and particles must be considered in choosing filter fabrics. **Scrubbers** that spray water or solutions into stack gas are employed to literally wash particles out of gas. In some cases these are operated with a minimal amount of water, which evaporates, so that a solid material is collected. One of the most effective means of particle control consists of **electrostatic precipitators**. These devices use a very high voltage to impart a negative charge onto particles, and the particles are attracted to and collect on the positively charged walls of the precipitator.

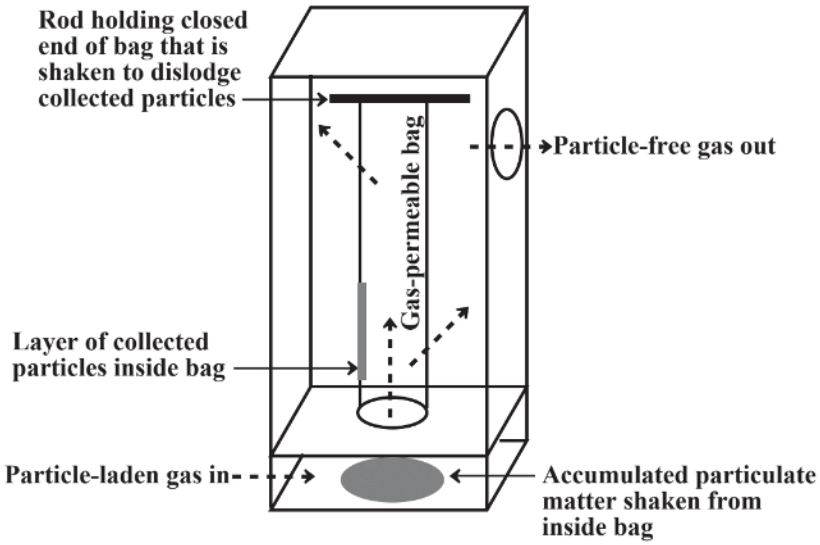


Figure **16.6.1**. Illustration of a unit of a single bag in a baghouse for removal of particulate matter from a gas stream. The flow of gas is illustrated by the dashed lines. A bag is mounted on an opening in a plate such that particle-laden gas is forced up through the inside of the cylindrical bag and is filtered through the bag walls, then flows out as filtered gas. The particulate matter accumulates on the inside of the bag. Periodically, the gas flow through the bag is stopped and the rod suspending the closed upper end of the bag is shaken to dislodge particulate matter from the bag walls, which is collected at the bottom of the apparatus.

In keeping with the practice of sustainability, the particulate matter such as that collected by a fabric filter in a baghouse may be used for various purposes. Typically, particulate matter from lead or zinc smelting operations is recycled back into the metal recovery process. Lime kiln dust is often used as agricultural lime. Some kinds of coal fly ash could be used as a source of aluminum if aluminum ore (bauxite) becomes scarce. Sulfur is a valuable raw material required in the manufacture of sulfuric acid, one of the largest volume chemicals made. Hydrogen sulfide, H_2S, occurs in large quantities in natural gas, such as that produced in the Canadian province of Alberta. This hydrogen sulfide must be removed from the natural gas. Rather than presenting a pollution problem, it is converted to elemental sulfur, then used to make sulfuric acid. Another green chemistry approach to the reclamation of waste sulfur is practiced in Kalundborg, Denmark. The huge coal-fired power plant in Kalundborg uses lime scrubbing to remove sulfur dioxide from stack gas. The calcium sulfite product of this process is oxidized

$$CaSO_3 + \frac{1}{2}\,O_2 + 2\,H_2O \rightarrow CaSO_4 \cdot 2H_2O$$

to generate gypsum, $CaSO_4 \cdot 2H_2O$. This mineral is then used to make wallboard, thus solving a pollution problem from the production of spent lime and a raw materials problem arising from the need for gypsum to make wallboard needed for building construction.

Controlling Photochemical Smog

As described in the previous section, smog is basically a chemical problem which would indicate that it should be amenable to chemical solutions. Indeed, the practice of green chemistry and the application of the principles of industrial ecology have helped to reduce smog. This is due in large part to the fact that a basic premise of green chemistry is to avoid the generation and release of chemical species with the potential to harm the environment. The best way to prevent smog formation is to avoid the release of nitrogen oxides and organic vapors that enable smog to form. At an even more fundamental level, measures can be taken to avoid the use of technologies likely to release such substances, for example, by using alternatives to polluting automobiles for transportation.

Initially, serious efforts to control pollution were based on a **command and control** approach, which specifies maximum concentration guideline levels of substances that can be allowed in the atmosphere (or water) and places limits on the amounts or concentrations of pollutants that can be discharged in waste streams. Command and control efforts to diminish pollution have resulted in implementation of various technologies to remove or neutralize pollutants in potential waste streams and stack gases. These are so-called "end-of-pipe" measures. As a result, numerous techniques, such as chemical precipitation of water pollutants, neutralization of acidic pollutants, stack gas scrubbing, and waste immobilization have been developed and refined to deal with pollutants after they are produced.

The evolution of automotive pollution control devices to reduce smog provides an example of how green chemistry can be used to reduce pollution. The first measures taken to reduce hydrocarbon and nitrogen oxide emissions from automobiles were very much command-and-control measures. These primitive measures implemented in the early 1970's did reduce emissions, but with a steep penalty in fuel consumption and in driving performance of vehicles. However, over the last three decades, the internal combustion automobile engine has evolved into a highly sophisticated, computer-controlled machine that generally performs well, emits few air pollutants, and is highly efficient. This change has required an integrated approach involving reformulation of gasoline. The first major change was elimination from gasoline of tetraethyllead, an organometallic compound that poisoned automotive exhaust catalysts (and certainly was not good for people). Gasoline was also reformulated to eliminate excessively volatile hydrocarbons and unsaturated hydrocarbons (those with double bonds between carbon atoms) that are especially reactive in forming photochemical smog. Many scientists have suggested that pumping gas at night could further reduce photochemical ozone formation by limiting the amount of exposure VOCs have with sunlight.

In addition to the reformulation of gasoline, every new vehicle sold in the United States must include a catalytic converter to reduce photochemical emissions. Catalytic converters force CO and incompletely combusted hydrocarbons to react with a metal catalyst, typically platinum, to produce CO_2 and H_2O. Additionally, catalytic converters reduce nitrogen oxides from exhaust gases into O_2 and N_2, eliminating the cycle of ozone formation.

An even more drastic approach to eliminating smog-forming emissions is the use of electric automobiles that do not burn gasoline. These vehicles certainly do not pollute as they are being driven, but they suffer from the problem of a very limited range between charges and the need for relatively heavy batteries that utilize rare metals. However, hybrid automobiles using a small gasoline or diesel engine that provides electricity to drive electric motors propelling the automobile and to recharge relatively smaller batteries can largely remedy the emission and fuel economy problems with automobiles. The internal combustion engine on these vehicles runs only as it is needed to provide power and, in so doing, can run at a relatively uniform speed that provides maximum economy with minimum emissions. Additionally, scientists are looking at new ways of producing batteries that are lighter, last longer, and use components that are more environmentally friendly to obtain.

Another approach that is being used on vehicles as large as buses, that have convenient and frequent access to refueling stations, is the use of fuel cells that can generate electricity directly from the catalytic combination of elemental hydrogen and oxygen, producing only harmless water as a product. There are also catalytic processes that can generate hydrogen from liquid fuels, such as methanol, so that vehicles carrying such a fuel can be powered by electricity generated in fuel cells.

Clean Air Acts

Much of the driving force for reducing particulate emissions and the components of photochemical smog has been due to the Clean Air Act of 1970 and its amendments in 1977 and 1990. These acts have set emission limits for different industries and motor vehicles in the United States and allow for their enforcement. Although much maligned, various pollution control measures implemented in response to these acts have reduced wastes and improved environmental quality (Figure **16.6.2**). Regulation-based pollution control has clearly been a success and well worth the expense and effort. However, it is much better to prevent the production of pollutants rather than having to deal with them after they are made. This was recognized in United States with the passage of the 1990 Pollution Prevention Act, which recognized that source reduction is fundamentally different and more desirable than waste management and pollution control. This act explicitly states that, wherever possible, wastes are not to be generated and their quantities are to be minimized. The means for accomplishing this objective can range from very simple measures, such as careful inventory control and reduction of solvent losses due to evaporation, to much more sophisticated and drastic approaches including complete redesign of manufacturing processes with waste minimization as a top priority.

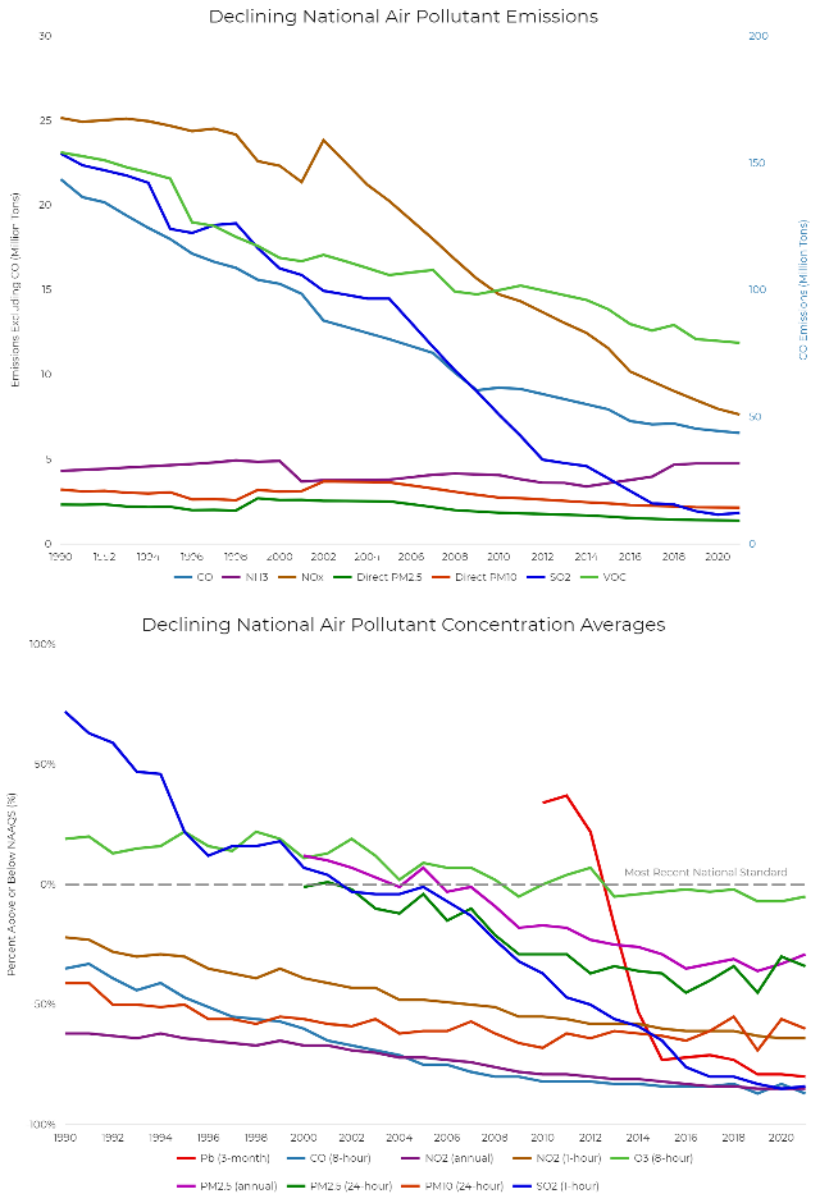


Figure **16.6.2**: Declining air pollutant emissions and the resultant air pollutant concentrations since the passage of the 1990 Amendments to the Clean Air Act of 1970 and the 1990 Pollution Prevention Act. NAAQS are National Ambient Air Quality Standards. (US Environmental Protection Agency, www.epa.gov)

The net result of these legislations has been a 78 percent drop in the six common pollutants between 1970 and 2020. The emission reductions have led to dramatic improvements in the quality of the air that we breathe and have enabled many areas of the country to meet national air quality standards set to protect public health and the environment. This means that we breathe less pollution and face lower risks of premature death and other serious health effects. In addition, environmental damage has been reduced resulting in the improved health of plants, soils, and aquatic life and increased crop and timber yields. New technologies have been developed which have helped to control costs in addition to reducing emissions. The increased health and new technologies have resulted in increasing gross domestic product while the pollutants have decreased (Figure **16.6.3**). Studies have estimated that the benefits of reducing pollution exceed the costs by a factor of more than 30-to-1.

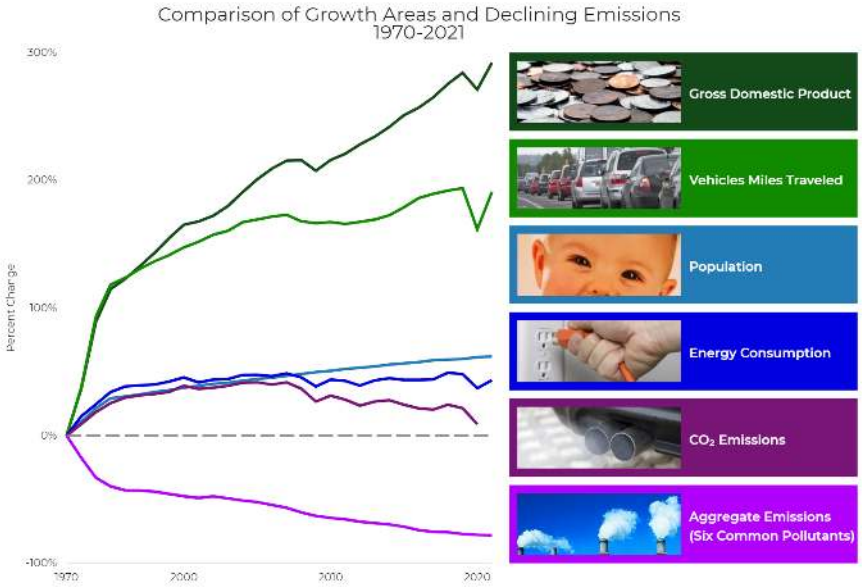


Figure **16.6.3**: Some of the benefits that have increased as a result of pollution reduction since the implementation of Clean Air Acts. (US Environmental Protection Agency, www.epa.gov)

Summary

- The first efforts to limit air pollution focused on controlling particle emissions. There are a number of different methodologies which have proven effective at preventing particulates from being released to the atmosphere.
- Particulate matter that is removed from gas emissions may be used in other areas, providing a secondary benefit to the effort required to clean the gases.
- A major contributor to the lowering of emissions is the advancements made in the internal combustion engine in cars and gasoline reformulation. These have resulted in cleaner burning vehicles with better gas mileage. The development of hybrid and electric vehicles will help to lower emissions even more.
- The Clean Air Acts have required many industries to take a close look at how they produce and dispose of potential pollutants and minimize their generation and release. This has resulted in a 78% drop in the six common pollutants from 1970 to 2020 while overall productivity has increased over the same time period.

Contributors and Attributions

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16.7: Global Warming

Learning Objectives

- Define global warming, climate change, the greenhouse effect, and greenhouse gases.
- Identify the gases present in the troposphere and stratosphere and their effects on life on Earth.
- Identify sources of greenhouse gases in the atmosphere and their levels over time.
- Discuss how the temperature of the Earth has varied over time and since 1900.

We hear a lot in the news about climate change and global warming. However, the terminology can sometimes be a bit confusing. How often have you heard someone ask why we are experiencing something like a polar vortex if global warming is happening? In order to make sense of what is happening to our planet, it is important to understand the two terms and the factors that influence them. It is only then that one can make informed decisions on the severity of the issue and how best to mitigate the effects. **Global warming** refers to the increase in the average temperature of the Earth's atmosphere due to elevated greenhouse gas concentrations, heightening the greenhouse effect. **Climate change** includes both global warming, driven by human-induced emissions of greenhouse gases, and the resulting large-scale shifts in weather patterns. Though there have been previous periods of climatic change, since the mid-20th century humans have had an unprecedented impact on Earth's climate system and caused change on a global scale.

The Atmosphere and the Greenhouse Effect

The atmosphere is the air around and above us. We know we must have air to breathe. A human deprived of air's life-giving oxygen for just a brief time will lose consciousness, and within a few minutes will die. But air is far more than just a source of oxygen. That is because it protects Earth's organisms in ways that are absolutely essential for their existence. One major protective function is to act as a blanket to keep us warm. It does that by reabsorbing some of the infrared radiation which Earth receives from the sun and radiates back to space. By delaying the exit of this energy into outer space, the average temperature of Earth's surface remains at about 15° C (59° F) at sea level, though it can be much colder at certain times and places and significantly warmer at others. Without this warming effect, plants could not grow and most other known organisms could not exist. The second protective function of the atmosphere is absorption of very short wavelength ultraviolet solar radiation. Were this radiation to reach our level, it would tear apart biomolecules, making it impossible for most life forms to exist.

Although one might get the impression that the atmosphere is very thick, it is "tissue thin" compared to Earth's diameter. The altitude at which high-flying jet aircraft cruise marks the upper limit of the lowest of several layers of the atmosphere, the **troposphere**, which extends from sea level to about 11 km (Figure **16.7.1**). As anyone who has driven to the top of Pike's Peak or some other mountain knows, the troposphere gets cooler with increasing altitude, from an average temperature of 15° C at sea level to an average of -56° C (-69° F) at 11 km. Above the layer of the troposphere, however, atmospheric temperature *increases* to an average of -2° C (35.6° F) at 50 km altitude. The layer above the troposphere is the **stratosphere**, which is heated by the absorption of intense ultraviolet radiation from the sun (Figure **16.7.2**). There is virtually no water vapor in the stratosphere, and it contains ozone (O_3) and O atoms as the result of ultraviolet radiation acting upon stratospheric O_2. Beyond the stratosphere are layers called the mesosphere and thermosphere, but they are relatively less important in our discussion of the atmosphere.

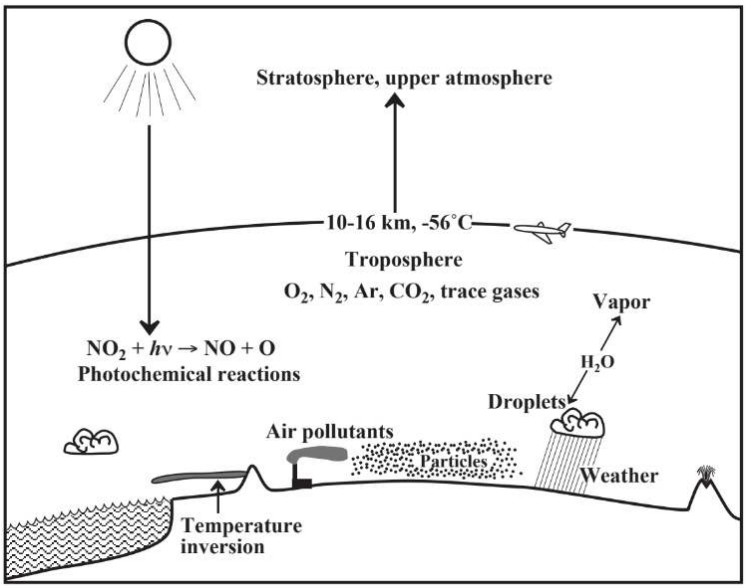


Figure **16.7.1**. The troposphere is the very thin layer of the atmosphere closest to Earth, containing most of the atmosphere's air and water vapor. It is the source of oxygen, carbon dioxide, nitrogen, and water used by living organisms and as raw materials for manufacturing. With the important exception of stratospheric ozone destruction, it is where most air pollution phenomena occur.

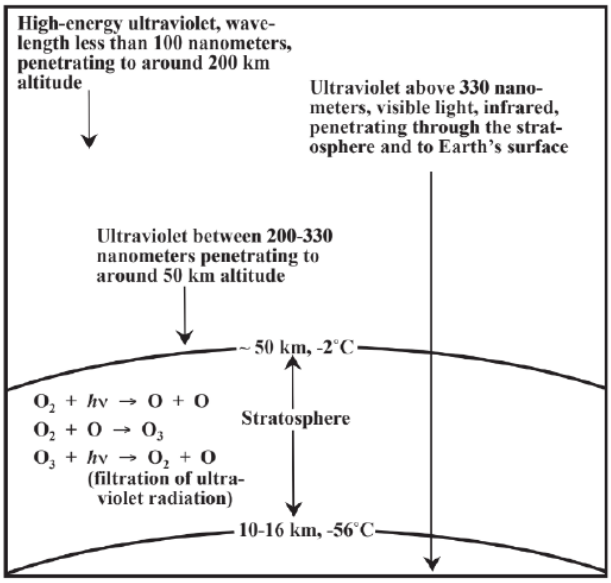


Figure **16.7.2**. The upper atmosphere, including the stratosphere and regions beyond, is particularly important in the absorption of radiation that would make life impossible if it reached Earth's surface. The layer of ozone, O_3, in the stratosphere is of the utmost importance and one that is subject to damage from anthropogenic species released into the atmosphere.

There is an enormous amount of energy reaching Earth's atmosphere from the sun as sunlight. As shown in Figure **16.7.3**, some of the incoming energy reaches Earth's surface, some is absorbed in the atmosphere, warming it, and some is scattered back to space. The energy that reaches the Earth's surface, primarily as light at a maximum intensity of 500 nanometers in the visible region, must be released back into the atmosphere, which it does as infrared radiation (with maximum intensity at about 10 micrometers (μm), primarily between 2 μm and 40 μm). Water molecules, carbon dioxide, methane, and other minor species in the atmosphere absorb some of the outbound infrared radiation, which eventually is all radiated to space. This temporary absorption of infrared radiation warms the atmosphere — a **greenhouse effect**. The term greenhouse effect is a slight misnomer in the sense that physical greenhouses warm via a different mechanism. The greenhouse effect as an atmospheric mechanism functions through radiative heat loss while a traditional greenhouse as a built structure blocks convective heat loss. The result, however, is an increase in temperature in both cases.

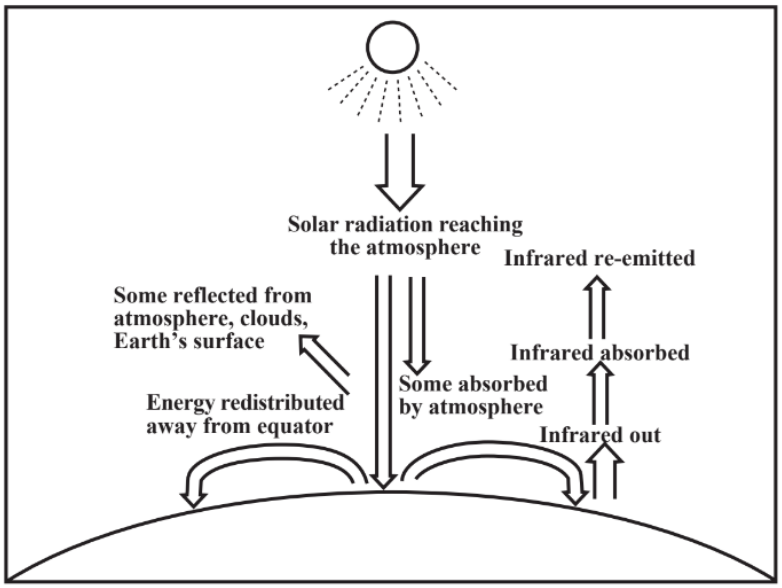


Figure **16.7.3**. Some of the solar energy reaching the top of the atmosphere penetrates to Earth's surface, some is absorbed in the atmosphere, which warms it, and some is scattered by the atmosphere and from Earth's surface. Outgoing energy is in the infrared region. Some of it is temporarily absorbed by the atmosphere before being radiated to space, causing a warming (greenhouse effect). The equatorial regions receive the most energy, part of which is redistributed by warm air masses and latent heat in water vapor away from the equator.

The fraction of electromagnetic radiation from the sun that is reflected by Earth's surface varies with the nature of the surface. The percentage reflected is very important because it determines how effective incoming radiation is in warming the surface and it is expressed as **albedo**. Freshly plowed black topsoil has a very low albedo of only about 2.5%. This means that only 2.5% of the radiation is reflected and the rest is absorbed. In contrast, the albedo of a covering of fresh snow is about 90%. The anthrosphere affects albedo. One example of this is in cultivating land, turning over relatively high albedo grass and covering it with exposed black soil that absorbs light energy very strongly. Another is covering of large areas of Earth's surface with asphalt paving which reflects sunlight poorly.

Earth's natural greenhouse effect is critical to supporting life and initially was a precursor to life moving out of the ocean onto land. Human activities, mainly the burning of fossil fuels and clear cutting of forests, have increased the greenhouse effect and caused global warming.

Greenhouse Gases and Global Warming

The maintenance of Earth's heat balance to keep temperatures within limits conducive to life is very complex and not well understood. Geological records show that in times past, Earth was sometimes relatively warm and that at other times there were ice ages in which much of Earth's surface was covered by ice a kilometer or two thick. The differences in average Earth temperature between these extremes and the relatively temperate climate conditions that we now enjoy were only a matter of a few degrees. It is also known that massive volcanic eruptions and almost certainly hits by large asteroids have caused cooling of the atmosphere that lasted for a year or more. There is now concern that anthropogenic gas emissions, particularly of carbon dioxide from fossil fuel combustion, may be having a warming effect upon the atmosphere.

A **greenhouse gas** (sometimes abbreviated **GHG**) is a gas that absorbs and emits radiant energy within the thermal infrared range, causing the greenhouse effect. It's important to realize that water vapor (H_2O) is also a greenhouse gas. While humans have little direct impact on water vapor concentrations in the atmosphere, it is still an essential component of the natural greenhouse effect that occurs in our atmosphere. The four major categories of greenhouse gases that have been impacted by humans the most are carbon dioxide (CO_2), methane (CH_4), nitrous oxide (N_2O), and synthetic fluorinated gases including hydrofluorocarbons (HFCs), perfluorocarbons (PFCs), and sulfur hexafluoride (SF_6). These will be discussed below and a numeric comparison is provided in Table **16.7.1**.

Carbon dioxide (CO_2) is the greenhouse gas responsible for most of the human-caused climate change in our atmosphere. It has the highest concentration in the atmosphere of any of the greenhouse gases that we'll discuss here. Remember that CO_2 is a direct product of both combustion and cellular respiration, causing it to be produced in great quantities both naturally and anthropogenically. Any time biomass or fossil fuels are burned, CO_2 is released. Major anthropogenic sources include: electricity production from coal-fired and natural gas power plants, transportation, and industry. To get an idea of how CO_2 concentration has changed over time, watch Video **16.7.1**. This video contains atmospheric CO_2 concentrations measured directly, dating back to 1958, as well as atmospheric CO_2 concentrations measured indirectly from ice core data, dating back to 800,000 BCE. Ice core data shows us that the atmospheric CO_2 concentration never exceeded 300 ppm before the industrial revolution. By 1990, a quantity of over seven billion tons of carbon (equivalent to 26 billion tons of carbon dioxide when the weight of the oxygen atoms are also considered) was being emitted into the atmosphere every year, much of it from industrialized nations. As of early 2022, the current atmospheric CO_2 concentration is over 400 ppm. Similar to the action of the naturally existing greenhouse gases, any additional greenhouse gases lead to an increase in the surface temperature of the Earth.



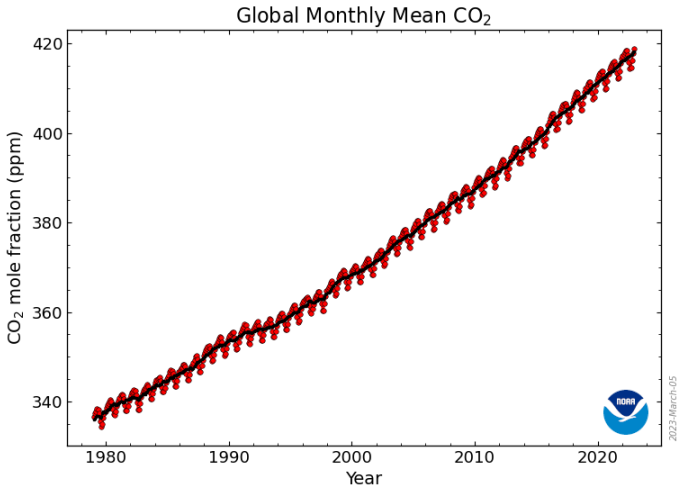
Video **16.7.1**: History of atmospheric CO_2, from 800,00 years ago to January 2022. (National Oceanic and Atmospheric Administration)

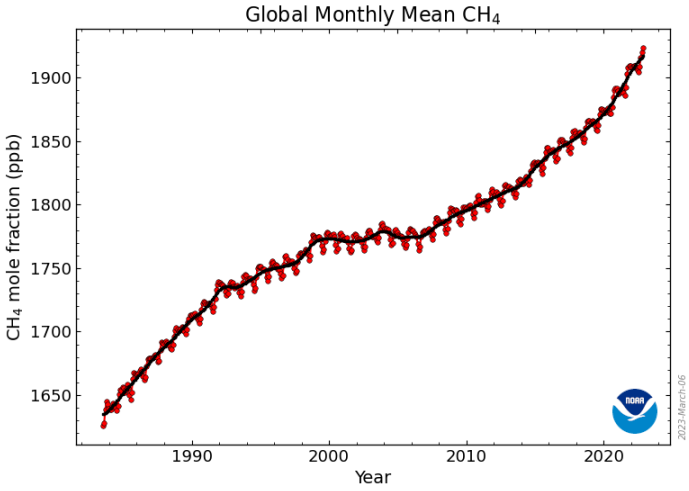
While CO_2 is produced by aerobic cellular respiration, gases such as **methane** (CH_4) and **nitrous oxide** (N_2O) are often the products of anaerobic metabolisms. Agriculture is a major contributor to CH_4 emissions. In addition to anaerobic bacteria, methane is also a significant component of natural gas, and is commonly emitted through the mining and use of natural gas and petroleum, in addition to coal mining. Finally, landfills contribute significantly to CH_4 emissions, as the waste put into the landfill largely undergoes anaerobic decomposition as it is buried under many layers of trash and soil. Natural sources of CH_4 include swamps, wetlands, and volcanoes.

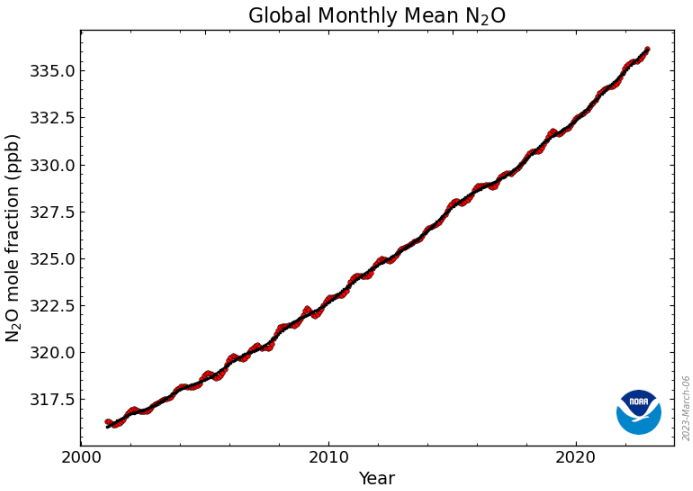
The vast majority of N_2O production by humans comes from agricultural land management. While some N_2O is naturally emitted to the atmosphere from soil as part of the nitrogen cycle, human changes in land management, largely due to agricultural practices, have greatly increased N_2O emissions. Some N_2O is also emitted from transportation and industry.

One class of greenhouse gas chemicals that has no natural sources is the **fluorinated gases**. These include HFCs, PFCs, and SF_6, among others. Because these are synthetic chemicals that are only created by humans, these gases were essentially non-existent before the industrial revolution. These synthetic gases are used for a wide variety of applications, from refrigerants to semiconductor manufacturing, and propellants to fire retardants. They tend to have a long lifetime in the atmosphere, as seen in Table **16.7.1**. Some of these chemicals, as well as the older **chlorofluorocarbons** (CFCs), have been phased out by international environmental legislation under the Montreal Protocol. Due to their long lifespan, many of these now banned CFCs remain in the atmosphere. Newer chemical replacements, such as HFCs, provide many of the same industrial applications, but unfortunately have their own environmental consequences.

Due to their relatively high concentrations in the atmosphere compared to synthetic gases, CO_2, CH_4, and N_2O, are responsible for most of the human-caused global climate change over the past century. Figure **16.7.4** shows the increases in all four categories of gases for the past 20 years or longer.


Global Monthly Mean CO_2
CO_2 mole fraction (ppm)
420
400
380
360
340
1980
1990
2000
2010
2020
Year
2023-March-05


Global Monthly Mean CH_4
CH_4 mole fraction (ppb)
1900
1850
1800
1750
1700
1650
1990
2000
2010
2020
Year
2023-March-06


Global Monthly Mean N_2O
N_2O mole fraction (ppb)
335.0
332.5
330.0
327.5
325.0
322.5
320.0
317.5
2000
2010
2020
Year
2023-March-06

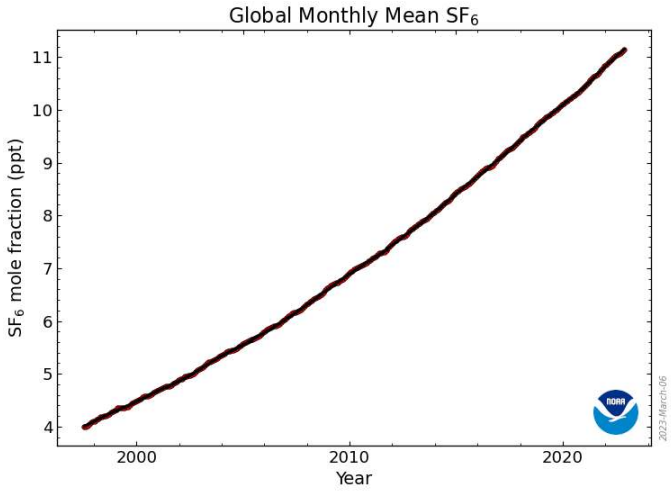


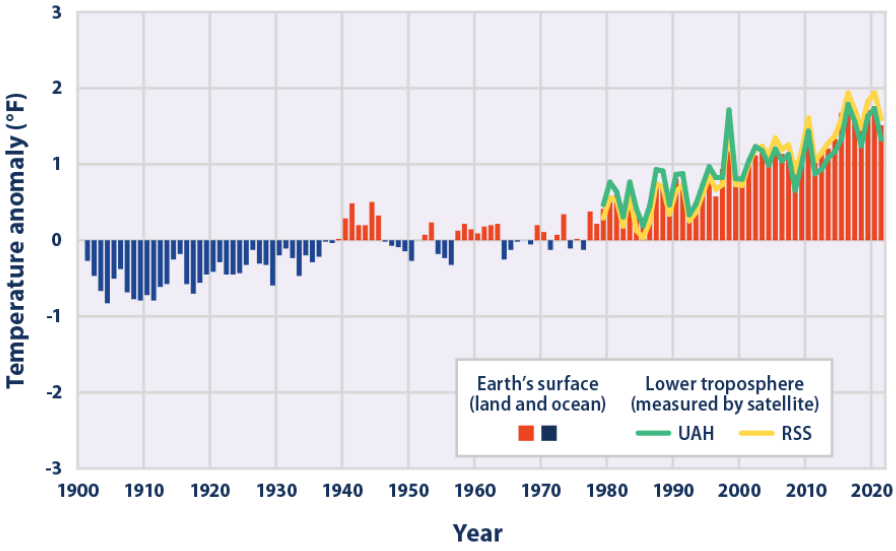
Figure 16.7.4: Global monthly averages for greenhouse gas concentrations expressed as ppm (parts per million), ppb (parts per billion), or ppt (parts per trillion). Samples were taken from a globally distributed network of air sampling sites. Red circles show the monthly average centered on the middle of each month. Black lines show the long term trend with seasonal cycles removed. (National Oceanic and Atmospheric Administration)

Just as greenhouse gases differ in their sources and their residence time in the atmosphere, they also differ in their ability to produce the greenhouse effect. This is measured by the **global warming potential**, or GWP, of each greenhouse gas. The GWP of a greenhouse gas is based on its ability to absorb and scatter energy, as well as its lifetime in the atmosphere. Since CO_2 is the most prevalent greenhouse gas, all other greenhouse gases are measured relative to it. As the reference point, CO_2 always has a GWP of 1. Note the very high GWP values of the synthetic fluorinated gases in Table 16.7.1. This is largely due to their very long residence time in the atmosphere. Also note the higher GWP values for CH_4 and N_2O compared to CO_2.

Table 16.7.1. Comparison of Greenhouse Gases.

Greenhouse gas	Chemical formula or abbreviation	Lifetime in atmosphere	Global warming potential (100-year)
Carbon dioxide	CO_2	Variable	1
Methane	CH_4	12 years	28-36
Nitrous oxide	N_2O	114 years	298
Hydrofluorocarbons	Abbreviation: HFCs	1-270 years	12-14,800
Perfluorocarbons	Abbreviation: PFCs	2,600-50,000 years	7,390
Sulfur hexafluoride	SF_6	3,200 years	22,800

The concern with increasing concentrations of greenhouse gases in the atmosphere is that it will lead to — indeed *is* leading to — an excess of a good thing, warming of the global atmosphere. Is Earth warming? Sophisticated computer models indicate that it is, backed by evidence from increasingly accurate temperature records over more than 100 years (Figure 16.7.5). These temperature records have been especially accurate over the last several decades because they have been read over all Earth's surface by satellite. Worldwide, 2016 was the warmest year on record and 2020 was the second-warmest. The decade from 2012 to 2021 was the warmest decade on record since measurements with thermometers began. While average surface temperature both globally and within the 48 contiguous states has risen at an average rate of 0.17° F per decade since 1901, the United States has warmed at a faster rate than the global rate since the late 1970s.



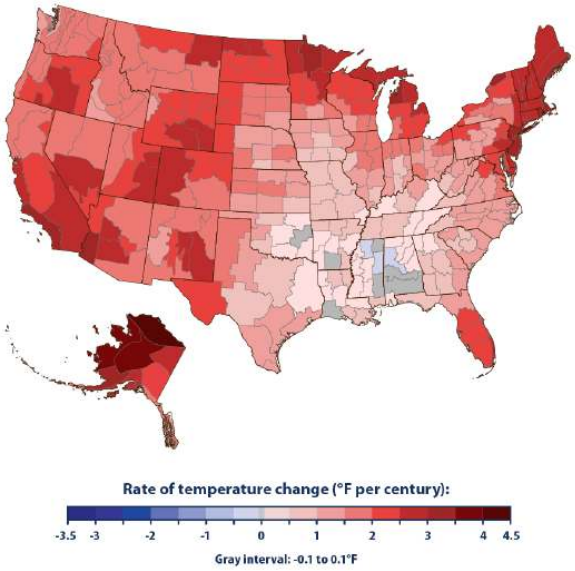


Figure **16.7.5**: The first chart shows worldwide temperatures from 1901 to 2021 using the 1901-2000 average as a baseline for depicting change. Choosing a different baseline period would not change the shape of the data over time. "UAH" and "RSS" represent two different methods of analyzing the original satellite measurements. The second figure shows how average air temperatures have changed in different parts of the United States since 1901 for the contiguous 48 states and since 1925 for Alaska. (NOAA via epa.gov)

Summary

- The atmosphere is made up of different layers of gases, each of which has a unique function in making life on Earth possible. The stratosphere contains ozone which prevents damaging UV radiation from reaching the surface. The troposphere contains a variety of gases which keep the Earth at a moderate temperature to sustain life.
- Gases that trap heat in the atmosphere are called greenhouse gases.
- The greenhouse effect is the process by which radiation from a planet's atmosphere warms the planet's surface to a temperature above what it would be without this atmosphere.
- Global warming refers to the increase in the average temperature of the Earth's atmosphere due to elevated greenhouse gas concentrations, heightening the greenhouse effect.

- Greenhouse gases differ in their ability to produce the greenhouse effect as measured by the global warming potential, or GWP. A molecule of CH_4 is about 28 times more effective than one of CO_2 at absorbing infrared radiation, while N_2O is 298 times more effective . The synthetic fluorinated gases are even much higher (in the thousands). This is largely due to their very long residence time in the atmosphere.

Contributors and Attributions

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- Libretext: Introduction to Environmental Science (Zendher et al.)
- US EPA
- National Oceanic and Atmospheric Administration
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16.8: Climate Change – An Effect of Global Warming

Learning Objectives

- Describe how water is able to redistribute energy around the Earth.
- Describe how changes in temperature in one area of the Earth can cause changes in precipitation in others.
- Describe how changes in temperature can cause sea levels to rise and the effects of that rise.
- Describe how elevated levels of CO_2 in the atmosphere contribute to the acidification of the oceans.

The most important aspect of the atmosphere for humans and other living beings is **climate**. Climate consists of long-term trends in weather and varies a lot over Earth's surface. For example, the climate in desert regions of the world may be relatively hot and dry, but the weather in such regions may at times produce torrential rainfall or frigid temperatures. Of all the things that humans might do to irreversibly and catastrophically harm Mother Earth, the only home that they have or ever can have, the potential to change climate is the most serious. It is well established that climate has changed markedly in past times, most notably during several ice ages that each lasted for thousands of years. Volcanic eruptions noted in historical accounts have caused temporary cooling of the atmosphere and widespread hunger when summer growing seasons were cut short as a result. Tree-ring data indicate centuries-long droughts in parts of the world in the past, which resulted in the decline of some civilizations in the present-day U.S. Southwest, for example. Although these were natural events, recent weather data indicate changes in microclimate due to human activities. For example, some urban areas in Southeast Asia have become darker in recent decades due to particulate air pollution.

Weather refers to relatively short term variations in the state of the atmosphere as expressed by temperature, cloud cover, precipitation, relative humidity, atmospheric pressure and wind. Weather is driven by redistribution of energy in the atmosphere. A particularly important aspect of this redistribution is the energy released when precipitation forms. Earth receives solar energy most directly at the equator, so equatorial regions are warmer than regions farther north and south. A significant fraction of this energy moves away from the equator. This is largely done by **convection** in which heat is carried by masses of air. However, heat can also be carried as **latent heat** in the form of water vapor. The **heat of vaporization** of water is 2,259 joules per gram (J/g) meaning that 2,259 joules of heat energy are required to evaporate a gram of water without raising its temperature. This is a very high value, meaning that the evaporation of ocean water by solar energy falling on it in warmer regions absorbs an enormous amount of heat to form water vapor. This vapor may be carried elsewhere and condense to form rainfall. The heat energy released as the vapor condenses raises the temperature of the surrounding atmosphere.

As an example, heat energy from sunlight and from hot masses of air is converted to latent heat by the evaporation of ocean water off the west coast of Africa. Prevailing winds drive masses of air laden with water vapor westward across the ocean. Rainfall forms, releasing the energy from the latent heat of water and warming the air mass. The hot mass of air that results rises, creating a region of low pressure into which air flows in a circular manner. This can result in the formation of a whirling mass of air in the form of a hurricane that may strike Puerto Rico, Cuba, Florida, or other areas thousands of miles from the area where the water was originally evaporated from the ocean.

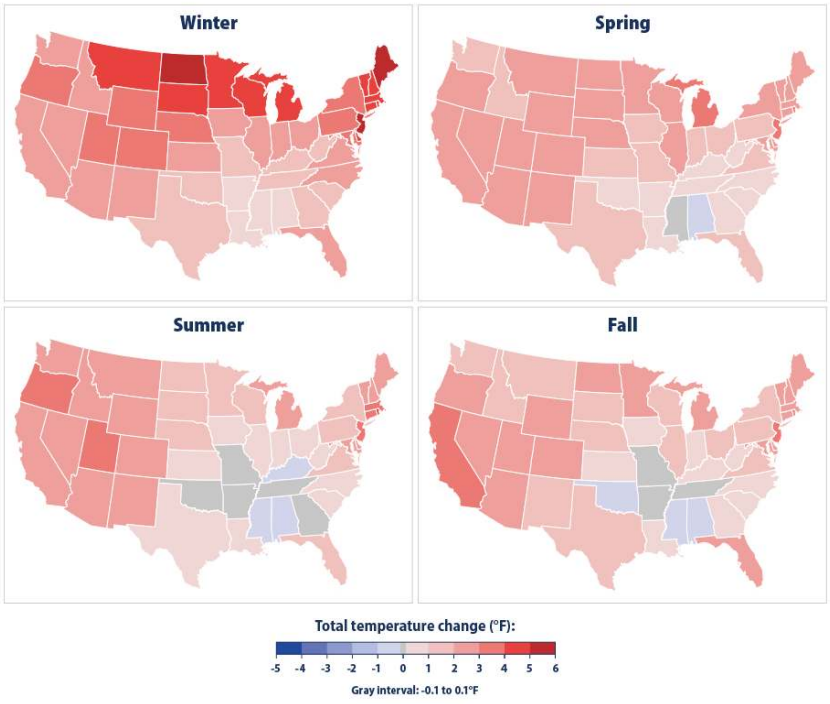
When the long-term trends in weather begin to significantly shift, the climate is said to change. Of all the things that humans may be doing that could change climate, emissions of large quantities of gases that cause atmospheric warming are the most serious. Although there could be some benefits of mild global warming, the net effect would almost certainly be bad, perhaps catastrophic. Climate models (which have been impressively accurate since they began in the 1970s) predict an average global temperature increase of 1.5–5° C by 2100 (Video**16.8.1**). That does not sound like much, but it is about as much as the temperature increase that occurred from the last ice age until now. As the global temperature warms, the types of energy redistribution that create hurricanes and other weather phenomena are predicted to happen more frequently. Especially if the warming is toward the high side of the projected range, this would greatly affect climate and rainfall. Some of these effects are described below.

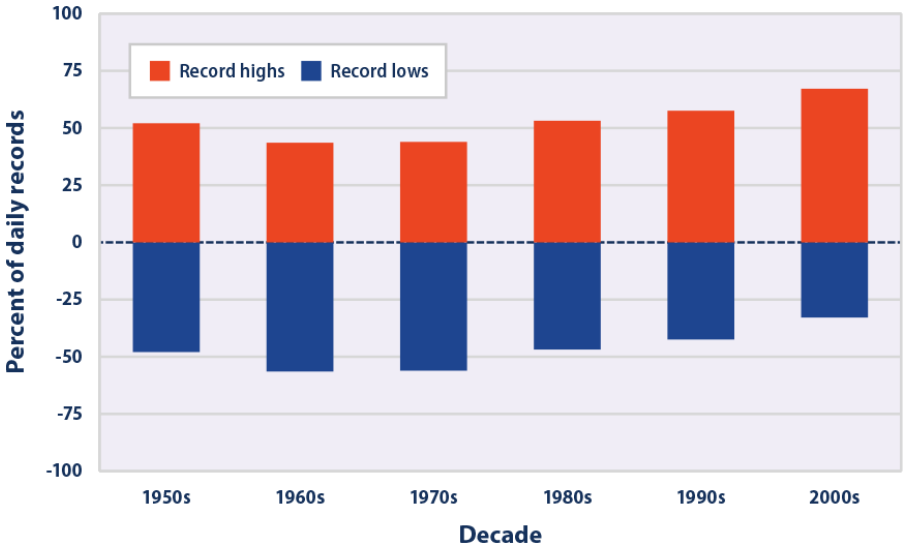


Video 16.8.1: An animation of a GISS (Goddard Institute for Space Studies) climate model simulation made for the United Nations, showing five-year averaged surface air temperature anomalies in degrees Celsius from 1880 to 2100. The temperature anomaly is a measure of how much warmer or colder it is at a particular time and place than the long-term mean temperature, defined as the average temperature over the 30-year base period of 1951 to 1980. Blue areas represent cool areas and yellow and red areas represent warmer areas. The number in the upper right corner represents the global mean anomaly. (NASA's Goddard Institute for Space Studies via climate.nasa.gov)

Temperature and Precipitation

Temperature and **precipitation** are the two most direct impacts on the Earth's climate due to climate change. As we saw in section 16.7, greenhouse gases in the atmosphere trap and prevent heat from leaving the Earth. Thus, increases in the levels of these gases cause more heat to be trapped and increase in the overall temperature of the Earth. You may have noticed in Figure 16.7.5 and Video 16.8.1 that the areas of hot and cold temperatures were not uniform across the United States or the globe. Why would this be the case? Again, this is due to the redistribution of energy in the Earth's atmosphere. Shifting air masses will cause temperatures to increase or decrease in different ways. In the US, the temperature is increasing more in the western states, the north, and in several areas along the Gulf and Atlantic coasts (Figure 16.8.1). This is due to an increase in the number of record hot days and a decrease in the lows at night. For northern states, this has been especially prevalent over the winter months. For all 48 contiguous states, these shifts in temperatures can bring about a variety of results, both good and bad. It has resulted in the fall frost beginning later in the year and ending earlier in the spring, resulting in a longer growing season but possibly affecting the types of crops which may be grown. Warmer winters influence the amount of snow that falls and accumulates which threatens water supplies in western states.


Winter
Spring
Summer
Fall
Total temperature change (°F):
-5 -4 -3 -2 -1 0 1 2 3 4 5 6
Gray interval: -0.1 to 0.1°F


Record highs
Record lows
Percent of daily records
100
75
50
25
0
-25
-50
-75
-100
1950s
1960s
1970s
1980s
1990s
2000s
Decade

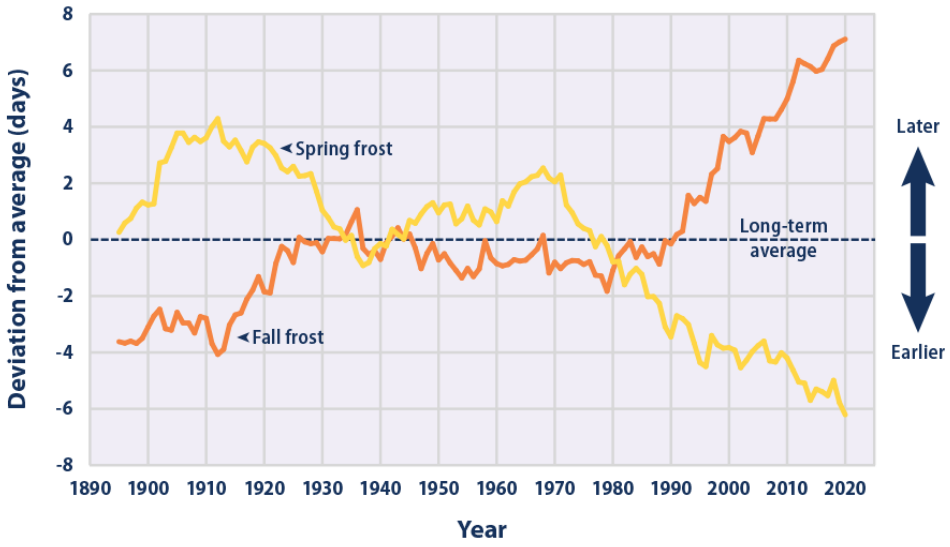


Figure 16.8.1: The top figure shows the total change in the average seasonal temperature for each of the 48 contiguous states in degrees Farenheit from 1896 to 2021. The total change in temperature was calculated from the long-term average rate of change. The middle figure shows the percentage of daily temperature records (highs are red, lows are blue) set at weather stations across the contiguous 48 states by decade. The bottom figure shows the timing of the last spring frost and the first fall frost compared with a long-term average. Choosing a different long-term average for comparison would not change the shape of the data over time. (NOAA via epa.gov)

Increases in greenhouse gases causes increases in temperature. But why does it also impact precipitation patterns? As was previously discussed, water vapor is an important component of the Earth's atmosphere. As the air in the troposphere warms and cools, the amount of water vapor that it holds changes dramatically. In the southeastern US, they have very hot and humid summers. The high summer humidity in this region is possible because of the increased capability warm air has to hold water vapor. Simply put, warmer air can hold more water than cooler air. As air cools, its ability to hold water vapor decreases, and any excess water will leave the air as liquid water. A great example of this is the formation of dew on surfaces overnight. During the day, the temperature is warmer than it is at night, and the air has a relatively high holding capacity for water vapor. When the sun sets, the air cools, decreasing its capacity to hold water vapor. That extra water must go somewhere, and it does that by accumulating on surfaces. Similarly, when warm and cool air fronts collide, the chances for rain and thunderstorms increase. Furthermore, an increase in temperature enhances evaporation occurring at the Earth's surface. This increased evaporation leads to greater concentrations of water vapor in the atmosphere which can lead to increased precipitation.

Changes in precipitation occur due to a variety of factors, including changes in atmospheric water vapor content due to changing temperature, as discussed above. Also at play is the heightened **evaporation** rate of water on Earth's surface under warmer temperatures. More evaporation leads to more precipitation. Finally, shifts in wind patterns impact the distribution of precipitation events. As you can see in Figure 16.8.2, there are some areas of the globe that are expected to have an increase in precipitation, while others are expected to have a dramatic decrease. Some major population centers projected to have a moderate to severe precipitation increase include (population estimates of the metropolitan area given in parentheses): New York, United States (20.1 million); Bogotá, Colombia (12.1 m.); and Manila, Philippines (11.9 m.). What sort of challenges might these cities face in the future as they deal with this change in their climate?

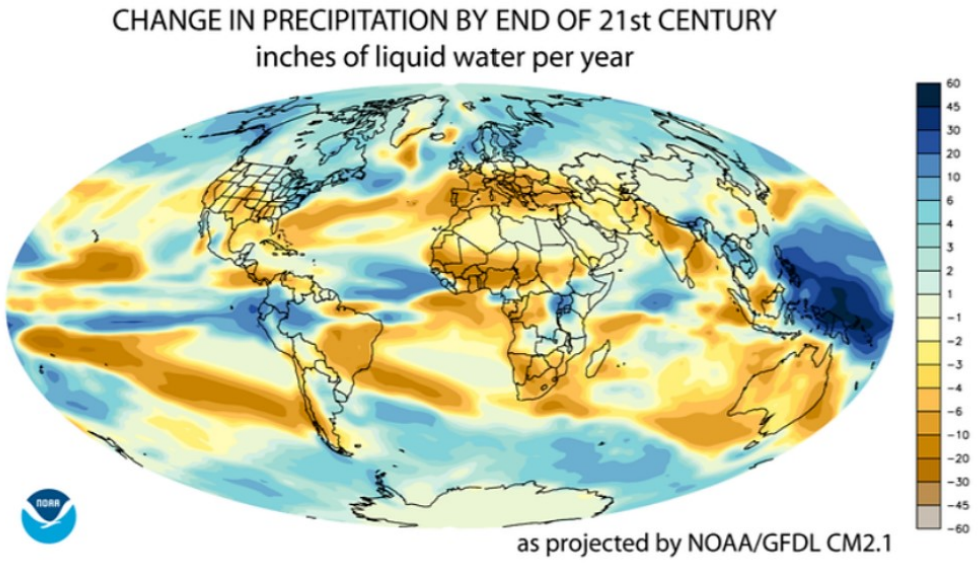


Figure **16.8.2**. The plotted precipitation differences were computed as the difference between the 2081 to 2100 20-year averages minus the 1951 to 2000 50-year average. Blue areas project increases in precipitation; brown areas project decreases.

In contrast, many more major metropolitan areas are projected to have a moderate to severe precipitation decrease (droughts) by the end of the 21st century. These include Delhi, India (21.8 m.); Lagos, Nigeria (21 m.); São Paulo, Brazil (20.9 m.); Kolkata, India (14.6 m.); Istanbul, Turkey (14.4 m.); Los Angeles, United States (13.3 m.); Rio de Janeiro, Brazil (12 m.); Paris, France (12 m.); and Lahore, Pakistan (11.3 m.). The largest challenge that these areas are likely to face is a dwindling water supply for drinking and agriculture. More details on challenges faced by societies to supply clean, reliable water to their populations and farms will be addressed in Section 16.11 and 16.12.

Additional challenges may be felt by all areas of the world with regard to changes in the seasonality or timing of precipitation, as well as the form in which precipitation falls (e.g., mist or downpour; rain, ice, or snow). All of these factors affect the availability of soil water for plants, the flow of rivers and streams, and the overall accessibility of water worldwide. Furthermore, scientists predict an increase in the number and severity of storms as climate change progresses.

Sea Level Rise

While we know that water continuously cycles around the world and that the overall quantity of water on Earth will not change due to global climate change, the distribution of this water is changing. In particular, oceans are increasing in volume while land ice stores (such as glaciers) are decreasing. This contributes to an increase in sea level worldwide (Figure **16.8.3**).

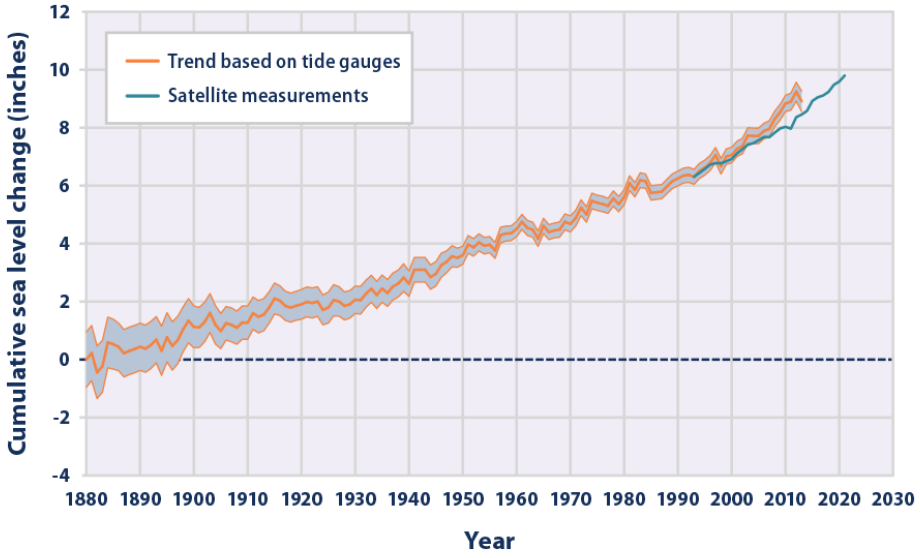


Figure 16.8.3: Cumulative changes in sea level for the world's oceans since 1880. The sea level change refers to the height of the ocean surface, regardless of whether nearby land is rising or falling. The shaded band shows the likely range of values based on the number of measurements collected and the precision of the methods used. (NOAA via epa.gov)

From the data in Figure 16.8.3, we see that sea level has increased at an average of 0.06 inches (0.15 cm) per year over the time period shown above. Most of this rise, however, has occurred within the most recent decades. The rate of increase has gone up to between 0.11 to 0.14 inches (0.28 to 0.36 cm) per year since 1993. There are two forces causing sea level to rise, both caused by climate change. First, the increased global temperature has caused increased ice melting in many regions of the globe. Melting **land ice** (such as the glacier shown in Figure 16.8.4) contributes to sea level rise because water that used to be stored in ice sitting on top of land becomes running water which reaches the ocean through **runoff.** We also observe **sea ice** melting. Sea ice, such as the ice that covers the arctic regions of the Northern Hemisphere, has no land underneath it. When it melts, the water stays in the same locations, and the overall sea level does not change.



Figure **16.8.4**: Photographs of McCall Glacier, Alaska, in 1958 and 2003. (Post, A. and Nolan, M. via epa.gov)

The second factor that influences sea level rise is a phenomenon called thermal expansion. Due to the physical properties of water, as water warms, its density decreases. A less dense substance will have fewer molecules in a given area than a more dense substance (Section 2.9). This means that as the overall temperature of the oceans increases due to global climate change, the same amount of water molecules will now occupy a slightly larger volume. This may not seem significant, but considering the 1.3 billion trillion liters (264 billion gallons) of water in the ocean, even a small change in density can have large effects on sea level as a whole.

Scientists have already documented sea level rise in some areas of the world, including one familiar to most of us: the United States. Figure **16.8.5** shows the relative rise in sea level along the coasts of the US. The change is particularly drastic along the eastern coast. Figure **16.8.6** depicts the measured land area lost due to this increasing sea level since 1996. Note that the Southeast (defined here as the Atlantic coast of North Carolina south to Florida) is particularly susceptible to land area loss due to the gently sloping nature of the coastline. Moving northward into the Mid-Atlantic States (defined here as Virginia north to Long Island, New York), coastal habitats tend to have a steeper geography, which protects against some losses.

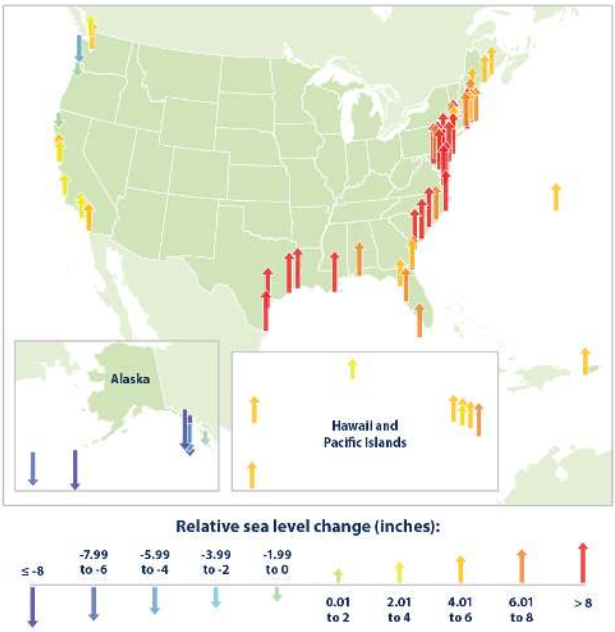


Figure **16.8.5**: Cumulative changes in relative sea level from 1960 to 2021 at tide gauge stations along the US coast. Relative sea level reflects changes in sea level as well as land elevation. (NOAA via epa.gov)

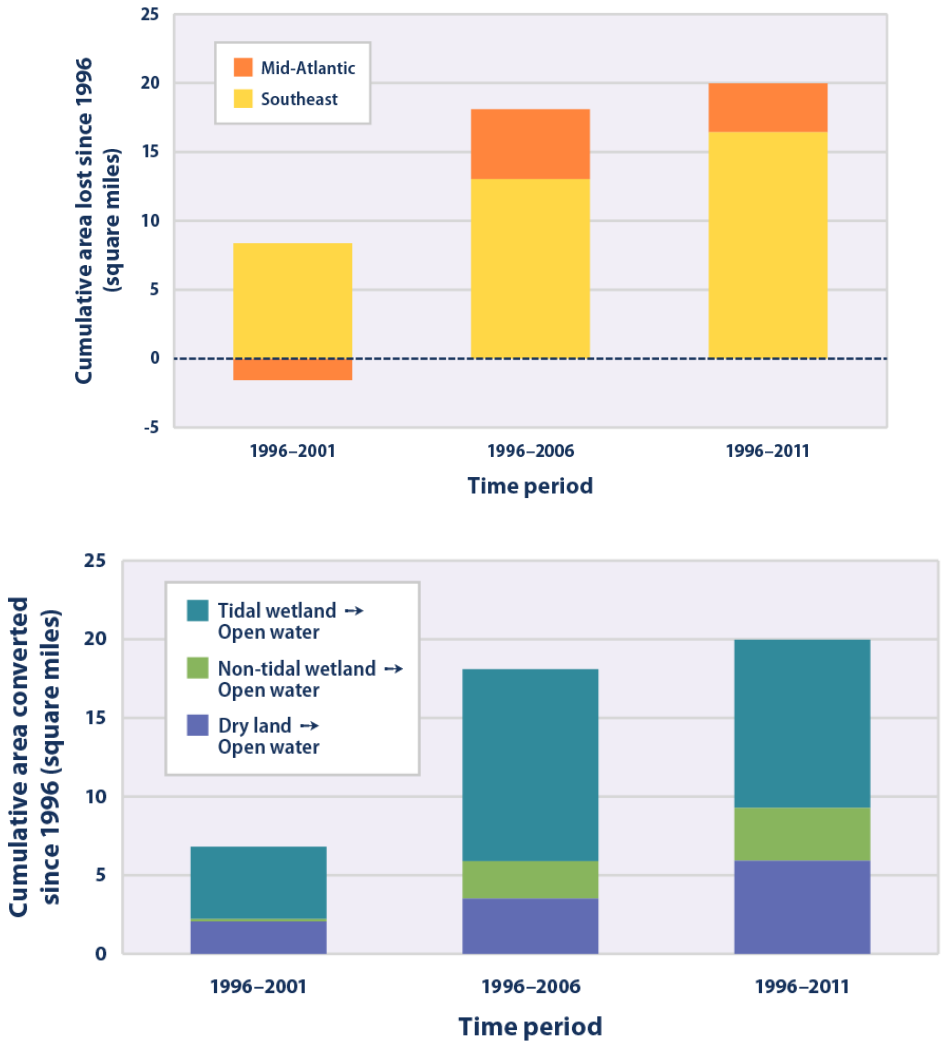


Figure **16.8.6**: The first graph shows the net amount of land converted to open water along the Atlantic coast during three time periods. The second graph breaks down the types of land that have been converted. (NOAA via epa.gov)

Although the rise in sea level will produce ecological effects that will be felt in the United States, catastrophic loss of life, property, or livelihood is not projected to happen for some time. This is due in part to large investments that we have made in infrastructure to protect our cities and farmlands. This is not the case in many areas of the world. The impacts of sea level rise on less industrialized nations such as Bangladesh, Maldives, Kiribati, and Fiji should be of concern to all nations and are being considered by groups such as the United Nations.

Ocean acidification

Dissolved CO_2 is essential for many organisms, including shell-building animals and other organisms that form a hard coating on their exterior (e.g., shellfish, corals, Haptophyte algae). This hard coating is built out of aragonite, a mineral form of calcium carbonate, $CaCO_3$. These organisms rely on the formation of carbonate ions, CO_3^{2-}, from dissolved CO_2 through a natural, chemical reaction that occurs in the oceans and lakes of the world. This takes place through a two-step reaction, where bicarbonate (HCO_3^-) is formed as an intermediate, and hydrogen ions (H^+) are generated (Equations **16.8.1** and **16.8.2**).

$$CO_2 + H_2O \leftrightarrow H^+ + HCO_3^- \tag{16.8.1}$$

$$HCO_3^- \leftrightarrow H^+ + CO_3^{2-} \tag{16.8.2}$$

As you can see, both Equations **16.8.1** and **16.8.2** each produce one H^+ . This is significant to water chemistry because an increase in H^+ concentration means a decrease in the pH of the water. As we learned (or will learn) in Section 15.7, a lower pH means that the liquid is more acidic. Thus, an increase in CO_2 in the atmosphere causes additional CO_2 to be dissolved in the ocean. This means that more CO_2 in the atmosphere leads to more acidic ocean environments. Unfortunately for shell-building animals, the buildup of H^+ in the more acidic ocean environment blocks the absorption of calcium and CO_3^{2-} and makes the formation of aragonite more difficult. An aragonite deficit is already being documented in many of the world's oceans, as shown in Figure **16.8.7**.

The increasing acidity of the world's oceans is resulting in habitat changes across the globe. This is only expected to worsen as atmospheric CO_2 levels continue to increase. Many organisms, including the corals that are the foundation species of the beautiful coral reefs, are very sensitive to changes in ocean pH. Scientists have documented cases of ecosystem destruction through coral bleaching, caused by the effects of climate change including ocean acidification and increased temperature. For more information, visit the NOAA Coral Reef Conservation Program website: coralreef.noaa.gov/issues/.

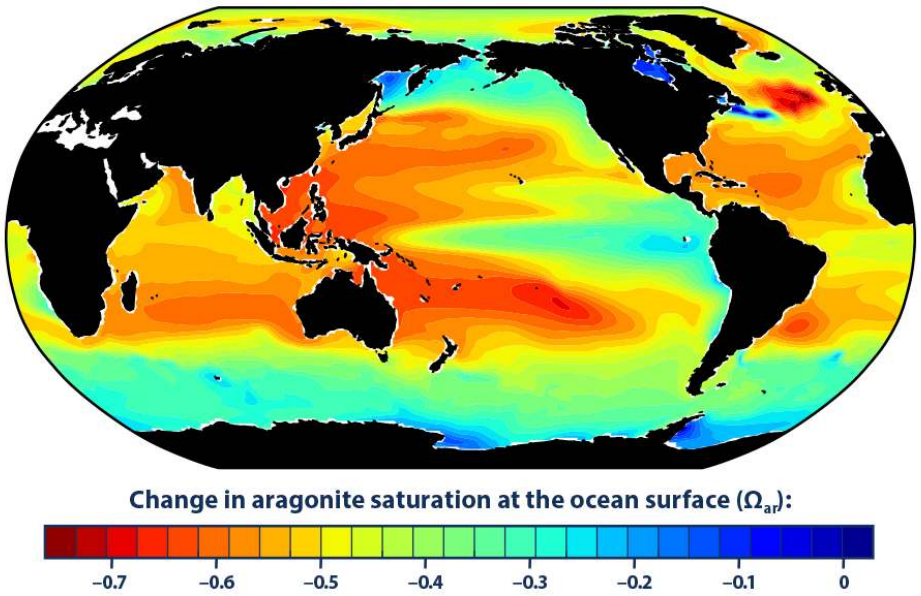


Figure **16.8.7**: This map shows changes in the aragonite saturation level of ocean surface waters between the 1880s and the most recent decade (2006–2015). Aragonite is a form of calcium carbonate that many marine animals use to build their skeletons and shells. The lower the saturation level, the more difficult it is for organisms to build and maintain their skeletons and shells. A negative change represents a decrease in saturation. (Woods Hole Oceanographic Institution via epa.gov)

Summary

- Climate change includes both global warming driven by human-induced emissions of greenhouse gases and the resulting large-scale shifts in weather patterns.
- Changes in temperature and precipitation occur with global warming because of the ability of water to absorb and release large quantities of heat as it evaporates and then condenses as rain. Changes to these atmospheric seas due to increased global temperatures causes certain areas to have warmer temperatures or greater precipitation while others have less.
- Melting land ice and thermal expansion of water causes sea levels to rise. In low-lying areas, this is removing wetlands that provide a buffer to flooding.
- Increased levels of CO_2 in the atmosphere leads to higher dissolved levels in oceans. The reaction between CO_2 and water leads to the formation of carbonic acid, resulting in the acidification of the oceans. When oceans become too acidic, shell-building animals are unable to absorb Ca^{2+} and CO_3^{2-} to form their exoskeletons.

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16.9: Climate Mitigation Strategies

Learning Objectives

- Identify the difference between adaptation and mitigation in dealing with climate change.
- Discuss different adaptation strategies that can be put in place to deal with a changing environment.
- Describe alternative energy sources that can be put in place to mitigate carbon dioxide emissions.
- Describe how carbon sequestration can lower carbon dioxide emissions.
- Identify different strategies that you as an individual can put in place to reduce greenhouse gas emissions.

While the situation surrounding global climate change is in serious need of our attention, it is important to realize that many scientists, leaders, and concerned citizens are making solutions to climate change part of their life's work. The two solutions to the problems caused by climate change are **adaptation** and **mitigation**, and we will likely need a combination of both in order to prosper in the future.

Adaptation strategies

We know that climate change is already occurring, as we can see and feel the effects of it. For this reason, it is essential to be able to adapt to our changing environment. This means that we must change our behaviors in response to the changing environment around us. Adaption strategies will vary greatly by region, depending on the largest specific impacts in that area. For example, in the city of Delhi, India, a dramatic decrease in rainfall is projected over the next century. This city will likely need to implement policies and practices relating to conservation of water such as rainwater harvesting, water re-use, and increased irrigation efficiency. Rain-limited cities near oceans, such as Los Angeles, California may choose to use desalination to provide drinking water to their citizens. **Desalination** involves taking the salt out of seawater to make it potable.

Cities with low elevations near oceans may need to implement adaptation strategies to cope with rising sea levels, from seawalls and levees to relocation of citizens. One adaptation strategy gaining use is the creation or conservation of **wetlands**, which provide natural protection against storm surges and flooding.

Green chemistry, biochemistry, and biology can be used to deal with global warming when it occurs. Crops, fertilizers, and pesticides can be developed that enable plants to grow under the drought conditions that will follow global warming in some locations. Another approach is the development of salt-tolerant crops that can be grown on soil irrigated with saline water, where fresh water supplies are limited.

Mitigation Strategies - Energy Usage

In general, a strategy to mitigate climate change is one that reduces the amount of greenhouse gases in the atmosphere or prevents additional emissions. As noted in the previous section, the most significant of these is carbon dioxide. One way to reduce the release of carbon dioxide is through the use of alternative energy sources.

Until about 1800 the sun provided virtually all of the energy used by humans. Biomass produced by photosynthesis was burned for heat or used to feed horses, oxen, and even humans who provided muscle power. Because the carbon dioxide generated by this burning was replacing that used to grow the plants, there was no net increase in atmospheric carbon dioxide levels. Wind resulting from uneven solar heating of air masses powered sailing ships and windmills. The solar powered hydrologic cycle produced running water that drove water wheels.

Around 1800 the dramatically increased use of coal began the era of fossil fuel energy sources. This use was enabled by the invention of the steam engine, which provided abundant and reliable power for stationary sources and locomotives and the power used for mining the coal. By 1900 coal was the dominant energy source in industrialized societies, but petroleum, especially well suited as fuel for the newly developed internal combustion engine, began a century of rapid development, becoming the favored fuel for transportation. Often encountered in wells drilled for petroleum, natural gas also developed during the 1900s, predominantly as a fuel for non-transportation needs. During this century hydroelectric power took on a significant share of electrical power production and remains a substantial renewable energy source. By around 1975, nuclear energy was supplying significant amounts of electricity and has maintained an appreciable share worldwide until the present. Miscellaneous renewable

sources including geothermal and, more recently, solar and wind energy are making increasing contributions to total energy supply. Biomass still contributes significantly to the total of the sources of energy used and will become even more important as emphasis on renewable sources increases.

Despite so many different sources of energy that are available to us, most are not directly useful and must be converted to other forms. Therefore, much of what is done with energy involves changing it from one form to another. This is not always done efficiently using current technologies. If those efficiencies can be improved, less energy needs to be generated to do the same work. This has been seen in heat engines such as automobile engines and gas turbines by higher combustion temperatures made possible by improved materials and heat-resistant lubricating oils. Computerized design and operation of engines enabling optimum ignition timing, valve timing, and fuel injection have made possible extremely efficient engines with vastly improved gas mileage.

Much of the lost energy comes in the form of heat. In stationary power plants much of this energy can be reclaimed for heating buildings or chemical processes with **combined power cycles** as illustrated in Figure 16.9.1. Typically, in combined power cycle installations, gas or fuel oil is burned in a turbine engine that is much like the engine of a turboprop airplane, and the rotating shaft of this engine is coupled to a generator to produce electricity. The hot exhaust gases from the combustion turbine can be injected into a boiler where their heat turns liquid water to steam. This steam can be run through a steam turbine coupled to a generator to produce more electricity. Steam leaving the steam turbine still contains a lot of heat, and can be conveyed to homes and other structures for heating. The water condensed from this steam is pure and is recycled to the boiler, thus minimizing the amount of makeup boiler feedwater, which requires expensive treatment to make it suitable for use in boilers. Such a system as the one described is in keeping with the best practice of industrial ecology. Heating with steam that has been through a steam turbine, a concept known as **district heating**, is commonly practiced in Europe (and many university campuses in the U.S.) and can save large amounts of fuel otherwise required for heating.

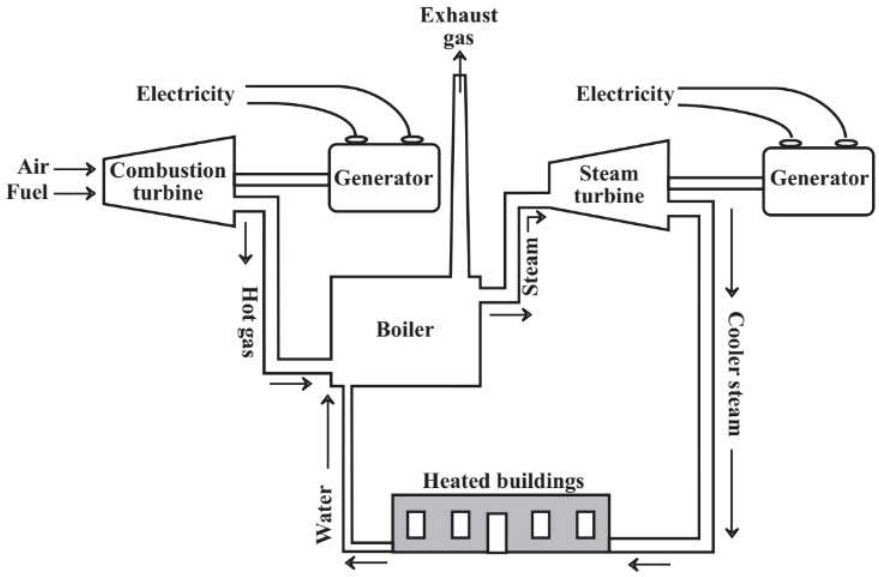


Figure 16.9.1: A combined power cycle in which natural gas or oil is first used to fire a gas turbine connected to an electrical generator. The hot exhaust gases from this turbine are fed to a boiler to produce steam, which drives a steam turbine, also connected to a generator. The still hot exhaust steam from the steam turbine is distributed to residential and commercial buildings for heating. The water condensed from the steam used to heat buildings is returned to the power plant to generate more steam, which conserves water and also prevents the necessity of treating more water to the high standards required for a boiler.

Part of what makes the US and other industrialized countries so dependent on fossil fuels is the relative ease by which they are obtained. One only needs to drill a hole in the ground and transportation can occur through existing pipelines and tankers to refineries that are already in existence. However, energy has been wasted to a shocking extent by the US and some other industrialized nations. This points to the opportunity afforded by energy conservation as the most effective means of providing adequate energy while reducing greenhouse gas emissions. The potential for energy conservation is illustrated by the fact that some nations that have living standards near or even exceeding those of the US have much lower energy consumption per capita.

The economic sector with the greatest potential for increased energy efficiency is transportation. Private automobiles and commercial aircraft are only about one-third as efficient as trains and buses in moving people. Movement of freight by truck requires 5-6 times as much energy as transport by train. Furthermore, electrified railways are much better adapted to using renewable sources of energy than are private automobiles, trucks, and aircraft.

As illustrated in Figure **16.9.2**, US automobile fuel economy increased impressively from the first "energy crisis" of the 1970s until about 1990. This was achieved along with much lower emissions of exhaust pollutants. Unfortunately, the trend toward better fuel economy, which if continued would have meant average mileage figures of at least 40 miles per gallon (mpg) by 2010, stopped moving up with increased popularity of oversized vehicles, especially the "sport utility vehicle" type. In 2007 the US Congress passed legislation mandating improved fuel economy for automobiles sold in the US. An average of 40 mpg is readily achievable without significantly compromising safety and comfort and has the added benefit of substantially reduced emissions of the greenhouse gas carbon dioxide.

Hybrid vehicles that are propelled by an electric motor connected to a battery rechargeable with a small internal combustion engine now deliver impressive fuel economy. The improvement is especially pronounced for stop-and-go driving in traffic where at least 50% better mileage figures have been achieved. In operation, the main battery in the vehicle, which is much larger than a conventional automobile battery, but smaller than one required for an all-electric vehicle, is kept charged by the onboard internal combustion engine coupled to a generator. A contribution is also made by regenerative braking that generates electricity during braking. When the vehicle is slowing, coasting downhill, or stopped, the internal combustion engine turns off which saves fuel.

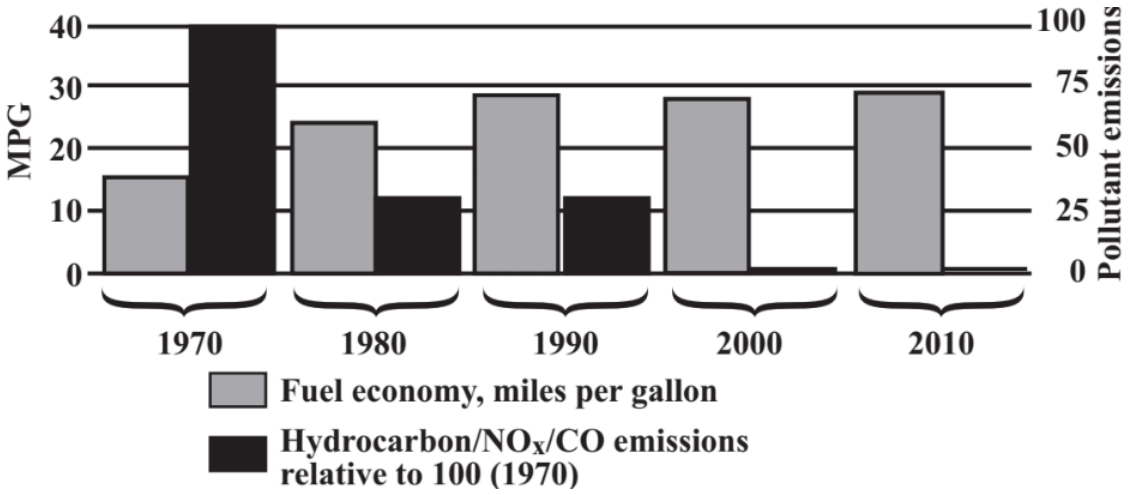


Figure **16.9.2**: U.S. auto fleet fuel economy and emissions over 4 decades, showing a trend of generally higher fuel economy with greatly reduced emissions.

Until recently, the battery of choice for hybrid vehicles has been the nickel-metal-hydride battery. As of 2010, development was actively underway on hybrid vehicles using lithium ion batteries which hold a relatively greater charge per unit mass. These batteries can be charged from an external source of electricity providing 40-50 km of driving range before the onboard internal combustion engine has to engage. Charging the battery from an external source is much less expensive than using an internal combustion engine, affording rather spectacular fuel economy.

One advantage of electrical energy, whether in driving cars or in the home, is that there are multiple ways of generating the electricity. While many electrical power plants in the US still use coal to generate electricity, there are a number of other options. Some areas use nuclear power plants which use nuclear power as a heat source to produce steam used to generate mechanical and then electrical energy. While this source of energy does not release any greenhouse gases, it does generate high-level nuclear wastes which must be dealt with for centuries.

Other energy types--considered to be renewable energy sources--are solar, wind, and water energy. Solar energy can be used simply for heating. It can also be converted directly to electricity using photovoltaic cells. While solar energy obviously cannot be collected at night, the use of photovoltaic cells in combination with storage devices will allow for a continuous supply of electricity. Wind and water energy are among the oldest sources of energy used by mankind, such as wind used to propel sailing ships and

waterwheels used for centuries to grind grain. They are also among the newest sources of energy with technologically advanced wind turbines and ingenious devices that capture the energy of moving water in ocean tides.

Wind power is undergoing rapid growth in a number of countries and has become competitive in cost with more conventional sources in some areas. In parts of Europe, California, Wyoming, and other locations, the sight of wind-powered generators mounted on towers has become common (Figure **16.9.3**). World wind power capacity increased by 31% in 2009 alone, reaching a capacity of 158 gigawatts (GW). During 2009, China's wind power capacity doubled from 12 GW to 25 GW and the US capacity grew by 10 GW to 35 GW total capacity. This gave the US the largest wind power capacity of any nation in the world, though it is rapidly being overtaken by China.

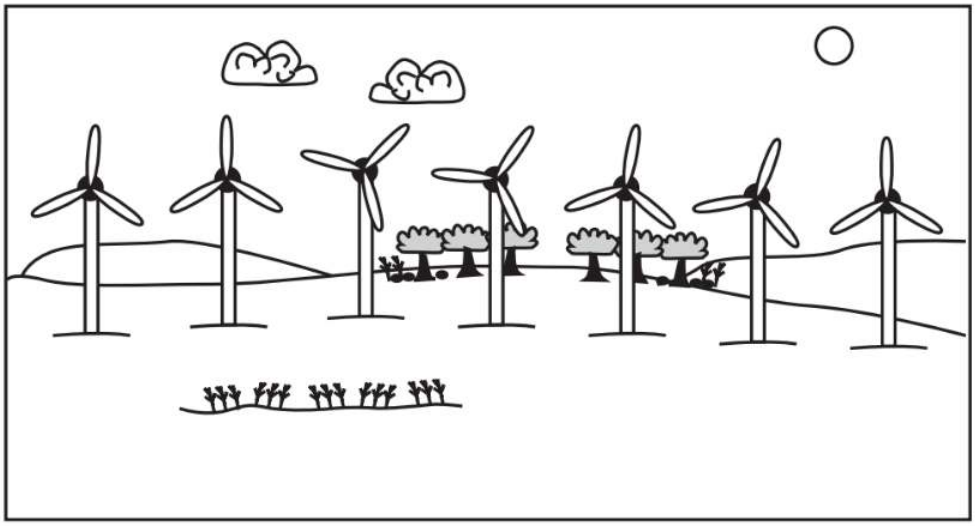


Figure **16.9.3**: Wind-powered electrical generators mounted on towers are becoming increasingly common sights in the world in areas where consistent wind makes this nonpolluting source of renewable energy practical.

Modern wind turbines are generally large and sophisticated machines with diameters of 40 to 50 meters and rated between 0.5 to 2 megawatts (MW). The largest turbines available as of 2010 were rated at 7 MW. Wind turbines are designed to operate consistently at varying wind speeds, to respond to changing wind direction, and to operate over a wide range of temperatures. Provision is made for electrical resistance heating of turbine blades in cold climates where ice accumulation is likely.

Offshore locations with turbines firmly anchored to the sea bed offer several advantages for production of wind power including generally consistent winds in coastal areas and lack of conflict with uses of land. The largest offshore wind power project to date began operation off the coast of southeast England in September 2010. This facility operated by Vattenfall, a Swedish energy company, has 100 turbines each 115 meters tall, and produces up to 300 megawatts of electricity, enough to power 200,000 homes.

A renaissance in waterpower occurred in the late 1800s when it became practical to drive electrical generators with water. Starting with the first hydroelectric plant in the U.S. on the Fox River near Appleton, Wisconsin, in 1882, hydroelectric power driven by sophisticated power turbines grew rapidly in the U.S. and throughout the world. By 1980 hydroelectric power composed 25% of world electricity production and 5% of total world energy generation. Norway generates essentially all of its electric power and about 50% of its total energy from hydroelectric sources.

The morphology of the geosphere has a strong influence on the potential for hydroelectric power generation, which is favored by mountainous terrain and large river valleys. Abundant rainfall and snowmelt are also important factors. China has about 1/10 of the world's potential for hydroelectric energy and its enormous Three Gorges installation on the Yangtze River is the world's largest. The reservoir for this power source has been impounded with a 185-meter high, 1,300-m long dam located at the end of a number of steep canyons holding a body of water that extends for 630 km with an average width of 1.2 km. When fully operational, this massive installation will have 32 generating units and a capacity of 22.5 gigawatts, equivalent to 22 large coal-fired power plants.

With the hydrologic cycle continuously pumping water into it, hydroelectric power is certainly sustainable, and prevents release of greenhouse gases. Bodies of water impounded to provide power can serve as water supplies for agriculture, municipalities, and industries. The potential exists to practice aquaculture in reservoirs by raising fish and freshwater shrimp(prawns). On the negative

side, the development of hydroelectric power can present some serious environmental problems. In the modern era, construction of a large power reservoir displaces significant numbers of people (more than 1 million for China's Three Gorges project), alters river flow, changes aquatic ecology, and fills once scenic valleys with water. In several significant cases dams have been removed from rivers to restore their valleys to their former state.

For those instances where electrical energy is not practical, new technologies are making it possible to use biomass, or the liquid and gaseous fuels made by processing it, for combustion. This is especially promising for transportation. Again, biomass consists of photosynthetically generated materials such as wood, charcoal, forestry residues, agricultural by-products, rapidly growing grasses, livestock manure, and algae. Since these compounds are made by removing carbon dioxide from the atmosphere during photosynthesis, the carbon dioxide generated during their combustion does not lead to a net increase in the levels in the atmosphere. In addition, there tends to be very little sulfur dioxide in biofuels and the ash contains nutrients that can be returned to the soil without adding the heavy metals that is a problem with coal ash. Today, more than 14% of the world's primary energy comes from biofuels, especially wood and charcoal.

Despite concerns that not enough biomass can be grown to produce fuel and that it detracts from food supplies, it should be noted that about 150 billion metric tons of biomass are produced in the world each year by photosynthesis, mostly from uncontrolled plant growth. Corn, the most productive common field crop, produces about 4 metric tons per acre of dry biomass annually (including stalks, leaves, husks, and corncobs). Switchgrass, a prolific producer of biomass, typically generates 11-12 tons of biomass per acre per year (there are 640 acres in a square mile of land). About 6% of the biomass generated globally each year would be equivalent to the world's demand for fossil fuels. Cultivation for fuel biomass of 6-8% of the land area of the 48 contiguous states would provide energy equivalent to annual U.S. consumption of petroleum and natural gas. Furthermore, only a small fraction of widely grown grain crops goes into grain; the rest is plant biomass, much of which could be used for energy production. And the U.S. has vast areas of underutilized land that could be devoted to the cultivation of energy-yielding plants. Much of this neglected, erosion-prone land would benefit from the cultivation of perennial plants that could be harvested for energy and regrow from roots left in the ground, thus lowering water and wind erosion.

In order for some of these alternative energy sources to become viable options in large-scale generation of energy and electricity, investments do need to be made in technology and infrastructure to support it. One area where governments can help with this is by providing incentives to citizens and companies for investing in renewable energy sources. This is one type of institutionalized mitigation strategy. While the research may be initially expensive, we often develop technologies that have major impacts in other sectors of the economy as well, such as improved batteries that can be used for other types of devices. Thus, it is imperative that initiatives be taken now instead of waiting for others to develop the technology and make the US dependent on them for its energy needs.

Governmental regulations regarding fuel efficiency of vehicles is another example of an institutionalized mitigation strategy already in place in the United States and in many other countries around the world. Unlike some other countries, there are no **carbon taxes** or charges on burning fossil fuels in the United States. This is another governmental mitigation strategy that has been shown to be effective in many countries including India, Japan, France, Costa Rica, Canada, and the United Kingdom.

Mitigation Strategies - Carbon Capture and Sequestration

In situations where carbon dioxide must be released on a large scale, technology can also be harnessed to mitigate climate change. One strategy for this is the use of **carbon capture and sequestration** (CCS). Through CCS, 80-90% of the CO_2 that would have been emitted to the atmosphere from sources such as a coal-fired power plant is instead captured and then stored deep beneath the Earth's surface. The CO_2 is often injected and sequestered hundreds of miles underground into porous rock formations sealed below an impermeable layer, where it is stored permanently (Figure **16.9.4**).

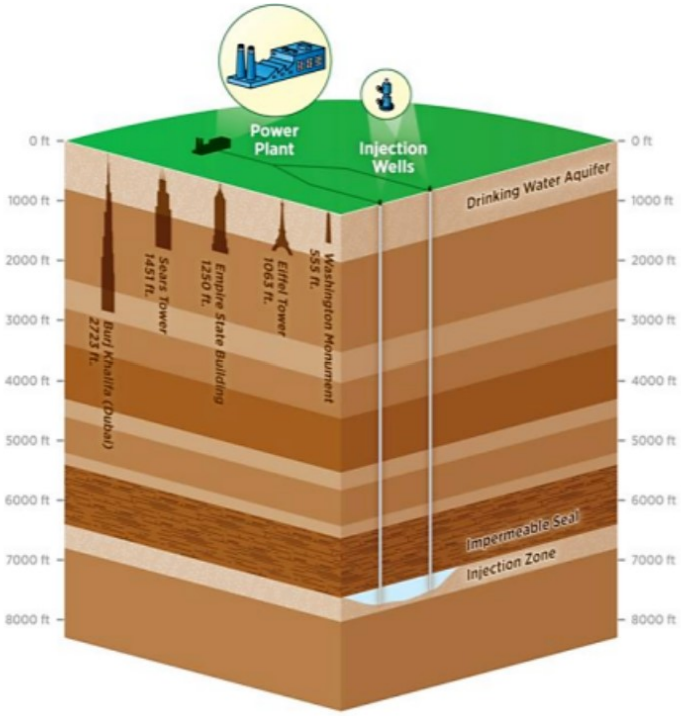


Figure **16.9.4**: Carbon capture and sequestration schematic with landmarks shown to scale for depth reference. (US EPA)

Another option when burning coal is that the hot carbon from the burning coal can be reacted with oxygen and water to produce elemental hydrogen and carbon dioxide. The net reaction for this production is the following:

$$2\,C_{(s)} + O_{2(g)} + 2\,H_2O_{(g)} \rightarrow 2\,CO_{2(g)} + 2H_{2(g)} \tag{16.9.1}$$

The hydrogen generated can be used as a pollution-free fuel in fuel cells or combustion engines. The carbon dioxide can be pumped deep underground. A side benefit of this approach is that in some areas carbon dioxide pumped underground can be used to recover additional crude oil from depleted oil-bearing formations.

Scientists are also looking into the use of soils and vegetation for carbon storage potential. Proper management of soil and forest ecosystems has been shown to create additional carbon sinks for atmospheric carbon, reducing the overall atmospheric CO_2 burden. Increasing soil carbon further benefits communities by providing better-quality soil for agriculture and cultivation.

Green chemistry can also be applied in the prevention of release of greenhouse gases other than carbon dioxide, especially ultrastable volatile compounds that have a high greenhouse gas potential. An excellent example of green chemistry has been the replacement of chlorofluorocarbons (Freons) with analogous compounds having at least one C-H bond, that are rather readily destroyed in the troposphere. Although this was done to prevent destruction of stratospheric ozone by chlorofluorocarbons, it has also been useful to reduce greenhouse warming. Both kinds of compounds act as greenhouse gases, but those with at least one C-H bond last for much shorter times during which they are available to absorb infrared radiation.

Another approach is to limit the emissions of methane, CH_4. Large quantities of methane are released by anoxic bacteria growing in flooded rice paddies. By developing strains of rice and means of cultivation that enable the crop to be grown on unflooded soil, this source of methane can be greatly reduced. Methane collection systems placed in municipal waste landfills can prevent the release of methane from this source and provide a source of methane fuel.

Adaptation, Mitigation and You

While we may look to institutions to solve global warming and the resultant climate change problems, we each have choices that we make every day that can influence greenhouse gas emissions or help with adaptation. When possible, use energy efficient appliances, lighting, heating and cooling equipment, and other products. Heat and cool your home smartly to reduce your energy needs. If financially feasible, consider rooftop solar or other self-supplied green power. Encourage your employer to do the same for larger impact.

Transportation is another area where we have an impact on the environment. Look for ways to use greener transportation such as biking, walking, carpooling, and public transportation. If you do need to drive your own vehicle, choose energy efficient vehicles, ideally hybrid or electric with energy-saving electric vehicle chargers. In addition, drive smarter. Improve your fuel economy by going easy on the brakes and gas, using cruise control, and keeping your vehicle well-maintained. Limiting the number of trips that you need to take by grouping errands and teleworking, if available, can also reduce fuel use and greenhouse gas emissions.

While the area in which you live may not be experiencing water shortages, it does take a lot of energy to pump, treat, and heat water. Thus, conserving water reduces greenhouse emissions as well. Again, look for energy efficient clothes washers and dishwashers. Check for water leaks in your home and workplace and turn off running faucets. Install a low-flow showerhead and take shorter showers. If you have a yard, develop a water-smart landscape suitable to your climate. Choose drought-resistant plants and be cognizant of when and how much to water.

As mentioned, waste contributes to greenhouse gases through the methane gas produced during the breakdown of wastes in landfills. Thus, the most effective to reduce waste emissions is to not create it in the first place by practicing the three R's: reduce, reuse, and recycle. In addition, by buying used items you also reduce the emissions created by producing the new materials.

While these may seem like small things, they do add up, especially if adopted by large community populations. Take a moment to identify ways that you personally can be involved in the mitigation of or adaptation to climate change. What changes can you make in your own life to prevent excess carbon emissions? You can calculate your carbon footprint using the EPA's carbon footprint calculator.

Summary

- Adaptation strategies involve changing behaviors based on changes to the environment which have already occurred.
- Most mitigation strategies are efforts to minimize the amount of carbon dioxide which is released into the atmosphere. The best way to accomplish this is through improved energy efficiency and through alternative energy sources such as wind, solar, water, and biomass.
- In addition to reducing the amount of carbon dioxide generated, the gas can be sequestered by pumping it deep into the Earth or by properly managing forests and soil to provide carbon sinks.
- Green chemistry can be utilized to develop alternative compounds which do not have as detrimental of an effect on the atmosphere as ones used in the past.
- Everyone can do their part to reduce greenhouse gas emissions in the choices that we make every day.

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16.10: Water

Learning Objectives

- Relate the structure of water to some of its unique properties.
- Describe the different sources of fresh water available for use.
- Identify the different ions that can be present in water.
- Describe what makes water hard and how that hardness can affect its used.

Water composes the hydrosphere, which is described in Section 16.2. This section enlarges upon that discussion and the crucial role of water in the environment. Here the term **natural water** is used in reference to water that occurs in the environment in comparison to water in the anthrosphere, such as water in municipal water distribution systems. Water has a special place in living organisms and the environment. The quality and availability of water are of the utmost importance to humans and the environment. Although scarce and badly polluted in many parts of the world, water is arguably the most recyclable of the substances that compose Earth's green capital and it is accurately described as the ultimate green substance.

The chemical formula of water, H_2O, is probably the best known of all compounds. This simple formula represents a substance that is unique and complex in its behavior. These special properties are due to the molecular structure of the H_2O molecule described in previous chapters of this book and represented in Figure **16.10.1**. There are two bonds between the O and H atoms and two sets of lone pairs on the O. The distribution of these lone pairs and bonds as far apart as possible around the O atom results in the two H-O bonds being located at an angle rather than a straight line. The side of the molecule with the two H atoms has a partial positive charge and the side with the two non-bonding pairs has a partial negative charge, so the molecule is polar. This polarity and the ability of the H atoms on one molecule to form hydrogen bonds to O atoms on other molecules determine the remarkable chemical and physical diversity of water.

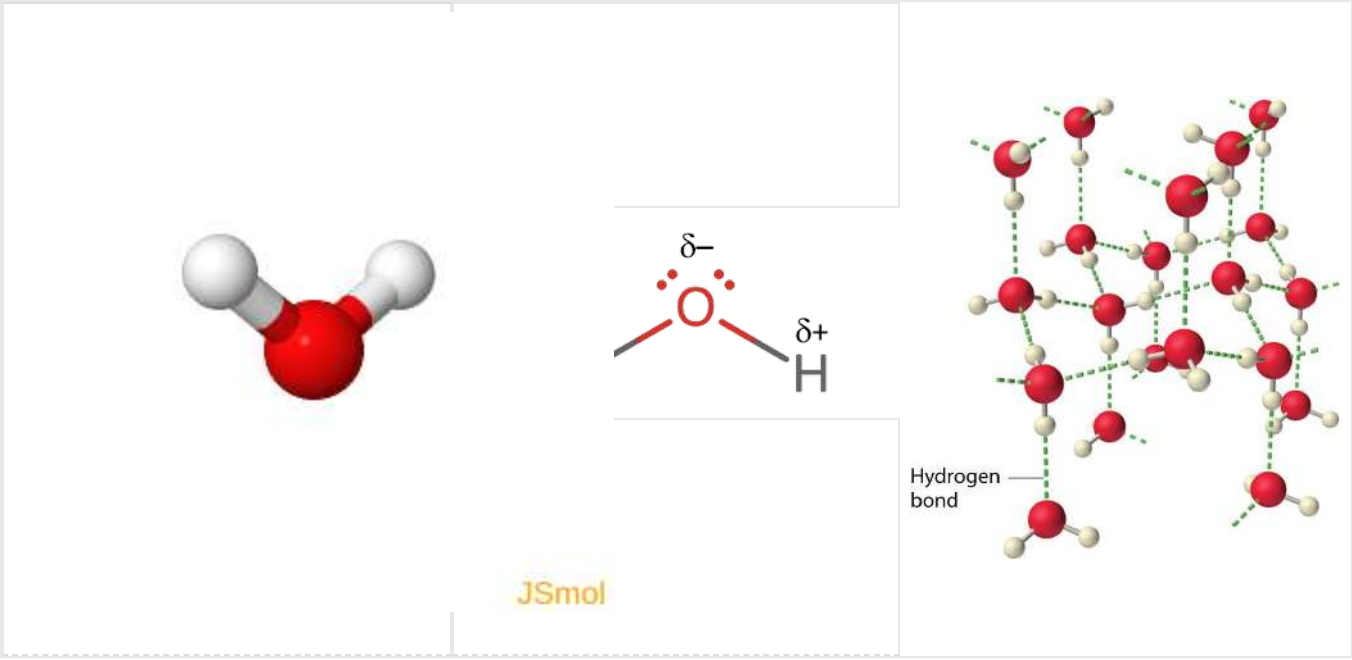


Figure **16.10.1**: Structure of a molecule of water (left), the polarity of water (center), and hydrogen bonding interactions in solid water (right).

Especially because of their hydrogen bonding capability, water molecules are strongly attracted to each other. This means that a large amount of heat energy must be added to a mass of water to enable the molecules to move more rapidly as the temperature is raised. This gives water a very high **heat capacity**. A very large amount of energy must be added to a mass of ice to break the hydrogen bonds holding the molecules in place in the solid as it melts and an equally large amount of heat energy is released when liquid water freezes. Thus water has a very high **latent heat of fusion**. Even more energy per unit mass is required to convert liquid

water to vapor (steam) and an equal amount of energy is released when water vapor condenses to liquid. This means that water has a very high **heat of vaporization**.

The ability of water to absorb, release, and store heat is crucial to its role in the environment and its practical uses. Water's high heat capacity stabilizes temperatures of organisms and geographical areas. Steam produced in a boiler can be transferred through insulated pipes to remote locations and condensed to release heat. The heat released when atmospheric water condenses warms the surrounding air and is the driving force behind tropical storms. Europe owes its relatively mild weather, despite its northern latitudes, to heat carried by water across the North Atlantic Ocean from the Gulf of Mexico. As the water releases heat and cools along the European coasts, its density increases and it flows at lower ocean depths back to the Gulf of Mexico to repeat the cycle. Water's high latent heat of fusion stabilizes temperatures of bodies of water at water's freezing point (0° C).

In addition to those listed above, there are other unique and environmentally important qualities of water. It is an excellent solvent, especially for ionic substances, making it important in the transport of nutrients and wastes in the biosphere and in the dissolution, transport, and deposition of minerals in the geosphere. Water has a very high surface tension, a controlling factor in physiology and a property involved in formation of drops in rainfall. The fact that the maximum density of water occurs as a liquid at 4° C means that solid ice floats. If that were not the case, bodies of water in northern climates would become frozen solid with only a surface layer thawing during the summer. Water is largely transparent to visible light, which can penetrate the liquid to some depth and enable photosynthetic phytoplankton and some bottom-rooted plants in shallow water regions to carry out photosynthesis and produce the biomass that is the basis of aquatic food chains.

The Hydrosphere

The total water system surrounding the planet Earth is called the **hydrosphere**. It includes freshwater systems, oceans, atmosphere vapor, and biological waters. The Arctic, Atlantic, Indian, and Pacific oceans cover 71% of the Earth surface, and contain 97% of all water. Less than 1% is fresh water, and 2-3 % is ice caps and glaciers. The Antarctic Ice Sheet is almost the size of North America continent. These waters dominate our weather and climate, directly and indirectly affecting our daily lives. They cover 3.35×10^8 km^2. The four oceans have a total volume of 1.35×10^9 km^3.

Of the approximately 33 million km^3 of freshwater in the hydrosphere, 86.9% is in the solid form in snowpack, ice and glaciers; 12.0% is accessible water under ground (groundwater), 0.37% is in freshwater lakes, reservoirs, and ponds; 0.31% is in saline lakes; 0.19% is soil moisture; 0.19% is in biospheric organisms; 0.039% is in the atmosphere as vapor and cloud droplets; 0.011% is in wetlands; and only 0.0051% constitutes the water in all Earth's rivers, streams, and canals. Therefore, a remarkably small percentage of Earth's water is relatively available for use, a major factor in the sustainability of this crucial part of Earth's natural capital.

Groundwater

A body of rock which contains appreciable quantities of water is called an **aquifer**. Below the **water table**, the aquifer is filled (or saturated) with water. Above the water table is the unsaturated zone. Some regions have two or more water tables. These zones are usually separated by water-impermeable material such as boulder and clay. **Groundwater**, or the water in the rocks, can be brought to the surface by drilling below the water table and pumped out. The amount of water that can be pumped out depends on the structure of the aquifer. Little water is stored in tight granite layers, but large quantities of water are stored in limestone aquifer layers. In some areas, there are underground rivers.

Figure **16.10.2** illustrates the pathway of groundwater into aquifers from precipitation falling on a watershed and infiltrating the aquifer through a recharge zone. The zone of saturation consists of the geospheric rock and soil layers in which all the pores are filled with water, the top of which defines the water table. The fraction of the aquifer formation consisting of pores is the **porosity** and the ability of the formation to allow movement of water is its **permeability**. Generally, high permeability and high porosity are desirable properties of aquifer formations that allow relatively large amounts of water to be withdrawn through water wells drilled into the aquifers. The quality of groundwater is strongly affected by its contact with mineral formations in the geosphere. Infiltration through rock and sand purifies water by filtering out microorganisms and suspended solids. Contact with geospheric solids largely determines the chemical composition of water by adding desirable levels of dissolved Ca^{2+} ion and alkalinity and sometimes undesirable solutes such as sodium chloride.

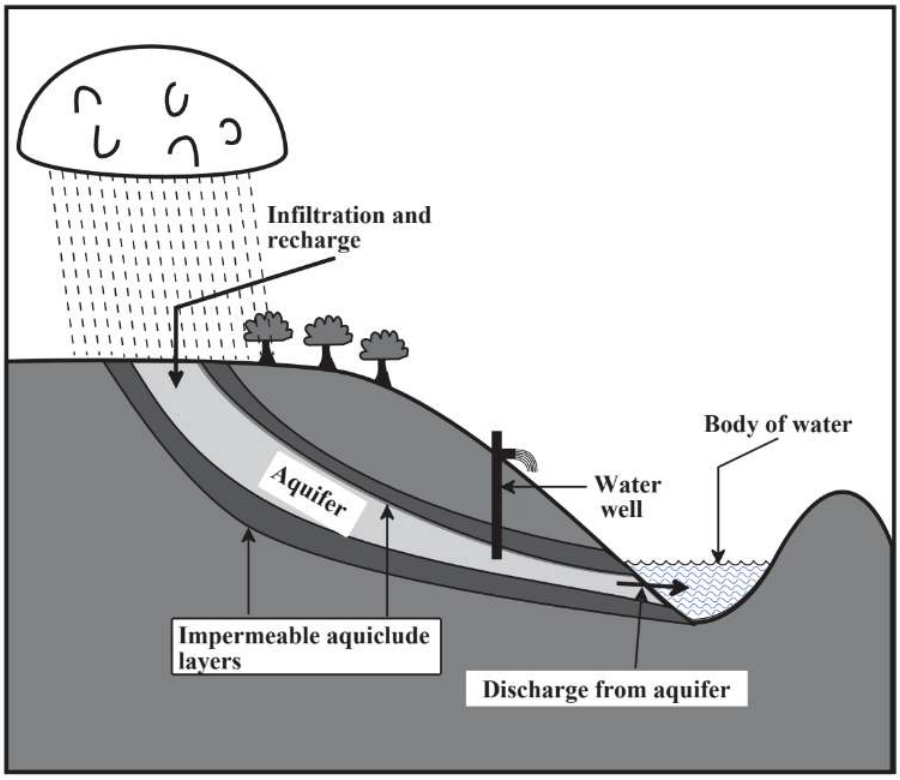


Figure **16.10.2**: Illustration of groundwater in aquifers in the geosphere. Groundwater enters aquifers from surface water in recharge zones and is discharged through springs or pumped out through wells.

Surface Streams and Standing Bodies of Water

A major source of water is surface water that flows in streams and rivers. The areas of land upon which precipitation falls to provide flows of surface water is called the **watershed**. Insofar as water utilization is concerned, watersheds constitute one of the most important connections between the geosphere and the hydrosphere; the nature of the watershed largely determines the quantity and quality of water available for human use. A good watershed retains water for a significant length of time, which reduces flooding, allows for a steady flow of runoff water, and maximizes recharge of water into groundwater reservoirs (aquifer recharge, Figure **16.10.2**). Measures employed to enhance watershed quality include minimization of cultivation and forest cutting, especially on steeply sloping sections of watershed, construction of terraces and waterways planted with grass on cultivated land to minimize erosion and accumulation of sediment, construction of small impoundments on feeder streams of the watershed, and of course minimization of pollution such as from herbicides.

An important aspect of streams that affects water utilization and quality is their ability to mobilize sedimentary materials through erosion, transport materials along with stream flow, and deposit them as solids. Normally it is desirable for streams used as a source of water to have minimal sedimentary material, which has to be removed during water treatment.

Most rivers, once free flowing and unimpeded by human intervention, have been modified by humans to generate power, for water supply, and reduce effects of flooding. Beginning with impoundment of water from the Owens Valley in northern California and extending to the Feather, Sacramento, and Colorado rivers, Los Angeles' voracious thirst for water is served by a vast system of dams, canals, and tunnels. Diversion of water from the Owens valley has turned a formerly productive agricultural area into one of limited agricultural use. Many of the adverse effects of water diversion have resulted from dams built to confine rivers. As a result of dam construction, many once beautiful river valleys have been covered with water and productive farmland in river valleys has been lost. For example, the beautiful Hetch-Hetchy Valley in Yosemite National Park in California has been inundated since the early 1900s by construction of a dam built to provide water and hydroelectric power to San Francisco. Serious proposals are now being considered to remove the dam and restore the valley to its former beauty.

Much of the water that humans use comes from standing bodies of water including natural lakes and reservoirs constructed by placing dams on rivers. **Wetlands** are bodies of water shallow enough to support the growth of bottom-rooted plants. **Estuaries**

form where fresh river water flows into oceans. They have unique physical, chemical, and biological properties because of the mixing of fresh water and saltwater. Wetlands and estuaries are important breeding grounds for a number of organisms and it is crucial that they be preserved. Many wetland areas have been drained for agricultural land and an important effort in conservation is their preservation and restoration.

Artesian Wells

Artesian aquifers are those confined by dense clay or shale such that water flows naturally from an **artesian well** drilled into them. The name comes from the Roman city of Artesum at the site of the French town of Artois known for the free-flowing water from wells dug in the Middle Ages. Artesian wells are often highly prized sources of water. Some modern day bottling companies have leased or purchased sites of artesian wells and advertise their product as artesian water, although it does not differ in quality from pumped well water. Many artesian wells have lost their free-flowing qualities because of depletion of their water resource.

Water Availability

Fresh surface water and groundwater are the major sources for human use. Desalinated seawater is a growing source of fresh water in some areas that do not have access to this valuable resource. Most seawater desalination plants are in the Middle East, an arid region with generally abundant sources of energy that can be used for desalination. The world's largest desalination plant is the Jebel Ali multi-stage flash distillation unit in the United Arab Emirates that can produce 300 million cubic meters of potable water per year. There are significant resources of brackish (saline) groundwater throughout the world. In many cases this water has a lower salt content than seawater making it easier and more economical to desalinate.

A major concern regarding water's availability is its uneven distribution with respect to time and location. Some areas experience tremendous rainfall and flooding during wet (monsoon) seasons with dry conditions in between. Whereas these fluctuations are predictable and compensation may be made for them, greater problems occur with long-term droughts. Regions of Africa are periodically afflicted with droughts that last for several years, killing crops and animals and inflicting great hardship on the people in the region. Evidence from tree rings indicates a pre-Columbian drought of almost three centuries duration in what is now the southwestern US. Uneven geographic distribution of water occurs throughout the world and is illustrated for the continental US in Figure **16.10.3**. It is seen that precipitation is generally adequate in the eastern part of the country, although damaging droughts do occur in this region. However, the western continental US has a shortage of precipitation with essentially permanent drought conditions in some regions including Nevada, Arizona, and southern California. The problem is exacerbated by the popularity of these regions as areas in which people want to live and the demands that they put on limited water resources.

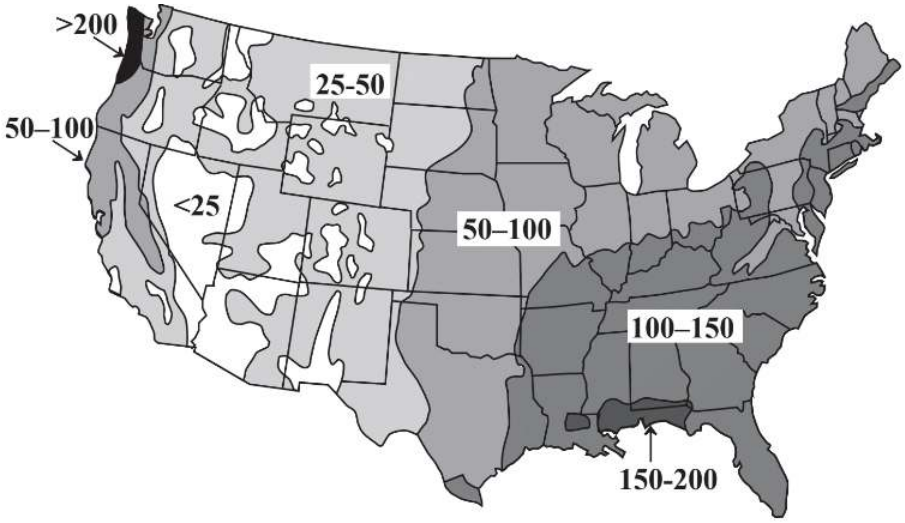


Figure **16.10.3**: Distribution of average annual rainfall in the continental United States, amounts in centimeters per year.

Common Ions Present in Natural Water

As stated above, the fact that water molecules are highly polar allows water to dissolve a variety of ionic compounds. Minerals found in rocks and soil usually dissolve in natural water bodies such as lakes, rivers, springs, and underground waterways (groundwaters). Calcium carbonate, $CaCO_3$, is one of the most common inorganic compounds in the Earth's crust. It is the ingredient for both calcite and aragonite. These two minerals have different crystal structures and appearances (Figure **16.10.4**).



Figure **16.10.4**: Crystals of calcite on dolomite from the Sweetwater Mine in Missouri (left, Rob Lavinsky, iRocks.com, via Wikimedia Commons) and aragonite from the Salsigne Mine in France (right, Didier Descouens, via Wikimedia Commons).

Dust particles and ions present in the air are nucleation centers for water drops. Thus, waters from rain and snow also contain cations such as Ca^{2+}, Mg^{2+}, Na^{+}, K^{+}, and NH_4^{+}. These cations are balanced by anions like HCO_3^{-}, SO_4^{-}, NO_2^{-}, Cl^{-}, and NO_3^{-}. The pH of rain is between 5.5 and 5.6 because of the presence of some of these ions. Rain and snow waters eventually become river or lake waters. When the rain or snow waters fall, they interact with vegetation, top soil, bed rock, river bed and lake bed, dissolving whatever is soluble. Bacteria, algae, and water insects also thrive. The most common ions in lake and river waters are the same as those present in rainwater, but at higher concentrations. The pH of these waters depends on the river bed and lake bed. All of these particles and ions can eventually wind up in the oceans where the ion concentration is generally greater than in fresh water sources. Table **16.10.1** lists the major ions present in seawater. The composition does vary depending on region, depth, latitude, and water temperature. Waters at the river mouths contain less salt. If the ions are utilized by living organism, its contents vary according to the populations of organisms.

Table **16.10.1**: Ions in sea water

Species	Cl^-	Na^+	SO_4^{2-}	Mg^{2+}	Ca^{2+}	K^+	HCO_3^-	Br^-	Sr^{2+}	BO_4^{3-}	F^-	H_4SiO_4	H^+
mg/kg	10,760	2,710	2,710	1,290	411	399	142	67	8	4.5	1.3	0.5-10	10−8.35

Hard Water

Waters containing Ca^{2+} and Mg^{2+} ions are usually called **hard water**. Hard water is high in dissolved minerals and, while other metal cations may be present in the water as well, measuring the concentration of Ca^{2+} and Mg^{2+} in the water will give a useful indication of how much dissolved mineral is present. General guidelines for classification of waters are: 0 to 60 mg/L as calcium carbonate is classified as soft; 61 to 120 mg/L as moderately hard; 121 to 180 mg/L as hard; and more than 180 mg/L as very hard. Water hardness varies throughout the United States (Figures **16.10.5** and **16.10.6**).

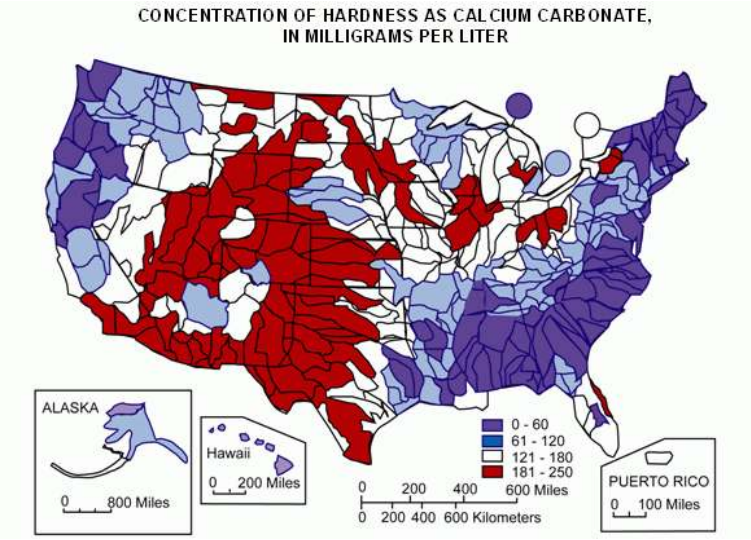


Figure **16.10.5**: Mean hardness as calcium carbonate at National Stream Quality Accounting Network (NASQAN) water-monitoring sites during the 1975 water year and edited in 2005. Colors represent streamflow from the hydrologic-unit area. Although the map illustrates data from 1975, these data have been found to be accurate and useful in current assessments. (Public Domain via usgs.gov)

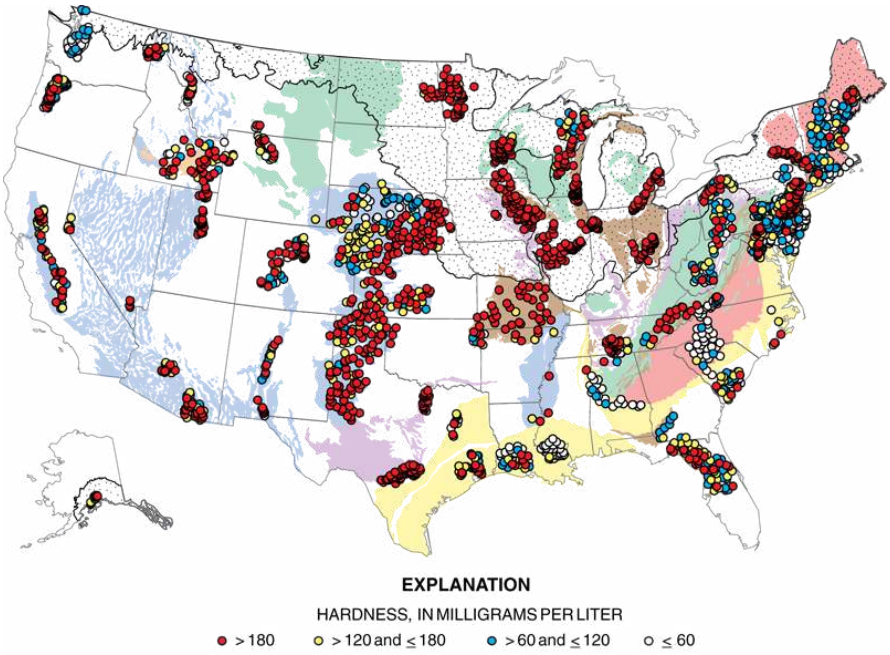


Figure 16.10.6: Results from a study of water hardness for about 2,100 domestic wells across the US by the National Water-Quality Assessment (NAWQA) Project. (Public Domain via usgs.gov)

For those with hard water, the metal cations present can have unpleasant consequences. These cations, especially Ca^{2+}, tend to react with soap to form soap scum. This can be felt as a residue when washing hands and spots or film on glasses after going through a dishwasher. It also means that more soap is usually needed in order to do the same amount of cleaning. In addition, when hard water is heated, solid calcium carbonate can form as scale in water heaters and boilers or as deposits in pipes. Both of these effects lower the efficiency and life span of the equipment.

Summary

- Water's polarity and hydrogen bonding capability allow it to absorb large quantities of heat while being heated and when undergoing phase changes.
- Groundwater is often contained in aquifers between layers of rock. This water is able to flow based on the porosity of the ground and its permeability and can be tapped using wells to provide the water for consumption.
- Surface water found in streams and lakes is also available for use.
- Changing precipitation is causing some areas to have excess water and other areas to experience drought. Communities need to find ways to adapt to water needs to ensure that there is adequate potable water for everyone.
- As water flows through soil, it can dissolve various minerals resulting in a number of different metal cations being dissolved in the water. The most prevalent of these ions are Ca^{2+} and Mg^{2+}. The amount of these ions in the water is referred to as the hardness of the water.

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16.11: Groundwater Contamination

Learning Objective

- Define groundwater contamination.
- Describe the impact of human activities on water quality.
- List the various groundwater contaminants and their sources.

Water is a resource that is required for all life, including human life. In 2010, the UN General Assembly explicitly recognized the human right to water and sanitation. For water to be useful for drinking and irrigation, it must not be polluted beyond certain thresholds. According to the World Health Organization, 2 billion people lived in water-stressed communities without safely managed drinking-water services in 2020. At the same time, over 1.7 billion people still did not have basic sanitation services, which is defined as having access to a public sewage system, septic tank, or even a simple pit latrine. Each year approximately 1.7 million people die from diarrheal diseases associated with unsafe drinking water, inadequate sanitation, and poor hygiene. Almost all of these deaths are in developing countries, and around 90% of them occur among children under the age of 5. Compounding the water crisis is the issue of social justice; poor people more commonly lack clean water and sanitation than wealthy people in similar areas. Globally, improving water safety, sanitation, and hygiene could prevent up to 9% of all disease and 6% of all deaths.

In addition to the global waterborne disease crisis, chemical pollution from agriculture, industry, cities, and mining threatens global water quality. In Gallup public polls conducted over the past decade, Americans consistently put water pollution and water supply as the top environmental concerns over issues such as air pollution, deforestation, species extinction, and global warming.

Groundwater contamination (also called groundwater pollution) occurs when pollutants are released to the ground and make their way down into groundwater. This type of water pollution can also occur naturally due to the presence of a minor and unwanted constituent, contaminant, or impurity in the groundwater, in which case it is more likely referred to as contamination rather than pollution.

How Does Groundwater Become Contaminated?

Water pollution is the contamination of water by an excess amount of a substance that can cause harm to human beings and/or the ecosystem. The level of water pollution depends on the abundance of the pollutant, the ecological impact of the pollutant, and the use of the water. Pollutants are derived from biological, chemical, or physical processes. Although natural processes such as volcanic eruptions or evaporation sometimes can cause water pollution, most pollution is derived from human, land-based activities. Water pollutants can move through different water reservoirs as the water carrying them progresses through stages of the water cycle.

Pollutants enter water supplies from **point sources**, which are readily identifiable and relatively small locations, or **nonpoint sources**, which are large and more diffuse areas (Figure 16.11.1). Point sources of pollution include animal factory farms that raise a large number and high density of livestock such as cows, pigs, and chickens. Also included are pipes from factories or sewage treatment plants. Combined sewer systems that have a single set of underground pipes to collect both sewage and storm water runoff from streets for wastewater treatment can be major point sources of pollutants. During heavy rain, storm water runoff may exceed sewer capacity, causing it to back up and spilling untreated sewage directly into surface waters.





Figure **16.11.1**: Point and nonpoint sources of water pollution. (AnonymousUnknown author, CC0, via Wikimedia Commons and Dwight Burdette, via Wikimedia Commons, respectively)

If the pollutant enters an aquifer, it often creates a contaminant plume within the aquifer. Movement of water and dispersion within the aquifer spreads the pollutant over a wider area. Its advancing boundary, often called a plume edge, can intersect with groundwater wells or move into surface water through seeps and springs, making the water supplies unsafe for humans and wildlife. The movement of the plume, called a plume front, may be analyzed through a hydrological transport model or groundwater model. Analysis of groundwater pollution may focus on soil characteristics and site geology, hydrogeology, hydrology, and the nature of the contaminants.

Types of Contamination

Several kinds of pollutants are detrimental to water that is a source of water supply. A variety of impurities can affect **taste**, **odor**, and **color**. **Sediments** require removal and sometimes special measures are needed to coagulate and settle colloidal suspended matter. Acidity, alkalinity, and salinity at excessive levels are water pollutants.

Several trace metals occur naturally in certain rock formations and can enter in the environment from natural processes such as weathering. However, industrial activities such as mining, metallurgy, solid waste disposal, paint and enamel works, etc. can lead to elevated concentrations of toxic metals including lead, cadmium, and chromium. These contaminants have the potential to make their way into groundwater. The migration of metals (and metalloids) in groundwater will be affected by several factors, in particular by chemical reactions which determine the partitioning of contaminants among different phases and species. Thus, the mobility of metals primarily depends on both the pH of the groundwater and the presence of ions in it with which they can react.

In some cases **radionuclides** occur as water pollutants. Leakage of radioactive tritium, a form of hydrogen, has gotten into water from reactors. The radionuclide of most concern usually is radium from natural sources and a number of groundwater sources of water have been discontinued because of the presence of radium.

Arsenic and fluoride

Arsenic and fluoride have been recognized by the World Health Organization (WHO) as the most serious inorganic contaminants in drinking-water on a worldwide basis. Inorganic arsenic is the most common type of arsenic in soil and water. The metalloid arsenic can occur naturally in groundwater, as seen most frequently in Asia, including in China, India and Bangladesh. In the Ganges Plain of northern India and Bangladesh, severe contamination of groundwater by naturally occurring arsenic affects 25% of water wells in the shallower of two regional aquifers. Groundwater in these areas is also contaminated by the use of arsenic-based pesticides. Arsenic in groundwater can also be present where there are mining operations or mine waste dumps that will leach arsenic.

Natural fluoride in groundwater is of growing concern as deeper groundwater is being used, "with more than 200 million people at risk of drinking water with elevated concentrations." Fluoride can especially be released from acidic volcanic rocks and dispersed volcanic ash when water hardness is low. High levels of fluoride in groundwater is a serious problem in the Argentinean Pampas, Chile, Mexico, India, Pakistan, the East African Rift, and some volcanic islands (Tenerife). In areas that have naturally occurring high levels of fluoride in groundwater which is used for drinking water, both dental and skeletal fluorosis can be prevalent and severe.

Nitrate

Nitrate is the most common chemical contaminant in the world's groundwater and aquifers. In some low-income countries, nitrate levels in groundwater are extremely high, causing significant health problems. It is also stable (it does not degrade) under high oxygen conditions. Elevated nitrate levels in groundwater can be caused by on-site sanitation, sewage sludge disposal and agricultural activities. It can therefore have an urban or agricultural origin.

Nitrate levels above 10 mg/L (10 ppm) in groundwater can cause "blue baby syndrome" (acquired methemoglobinemia). Drinking water quality standards in the European Union stipulate less than 50 mg/L for nitrate in drinking water. However, the linkages between nitrates in drinking water and blue baby syndrome have been disputed in other studies. The syndrome outbreaks might be due to other factors than elevated nitrate concentrations in drinking water.

Biochemical Oxygen Demand

Some pollutants are picked up by water as a natural consequence of its use in municipal water systems, which take in potable water fit to drink and discharge wastewater. One of the main components of this wastewater consists of sewage including human wastes, food wastes, and other substances that get into water through bathrooms and kitchens. A major component of these water pollutants consists of biodegradable organic matter, abbreviated here as $\{CH_2O\}$ and commonly called **biochemical oxygen demand**, **BOD**, which rapidly depletes receiving water of its dissolved oxygen by the following microbially-mediated process:

$$\{CH_2O\} + O_2 \rightarrow CO_2 + H_2O \tag{16.11.1}$$

Oxygen-demanding waste is an extremely important pollutant to ecosystems. Most surface water in contact with the atmosphere has a small amount of dissolved oxygen which is needed by aquatic organisms for cellular respiration. Too much decaying organic matter in water is a pollutant because it removes oxygen from water, which can kill fish, shellfish, and aquatic insects. An unpolluted water body with respect to BOD is a turbulent river that flows through a natural forest. Turbulence continually brings water in contact with the atmosphere where the O_2 content is restored. The dissolved oxygen content in such a river ranges from 10 to 14 ppm O_2, BOD is low, and clean-water fish such as trout thrive. A polluted water body with respect to oxygen is a stagnant deep lake in an urban setting with a combined sewer system. This system favors a high input of dead organic carbon from sewage overflows and limited chance for water circulation and contact with the atmosphere. In such a lake, the dissolved O_2 content is ≤5 ppm O_2, BOD is high, and low O_2-tolerant fish, such as carp and catfish, dominate.

Excessive plant nutrients, particularly nitrogen (NH_4^+, NO_3^-) and phosphorous (as phosphates $H_2PO_4^-$, HPO_4^{2-}), are pollutants closely related to oxygen-demanding waste. Aquatic plants require about 15 nutrients for growth, most of which are plentiful in water. Nitrogen and phosphorus are called **limiting nutrients**, however, because they usually are present in water at low concentrations and therefore restrict the total amount of plant growth. This explains why nitrogen and phosphorus are major

ingredients in most fertilizer. High concentrations of these elements from human sources (mostly agricultural and urban runoff including fertilizer, sewage, and phosphorus-based detergent) can cause cultural **eutrophication**, which leads to the rapid growth of aquatic producers, particularly algae. Thick mats of floating algae or rooted plants lead to a form of water pollution that damages the ecosystem by clogging fish gills and blocking sunlight. A small percentage of algal species produce toxins that can kill animals, including humans. Exponential growths of these algae are called **harmful algal blooms**. When the prolific algal layer dies, it becomes oxygen-demanding waste, which can create very low O_2 concentrations in the water (< 2 ppm O_2), a condition called **hypoxia.** This results in a **dead zone** because it causes death from asphyxiation to organisms that are unable to leave that environment. An estimated 50% of lakes in North America, Europe, and Asia are negatively impacted by cultural eutrophication. In addition, the size and number of marine hypoxic zones have grown dramatically over the past 50 years including a very large dead zone located offshore Louisiana in the Gulf of Mexico. Cultural eutrophication and hypoxia are difficult to combat, because they are caused primarily by nonpoint source pollution, which is difficult to regulate, and nitrogen and phosphorus, which are difficult to remove from wastewater.

Organic compounds

A variety of organic compounds occur as water pollutants. These include industrial chemicals, petroleum products (especially significant in light of the notorious 2010 leak from the Deepwater Horizon oil spill that dumped as much as 5 million gallons of crude oil into the Gulf of Mexico and caused billions of dollars in pollution damage to the Gulf and its shore areas), polychlorinated biphenyls (PCBs, now banned from manufacture but still sometimes encountered, especially in sediments), and dioxin. Polyfluoroalkyls (PFAs) and volatile organic compounds (VOCs) are a dangerous contaminant of groundwater that come from industry. They are generally introduced to the environment through careless industrial practices. Many of these compounds were not known to be harmful until the late 1960s and it was some time before regular testing of groundwater identified these substances in drinking water sources.

Organic pollutants can also be found in groundwater as insecticides and herbicides. As with many other synthetic organic compounds, most pesticides have very complex molecular structures. This complexity determines the water solubility, adsorption capacity, and mobility of pesticides in the groundwater system. Thus, some types of pesticides are more mobile than others so they can more easily reach a drinking-water source.

Pharmaceuticals

Since the 1990s, water contamination by pharmaceuticals has been an environmental issue of concern. The use of pharmaceuticals and personal care products (PPCPs) is on the rise with an estimated increase from 2 billion to 3.9 billion annual prescriptions between 1999 and 2009 in the United States alone. Popular pharmaceuticals such as antibiotics, anti-inflammatories, antidepressants, decongestants, tranquilizers, etc. are normally found in treated wastewater (Figure 16.11.2). This wastewater is discharged from the treatment facility and often makes its way into the aquifer or other source of surface water used for drinking water. Trace amounts of pharmaceuticals in both groundwater and surface water are far below what is considered dangerous or of concern in most areas, but it could be an increasing problem as population grows and more reclaimed wastewater is utilized for municipal water supplies.



Figure **16.11.2** Illicit drugs such as ecstasy (above) can be found in waterways. (Willy turner, Public domain, via Wikimedia Commons)

The major route for pharmaceutical residues to reach the aquatic environment is most probably by excretion from patients undergoing pharma treatment. Since many pharmaceutical substances are not metabolized in the body, they may be excreted in biologically active form, usually via the urine. Furthermore, many pharmaceutical substances are not fully taken up from the intestine (following oral administration in patients) into their blood stream. The fraction not taken up into the blood stream will remain in the gut and eventually be excreted via the feces. Hence, both urine and feces from treated patients contain pharmaceutical residues. Between 30 and 90% of the orally administered dose is generally excreted as active substance in the urine. Figure **16.11.3** illustrates how PPCPs then enter into the aquatic environment through this type of individual human activity.

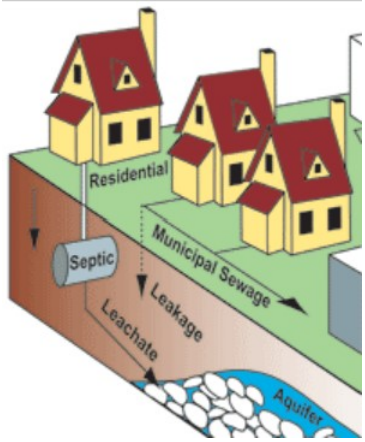


Figure **16.11.3** Methods of PPCP entry into the environment from residential homes via septic and sewage systems.

An additional source to environmental pollution with pharmaceuticals is improper disposal of unused or expired drug residues (pharma residues), either by individual consumers or larger facilities. In 2009, an investigative report by Associated Press concluded that U.S. manufacturers had legally released 271 million pounds of compounds used as drugs into the environment, 92% of which was the industrial chemicals phenol and hydrogen peroxide, which are also used as antiseptics. It could not distinguish between drugs released by manufacturers as opposed to the pharmaceutical industry. It also found that an estimated 250 million pounds of pharmaceuticals and contaminated packaging were discarded by hospitals and long-term care facilities. In Europe, the input of pharmaceutical residues via domestic (individual) waste water is estimated to be around 80%, whereas 20% is coming from hospitals.

Pharmaceuticals used in veterinary medicine, or as additives to animal food, pose a different problem since they are excreted into soil or possibly open surface waters. It is well known that such excretions may affect terrestrial organisms directly, leading to extinction of exposed species (e.g. dung-beetles). Lipid-soluble pharma residues from veterinary use may bind strongly to soil

particles, with little tendency to leak out to ground water or to local surface waters. More water-soluble residues may be washed out with rain or melting snow and reach both ground water and surface water streams.

Once in the water, pharma residues can have diverse, subtle effects on organisms, although research is still limited. Because of the high solubility of most PPCPs, aquatic organisms are especially vulnerable to their effects. The increased presence of estrogen and other synthetic hormones in waste water due to birth control and hormonal therapies has been linked to increased feminization of exposed fish and other aquatic organisms. The chemicals within these PPCP products could either affect the feminization or masculinization of different fishes, therefore affecting their reproductive rates. While the full effects of most PPCPs on the environment are not understood, there is concern about the potential they have for harm because they may act unpredictably when mixed with other chemicals from the environment or concentrate in the food chain. Additionally, some PPCPs are active at very low concentrations and are often released continuously in large or widespread quantities.

Water Borne Diseases

Waterborne diseases are conditions caused by pathogenic micro-organisms that are transmitted in water. These diseases can be spread while bathing, washing, drinking water, or by eating food exposed to contaminated water. While diarrhea and vomiting are the most commonly reported symptoms of waterborne illness, other symptoms can include skin, ear, respiratory, or eye problems. Waterborne diseases are impacted by a country's economy and also impact the economy by being costly to deal with.

The lack of proper sanitation measures, as well as improperly placed wells, can lead to drinking water contaminated with pathogens carried in feces and urine. Such fecal-oral transmitted diseases include typhoid, cholera and diarrhea. Microorganisms causing diseases that characteristically are waterborne prominently include protozoa and bacteria, many of which are intestinal parasites, or invade the tissues or circulatory system through walls of the digestive tract. Various other waterborne diseases are caused by viruses.

Deep, confined aquifers are usually considered the safest source of drinking water with respect to pathogens. Pathogens from treated or untreated wastewater can contaminate certain types of aquifers, especially shallow ones.

The table below shows water-borne diseases that can result from viruses, bacteria, and parasites. In some cases, vaccines are available. When eating, drinking, or swimming, it is important to be aware of how you could be affected by these pathogens. In addition, proper handling of foods and beverages could reduce your risk of developing one or more of the following health problems.

Table **16.11.1** Pathogens That Cause Waterborne Illnesses.

Pathogen Name	Pathogen Type	Source	Health problem	Prevention/Treatment
Giardia	Parasite	Fecal contamination and uncooked food	Vomiting, diarrhea, and cramps	Medication afterward
Cryptosporidium	Parasite	Fecal contamination	Vomiting, diarrhea, fever, and cramps	Medication afterward
Typhoid	Bacteria	Fecal contamination	High fever, stomach pains, headache, and rash	Vaccination/Antibiotics
E. coli	Bacteria	Fecal contamination	Diarrhea and cramps	Fluids
Legionella	Bacteria	Found naturally in heated water	Causes Legionnaires (a type of pneumonia)	Medications afterward
Cholera	Bacteria	Related to fecal contamination or undercooked or raw shellfish	Diarrhea	Vaccine/Rehydration, antibiotics, and Zinc
Hepatitis A	Virus	Contaminated food and water	Vomiting, dark urine, and yellowing of the eyes.	Vaccination/Fluids
Polio	Virus	Fecal contamination	Flu symptoms, paralysis	Vaccination

Figure **16.11.4** shows how a person might contract Giardiasis from *giardia*, a parasite. This particular pathogen can live in a body up to six months. Once detected through a stool sample, a patient can be prescribed specific antibiotics like Flagyl to treat the infection. Unfortunately, there is no vaccine for preventing Giardiasis.

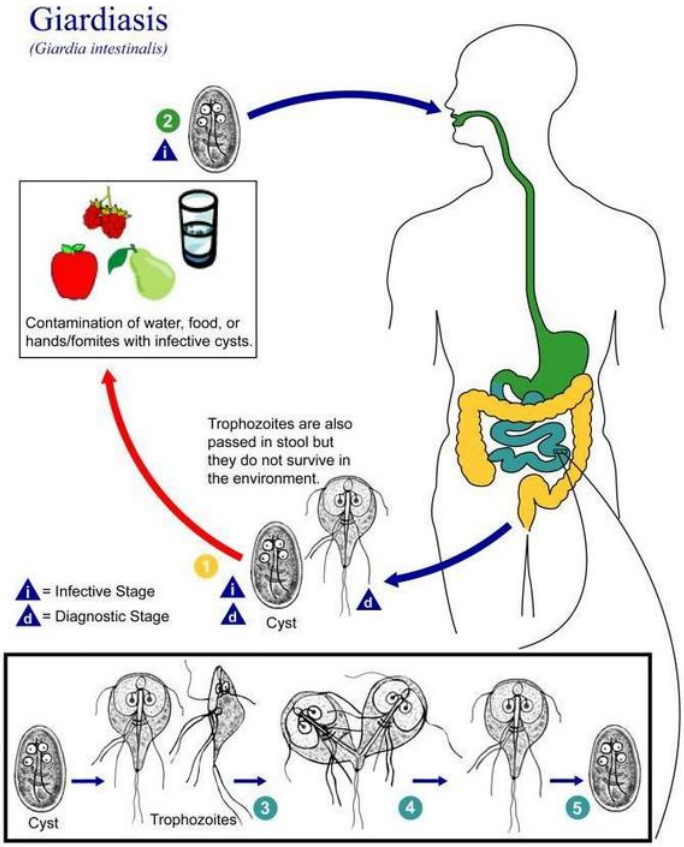


Figure **16.11.4** The life cycle of Giardia.

Summary

- Water pollution is the contamination of water by an excess amount of a substance that can cause harm to human beings and/or the ecosystem.
- Pollutants enter water supplies from point sources, which are readily identifiable and relatively small locations, or nonpoint sources, which are large and more diffuse areas.
- Contaminants found in groundwater cover a broad range of physical, inorganic chemical, organic chemical, bacteriological, and radioactive parameters.
- Organic matter, as well as phosphates and nitrates from human and farm animal waste, support the growth of algae and microorganisms, including bacteria.

- Eutrophication is an increase in the concentration of chemical nutrients in an ecosystem to an extent that increases the primary productivity of the ecosystem. Depending on the degree of eutrophication, subsequent negative environmental effects such as anoxia (oxygen depletion) and severe reductions in water quality may occur, affecting fish and other animal populations.
- A contaminant of growing concern is pharmaceuticals which enter waste water streams either through improper disposal of expired prescriptions or through excrement.
- A variety of diseases can be found in water due to a lack of proper sanitation measures or improper placement of wells.

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16.12: Water Purification

Learning Objectives

- Describe ways in which hard water may be softened.
- Identify steps that can be taken to minimize the release of contaminants into groundwater.
- Describe ways in which contaminants can be removed from municipal water sources.
- Suggest ways to provide potable water for those in areas currently without clean water.

We saw in Section 16.10 that natural water often contains a variety of metal ions that it picks up from dissolving minerals as it makes its way through the ground. These metal ions can affect the taste and color of the water. They can also cause build up in pipes and lessen the effectiveness of soaps. In addition, Section 16.11 described a variety of pollutants which can be present in the water, many of which can have adverse health effects. In order for water to be safe for drinking, it often needs to undergo some sort of purification process. The type of process used depends on the compounds which need to be removed from the water and the amount of water being purified.

Softening Water

In an effort to mitigate the effects of hard water, many choose to treat the water in order to "soften" it. Softening water simply means removing the cations such as Ca^{2+} and Mg^{2+} which tend to precipitate out and replace them with more soluble ions like Na^{+}. There are a number of ways to accomplish this depending on the scale required and the amount one is willing to spend.

For larger scale municipal water systems, **lime-soda softening** can be used to remove Ca^{2+} and Mg^{2+} by adding a calculated amount of hydrated lime, $Ca(OH)_2$, and sodium carbonate, Na_2CO_3. The lime will react with Ca^{2+} ions:

$$\mathrm{Ca^{2+} + 2\,HCO_3^- + Ca(OH)_{2(s)} \rightarrow 2\,CaCO_{3(s)} + 2\,H_2O} \tag{16.12.1}$$

Adding more lime causes the pH of water to increase, and as a result, magnesium ions are removed by the reaction:

$$\mathrm{Mg^{2+} + Ca(OH)_{2(s)} \rightarrow Mg(OH)_{2(s)} + Ca^{2+}} \tag{16.12.2}$$

The extra calcium ions can be removed by the addition of the sodium carbonate, also known as soda ash.

$$\mathrm{Na_2CO_3 \rightarrow 2\,Na^{+} + CO_3^{2-}} \tag{16.12.3}$$

$$\mathrm{Ca^{2+} + CO_3^{2-} \rightarrow CaCO_{3(s)}} \tag{16.12.4}$$

Due to the number of solid products formed, this type of treatment will often have a series of ponds for settling out solids and filters for removing them. These solids can then be used for other purposes such as the formation of cement. The water tends to be quite basic after the process is completed and thus the pH must be lowered after treatment before continuing on to homes and businesses.

Many households will choose to have **water softeners** to soften the municipal water or well water that comes into their homes. Today, most water softeners are using zeolites and employing ion exchange techniques to soften hard water. Zeolites are a group of hydrated crystalline aluminosilicates found in certain volcanic rocks. The tetrahedrally coordinated aluminum and silicon atoms form AlO_4 and SiO_4 tetrahedral groups. They interconnect to each other sharing oxygen atoms, forming cage-type structures as shown in Figure **16.12.1**. Whatever kind is used, the crystal structure of zeolites contains large cages. The cages are connected to each other forming a framework with many cavities and channels. Both positive and negative ions can be trapped in these cavities and channels.

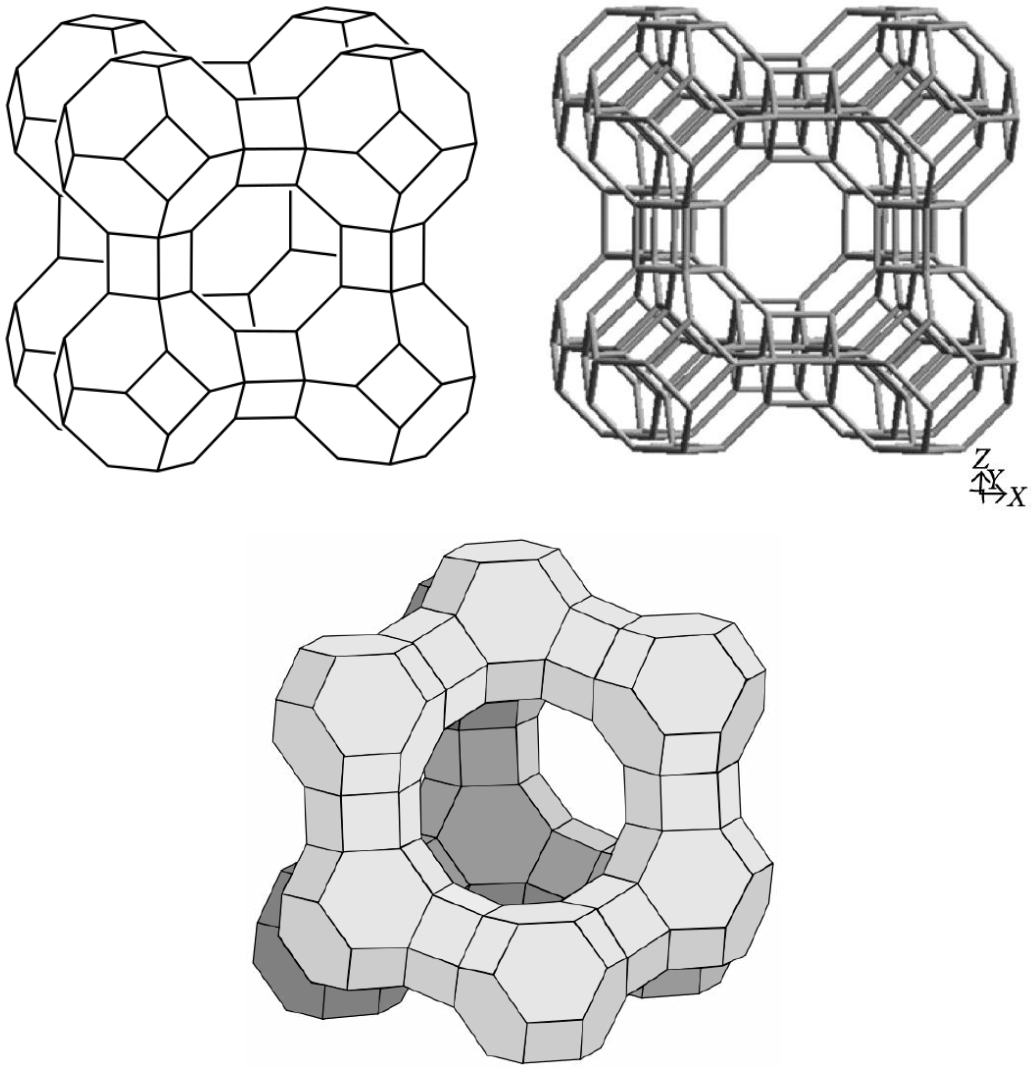


Figure **16.12.1**: Different zeolite cage structures used to trap ions. The top two show the surfaces and bonding of zeolite-A (Nao1958, via Wikimedia Commons and Ángel Berenguer Murcia, via Wikimedia Commons, respectively). The bottom shows the surfaces of faujasite (Roland.chem, CC0, via Wikimedia Commons).

For each oxygen that is not shared in the AlO_4 and SiO_4 tetrahedral groups, a negative charge is left on the group. These negative charges are balanced by trapping alkali metal and alkaline earth metal ions. When more cations are trapped, hydroxide and chloride ions will remain in the cavities and channels of the zeolites.

To prepare a zeolite for water treatment, they are soaked in concentrated NaCl solution. The cavities trap as many sodium ions as they can accommodate. After the treatment, the zeolite is designated as Na-zeolite. Then the salt solution is drained, and the zeolite is washed with water to eliminate the extra salt. When hard water flows through them, calcium and magnesium ions will be trapped by the Na-zeolite. For every Ca^{2+} or Mg^{2+} trapped, two Na^+ ions are released. The treated water contains a rather high concentration of Na^+ ions, but low concentrations of Mg^{2+} and Ca^{2+}. Thus, zeolite ion exchange converts hard water into soft water.

Preventing the Release of Contaminants

One way to keep groundwater free from contaminants is to prevent them from reaching the aquifers. When spills have happened, either the companies responsible or the government will work to contain the pollutant. If it has entered into the soil, the dirt may be dug up and removed to prevent water from flowing through it and leaching out the contaminant. The soil is then treated to remove the waste before being returned to the ground. Spills that occur in the water can have absorbant materials applied to prevent them from spreading through the water system.

Contaminants that enter through wastewater streams can be minimized by decreasing the amounts that are discharged. Proper destruction of pharma residues (unused or expired medications) should yield products without any pharmaceutical or ecotoxic activity. To facilitate this, European countries have take-back systems in place for such residues (although they are not always utilized to full extent) while in the US only voluntary initiatives on a local basis exist. Incineration at a high temperature (>1000 degrees Celsius) is considered to destroy all pharmacological activity, but even following such incineration residual ashes from the incineration should be properly taken care of. Though most of the waste goes to incineration and people are asked to throw unused or expired pharmaceuticals into the household waste, investigations in Germany showed that up to 24% of liquid pharmaceuticals and 7% of tablets or ointments are disposed of via the toilet or sink and may thus wind up in wastewater treatment.

Treatment of Contaminated Groundwater

With the possibility of so many different contaminates in groundwater, it is imperative that it is treated to make it safe for consumption. For many municipal systems, a key step is filtration or reverse osmosis to create purified water and to concentrate the impurities in a waste stream. If the waste stream contains heavy metals, it can then be reacted with specific chemicals to cause the metals to precipitate out and then they can be isolated by sedimentation. If large organic contaminants are present and trapped by the filtration, they can then be incinerated, whether they are VOCs or pharmaceuticals. Pathogens can be broken down or killed using chlorination or ozone at any point along the treatment process.

For those who live in areas without water treatment systems, some sort of home filtration system is necessary. If water cannot be filtered, boiling can kill some pathogens. However, it will not remove things like heavy metals. There are a number of organizations who focus on providing families with filtration systems or communities with deep wells so that the soil can filter out some of the contaminants. This is especially useful in areas where sewage and other waste waters are released close to the source of drinking water.



Video **16.12.1**: The Coalition for Global Community Health is working within existing social structures in Belén, Iquitos, Peru to uphold the human rights of the community members. We speak directly with community members in an open forum to learn about their needs, desires, and ideas for creating an opportunity to change their communities for the better.

Summary

- Water softening exchanges Ca^{2+} and Mg^{2+} ions with Na^{+} ions. The Na^{+} ions are more soluble and therefore mitigate the effects of hard water.
- When pollutants are found in the ground, large scale, expensive treatment is often required in order to prevent the pollutants from entering the groundwater.
- Municipal water treatment facilities use filtration and reverse osmosis to remove contaminants from water that is then distributed to the community.
- Safe water is a right of all people. Many organizations are working hard to provide clean wells and filtration systems for those who live in areas with contaminated water supplies.

Contributors and Attributions

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16.13: Acid Rain

Learning Objective

- Describe how acid rain is formed.
- Describe the effects of acid rain on the environment and infrastructure.

Acid rain is a term referring to a mixture of wet and dry deposition (deposited material) from the atmosphere (Figure **16.13.1**) containing higher than normal amounts of nitric and sulfuric acids. The precursors, or chemical forerunners, of acid rain formation result from both natural sources, such as volcanoes and decaying vegetation, and man-made sources, primarily emissions of sulfur dioxide (SO_2) and nitrogen oxides (NO_x) resulting from fossil fuel combustion. Acid rain occurs when these gases react in the atmosphere with water, oxygen, and other chemicals to form various acidic compounds. The result is a dilute solution of sulfuric acid and nitric acid.

$$2\,SO_2 + O_2 \rightarrow 2\,SO_3 \tag{16.13.1}$$

$$SO_3 + H_2O \rightarrow H_2SO_4 \tag{16.13.2}$$

$$2\,NO_2 + H_2O \rightarrow HNO_2 + HNO_3 \tag{16.13.3}$$

When sulfur dioxide and nitrogen oxides are released from power plants and other sources, *prevailing winds* blow these compounds across state and national borders, sometimes over hundreds of miles. Thus acid rain becomes a problem for everyone and not just those who live close to the sources.

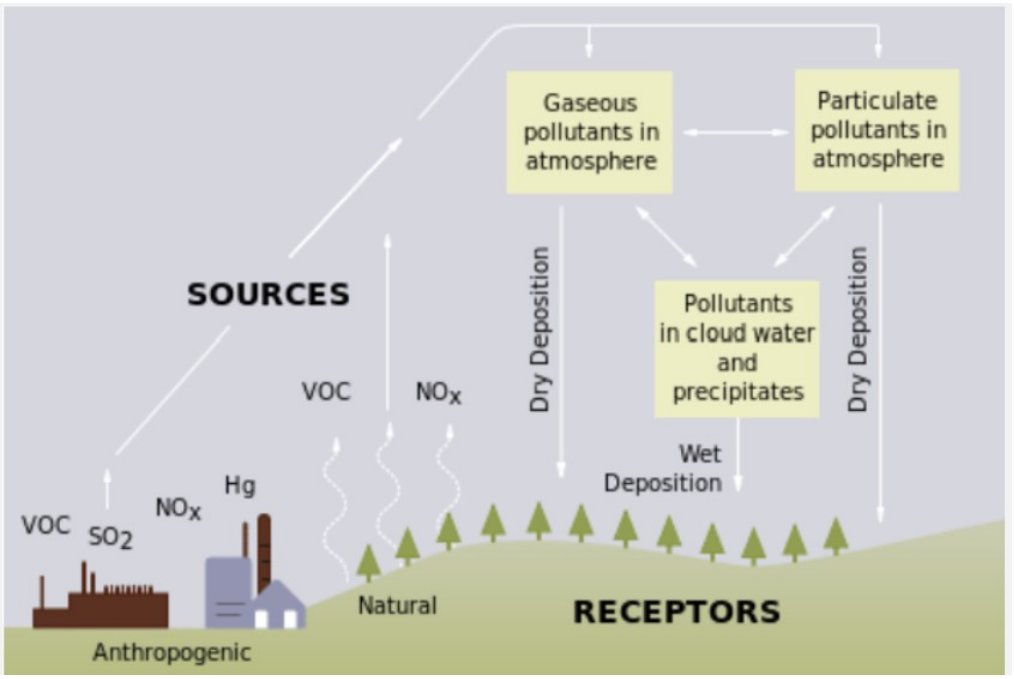


Figure **16.13.1** Processes involved in acid deposition.

Acid rain is measured using the pH scale. As discussed (or will be discussed) in Section 15.7, the lower a substance's pH, the more acidic it is. Pure water has a pH of 7.0. However, normal rain is slightly acidic because carbon dioxide (CO_2) dissolves into it forming dilute carbonic acid solution, giving the resulting mixture a pH of approximately 5.6 at typical atmospheric concentrations of CO_2. Acid rain falling in the US has a pH between 4.2 and 4.4, making it more than ten times more acidic.

Effects of Acid Rain

Acid rain causes **acidification** of lakes and streams and contributes to the damage of trees at high elevations (for example, red spruce trees above 2,000 feet) and many sensitive forest soils. In addition, acid rain accelerates the decay of building materials and

paints, including irreplaceable buildings, statues, and sculptures that are part of our world's cultural heritage. Prior to falling to the earth, sulfur dioxide (SO_2) and nitrogen oxide (NO_x) gases and their particulate matter derivatives—sulfates and nitrates—contribute to visibility degradation and harm public health.

The **ecological** effects of acid rain are most clearly seen in the aquatic, or water, environments, such as streams, lakes, and marshes. Most lakes and streams have a pH between 6 and 8, although some lakes are naturally acidic even without the effects of acid rain. Acid rain primarily affects sensitive bodies of water, which are located in watersheds whose soils have a limited ability to neutralize acidic compounds (called "buffering capacity"). Lakes and streams become acidic when the water itself and its surrounding soil cannot buffer the acid rain enough to neutralize it. In areas where buffering capacity is low, acid rain releases aluminum from soils into lakes and streams; aluminum is highly toxic to many species of aquatic organisms. In addition, different aquatic organisms have different sensitivities to acid. At pH 5, most fish eggs cannot hatch. At lower pH levels, some adult fish die. Even if a species of fish or animal can tolerate moderately acidic water, the animals or plants it eats might not. For example, frogs have a critical pH around 4. However, the mayflies they eat may not survive below pH 5.5.

Acid rain causes slower growth, injury, or death of forests as shown in Figure **16.13.2**. Of course, acid rain is not the only cause of such conditions. Other factors contribute to the overall stress of these areas, including air pollutants, insects, disease, drought, or very cold weather. In most cases, in fact, the impacts of acid rain on trees are due to the combined effects of acid rain and these other environmental stressors.



Figure **16.13.2**. Acid rain damage to a forest in the Czech Republic.

Nature is not the only area harmed by acid rain. Acid rain and the dry deposition of acidic particles contribute to the corrosion of **metals** (such as bronze). The damage that acid rain does to limestone and marble buildings and sculptures is due to a classic acid–base reaction. Marble and limestone both consist of calcium carbonate ($CaCO_3$), a salt derived from the weak acid H_2CO_3. The reaction of a strong acid with a salt of a weak acid goes to completion. Thus we can write the reaction of limestone or marble with dilute sulfuric acid as follows:

$$CaCO_{3(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + H_2O_{(l)} + CO_{2(g)} \tag{16.13.4}$$

Because $CaSO_4$ is sparingly soluble in water, the net result of this reaction is to dissolve the marble or limestone. These effects significantly reduce the societal value of buildings, bridges, and cultural objects such as statues, monuments, and tombstones (Figure **16.13.3**).



Figure 16.13.3. A gargoyle that has been damaged by acid rain.

Sulfates and nitrates that form in the atmosphere from sulfur dioxide (SO_2) and nitrogen oxide (NO_x) emissions contribute to **visibility impairment,** meaning we cannot see as far or as clearly through the air. The pollutants that cause acid rain—SO_2 and NO_x—damage **human health**. These gases interact in the atmosphere to form fine sulfate and nitrate particles that can be transported long distances by winds and inhaled deep into people's lungs. Fine particles can also penetrate indoors. Many scientific studies have identified a relationship between elevated levels of fine particles and increased illness and premature death from heart and lung disorders, such as asthma and bronchitis.

Acid Rain Program

In an effort to deal with the effects of acid rain, the Acid Rain Program (ARP) was established in 1995 as part of the Clean Air Act. It required major emission reductions of SO_2 and NO_x from the power sector. For SO_2, a permanent cap was set at 8.95 million tons, about one-half of the emissions from the power sector in 1980. NO_x reductions were achieved via a more traditional, rate-based regulatory system. Since the program began, the ARP has achieved significant emission reductions for both SO2 and NOx as seen in Figure 16.13.4.

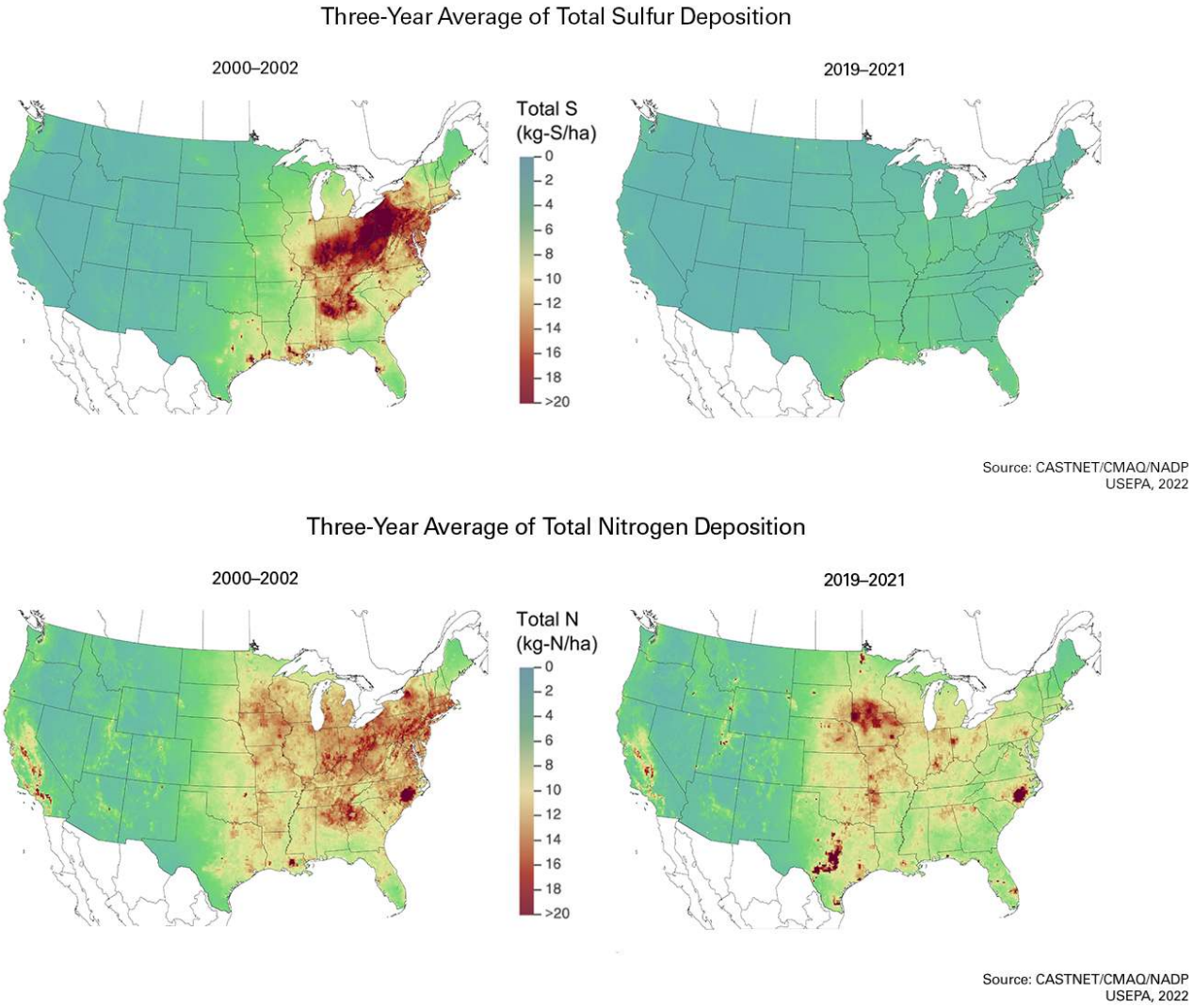


Figure **16.13.4**: Three year averages for total sulfur (top) and nitrogen (bottom) deposition in the US. (US EPA Clean Air Markets Division *Clean Air Status and Trends Network* (CASTNET) via epa.gov)

The ARP was the first national cap and trade program in the country and it introduced a system of allowance trading that uses market-based incentives to reduce pollution. This provides regulated sources with the flexibility to select the most cost-effective approach to reduce emissions and has proven to be a highly effective way to achieve emission reductions and thus attain environmental goals and improve human health.

Some measures can also be taken to mitigate the effects of acid rain, although these are very limited once the pollutant has formed. Some success has been achieved with treating acidified lakes with pulverized limestone to neutralize acid. Corrosion-resistant materials can be used in applications where exposure to acid rain is likely. Protective coatings, such as corrosion-resisting paint primers on metals, can be applied to materials likely to be exposed to acidic precipitation. But the best protection is to prevent formation and release of SO_2 and NO_x gases leading to acid rain formation by measures described earlier in this chapter.

Summary

- Acid rain is a term referring to a mixture of wet and dry deposition (deposited material) from the atmosphere containing higher than normal amounts of nitric and sulfuric acids.
- Acidic rainwater in the soil, streams, lakes, and marshes (and other bodies of water) can be harmful to trees, plants, and animals, especially aquatic plants and animals.
- Acid rain and the dry deposition of acidic particles contribute to the corrosion of metals (such as bronze) and the deterioration of paint and stone (such as marble and limestone).

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16.15: Exercises

Access to and use of the internet is assumed in answering some of the questions. Some questions are designed to promote inquiry and thought rather than just finding material in the text. So in some cases there may be several "right" answers. Therefore, if your answer reflects intellectual effort and a search for information from available sources, your answer can be considered to be "right."

16.1: Chemistry is Good...and Unavoidable

1. What is the definition of green chemistry?

Answer

The practice of chemistry in a manner that maximizes its benefits while eliminating, or at least greatly reducing, its adverse impacts.

16.2: Chemistry and the Environment

2. What is the definition of the environment?

Answer

Our surroundings which affect us and which we may affect.

3. What are the five spheres of the environment? What is each sphere comprised of?

Answer

Atmosphere: atmospheric gases within a few kilometers of sea level

Hydrosphere: Earth's water

Geosphere: all rocks and minerals

Biosphere: all living organisms

Anthrosphere: the part of the environment made and operated by humans

4. What is the definition of environmental chemistry?

Answer

The branch of chemistry that studies the sources, reactions, transport, effects, and fates of chemical species in water, soil, air, and living environments and the effects of technology on them.

5. What are the categories of environmental chemistry that correspond to each of the five spheres of the environment?

Answer

Atmospheric chemistry, aquatic chemistry, geochemistry, environmental biochemistry, and industrial ecology

16.3: The Air We Breathe

6. In what respect are elemental nitrogen and oxygen green elements?

Answer

They are naturally present in the atmosphere in large quantities.

7. Why is there essentially no atmospheric chemistry involving elemental nitrogen gas in the atmosphere?

Answer

Nitrogen is very stable and has a low reactivity. It can absorb excess energy from atmospheric chemical reactions.

8. Cite an atmospheric chemical condition or phenomenon that shows that the O_2 molecule is easier to break apart than the N_2 molecule.

Answer

It will react with sulfur dioxide and pollutant hydrocarbons through a hydroxyl radical. It is involved in making stratospheric ozone. It is consumed in burning hydrocarbons and chemical weathering of minerals.

16.4: Air Pollution

9. What is the definition of air pollution?

Answer

The introduction into the atmosphere of substances that have harmful effects on humans, other living organisms, and the environment.

10. What are two types of pollutants? Where do they come from?

Answer

Primary pollutants are released directly from the source into the air. Secondary pollutants are produced through reactions between primary polluatants and normal atmospheric compounds.

11. What are the two types of standards used to set limits for air pollutants? What is each of them based on?

Answer

Primary standards based on human health and secondary standards based on environmental and property damage

12. What are the six criteria pollutants monitored by the EPA?

Answer

Carbon monoxide, ground-level ozone, nitrogen dioxide, sulfur dioxide, lead, and particulate matter

13. Which of the six criteria pollutants is a secondary pollutant?

Answer

Ground-level ozone

14. Suggest why lead has become less of a problem as an atmospheric pollutant in recent years.

Answer

It is no longer allowed to be in gasoline.

15. What are two types of particulate material?

Answer

Coarse particles (2.5 to 10 micrometers) and fine particles (up to 2.5 micrometers)

16. Why are very small particles especially dangerous to breathe?

Answer

They are small enough to pass through the throat and nose and enter the lungs and possibly the blood.

17. What are some natural and human-made sources of volatile organic compounds?

Answer

Natural sources include plants and gut bacteria. Human-made sources paints, varnishes, cleaning agents, and fuels.

18. What is the major health effect of carbon monoxide?

Answer

It reduces the amount of oxygen that reaches the body's organs and tissues.

19. What are two classifications of outdoor pollutants? How do they differ?

Answer

Mobile and stationary sources. They differ in mobility. Mobile sources move from place to place such as cars. Stationary sources are in a fixed location such as a power plant.

20. What are two types of stationary sources? How do they differ?

Answer

Point sources and nonpoint sources. They differ in size. Point sources are larger emitters such as an industrial facility. Nonpoint sources are abundant and small such as individual homes in a community.

21. What is exhaust gas? What types of pollutants does it contain?

Answer

Exhaust gas is emitted as a result of combustion of fuels. As such, it will contain CO, NO_x, and possibly SO_2. It can include unreacted fuel vapors or particulate matter.

22. Based on Figure 16.4.3, rank the different pollutants emitted in 2017 from greatest to least.

Answer

CO > VOC > NO_x > NH_3 > PM_{10} > SO_2 > $PM_{2.5}$

23. Based on Figure 16.4.3, what is the largest anthropogenic source of each pollutant?

Answer

CO: highway vehicles

NH_3: industrial and other processes

NO_x: stationary fuel combustion and highway vehicles

$PM_{2.5}$: stationary fuel combustion and industrial and other processes

PM_{10}: industrial and other processes

SO_2: stationary fuel combustion

VOC: industrial and other processes

24. What is an example of a stationary fuel combustion source?

Answer

Coal-burning power plant

16.5: Photochemical Smog

25. What are the ingredients and conditions leading to the formation of photochemical smog?

Answer

The ingredients are nitrogen oxides from combustion products, oxygen gas in the atmosphere, and VOCs. The conditions necessary are dry, stagnant air that allows the gases to accumulate and provides UV light to initiate the reactions.

26. What is an important health effect of nitrogen dioxide? Why is it particularly important in atmospheric chemistry?

Answer

It aggravates respiratory diseases. It is important in atmospheric chemistry as it is a precursor and intermediate in the formation of photochemical smog.

27. What substances are found in a smoggy atmosphere?

Answer

Ozone in the main component. Oxygen-rich organic compounds. Nitrogen oxides. VOCs.

28. What are some harmful effects of smog?

Answer

Difficulty in breathing and irritation to eyes and respiratory tract. Damage to plants and materials attacked by oxidants. The formation of visibility-reducing atmospheric particles.

29. What is a serious air pollution phenomenon resulting from an atmospheric reaction of sulfur dioxide?

Answer

Industrial smog

30. What are the major constituents of industrial smog?

Answer

Soot, fly ash, sulfur dioxide, sodium chloride, and calcium sulfate particles

31. What is a secondary environmental problem that can arise from industrial smog?

Answer

Formation of acid rain

32. What is a method used to separate sulfur dioxide from furnace stack gas?

Answer

Scrubbers that react SO_2 with CaO to produce a solid precipitate that is collected

16.6: Clean Air

33. What are two major classes of atmospheric particles, or aerosols, based upon how they are produced?

Answer

Dispersion aerosols formed by grinding solids and dispersing dusts; condensation aerosols formed by the condensation of gases or vapors

34. List five means of controlling particle emissions.

Answer

Sedimentation, inertial mechanisms, baghouses, scrubbers, and electrostatic precipitators

35. What have been some of the effects of the Clean Air Acts on pollutant levels? What have been some benefits experienced in the US over the same time? You may wish to reference Figure 16.6.2 and Figure 16.6.3.

Answer

Answers can include percent declines for different pollutants, steady energy consumption, and increases in vehicular travel and GDP.

36. What are some of the ways that green chemistry can help prevent air pollution?

Answer

Answers can include scrubbers to eliminate SO_2 emissions, catalytic converters to minimize nitrogen oxides, limiting release of VOCs, reformulating gasoline to allow for better combustion, etc.

16.7: Global Warming

37. What is the definition of global warming?

Answer

The increase in the average temperature of the Earth's atmosphere due to elevated greenhouse gas concentrations.

38. What is the definition of climate change? How does it differ from global warming?

Answer

It includes global warming and the resulting large-scale shifts in weather patterns. It differs in that it looks at the effects of the increased temperatures on other areas of the Earth (like the weather).

39. What chemical species in the stratosphere is essential for life on Earth? Why is it important?

Answer

Ozone. It blocks UV light which would cause severe damage to biological species.

40. In what respect does the radiation by which Earth loses energy differ from that by which it gets energy from the sun?

Answer

The Earth gets it energy from the sun as UV and visible light. It radiates it back to the atmosphere as infrared radiation (heat).

41. What is the definition of albedo? What types of surfaces have beneficial albedos?

Answer

Albedo is the percentage of radiation that is reflected off of the Earth's surface and thus is less effective in warming the surface. Beneficial albedos have high percentages and lighter colors such as snow grass. Poor albedos have lower percentages and are darker colors such as dirt and asphalt.

42. What are two catastrophic events that could cause a sudden cooling of Earth’s atmosphere?

Answer

Volcanic eruptions and asteroid hits

43. In what respect is atmospheric carbon dioxide essential to life on Earth? Why may it end up being the “ultimate air pollutant”?

Answer

It is needed to provide the blanket which allows the Earth to be warm enough for life to exist. However, it is being produced in levels that is leading to temperature levels that will upset the balance of processes on Earth.

44. What is a greenhouse gas other than carbon dioxide that is produced by microorganisms?

Answer

Methane, and to a lesser extent nitrous oxide.

45. Rank the greenhouse gases from the greatest global warming potential to least.

Answer

Fluorinated gases (SF_6 > HFCs > PFCs) > N_2O > CH_4 > CO_2

46. Based on the information in Figure 16.7.5, which areas of the US have had the greatest rate of temperature change over the last century? Why do you think that is? Why has the US warmed at a faster rate than the global rate since the late 1970s?

Answer

Areas of greatest rate include the west, Florida, the Northeast and upper Midwest. Reasons will vary but could include increased manufacturing and weather patterns.

16.8: Climate Change: Effects of Global Warming

47. Distinguish between the terms weather and climate.

Answer

Weather is short term variations in the state of the atmosphere as seen in temperatures, cloud cover, precipitation, humidity, atmospheric pressure, and wind. Climate looks at long-term trends over the whole of Earth's surface.

48. How is water vapor involved in moving energy through the atmosphere?

Answer

Water absorbs significant amounts of heat energy as it evaporates. This gaseous water can then move the energy to other parts of the Earth and condense to form rain. As it condenses, it releases the energy as heat to warm the area where the rain falls.

49. What are some of the more harmful effects projected if global warming occurs to a significant extent?

Answer

Increased temperatures in many areas, changes in precipitation which can affect water availability, rising sea levels, and ocean acidification.

50. Based on the information in Figure 16.8.1, which season is most responsible for the increase in temperature in the upper Midwest? How is this manifested during the season?

Answer

Winter has the biggest increase in temperature. This is seen in fewer record lows, later fall frost and earlier last spring frost.

51. Based on the information in Figure 16.8.2, which areas of the US are expected to have an increase in precipitation? Which are expected to have a decrease in precipitation? What types of challenges could these changes cause economically?

Answer

Increases in precipitation are expected on the East Coast. Decreases are expected in the South, the lower Midwest, and the West Coast. Increased precipitation can affect manufacturing and infrastructure. Decreased precipitation can affect water availability, both for drinking and for agriculture which are prevalent in the lower Midwest and California.

52. What are two factors related to global warming which are causing the sea levels to rise?

Answer

Increased temperatures are causing land ice to melt and runoff to oceans. Increased temperatures also cause water to expand and take up a greater volume.

53. Based on the information in Figure 16.8.6, what is the main type of land that has been affected by sea level rise? Why is the loss of this type of land particularly harmful in the face of changing weather patterns?

Answer

Wetlands. Wetlands provide buffers to flooding which results from increased precipitation or hurricanes. They are also a habitat for many species of plants and animals.

16.9: Climate Mitigation Strategies

54. How do adaptation strategies work?

Answer

They provide ways to deal with changing environments. Examples would be conservation of water or improved methods of water purification.

55. How can green chemistry be used to implement adaptation strategies for dealing with climate change?

Answer

Answers can vary but may include development of fertilizers and crops that will grow under drought conditions or with saline water.

56. What contributed to the increased use of coal in the early 1800s?

Answer

The invention of the steam engine.

57. Describe a combined power cycle. How may it be tied with district heating?

Answer

A combined power cycle uses waste heat energy from a primary electricity generator (like coal or gas combustion) to generate additional electricity by heating water and using the steam in a steam turbine. The steam can then be used to heat buildings in the vicinity before returning to the boiler for reheating.

58. What are two major problems with reliance upon coal and petroleum for energy?

Answer

They generate substantial amounts of greenhouse gases and are not renewable energy sources.

59. Cite two examples of vastly increased efficiency of energy utilization that took place during the 1900s.

Answer

Increased automobile fuel economy as a result of the energy crisis in the 1970s and improved efficiency of engines in general. Combined power cycles.

60. What is the biggest problem with nuclear energy?

Answer

High-level nuclear wastes which last for centuries

61. Describe two ways to use solar energy.

Answer

Heating and converting to electricity with photovoltaic cells

62. Using internet resources for information, list some possible means for storing energy generated from solar radiation?

Answer

Answers will vary but may include a variety of storage batteries.

63. In what respects is wind both one of the oldest, as well as one of the newest, sources of energy?

Answer

Wind has been used for centuries for sailing and windmills. They are new based on technologically advanced wind turbines.

64. Why does the burning of biomass not contribute to an increase in greenhouse gases?

Answer

The carbon dioxide released during the burning is the same carbon dioxide that was captured during the biomass growth. Thus, there is a net zero change in the amount of carbon dioxide over time.

65. What are concerns in growing biomass for fuel? How can these concerns be addressed?

Answer

Not enough can be grown and its use as fuel will mean a food shortage. Only a small amount of the widely grown grain crops goes to food. A lot of the waste can be used for energy production. Biomass can be used which is not a food source. Cultivation of less than 10% of the land area of the US would be sufficient to provide for energy needs.

66. What are benefits of using biomass for energy and electricity generation?

Answer

Burning does not lead to global warming. Can use perennial plants to prevent erosion.

67. How can technology and green chemistry be used to mitigate climate change?

Answer

Carbon dioxide can be injected deep into the ground to capture it and prevent it from entering the atmosphere. Unreacted coal can be converted to carbon dioxide (which can be sequestered) and hydrogen gas which is a clean fuel. New compounds can be developed which do not lead to greenhouse gases.

68. Go to the EPA's Carbon Footprint Calculator. What is your carbon footprint? What are ways that you can minimize this footprint?

Answer

Answers will vary.

16.10: Water

69. What are some special properties of water related to its polarity and hydrogen bonding?

Answer

High heat capacity, high latent heat of fusion, high heat of vaporization, excellent solvent for ionic compounds, and high surface tension

70. Where is most of the freshwater on the Earth located?

Answer

Snowpack, ice and glaciers

71. What is an aquifer?

Answer

A body of rock which contains appreciable quantities of water

72. What are desirable properties of an aquifer that is used for water wells?

Answer

High permeability and high porosity

73. What is a watershed?

Answer

Areas of land upon which precipitation falls to provide flows of surface water

74. What is desalination? What are benefits and drawbacks to using desalination to provide potable water?

Answer

Desalination is a means of removing salt from sea water to make it potable. It is often done by reverse osmosis. Benefits are that it provides potable water, especially in areas near seas and oceans. Drawbacks are the expense of purification.

75. What are some minerals which are present in natural water?

Answer

Calcium carbonate from calcite and aragonite, magnesium salts

76. What makes water hard?

Answer

The presence of dissolved minerals resulting in the presence of metal cations, especially Ca^{2+} and Mg^{2+}.

77. What are some drawbacks of using hard water?

Answer

Minerals present in hard water can affect color and taste. Mineral deposits in pipes and heaters, lowering efficiency. Soap is less effective and scum can form.

78. Look up proposals to restore the Hetch-Hetchy Valley in Yosemite National Park to its former state. How might this affect water supply to parts of California? What might be some benefits of restoration of this valley to its former state?

Answer

Answers will vary.

79. Look up and explain the significance of the name Mulholland in relationship to water. How did Mulholland affect history?

Answer

Mulholland was a civil engineer who designed the water system for LA. Answers will vary.

16.11: Groundwater Contamination

80. According to the WHO, how many people lived in water-stressed communities in 2020?

Answer

2 billion

81. What is water pollution?

Answer

The contamination of water by an excess amount of a substance that can cause harm to human beings and/or the ecosystem.

82. What are examples of point sources of water pollution? What are examples of nonpoint sources of water pollution?

Answer

Point sources are small locations such as farms or pipes from factories or sewage treatment plants. Nonpoint sources are large and more diffuse areas such as agricultural runoff.

83. What are sources of biochemical oxygen demand as pollutants?

Answer

Sewage, food wastes, and fertilizer runoff

84. How does BOD act as a pollutant?

Answer

BOD removes oxygen from the water which kills fish and aquatic insects.

85. Why do some cities have fountains in the middle of ponds or small lakes?

Answer

The fountains agitate the water allowing for more contact with oxygen in the air and increasing its concentration in the water.

86. Tests can be performed on water that show the presence of biochemical demand (BOD) and other tests that react with organic matter to produce CO_2 can show total organic carbon (TOC). Applied to a particular sample of water, these two tests showed

relatively high TOC and relatively low BOD. What does this say about the nature of the organic pollutants in the water?

Answer

The organic pollutants are probably industrial chemicals which are not easily broken down by microbes utilizing oxygen.

87. Agricultural fertilizer normally adds nitrogen, phosphorus, and potassium to soil. Explain how fertilizer runoff into a body of water can lead eventually to increased biochemical oxygen demand pollution.

Answer

The fertilizer causes rapid growth of aquatic plants like algae. When the algae dies, it becomes a source of oxygen-demanding waste.

88. How do pharmaceuticals and personal care products enter groundwater and surface water?

Answer

They are excreted by patients and enter the waste water which then leaks into the groundwater. Some pharmaceuticals are improperly disposed of by manufacturers and consumers and it will runoff into groundwater.

89. Why are pharmaceuticals in groundwater an area of concern?

Answer

Low levels may affect the life cycles and reproductive cycles of various fish and aquatic organisms. Pollutants can become concentrated in larger species as various smaller ones are consumed.

90. By doing some search on the internet, gather information regarding the use of wastewater for irrigation. Is this a practice that is used and if so where does it usually take place? What are some of the benefits? What are some of the risks?

Answer

Answers will vary.

16.12: Water Purification

91. How does water softening work? What are ways it can be accomplished on large and small scales?

Answer

Water softening replaces "hard" ions such as Ca^{2+} and Mg^{2+} with "soft" ions like Na^{+}. The Na^{+} does not precipitate out of the water as easily and so the effects of hard water are reduced. Large scale softening can be accomplished with lime-soda softening. Small scale softening is done with water softeners.

92. Membrane filtration processes can be very effective in removing residual BOD from secondary wastewater effluent. What does this suggest regarding the nature of contaminants responsible for the BOD?

Answer

The contaminants are large enough to be contained by the membranes. Thus they are suspended solids and not dissolved solids.

93. Phosphate in the form of $H_2PO_4^-$ and HPO_4^{2-} ions is the substance usually removed from secondary sewage effluent to prevent excessive algal growth and eutrophication in receiving waters. Of several possible algal nutrients, why is phosphate chosen? Show with a chemical reaction the most common means of removal.

Answer

Phosphates tend to make insoluble solids. Thus it is relatively easy to precipitate the various phosphates out of the waste water.

16.13: Acid Rain

94. What is acid rain?

Answer

A mixture of wet and dry deposited material from the atmosphere with higher than normal amounts of nitric and sulfuric acids.

95. What is the typical pH of acid rain in the US?

Answer

Between 4.2 and 4.4.

96. What are five adverse effects from acid rain and its precursors?

Answer

Acidification of lakes and streams

Damage of trees and sensitive forest soils

Accelerates the decay of building materials and paints

Visibility degradation

Harmful to public health

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Thumbnail Chapter 17: Bar graph. (Алекс Арцибашев via Unsplash)

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17.1: Periodic Tables

Versions of Periodic Table below:

- PNG Image Version (Figure 17.1.1)
- Downloadable PDF Version (Figure 17.1.2)
- Interactive Periodic Table from PubChem (Figure 17.1.3)

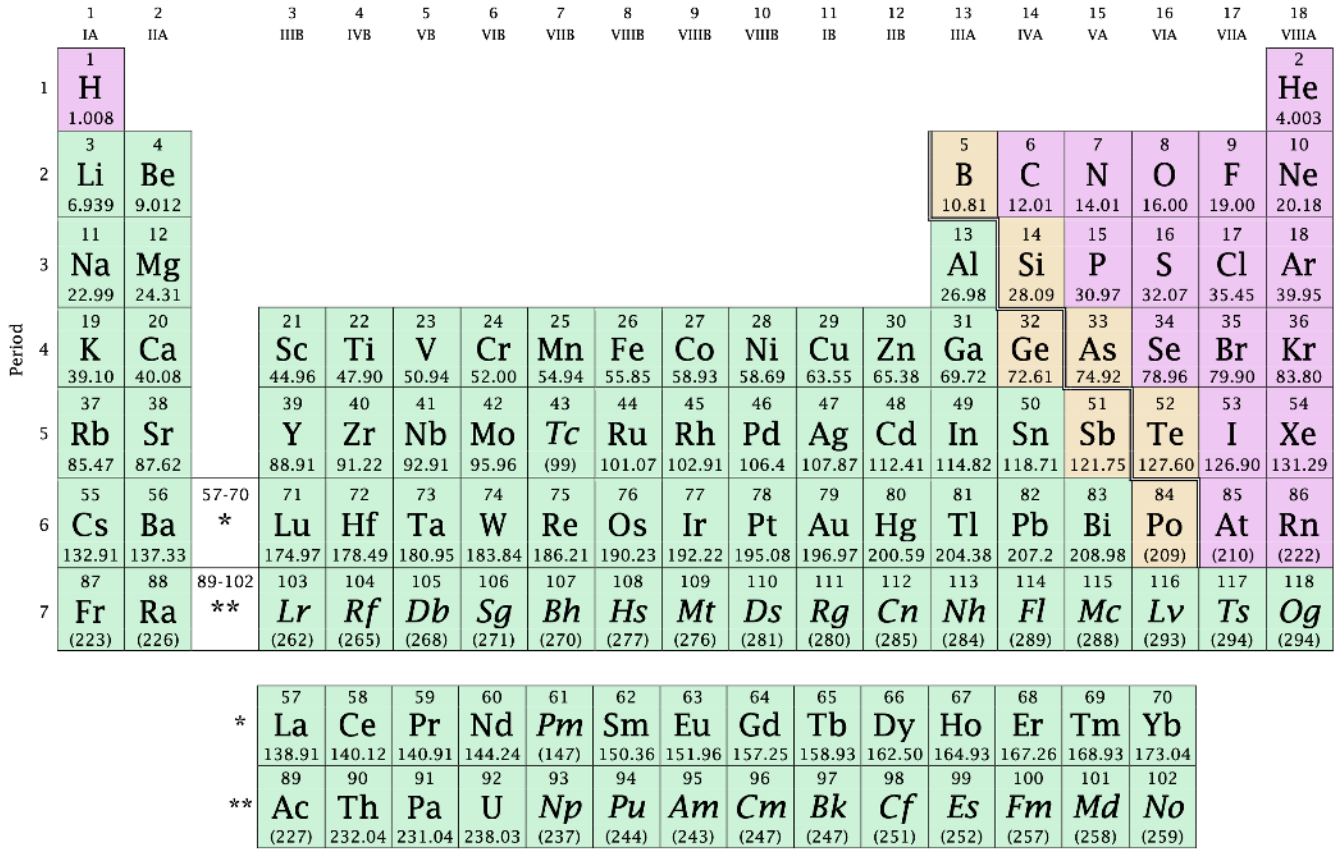


Figure **17.1.1**: PNG Version of the Periodic Table.

Figure 17.1.2: Downloadable PDF Version of the Periodic Table.

Periodic Table of Elements

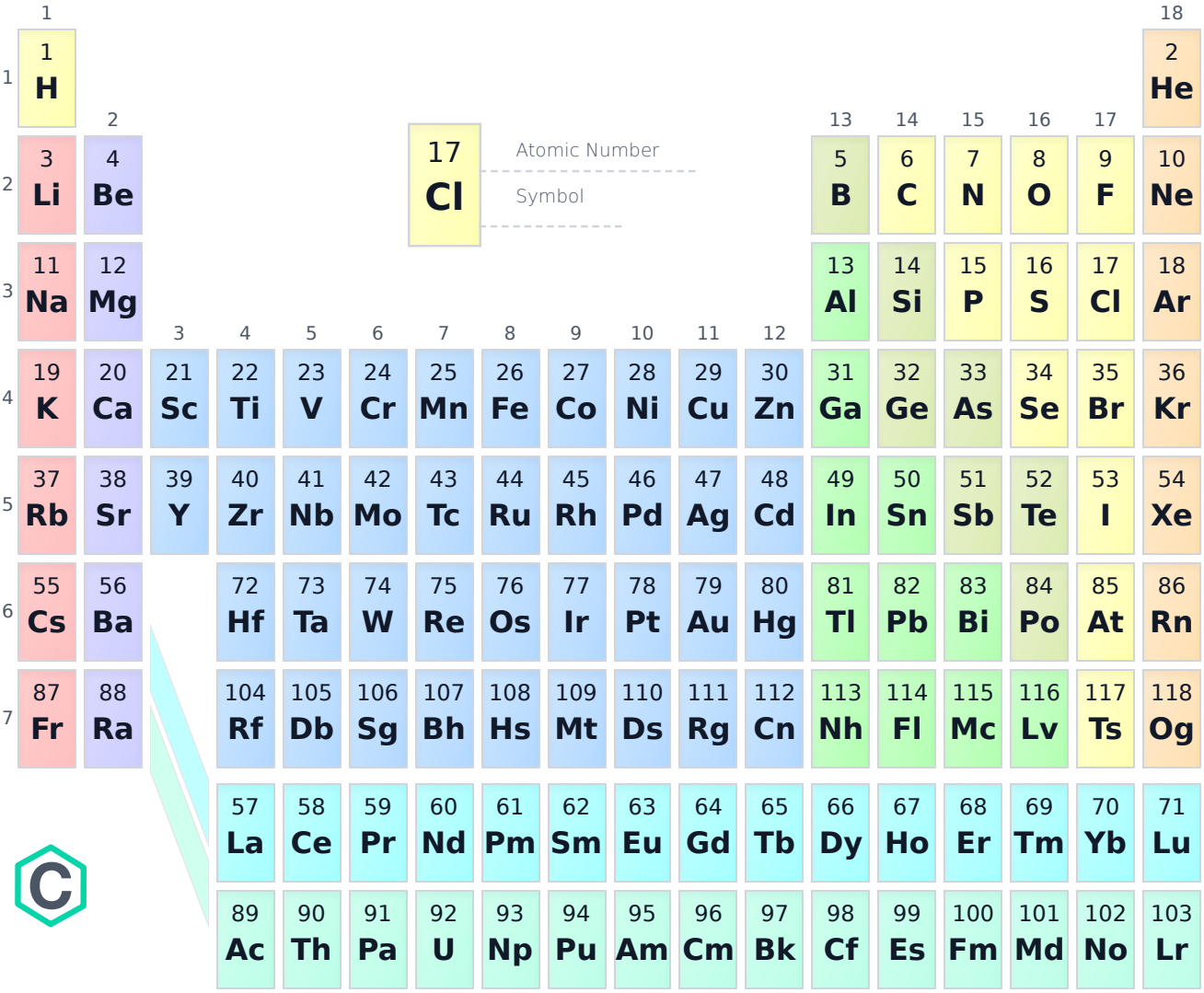


Figure 17.1.3: An interactive modern periodic table. The displayed property may be changed by using the dropdown menu to the left of helium, He. (Public Domain; PubChem via NIH)

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17.2: Conversions and Constants

Versions of Conversion Tables below:

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Common Chemistry Conversions

English to Metric Conversions

The mass, length, volume, and area conversions are given to 4 significant figures. The temperature conversions are exact.

Mass	Length	Volume	Area	Temperature
1 lb = 453.6 g	1 in. = 2.540 cm	1 fl oz = 29.57 mL	1 $in.^2$ = $(2.54\ cm)^2$ = 6.452 cm^2	$T_{°C} = \frac{5}{9}(T_{°F} - 32)$
1 oz = 28.35 g	1 ft = 30.48 cm	1 L = 1.057 qt	1 m^2 = $(3.281\ ft)^2$ = 10.76 ft^2	$T_{°F} = \frac{9}{5}T_{°C} + 32$
1 kg = 2.205 lb	1 m = 3.281 ft	1 gal = 3.785 L		$T_K = T_{°C} + 273.15$
1 metric ton = 1000 kg	1 mi = 1.609 km	1 in^3 = $(2.54\ cm)^3$ = 16.39 cm^3		

English to English Conversions. *All of these conversions are exact.*

Mass	Length	Volume	Area
1 lb = 16 oz	1 ft = 12 in.	1 cup = 8 fl oz	1 ft^2= $(12\ in)^2$ = 144 $in.^2$
1 ton = 2000 lb	1 yd = 3 ft	1 pt = 2 cups	1 mi^2 = 640 acres
	1 mi = 5280 ft	1 qt = 2 pt	
		1 gal = 4 qt	

Other Conversions

Energy	Pressure
1 cal = 4.184 J	1 atm = 760 mm Hg = 760 torr = 29.92 in. Hg
1 J = 1 $\frac{kg \cdot m^2}{s^2}$	1 atm = 14.7 psi = 101,325 Pa = 1.01325 bars
	1 Pa = 1 $\frac{kg}{m \cdot s^2}$

Constants

speed of light (in a vacuum)	$c = 2.998 \times 10^8$ m/s
Planck's constant	$h = 6.626 \times 10^{-34}$ J·s
electron mass proton mass neutron mass	$m_e = 9.109 \times 10^{-31}$ kg $m_p = 1.673 \times 10^{-27}$ kg $m_n = 1.675 \times 10^{-27}$ kg
Avogadro's number	$N_A = 6.0221367 \times 10^{23}$ particles/mol
Gas Constant	$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $= 8.315\ \frac{J}{mol \cdot K}$ $= 8.315\ \frac{kPa \cdot dm^3}{mol \cdot K}$
Faraday Constant	$F = 9.65 \times 10^4$ C/mol
Electronic charge	$e = 1.602 \times 10^{-19}$ C

Figure **17.2.1**: Common Chemistry Conversions. (Lance S. Lund)

Figure 17.2.2: Common Chemistry Conversions. (Lance S. Lund)

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