CHEM 110

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CHEM 110: Chemical Concepts

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TABLE OF CONTENTS

10: Radiation, Electromagnetic Waves, Biological Effects of UV Radiation

11: Climate and The Environment

Index

Glossary

Detailed Licensing

Detailed Licensing

Detailed Licensing

Licensing

A detailed breakdown of this resource's licensing can be found in ***Back Matter/Detailed Licensing****.*

CHAPTER OVERVIEW

1: The Chemical World

1.1: Introduction

Your alarm goes off and, after hitting “snooze” once or twice, you pry yourself out of bed. You make a cup of coffee to help you get going, and then you shower, get dressed, eat breakfast, and check your phone for messages. On your way to school, you stop to fill your car’s gas tank, almost making you late for the first day of chemistry class. As you find a seat in the classroom, you read the question projected on the screen: “Welcome to class! Why should we study chemistry?”

Figure 1.1.1: Chemical substances and processes are essential for our existence, providing sustenance, keeping us clean and healthy, fabricating electronic devices, enabling transportation, and much more. (credit “left”: modification of work by “vxla”/Flickr; credit “left middle”: modification of work by “the Italian voice”/Flickr; credit “right middle”: modification of work by Jason Trim; credit “right”: modification of work by “gosheshe”/Flickr)

Do you have an answer? You may be studying chemistry because it fulfills an academic requirement, but if you consider your daily activities, you might find chemistry interesting for other reasons. Most everything you do and encounter during your day involves chemistry. Making coffee, cooking eggs, and toasting bread involve chemistry. The products you use—like soap and shampoo, the fabrics you wear, the electronics that keep you connected to your world, the gasoline that propels your car—all of these and more involve chemical substances and processes. Whether you are aware or not, chemistry is part of your everyday world. In this course, you will learn many of the essential principles underlying the chemistry of modern-day life.

1.2: Scientific Method

Learning Objectives

- Identify the components of the scientific method.

Scientists search for answers to questions and solutions to problems by using a procedure called the scientific method. This procedure consists of making observations, formulating hypotheses, and designing experiments; which leads to additional observations, hypotheses, and experiments in repeated cycles (Figure $1.2.1$).

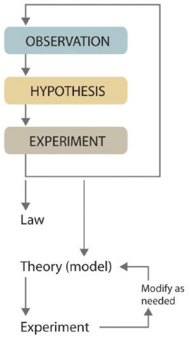

Figure $1.2.1$: The Steps in the Scientific Method.

Step 1: Make observations

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

Step 2: Formulate a hypothesis

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

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

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

Step 3: Design and perform experiments

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

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether or not the original hypothesis is valid. If the hypothesis is valid, the scientist can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified and requires further experimentation.

Step 5: Development into a law and/or theory

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

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

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

Because scientists can enter the cycle shown in Figure $1.2.1$ at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

Video: The Scientific Method by Teacher's Pet

Example 1.2.1

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

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

Solution

a. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.
b. This is a possible explanation for the origin of birds, so it is a hypothesis.
c. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.

d. The temperature is measured before and after a change is made in a system, so these are observations.
e. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

? Exercise 1.2.1

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

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

Answer a

experiment

Answer b

law

Answer c

theory

Answer d

hypothesis

Answer e

observation

Summary

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

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

Contributions & Attributions

- Wikipedia

1.2.1: Reading- Understanding Science

An Overview

To understand what science is, just look around you. What do you see? Perhaps your hand on the mouse, a computer screen, papers, ballpoint pens, the family cat, the sun shining through the window …. Science is, in one sense, our knowledge of all that — all the *stuff* that is in the universe, including the tiniest subatomic particles in a single atom of the metal in your computer's circuits, the nuclear reactions that formed the immense ball of gas that is our sun, and the complex chemical interactions and electrical fluctuations within your own body that allow you to read and understand these words. But science is not just a collection of knowledge. Just as importantly, science is also a reliable *process* by which we learn about all that stuff in the universe. And science is different from many other ways of learning because of the way it is done. Science relies on testing ideas with evidence gathered from the natural world. This website will help you learn more about science as a process of learning about the natural world and access the parts of science that affect your life.

From microbiology to microchips, it's all science. Photo credits: Frank Fox via Wikimedia, Flickr user NASA Goddard Space Flight Center, and Flickr user Nikki Pugh.

Science helps to satisfy the natural curiosity with which we are all born: Why is the sky blue? How did the leopard get its spots? What is a solar eclipse? With science, we can answer such questions without resorting to magical explanations. And science can lead to technological advances, as well as helping us learn about enormously important and useful topics, such as our health, the environment, and natural hazards. Without science, the modern world would not be modern at all. Still, we have so much to learn. Millions of scientists all over the world are working to solve different parts of the puzzle of how the universe works, peering into its nooks and crannies and deploying their microscopes, telescopes, and other tools to unravel its secrets.

Scientists are everywhere, unravelling the secrets of the universe. Photo credits: Flickr user IFPRI, Flickr user NASA Johnson, and Flickr user Alaska Region U.S. Fish & Wildlife Service.

Science is complex and multi-faceted, but the most important characteristics of science are straightforward:

- Science is a way of learning about what is in the natural world, how the natural world works, and how the natural world got to be the way it is. It is not simply a collection of facts; rather it is a path to understanding.
- Science focuses exclusively on the natural world and does not deal with supernatural explanations.
- Although scientists work in many different ways, all science relies on testing ideas by figuring out what expectations are generated by an idea and making observations to find out whether those expectations hold true.
- Accepted scientific ideas are reliable because they have been subjected to rigorous testing. But, as new evidence is acquired and new perspectives emerge, these ideas can be revised.
- Science is a community endeavor. It relies on a system of checks and balances, which helps ensure that science moves in the direction of greater accuracy and understanding. This system is facilitated by diversity within the scientific community, which offers a broad range of perspectives on scientific ideas.

To many, science may seem like an arcane, ivory-towered institution — but that impression is based on a misunderstanding of science. In fact:

- Science affects your life everyday in all sorts of different ways.
- Science can be fun and is accessible to everyone.

- You are probably already using scientific thinking in your everyday life – maybe without even knowing it.
- Anyone can “do” science by investigating questions scientifically.

Science doesn’t just take place in laboratories. You can have fun with and make use of science in everyday life. Photo credits: Flickr user ActiveSteve, Flickr user Tim Sackton, and Flickr user IndianaDunesNPS.

Source: https://undsci.berkeley.edu/understanding-science-101

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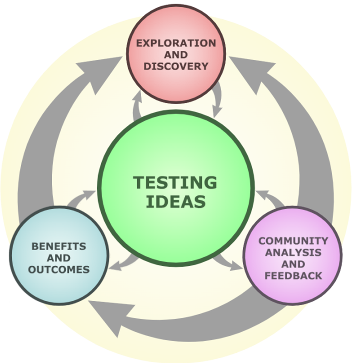
1.2.2: Reading - The Real Process of Science

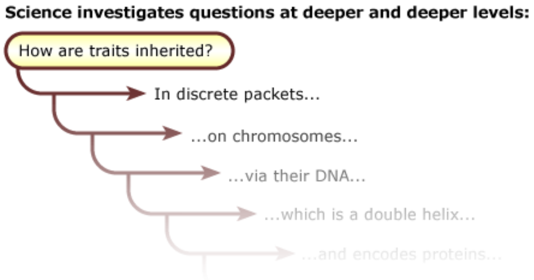
The process of science, as represented here, is the opposite of “cookbook” (to see the full complexity of the process, roll your mouse over each element). In contrast to the linear steps of the simplified scientific method, this process is non-linear:



Click this link for interactive image: https://undsci.berkeley.edu/understanding-science-101/how-science-works/the-real-process-of-science/

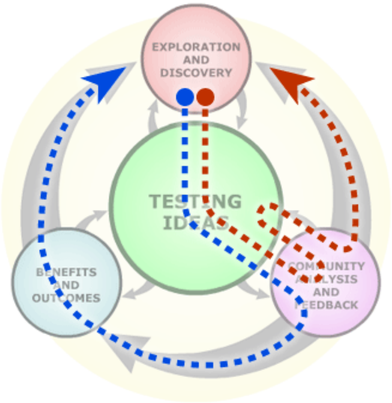
The process of science is iterative.

Science circles back on itself so that useful ideas are built upon and used to learn even more about the natural world. This often means that successive investigations of a topic lead back to the same question, but at deeper and deeper levels. Let’s begin with the basic question of how biological inheritance works. In the mid-1800s, Gregor Mendel showed that inheritance is particulate — that information is passed along in discrete packets that cannot be diluted. In the early 1900s, Walter Sutton and Theodor Boveri (among others) helped show that those particles of inheritance, today known as genes, were located on chromosomes. Experiments by Frederick Griffith, Oswald Avery, and many others soon elaborated on this understanding by showing that it was the DNA in chromosomes which carries genetic information. And then in 1953, James Watson and Francis Crick, again aided by the ideas of many others and using data collected by Rosalind Franklin, provided an even more detailed understanding of inheritance by outlining the molecular structure of DNA. Still later in the 1960s, Marshall Nirenberg, Heinrich Matthaei, and others built upon this work to unravel the molecular code that allows DNA to encode proteins. And it doesn’t stop there. Biologists have continued to deepen and extend our understanding of genes, how they are controlled, how patterns of control themselves are inherited, and how they produce the physical traits that pass from generation to generation.



The process of science is not predetermined.

Any point in the process leads to many possible next steps, and where that next step leads could be a surprise. For example, instead of leading to a conclusion about tectonic movement, testing an idea about plate tectonics could lead to an observation of an unexpected rock layer. And that rock layer could trigger an interest in marine extinctions, which could spark a question about the dinosaur extinction — which might take the investigator off in an entirely new direction.



At first this process might seem overwhelming. And it is, a bit. Even within the scope of a single investigation, science may involve many different people engaged in all sorts of different activities in different orders and at different points in time — science is simply much more dynamic, flexible, unpredictable, and rich than many textbooks represent it as. But don't panic! The scientific process may be complex, but the details are less important than the big picture…

Source: https://undsci.berkeley.edu/understanding-science-101/how-science-works/the-real-process-of-science/

1.3: Correlation vs causation

Video: **Correlation and Causation**

Reading: Why Coffee definitely either kills you or makes you live forever by Emily Oster

Attribution:

Article: Oster, E. (n.d.). *Coffee definitely either kills you or makes you live forever.* ParentData. Retrieved January 24, 2025, from https://parentdata.org/coffee-definitely-either-kills-you-or-makes-you-live-forever/

Video: "CRITICAL THINKING - Fundamentals: Correlation and Causation" YouTube, Wireless Philosophy, Mar 10, 2017, https://www.youtube.com/watch?v=U-_f8RQIIiw

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CHAPTER OVERVIEW

2: Units and Measurements

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.

Figure $2.1.1$: Measuring Blood Pressure. A nurse or a doctor measuring a patient's blood pressure is taking a measurement. (GFDL; Pia von Lützau).

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

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, and the unit is a dozen 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.

Key Take Away

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

2.2: Scientific Notation

Learning Objectives

- Express a large number or a small number in scientific notation.
- Carry out arithmetical operations and express the final answer in scientific notation

Chemists often work with numbers that are exceedingly large or small. For example, entering the mass in grams of a hydrogen atom into a calculator would require a display with at least 24 decimal places. A system called **scientific notation** avoids much of the tedium and awkwardness of manipulating numbers with large or small magnitudes. In scientific notation, these numbers are expressed in the form

$$N \times 10^n$$

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

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

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

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases and vice versa. The application of this rule is illustrated in Example 2.2.1.

✓ Example 2.2.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

a. 637.8
b. 0.0479
c. 7.86
d. 12,378
e. 0.00032
f. 61.06700
g. 2002.080
h. 0.01020

Solution

Solutions to Example 2.2.1

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, n = 2.	6.378×10^2
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, n = −2.	4.79×10^{-2}
c	This is usually expressed simply as 7.86. (Recall that $10^0 = 1$.)	7.86×10^0
d	Because the decimal point was moved four places to the left, n = 4.	1.2378×10^4
e	Because the decimal point was moved four places to the right, n = −4.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, n = 1.	6.106700×10^1
g	Because the decimal point was moved three places to the left, n = 3.	2.002080×10^3
h	Because the decimal point was moved two places to the right, n = -2.	1.020×10^{-2}

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N. Example 2.2.2 illustrates how to do this.

Example 2.2.2: Expressing Sums and Differences in Scientific Notation

Carry out the appropriate operation and then express the answer in scientific notation.

a. $(1.36 \times 10^2) + (4.73 \times 10^3)$
b. $(6.923 \times 10^{-3}) - (8.756 \times 10^{-4})$

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) =$ $(1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) =$ $(0.136 + 4.73) \times 10^3) = 4.87 \times 10^3$. Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66 × 10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.87×10^3
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) =$ $(6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) =$ $(69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}$. In converting 60.474 × 10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.047×10^{-3}

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n. Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example 2.2.3.

Example 2.2.3: Expressing Products and Quotients in Scientific Notation

Perform the appropriate operation and express your answer in scientific notation.

a. $(6.022 \times 10^{23})(6.42 \times 10^{-2})$
b. $\dfrac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}}$
c. $\dfrac{(6.63 \times 10^{-34})(6.0 \times 10)}{8.52 \times 10^{-2}}$

Solution

Solution to Example 2.2.3

	Explanation	
a	In multiplication, we add the exponents: $(6.022 \times 10^{23})(6.42 \times 10^{-2}) = (6.022)(6.42) \times 10^{[23+(-2)]} = 38.7 \times 10^{21}$ In converting 38.7×10^{21} to scientific notation, n has become more positive by 1 because the value of N has decreased.	3.87×10^{22}
b	In division, we subtract the exponents: $\frac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}} =$ $\frac{1.67}{9.12} \times 10^{[-24-(-28)]} = 0.183 \times 10^4$ In converting 0.183×10^4 to scientific notation, n has become more negative by 1 because the value of N has increased.	1.83×10^3
c	This problem has both multiplication and division: $\frac{(6.63 \times 10^{-34})(6.0 \times 10)}{(8.52 \times 10^{-2})} =$ $\frac{39.78}{8.52} \times 10^{[-34+1-(-2)]}$	4.7×10^{-31}

Video: **Scientific Notation by Mathantics**

- **2.2: Scientific Notation - Writing Large and Small Numbers** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

2.3: Significant Figures

Learning Objectives

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

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 below, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 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 at 2.553 cm as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

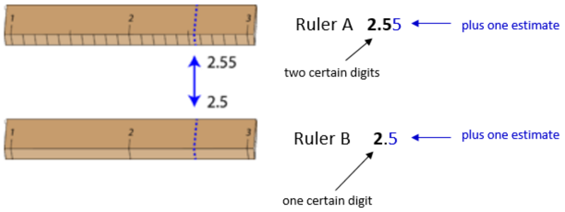


Figure $2.3.1$: Measurement with two different rulers.
Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

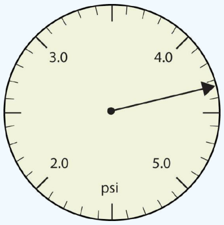
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 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 blue arrow).

With either ruler, it is clear that the length of the object is between 2 and 3 cm. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as 2.5 cm. However, another person may judge that the measurement is 2.4 cm or perhaps 2.6 cm. While the 2 is known for certain, the value of the tenths digit is uncertain.

The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as 2.55 cm. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be 2.54 cm or 2.56 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

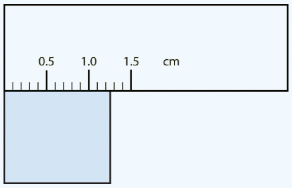
Example 2.3.1: Reporting Measurements to the Proper Number of Significant Figures

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



a.

b.



Ruler measuring a rectangle in units of centimeters, with the rectangle's edge between 1.2 and 1.3 cm marks

Solutions

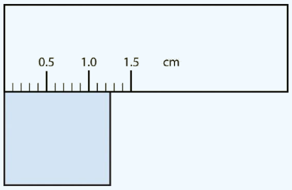
Solutions to Example 2.3.1

Explanation	Answer

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

? Exercise 2.3.1

What would be the reported width of this rectangle?



Answer

1.25 cm

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 (middle zeros) are always significant.	• 39,004 has five significant figures. • 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	• 0.008 has one significant figure. • 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. 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.**	• 1400 is ambiguous. ◦ 1.4×10^3 has two significant figures. ◦ 1.40×10^3 three significant figures. ◦ 1.400×10^3 has four significant figures.
5. Trailing zeros 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. • 19.000 has five significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have infinitely many significant figures. If we have counted four objects, for example, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000…). Similarly, 1 foot (ft) is defined to contain 12 inches (in), so the number 12 in the following equation has infinitely many significant figures:

✓ Example 2.3.2

Give the number of significant figures in each. Identify the rule for each.

a. 5.87

b. 0.031
c. 52.90
d. 00.2001
e. 500
f. 6 atoms

Solution

Solution to Example 2.3.2

	Explanation	Answer
a	All three numbers are significant (rule 1).	5.87, three significant figures
b	The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1).	0.031, two significant figures
c	The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures
d	The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2).	00.2001, four significant figures
e	The number is ambiguous. It could have one, two or three significant figures.	500, ambiguous
f	The 6 is a counting number. A counting number is an exact number.	6, infinite

Video: **Significant figures by Khan Academy**

? Exercise 2.3.2

Give the number of significant figures in each.

a. 36.7 m
b. 0.006606 s
c. 2,002 kg
d. 306,490,000 people
e. 3,800 g

Answer a

three significant figures

Answer b

four significant figures

Answer c

four significant figures

Answer d

infinite (exact number)

Answer e

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

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.

2.4: Accuracy and Precision

Learning Objectives

- Define accuracy and precision
- Distinguish exact and uncertain numbers

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to know both the precision and the accuracy of their results. Measurements are said to be precise if they yield very similar results when repeated in the same manner. A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value. These characterizations can be extended to other contexts, such as the results of an archery competition (Figure $2.4.2$).

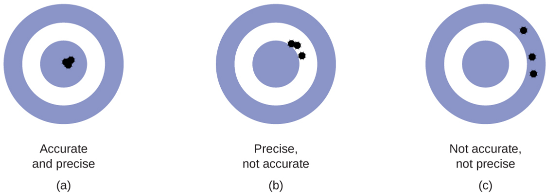


Figure $2.4.2$: *(a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.*

Suppose a quality control chemist at a pharmaceutical company is tasked with checking the accuracy and precision of three different machines that are meant to dispense 10 ounces (296 mL) of cough syrup into storage bottles. She proceeds to use each machine to fill five bottles and then carefully determines the actual volume dispensed, obtaining the results tabulated in Table $2.4.2$.

Table $2.4.2$: Volume (mL) of Cough Medicine Delivered by 10-oz (296 mL) Dispensers

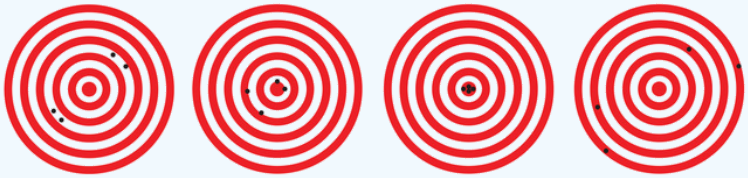
Dispenser #1	Dispenser #2	Dispenser #3
283.3	298.3	296.1
284.1	294.2	295.9
283.9	296.0	296.1
284.0	297.8	296.0
284.1	293.9	296.1

Considering these results, she will report that dispenser #1 is precise (values all close to one another, within a few tenths of a milliliter) but not accurate (none of the values are close to the target value of 296 mL, each being more than 10 mL too low). Results for dispenser #2 represent improved accuracy (each volume is less than 3 mL away from 296 mL) but worse precision (volumes vary by more than 4 mL). Finally, she can report that dispenser #3 is working well, dispensing cough syrup both accurately (all volumes within 0.1 mL of the target volume) and precisely (volumes differing from each other by no more than 0.2 mL).

Video: **What's the difference between accuracy and precision? - Matt Anticole**

✓ Example

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


(a) (b) (c) (d)

Which target shows

a. a precise, but inaccurate set of measurements?
b. a set of measurements that is both precise and accurate?
c. a set of measurements that is neither precise nor accurate?

Solution

a. Set **(a)** is precise, but inaccurate.
b. Set **(c)** is both precise and accurate.
c. Set **(d)** is neither precise nor accurate.

Summary

- Accuracy is a measure of how close a measurement is to the correct or accepted value of the quantity being measured.
- Precision is a measure of how close a series of measurements are to one another.

2.5: The Basic Units of Measurements

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

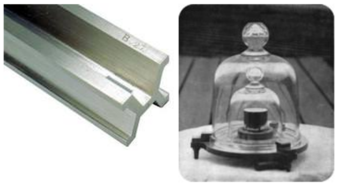


Figure 2.5.1: Meter standard (left) and Kilogram standard (right).

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 very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of 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 units are frequently encountered in chemistry. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

Unfortunately, the Metric System is Not Ubiquitous

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

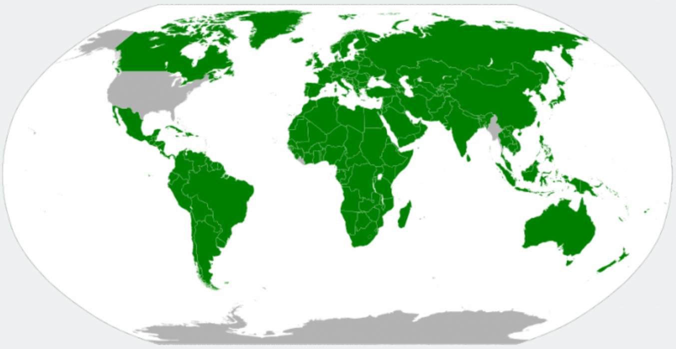


Figure 2.5.1: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

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

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 $3.5\ \text{mL}$.

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 $6.5\ \text{kg}$, this is easier than saying it weighs $6500\ \text{g}$ or $0.65\ \text{dag}$. All three are correct, but the kg units in this case make for a small and easily managed number. However, if a specific problem needs grams instead of kilograms, go with the grams for consistency.

✓ Example 2.5.1: Unit Abbreviations

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

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

Solutions

Solutions to Example 2.5.1

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

Video: **What are Physical Quantities? by Free Animated Education**

? 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

Answer b:

mg

Answer c:

ns

Answer d:

cL

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.

2.6: Measured Units vs. Derived Unit

Learning Objectives

- Explain the process of measurement and describe the three basic parts of a quantity.
- Describe the properties and units of length, mass, volume, density, temperature, and time.
- Recognize the common unit prefixes and use them to describe the magnitude of a measurement.
- Describe and calculate the density of a substance.
- Perform basic unit calculations and conversions in the metric and other unit systems.

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

The number in the measurement can be represented in different ways, including decimal form and scientific notation. For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as 2.98×10^5 kg. The mass of the average mosquito is about 0.0000025 kilograms, which can be written as 2.5×10^{-6} kg.

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

We usually report the results of scientific measurements in SI units, an updated version of the metric system, using the units listed in Table $2.6.1$. Other units can be derived from these base units. The standards for these units are fixed by international agreement, and they are called the International System of Units or SI Units (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964.

Table $2.6.1$: Base Units of the SI System

Property Measured	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K
electric current	ampere	A
amount of substance	mole	mol
luminous intensity	candela	cd

Sometimes we use units that are fractions or multiples of a base unit. Ice cream is sold in quarts (a familiar, non-SI base unit), pints (0.5 quart), or gallons (4 quarts). We also use fractions or multiples of units in the SI system, but these fractions or multiples are always powers of 10. Fractional or multiple SI units are named using a prefix and the name of the base unit. For example, a length of 1000 meters is also called a kilometer because the prefix *kilo* means "one thousand," which in scientific notation is 10^3 (1 kilometer = 1000 m = 10^3 m). The prefixes used and the powers to which 10 are raised are listed in Table $2.6.2$.

NG, equals 4 times ten to the negative 9, or 0.000000004 g. The prefix micro has the greek letter mu as its symbol and a factor of 10 to the negative sixth power. Therefore, 1 microliter, or mu L, is equal to one times ten to the negative 6 or 0.000001 L. The prefix milli has a lowercase M as its symbol and a factor of 10 to the negative third power. Therefore, 2 millimoles, or M mol, are equal to two times ten to the negative 3 or 0.002 mol. The prefix centi has a lowercase C as its symbol and a factor of 10 to the negative second power. Therefore, 7 centimeters, or C M, are equal to seven times ten to the negative 2 meters or 0.07 M O L. The prefix deci has a lowercase D as its symbol and a factor of 10 to the negative first power. Therefore, 1 deciliter, or lowercase D uppercase L, are equal to one times ten to the negative 1 meters or 0.1 L. The prefix kilo has a lowercase K as its symbol and a factor of 10 to the third power. Therefore, 1 kilometer, or K M, is equal to one times ten to the third meters or 1000 M. The prefix mega has an uppercase M as its symbol and a factor of 10 to the sixth power. Therefore, 3 megahertz, or M H Z, are equal to three times 10 to the sixth hertz, or 3000000 H Z. The prefix giga has an uppercase G as its symbol and a factor of 10 to the ninth power. Therefore, 8 gigayears, or G Y R, are equal to eight times 10 to the ninth years, or 800000000 G Y R. The prefix tera has an uppercase T as its symbol and a factor of 10 to the twelfth power. Therefore, 5 terawatts, or T W, are equal to five times 10 to the twelfth watts, or 5000000000000 W." data-quail-id="64" data-mt-width="1076">

Table 2.6.2: Common Unit Prefixes

Prefix	Symbol	Factor	Example
femto	f	10^{-15}	1 femtosecond (fs) = 1 × 10^{-15} s (0.000000000000001 s)
pico	p	10^{-12}	1 picometer (pm) = 1 × 10^{-12} m (0.000000000001 m)
nano	n	10^{-9}	4 nanograms (ng) = 4 × 10^{-9} g (0.000000004 g)
micro	µ	10^{-6}	1 microliter (µL) = 1 × 10^{-6} L (0.000001 L)
milli	m	10^{-3}	2 millimoles (mmol) = 2 × 10^{-3} mol (0.002 mol)
centi	c	10^{-2}	7 centimeters (cm) = 7 × 10^{-2} m (0.07 m)
deci	d	10^{-1}	1 deciliter (dL) = 1 × 10^{-1} L (0.1 L)
kilo	k	10^{3}	1 kilometer (km) = 1 × 10^{3} m (1000 m)
mega	M	10^{6}	3 megahertz (MHz) = 3 × 10^{6} Hz (3,000,000 Hz)
giga	G	10^{9}	8 gigayears (Gyr) = 8 × 10^{9} yr (8,000,000,000 Gyr)
tera	T	10^{12}	5 terawatts (TW) = 5 × 10^{12} W (5,000,000,000,000 W)

SI Base Units

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

Length

The standard unit of length in both the SI and original metric systems is the meter (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. A meter is about 3 inches longer than a yard (Figure 2.6.1); one meter is about 39.37 inches or 1.094

yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

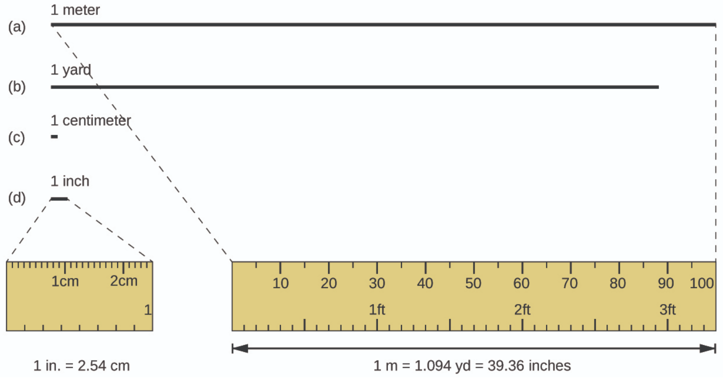


Figure 2.6.1: The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

A ruler is shown with various lengths of black line shown above it to compare the relative lengths of 1 inch, meter, centimeter, and yard.

Mass

The standard unit of mass in the SI system is the kilogram (kg). A kilogram was originally defined as the mass of a liter of water (a cube of water with an edge length of exactly 0.1 meter). It is now defined by a certain cylinder of platinum-iridium alloy, which is kept in France (Figure 1.4.2). Any object with the same mass as this cylinder is said to have a mass of 1 kilogram. One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (10^{-3} kg).

Figure 2.6.2: This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology).

Temperature

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

Time

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

Derived SI Units

We can derive many units from the seven SI base units. For example, we can use the base unit of length to define a unit of volume, and the base units of mass and length to define a unit of density.

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (Figure 2.6.3). The standard volume is a cubic meter (m^3), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

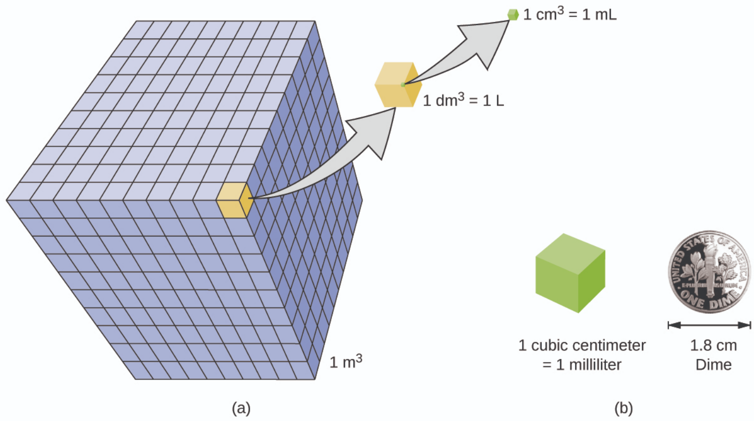


Figure 2.6.3: (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

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

Density

We use the mass and volume of a substance to determine its density. Thus, the units of density are defined by the base units of mass and length.

The density of a substance is the ratio of the mass of a sample of the substance to its volume. The SI unit for density is the kilogram per cubic meter (kg/m^3). For many situations, however, this as an inconvenient unit, and we often use grams per cubic centimeter (g/cm^3) for the densities of solids and liquids, and grams per liter (g/L) for gases. Although there are exceptions, most liquids and solids have densities that range from about 0.7 g/cm^3 (the density of gasoline) to 19 g/cm^3 (the density of gold). The density of air is about 1.2 g/L. Table 2.6.3 shows the densities of some common substances.

Table 2.6.3: Densities of Common Substances

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

While there are many ways to determine the density of an object, perhaps the most straightforward method involves separately finding the mass and volume of the object, and then dividing the mass of the sample by its volume. In the following example, the mass is found directly by weighing, but the volume is found indirectly through length measurements.

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

✓ Example 2.6.1

Calculation of Density Gold—in bricks, bars, and coins—has been a form of currency for centuries. In order to swindle people into paying for a brick of gold without actually investing in a brick of gold, people have considered filling the centers of hollow gold bricks with lead to fool buyers into thinking that the entire brick is gold. It does not work: Lead is a dense substance, but its density is not as great as that of gold, 19.3 g/cm^3. What is the density of lead if a cube of lead has an edge length of 2.00 cm and a mass of 90.7 g?

Solution

The density of a substance can be calculated by dividing its mass by its volume. The volume of a cube is calculated by cubing the edge length.

$$\text{volume of lead cube} = 2.00 \text{ cm} \times 2.00 \text{ cm} \times 2.00 \text{ cm} = 8.00 \text{ cm}^3$$

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{90.7 \text{ g}}{8.00 \text{ cm}^3} = \frac{11.3 \text{ g}}{1.00 \text{ cm}^3} = 11.3 \text{ g/cm}^3$$

(We will discuss the reason for rounding to the first decimal place in the next section.)

? Exercise 2.6.1

a. To three decimal places, what is the volume of a cube (cm^3) with an edge length of 0.843 cm?
b. If the cube in part (a) is copper and has a mass of 5.34 g, what is the density of copper to two decimal places?

Answer a

0.599 cm^3;

Answer b

8.91 g/cm^3

✓ Example 2.6.2: Using Displacement of Water to Determine Density

This PhET simulation illustrates another way to determine density, using displacement of water. Determine the density of the red and yellow blocks.

Solution

When you open the density simulation and select Same Mass, you can choose from several 5.00-kg colored blocks that you can drop into a tank containing 100.00 L water. The yellow block floats (it is less dense than water), and the water level rises to 105.00 L. While floating, the yellow block displaces 5.00 L water, an amount equal to the weight of the block. The red block sinks (it is more dense than water, which has density = 1.00 kg/L), and the water level rises to 101.25 L.

The red block therefore displaces 1.25 L water, an amount equal to the volume of the block. The density of the red block is:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.00 \text{ kg}}{1.25 \text{ L}} = 4.00 \text{ kg/L}$$

Note that since the yellow block is not completely submerged, you cannot determine its density from this information. But if you hold the yellow block on the bottom of the tank, the water level rises to 110.00 L, which means that it now displaces 10.00 L water, and its density can be found:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.00 \text{ kg}}{10.00 \text{ L}} = 0.500 \text{ kg/L}$$

Exercise 2.6.1

Remove all of the blocks from the water and add the green block to the tank of water, placing it approximately in the middle of the tank. Determine the density of the green block.

Answer

2.00 kg/L

Summary

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

Key Equations

- $\text{density} = \frac{\text{mass}}{\text{volume}}$

2.7: Conversion between Units with Conversion Factor (Single-step Conversion Problems)

Learning Objectives

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

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

Converting Between Units with Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. A simple conversion factor can convert meters into centimeters, or a more complex one can convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. Always remember that a conversion factor has to represent a fact; this fact can either be simple or more 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/sec. Either one of these can be used as a conversion factor depending on what type of calculation you are working with (Table $2.7.1$).

Table $2.7.1$: Conversion Factors from SI units to English Units

English Units	Metric Units	Quantity
1 ounce (oz)	28.35 grams (g)	*mass
1 fluid once (oz)	29.6 mL	volume
2.205 pounds (lb)	1 kilogram (kg)	*mass
1 inch (in)	2.54 centimeters (cm)	length
0.6214 miles (mi)	1 kilometer (km)	length
1 quarter (qt)	0.95 liters (L)	volume

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

Of course, there are other ratios which are not listed in Table $2.7.1$. They may include:

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

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

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

or

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

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

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

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

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

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

Performing Dimensional Analysis

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

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

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

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

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

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

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

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

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

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

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

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

For the answer to be meaningful, we have to *construct the conversion factor in a form that causes the original unit to cancel out*. Figure 2.7.1 shows a concept map for constructing a proper conversion.

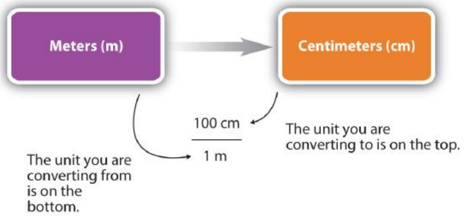


Figure 2.7.1: A Concept Map for Conversions. This is how you construct a conversion factor to convert from one unit to another. Converting meters to centimeters: the unit you are converting is on the bottom of the fraction, and the unit to convert to is on top

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 your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

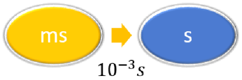
Significant Figures in Conversions

How do conversion factors affect the determination of significant figures?

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

Example 2.7.1

Steps for Problem Solving for Example 2.6.1 and 2.6.2

	Example 2.7.1	**Example 2.7.2**
Steps for Problem Solving	The average volume of blood in an adult male is 4.7 L. What is this volume in milliliters?	A hummingbird can flap its wings once in 18 ms. How many seconds are in 18 ms?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.7 L Find: mL	Given: 18 ms Find: s
List other known quantities.	$1\,mL = 10^{-3}L$	$1\,ms = 10^{-3}s$
Prepare a concept map and use the proper conversion factor.	 $\frac{1mL}{10^{-3}L}$	 $\frac{10^{-3}s}{1ms}$

Cancel units and calculate.	$4.7\ \cancel{\text{L}} \times \frac{1\ \text{mL}}{10^{-3}\ \ \cancel{\text{L}}} = 4,700\ \text{mL}$ or $4.7\ \cancel{\text{L}} \times \frac{1,000\ \text{mL}}{1\ \ \cancel{\text{L}}} = 4,700\ \text{mL}$ or 4.7 x 10^3 2SF, not ambiguous	$18\ \ \cancel{\text{ms}} \times \frac{10^{-3}\ \text{s}}{1\ \ \cancel{\text{ms}}} = 0.018\ \text{s}$ or $18\ \ \cancel{\text{ms}} \times \frac{1\ \text{s}}{1,000\ \ \cancel{\text{ms}}} = 0.018\ \text{s}$
Think about your result.	The amount in mL should be 1000 times larger than the given amount in L.	The amount in s should be 1/1000 the given amount in ms.

? Exercise 2.7.1

Perform each conversion.

a. 101,000 ns to seconds
b. 32.08 kg to grams
c. 1.53 grams to cg

Answer a:

$1.01000x10^{-4}s$

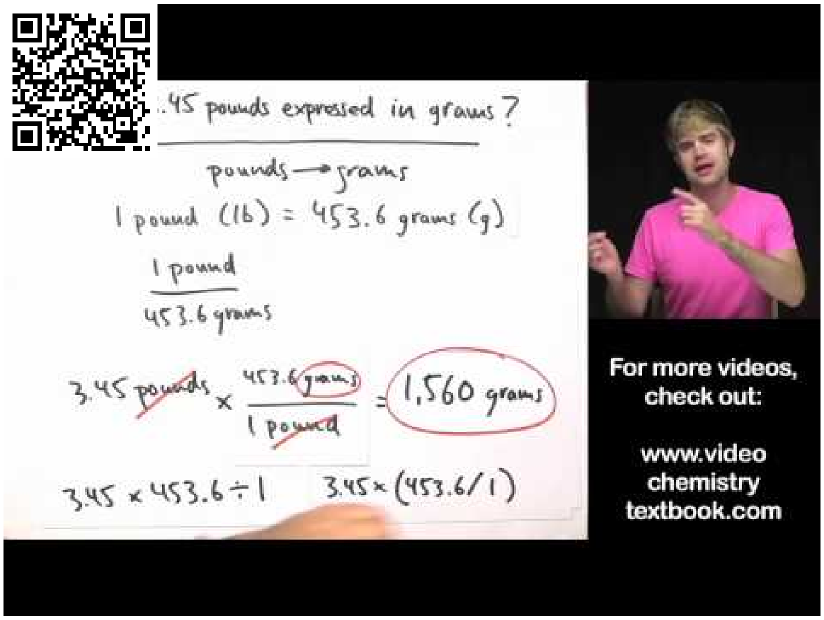
Answer b:

$3.208x10^{4}g$

Answer c:

$1.53x10^{2}cg$

Video: **Converting Units with Conversion Factors by Tyler Dewitt**



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.

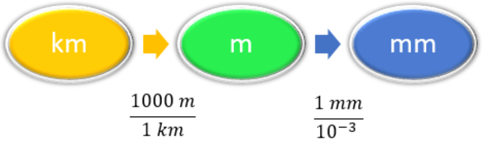
2.8: Conversion Between Units with Conversion Factor (Multi-step Conversion Problems)

Multiple Conversions

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

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

Concept Map



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

Calculation

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

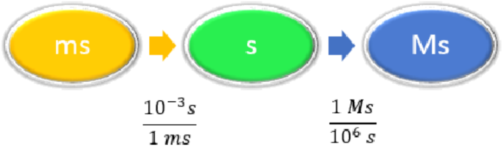
$$= 5.47 \times 10^{7}\ \text{mm}$$

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

✓ Example 2.8.1: Unit Conversion

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

Solution

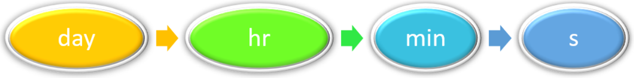
Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 58.2 ms Find: Ms
List other known quantities	$1 ms = 10^{-3} s$ $1 Ms = 10^{6} s$
Prepare a concept map.	 Convert milliseconds to seconds to microseconds: use conversion factors 0.001 second per millisecond and 1 microsecond per 1 million seconds
Calculate.	$58.2\ \cancel{\text{ms}} \times \frac{10^{-3}\ \cancel{\text{s}}}{1\ \cancel{\text{ms}}} \times \frac{1\ \text{Ms}}{1,000,000\ \cancel{\text{s}}} = 0.0000000582\ \text{Ms}$ $= 5.82 \times 10^{-8}\ \text{Ms}$ Neither conversion factor affects the number of significant figures in the final answer.

✓ Example 2.8.2: Unit Conversion

How many seconds are in a day?

Solution

Steps for Problem Solving	Unit Conversion
Identify the "given" information and what the problem is asking you to "find."	Given: 1 day Find: s
List other known quantities.	1 day = 24 hours 1 hour = 60 minutes 1 minute = 60 seconds

Steps for Problem Solving	Unit Conversion
Prepare a concept map.	 $\frac{24\text{hr}}{1\text{ day}}$ $\frac{60\text{min}}{1\text{ hr}}$ $\frac{60\text{ s}}{1\ min}$ Convert day to hour to minute to second: use conversion factors 24 hours per day, 60 minutes per hour, and 60 seconds per minute
Calculate.	$1\text{ d}\times\frac{24\text{ hr}}{1\text{ d}}\times\frac{60\text{ min}}{1\text{ hr}}\times\frac{60\text{ s}}{1\text{ min}}=86,400\text{ s}$

? Exercise 2.8.1

Perform each conversion in one multi-step calculation.

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

Answer a

$4.3007\times10^{-11} kg$

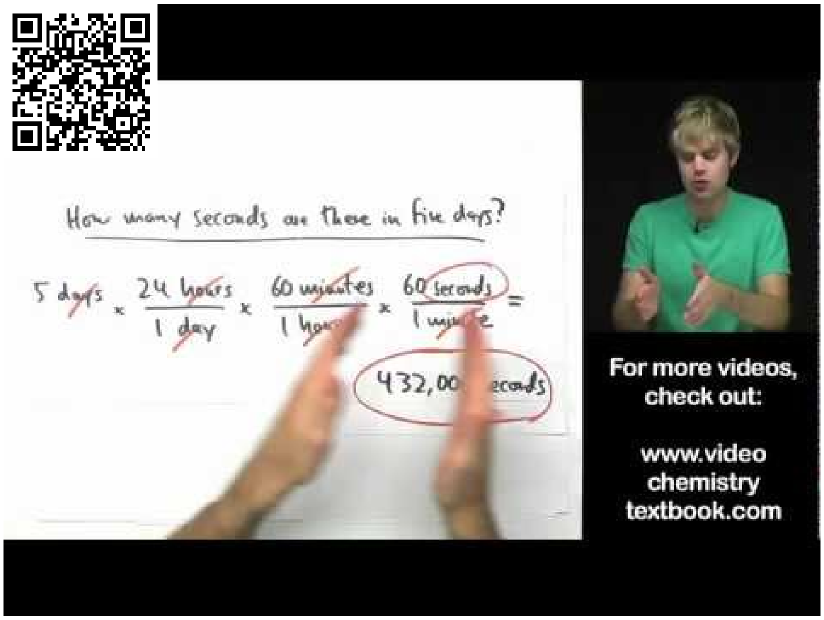
Answer b

$83.75\ ft$

Answer c

$19\ km$

Video: **Converting Units using Multiple Conversion Factors by Tyler Dewitt**



Career Focus: Pharmacist

A pharmacist dispenses drugs that have been prescribed by a doctor. Although that may sound straightforward, pharmacists in the United States must hold a doctorate in pharmacy and be licensed by the state in which they work. Most pharmacy programs require four years of education in a specialty pharmacy school. Pharmacists must know a lot of chemistry and biology so they can understand the effects that drugs (which are chemicals, after all) have on the body. Pharmacists can advise physicians on the selection, dosage, interactions, and side effects of drugs. They can also advise patients on the proper use of their medications, including when and how to take specific drugs properly. Pharmacists can be found in drugstores, hospitals, and other medical facilities. Curiously, an outdated name for pharmacist is *chemist*, which was used when pharmacists formerly did a lot of drug preparation, or *compounding*. In modern times, pharmacists rarely compound their own drugs, but their knowledge of the sciences, including chemistry, helps them provide valuable services in support of everyone’s health.


A woman consulting with a pharmacist. (Public Domain; Rhoda Baer via National Cancer Institute, an agency 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.

2.9: Convert between %, ppm, and ppb

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.

Introduction

Concentration is the measure of how much of a given substance is mixed with another substance. Solutions are said to be either dilute or concentrated. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution.

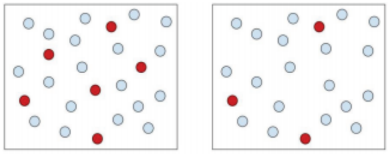


Figure $2.9.1$: The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. To make juice, you have to mix the frozen juice concentrate from inside these containers with three or four times the container size full of water. Therefore, you are diluting the concentrated juice. In terms of solute and solvent, the concentrated solution has a lot of solute versus the dilute solution that would have a smaller amount of solute.

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

There are many ways to express the concentration of a solution. In this section, we will consider mass percent, volume percent, mass-volume percent, and parts per. In the next section, we will look at molarity.

Mass Percent

There are several ways of expressing the concentration of a solution by using a percentage. The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\% \text{ m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving 25.0 g of sugar into 100.0 g of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0 \text{ g sugar}}{125.0 \text{ g solution}} \times 100\% = 20.0\% \text{ sugar}$$

Example 2.9.1

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$\% \text{ m/m} = \frac{36.5 \text{ g}}{355 \text{ g}} \times 100\% = 10.3\%$$

Exercise 2.9.1

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00 × 10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer

7.90 %

Using Mass Percent in Calculations

Sometimes you may want to make up a particular mass of solution of a given percent by mass and need to calculate what mass of the solute to use. Using mass percent as a conversion can be useful in this type of problem. The mass percent can be expressed as a conversion factor in the form $\frac{g \text{ solute}}{100 \text{ gsolution}}$ or $\frac{100 \text{ gsolution}}{g \text{ solute}}$

For example, if you need to make $3000.0 \text{ g}$of a 5.00%solution of sodium chloride, the mass of solute needs to be determined.

Solution

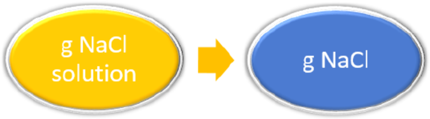
Given: 3000.0 g NaCl solution

5.00% NaCl solution

Find: mass of solute = ? g NaCl

Other known quantities: 5.00 g NaCl is to 100 g solution

The appropriate conversion factor (based on the given mass percent) can be used follows:



$$\frac{5.00 \; g \; NaCl}{100 \; g \; NaCl \; solution}$$

To solve for the mass of NaCl, the given mass of solution is multiplied by the conversion factor.

$$gNaCl = 3,000.0 \; g \; NaCl \; \cancel{solution} \times \frac{5.00 \; g \; NaCl}{100 \; g \; NaCl \; \cancel{solution}} = 150.0g \; NaCl$$

You would need to weigh out 150 g of NaCl and add it to 2850 g of water. Notice that it was necessary to subtract the mass of the NaCl (150 g) from the mass of solution (3000 g) to calculate the mass of the water that would need to be added.

? Exercise 2.9.1

What is the amount (in g) of hydrogen peroxide (H_2O_2) needed to make a 6.00 kg, 3.00 % (by mass) H_2O_2 solution?

Answer

180 g H_2O_2

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, %vol or (v/v)%:

$$\text{volume percentage} = \frac{\text{volume solute}}{\text{volume solution}} \times 100\% \tag{2.9.1}$$

✓ Example 2.9.3: Calculations using Volume Percentage

Rubbing alcohol (isopropanol) is usually sold as a 70%vol aqueous solution. If the density of isopropyl alcohol is 0.785 g/mL, how many grams of isopropyl alcohol are present in a 355 mL bottle of rubbing alcohol?

Solution

Per the definition of volume percentage, the isopropanol volume is 70% of the total solution volume. Multiplying the isopropanol volume by its density yields the requested mass:

$$355 \text{ mL solution}\left(\frac{70 \text{ mL isopropyl alcohol}}{100 \text{ mL solution}}\right)\left(\frac{0.785 \text{ g isopropyl alcohol}}{1 \text{ mL isopropyl alcohol}}\right) = 195 \text{ g isopropyl alcohol}$$

? Exercise 2.9.3

Wine is approximately 12% ethanol (CH_3CH_2OH) by volume. Ethanol has a molar mass of 46.06 g/mol and a density 0.789 g/mL. How many moles of ethanol are present in a 750-mL bottle of wine?

Answer

1.5 mol ethanol

Mass-Volume Percentage

“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 percent is a ratio of a solute’s mass to the solution’s volume expressed as a percentage. The specific units used for solute mass and solution volume may vary, depending on the solution. For example, physiological saline solution, used to prepare intravenous fluids, has a concentration of 0.9% mass/volume (m/v), indicating that the composition is 0.9 g of solute per 100 mL of 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, its concentration is usually given in milligrams of glucose per deciliter (100 mL) of blood (Figure 2.9.2).

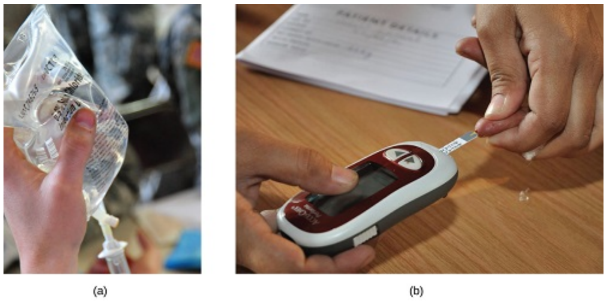


Figure 2.9.2: “Mixed” mass-volume units are commonly encountered in medical settings. (a) The NaCl concentration of physiological saline is 0.9% (m/v). (b) This device measures glucose levels in a sample of blood. The normal range for glucose concentration in blood (fasting) is around 70–100 mg/dL. (credit 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:

$$\text{ppm} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^6 \text{ ppm} \tag{2.9.2}$$

$$\text{ppb} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^9 \text{ ppb} \tag{2.9.3}$$

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 2.9.3).

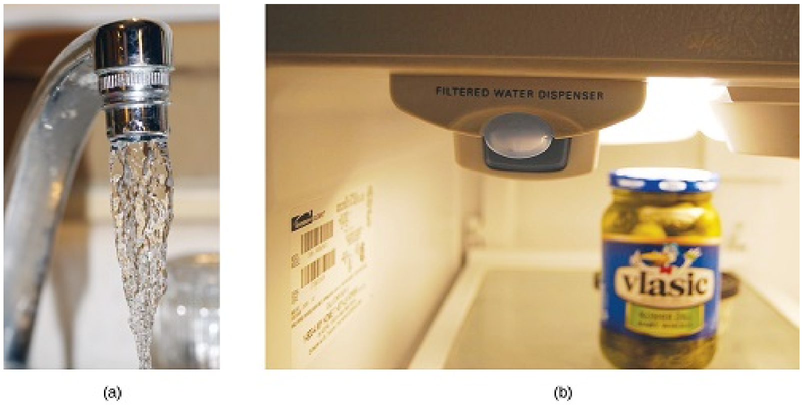


Figure 2.9.3: (a) In some areas, trace-level concentrations of contaminants can render unfiltered tap water unsafe for drinking and cooking. (b) Inline water filters reduce the concentration of solutes in tap water. (credit a: modification of work by Jenn Durfey; credit b: modification of work by “vastateparkstaff”/Wikimedia commons).

✓ Example 2.9.4: Parts per Million and Parts per Billion Concentrations

According to the EPA, when the concentration of lead in tap water reaches 15 ppb, certain remedial actions must be taken. What is this concentration in ppm? At this concentration, what mass of lead (μg) would be contained in a typical glass of water (300 mL)?

Solution

The definitions of the ppm and ppb units may be used to convert the given concentration from ppb to ppm. Comparing these two unit definitions shows that ppm is 1000 times greater than ppb (1 ppm = 10^3 ppb). Thus:

$$15\ \cancel{\text{ppb}} \times \frac{1\ \text{ppm}}{10^3\ \cancel{\text{ppb}}} = 0.015\ \text{ppm}$$

The definition of the ppb unit may be used to calculate the requested mass if the mass of the solution is provided. However, only the volume of solution (300 mL) is given, so we must use the density to derive the corresponding mass. We can assume the density of tap water to be roughly the same as that of pure water (~1.00 g/mL), since the concentrations of any dissolved substances should not be very large. Rearranging the equation defining the ppb unit and substituting the given quantities yields:

$$\text{ppb} = \frac{\text{mass solute}}{\text{mass solution}} \times 10^9\ \text{ppb}$$

$$\text{mass solute} = \frac{\text{ppb} \times \text{mass solution}}{10^9\ \text{ppb}}$$

$$\text{mass solute} = \frac{15\ \text{ppb} \times 300\ \text{mL} \times \frac{1.00\ \text{g}}{\text{mL}}}{10^9\ \text{ppb}} = 4.5 \times 10^{-6}\ \text{g}$$

Finally, convert this mass to the requested unit of micrograms:

$$4.5 \times 10^{-6}\ \text{g} \times \frac{1\ \mu\text{g}}{10^{-6}\ \text{g}} = 4.5\ \mu\text{g}$$

? Exercise 2.9.4

A 50.0-g sample of industrial wastewater was determined to contain 0.48 mg of mercury. Express the mercury concentration of the wastewater in ppm and ppb units.

Answer

9.6 ppm, 9600 ppb

Measures of Concentration: https://youtu.be/RjMGaUpkg8g

Summary

In addition to molarity, a number of other solution concentration units are used in various applications. Percentage concentrations based on the solution components' masses, volumes, or both are useful for expressing relatively high concentrations, whereas lower concentrations are conveniently expressed using ppm or ppb units. These units are popular in environmental, medical, and other fields where mole-based units such as molarity are not as commonly used.

Contributors and Attributions

2.10: Mole and Avagadro’s Number Conversions

Learning Objectives

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

We need a unit of amount that relates quantities of substances on a scale that we can interact with.

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

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

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

Example 2.10.1

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

Solution

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

$$2.76\ \ mol\ \cancel{H_2O} \times \frac{6.022 \times 10^{23} molecules\ H_2O}{mol\ \cancel{H_2O}} = 1.66 \times 10^{24} molecules\ H_2O$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24}\ molecules\ \cancel{H_2O} \times \frac{3\ atoms}{1\ molecule} = 4.99 \times 10^{24}\ atoms$$

Exercise 2.10.1

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

Answer

2.78×10^{22} molecules

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

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

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

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

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

$$1.50 \times 10^{25}\ \cancel{atoms\,C} \times \frac{12.0000\,g\,C}{6.022 \times 10^{23}\ \cancel{atoms\,C}} = 299\,g\,C$$

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

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

✓ Example 2.10.2: Sugar

What is the molar mass of sugar ($C_6H_{12}O_6$)?

Solution

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

Solutions to Example 5.3.2

6 C = 6 × 12.011	= 72.066
12 H = 12 × 1.0079	= 12.0948
6 O = 6 × 15.999	= 95.994
TOTAL	= 180.155 g/mol

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

? Exercise 2.10.2

What is the molar mass of $AgNO_3$?

Answer

169.87 g/mol

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

✓ Example 2.10.3

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

Solution

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

$$3.56\ mol\ \cancel{HgCl_2} \times \frac{271.49\,g\,HgCl_2}{mol\ \cancel{HgCl_2}} = 967\,g\,HgCl_2$$

? Exercise 2.10.3

What is the mass of 33.7 mol of H_2O?

Answer

607 g

✓ Example 2.10.4

How many moles of H_2O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H_2O as a conversion factor from mass to moles. The molar mass of water is

(1.0079 + 1.0079 + 15.999) = 18.015 g/mol.

However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0\ \ g\,\cancel{H_2O} \times \frac{1\,mol\,H_2O}{18.015\ g\,\cancel{H_2O}} = 13.32\,mol\,H_2O$$

? Exercise 2.10.4

How many moles are present in 35.6 g of H_2SO_4 (molar mass = 98.08 g/mol)?

Answer

0.363 mol

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

✓ Example 2.10.5

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

Solution

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

$$100\ \cancel{ml}\ ethanol \times \frac{0.789\,g}{\cancel{ml}} \times \frac{1\,mol}{46.08\ \cancel{g}} = 1.71\,mol\,ethanol$$

? Exercise 2.10.5

If the density of benzene, C_6H_6, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer

0.201 mol

Summary

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

CHAPTER OVERVIEW

3: Matter

3.1: Definition

Learning Objectives

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

We are all familiar with matter. The definition of **Matter** is anything that has mass and volume (takes up space). For most common objects that we deal with every day, 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 amounts 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 3.1.1: Everything from an ant, to a truck, to the earth, and even the entire galaxy is composed of matter. Images used with permission from Wikipedia (CC_SA-BY-3.0; credit High Contrast).

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 you 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".

Contributions & Attributions

- Wikipedia

3.2: States of Matter (Solid, Liquid and Gas)

Learning Objectives

- To describe the solid, liquid and gas phases.

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

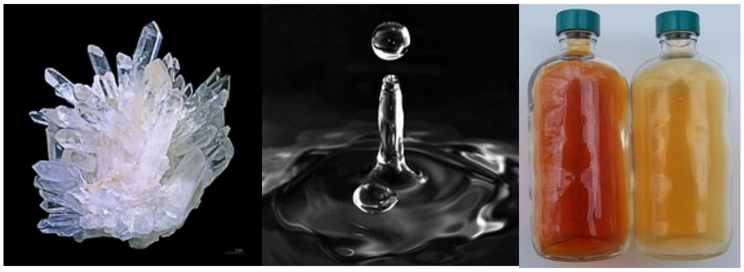


Figure $3.2.1$: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure $3.2.2$ 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.

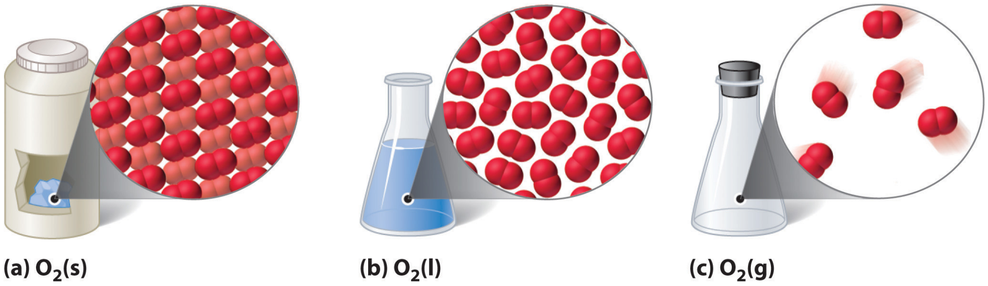
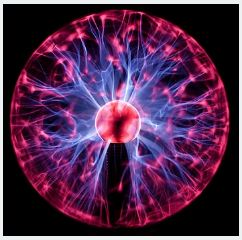


Figure $3.2.2$: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous 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 does not naturally occur on earth, so we will omit it from our study here.



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

Solids

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

Solids are defined by the following characteristics:

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

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

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 $3.2.3$. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

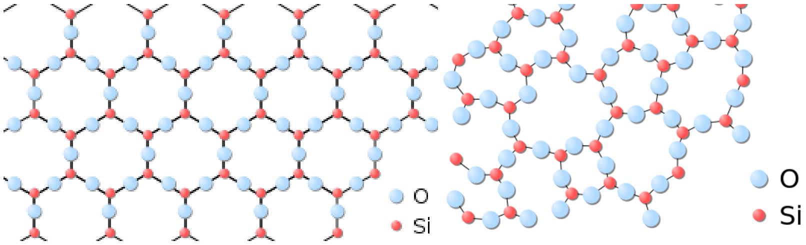


Figure $3.2.3$: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view 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.

Video $3.2.1$: Mercury boiling to become a gas.

If we heat liquid mercury 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.

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 1,000 or more. Gases have the following characteristics:

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

Table 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 $3.2.1$

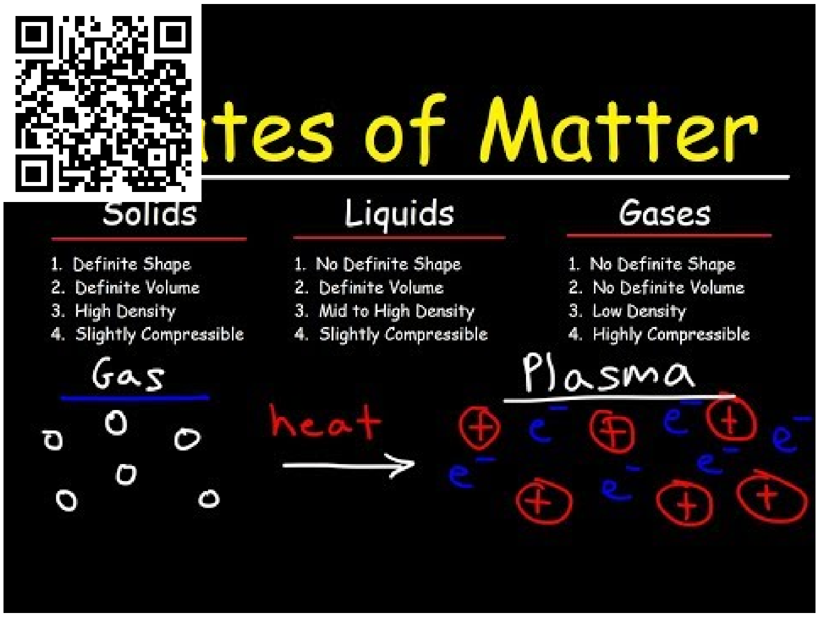
What state or states of matter does each statement, describe?

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.

Video: **States of Matter by The Organic Chemistry Tutor**



States of Matter

Solids
1. Definite Shape
2. Definite Volume
3. High Density
4. Slightly Compressible

Liquids
1. No Definite Shape
2. Definite Volume
3. Mid to High Density
4. Slightly Compressible

Gases
1. No Definite Shape
2. No Definite Volume
3. Low Density
4. Highly Compressible

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.

3.3: Classification of Matter (Pure Substances and Mixtures)

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.

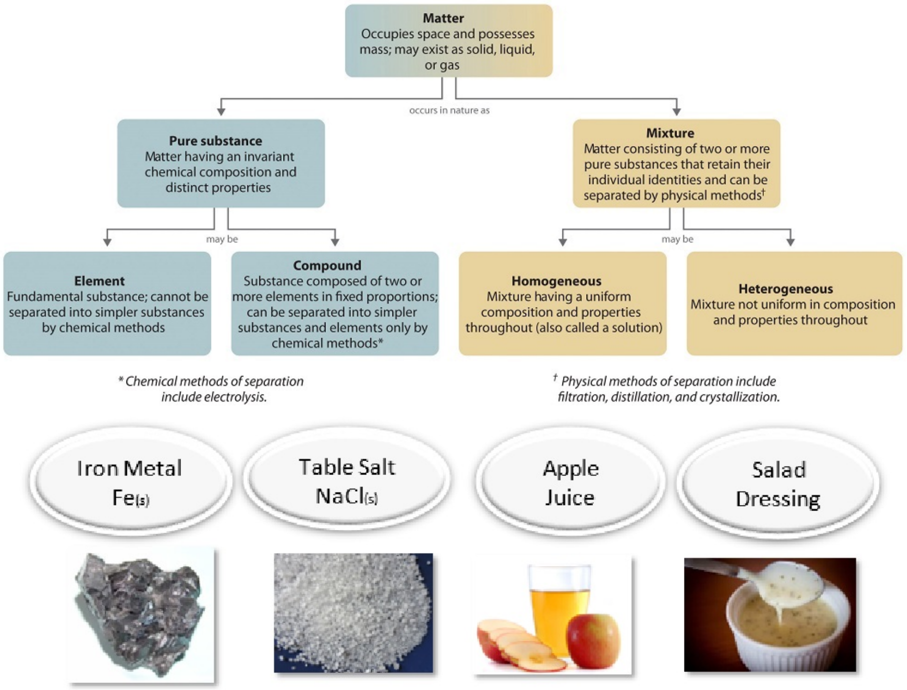


Figure 3.3.1: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

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

A **homogeneous** mixture is a mixture in which the composition is uniform throughout the mixture. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the 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

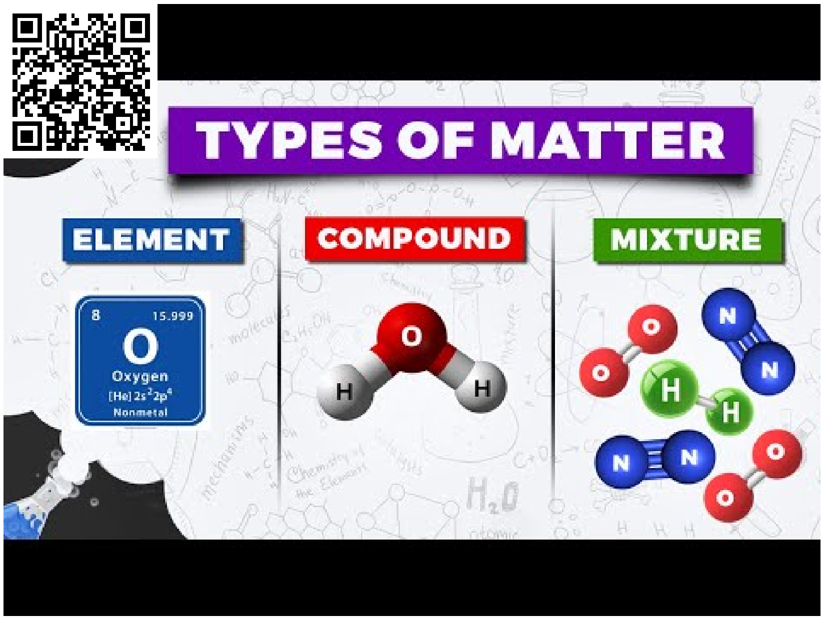
Strategy:

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

Solution

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

Video: **Types of Matter by Professor Dave Explains**



? 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 substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
d. Oxygen, a 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. Mixtures are physical combinations of two or more elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

Vocabulary

- Element: a substance that is made up of only one type of atom.
- Compound:a substance that is made up of more than one type of atom bonded together.
- Mixture: a combination of two or more elements or compounds which have not reacted to bond together; each part in the mixture retains its own properties.

Contributions & Attributions

-
- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook

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- **3.4: Classifying Matter According to Its Composition** by Henry Agnew, Marisa Alviar-Agnew, Stephen Lower is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

3.4: Properties of Matter (Chemical and Physical)

Learning Objectives

To separate 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 point, and boiling point.

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/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

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

Figure 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: Contrasting Physical and Chemical Properties

Physical Properties	Chemical Properties
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

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

? Exercise $3.4.1A$

Which of the following is a physical property of matter?

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

Answer

c

? Exercise $3.4.1B$

Which of the following is a chemical property?

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

Answer

a

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. 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.

3.5: Changes in Matter (Physical and Chemical)

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

Figure 3.5.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

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

Observations that help to indicate chemical change include:

- Temperature changes (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. A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
d. Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

Solution

a. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written: $H_2O\,(l) \rightarrow H_2O\,(g)$
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. The following reaction occurs: $Fe + O_2 \rightarrow Fe_2O_3$

c. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
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 corroded 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 some physical property, such as differences in their boiling points. Two of these separation methods are distillation and crystallization. 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. 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.

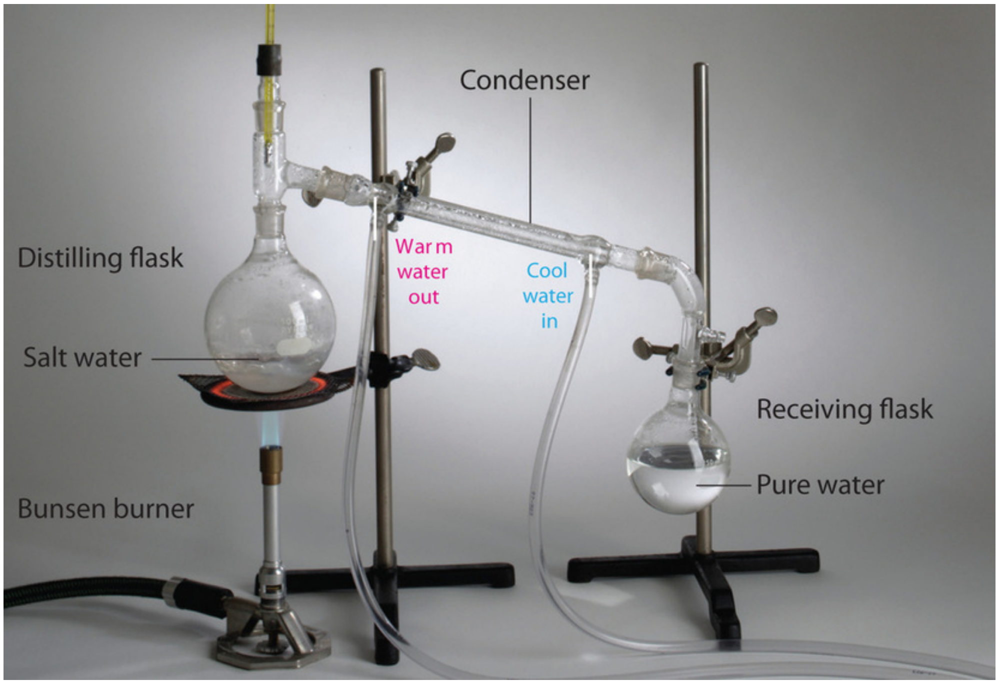


Figure $3.5.3$: *The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the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

Parts of a distillation setup: Bunsen burner, salt water in distilling flask, condenser with cool water in and warm water out, pure water in 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.

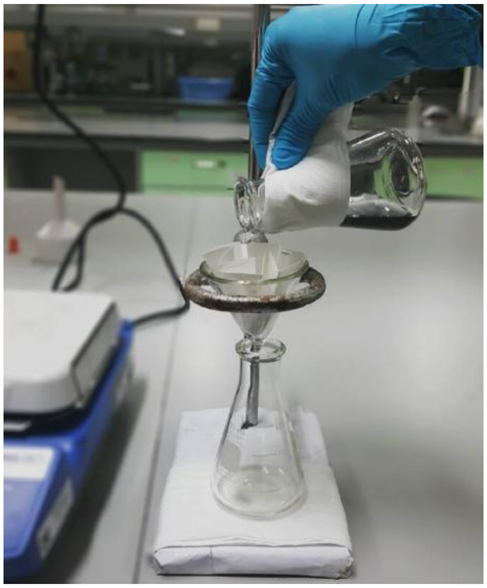


Figure 3.5.4*: Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).*

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.

Contributions & Attributions

-
- Boundless (www.boundless.com)

CHAPTER OVERVIEW

4: Atoms and Periodic Table

4.1: Indivisible - The Atomic Theory

Learning Objectives

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

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

Democritus and the Greek Philosophers

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

Figure 4.1.1: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

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

Greeks: "All Thought and No Action"

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

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

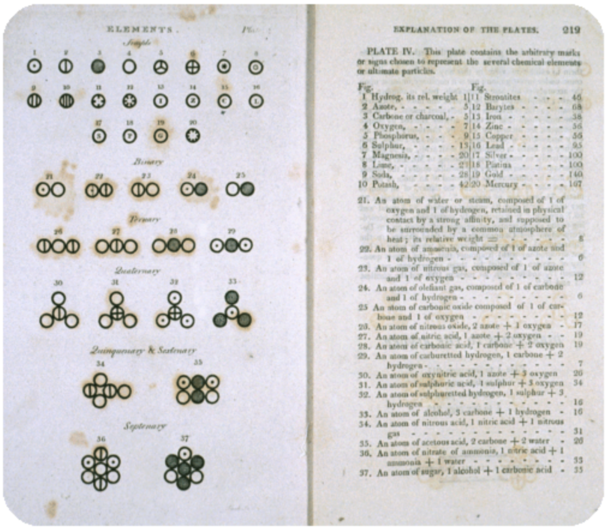
The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 4.1.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

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

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

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

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



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.

Fig.			Fig.		
1	Hydrog. its rel. weight	1	11	Strontites	46
2	Azote,	5	12	Barytes	68
3	Carbone or charcoal,	5	13	Iron	38
4	Oxygen,	7	14	Zinc	56
5	Phosphorus,	9	15	Copper	56
6	Sulphur,	13	16	Lead	95
7	Magnesia,	20	17	Silver	100
8	Lime,	23	18	Platina	100
9	Soda,	28	19	Gold	140
10	Potash,	42	20	Mercury	167

21. An atom of water or steam, composed of 1 of oxygen and 1 of hydrogen, retained in physical contact by a strong affinity, and supposed to be surrounded by a common atmosphere of heat; its relative weight = 8
22. An atom of ammonia, composed of 1 of azote and 1 of hydrogen 6
23. An atom of nitrous gas, composed of 1 of azote and 1 of oxygen 12
24. An atom of olefiant gas, composed of 1 of carbone and 1 of hydrogen 6
25 An atom of carbonic oxide composed of 1 of carbone and 1 of oxygen 12
26. An atom of nitrous oxide, 2 azote + 1 oxygen 17
27. An atom of nitric acid, 1 azote + 2 oxygen 19
28. An atom of carbonic acid, 1 carbone + 2 oxygen 19
29. An atom of carburetted hydrogen, 1 carbone + 2 hydrogen 7
30. An atom of oxynitric acid, 1 azote + 3 oxygen 26
31. An atom of sulphuric acid, 1 sulphur + 3 oxygen 34
32. An atom of sulphuretted hydrogen, 1 sulphur + 3 hydrogen 16
33. An atom of alcohol, 3 carbone + 1 hydrogen 16
34. An atom of nitrous acid, 1 nitric acid + 1 nitrous gas 31
35. An atom of acetous acid, 2 carbone + 2 water 26
36. An atom of nitrate of ammonia, 1 nitric acid + 1 ammonia + 1 water 33
37. An atom of sugar, 1 alcohol + 1 carbonic acid 35

Figure 4.1.2: 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 4.1.1, control of individual atoms can be use used create animations.

Video 4.1.1: 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.

- **4.2: Indivisible - The Atomic Theory** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

4.2: The Properties of 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. 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. Unlike protons and neutrons, which consist of smaller, simpler particles, electrons are fundamental particles that do not consist of smaller particles. They are a type of fundamental particle called leptons. All leptons have an electric charge of -1 or 0. Electrons are extremely small. The mass of an electron is only about 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.

Unlike protons and neutrons, which are located inside the nucleus at the center of the atom, electrons are found outside the nucleus. Because opposite electric charges attract one another, negative electrons are attracted to the positive nucleus. This force of attraction keeps electrons constantly moving through the otherwise empty space around the nucleus. The figure below is a common way to represent the structure of an atom. It shows the electron as a particle orbiting the nucleus, similar to the way that planets orbit the sun. However, this is an incorrect perspective, as quantum mechanics demonstrates that electrons are more complicated.

Figure 4.2.1: 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 one of three main particles that make up the atom. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one $(+1)$ and a mass of 1 atomic mass unit (amu), which is about 1.67×10^{-27} kilograms. Together with neutrons, they make up virtually all of the mass of an atom.

Neutrons

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 diagram above are labeled n^0. The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic mass unit (amu). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they are bound together with other neutrons and protons in the atomic nucleus.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons, and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Table 4.2.1: Properties of Subatomic Particles

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

Table $4.2.1$ gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit** (amu) is defined as one-twelfth of the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

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

How are these subatomic particles arranged in atoms?

How are these subatomic particles arranged in atoms? They are not arranged at random. Experiments by Ernest Rutherford in England in the 1910's pointed to a nuclear model with atoms that has the protons and neutrons in a central nucleus with the electrons in orbit about the nucleus. The relatively massive protons and neutrons are collected in the center of an atom, in a region called the **nucleus** of the atom (plural *nuclei*). The electrons are outside the nucleus and spend their time orbiting in space about the nucleus. (Figure 4.2.2). Because protons and neutrons are so massive compared to electrons, Table 4.2.1, nearly all of the mass of an atom is contained in the nucleus.

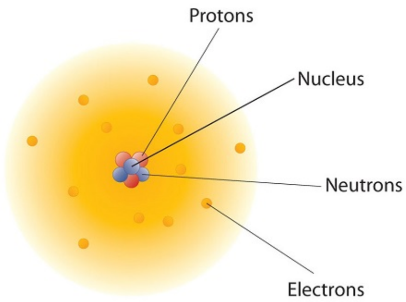
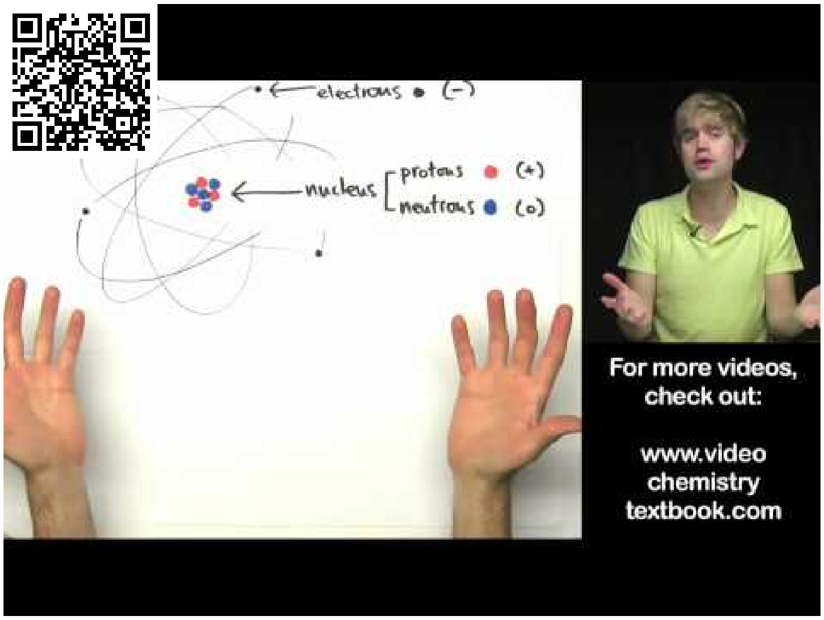


Figure 4.2.2: The Structure of the Atom. Atoms have protons and neutrons in the center, making the nucleus, while the electrons orbit the nucleus.

Video: **Basic Atomic Structure by Tyler DeWitt**



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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4.3: Elements - Defined by Their Number of Protons

Learning Objectives

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

It's important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up all other 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 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 $4.3.1$: It is difficult to find qualities that differ between each element, and to distinguish one element from another. Each element, however, does have 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).

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table $4.3.1$). If an atom has only one proton, we know that it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

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

Table $4.3.1$: Atoms of the First Six Elements

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

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

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table $4.3.1$). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (amu). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table $4.3.1$).

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

Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball). Counting the number of protons and neutrons tells scientists about the total mass of an atom.

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

An atom's mass number is very easy to calculate, provided that you know the number of protons and neutrons in an atom.

✓ Example 4.5.1

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

Solution

$(\text{number of protons}) = 2$ (Remember that an atom of helium always has 2 protons.)

$(\text{number of neutrons}) = 2$

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

$$\text{mass number} = \mathbf{2 + 2 = 4}$$

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 $4.3.2$.

Table $4.3.2$: *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

Chemical Symbol	Name	Latin Name
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

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.
- The mass number of an atom is the sum of the protons and neutrons in the atom.
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

4.4: Isotopes

Learning Objectives

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

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

An Example: Hydrogen Isotopes

Hydrogen is an example of an element that has isotopes. Three isotopes of hydrogen are modeled in Figure 4.4.1. Most hydrogen atoms have just one proton, one electron, and lack a neutron. These atoms are just called hydrogen. Some hydrogen atoms have one neutron as well. These atoms are the isotope named deuterium. Other hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

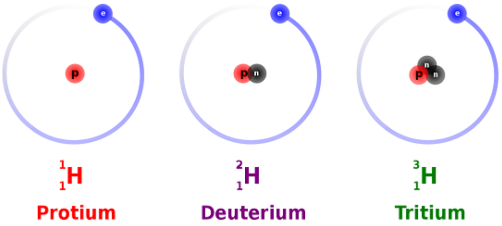


Figure 4.4.1: The three most stable isotopes of hydrogen: protium (A = 1), deuterium (A = 2), and tritium (A = 3). (CC SA-BY 3.0; Balajijagadesh via Wikipedia).

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

Example 4.4.1: Lithium Isotopes

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

Solution

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

a.

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

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

b.

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

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

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

Stability of Isotopes

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

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

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

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

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

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

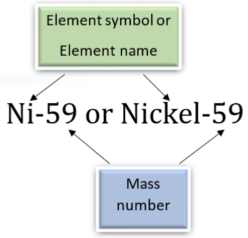
The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$${}^{59}_{28}\text{Ni}$$

$${}^{238}_{92}\text{U}$$

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

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



✓ Example 4.4.2: Potassium-40

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

Solution

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

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

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

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

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

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

✓ Example 4.4.3: Zinc-65

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

Solution

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

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

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

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

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

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

? Exercise 4.4.3

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

a. ${}^{60}_{27}\text{Co}$
b. Na-24

c. $^{45}_{20}\text{Ca}$
d. Sr-90

Answer a:

27 protons, 27 electrons, 33 neutrons

Answer b:

11 protons, 11 electrons, 13 neutrons

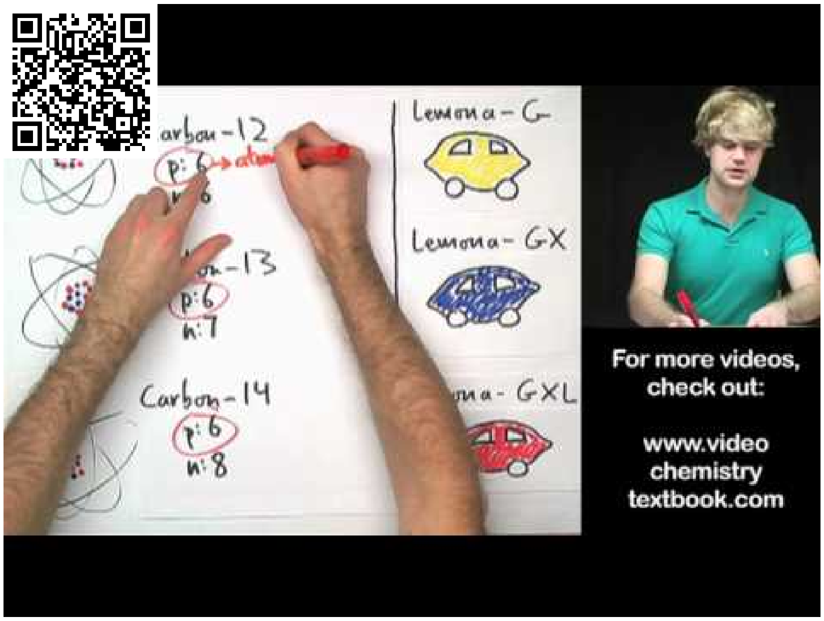
Answer c:

20 protons, 20 electrons, 25 neutrons

Answer d:

38 protons, 38 electrons, 52 neutrons

Video: **What are Isotopes? by Tyler DeWitt**



Summary

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

4.5: Periodic Table

Learning Objectives

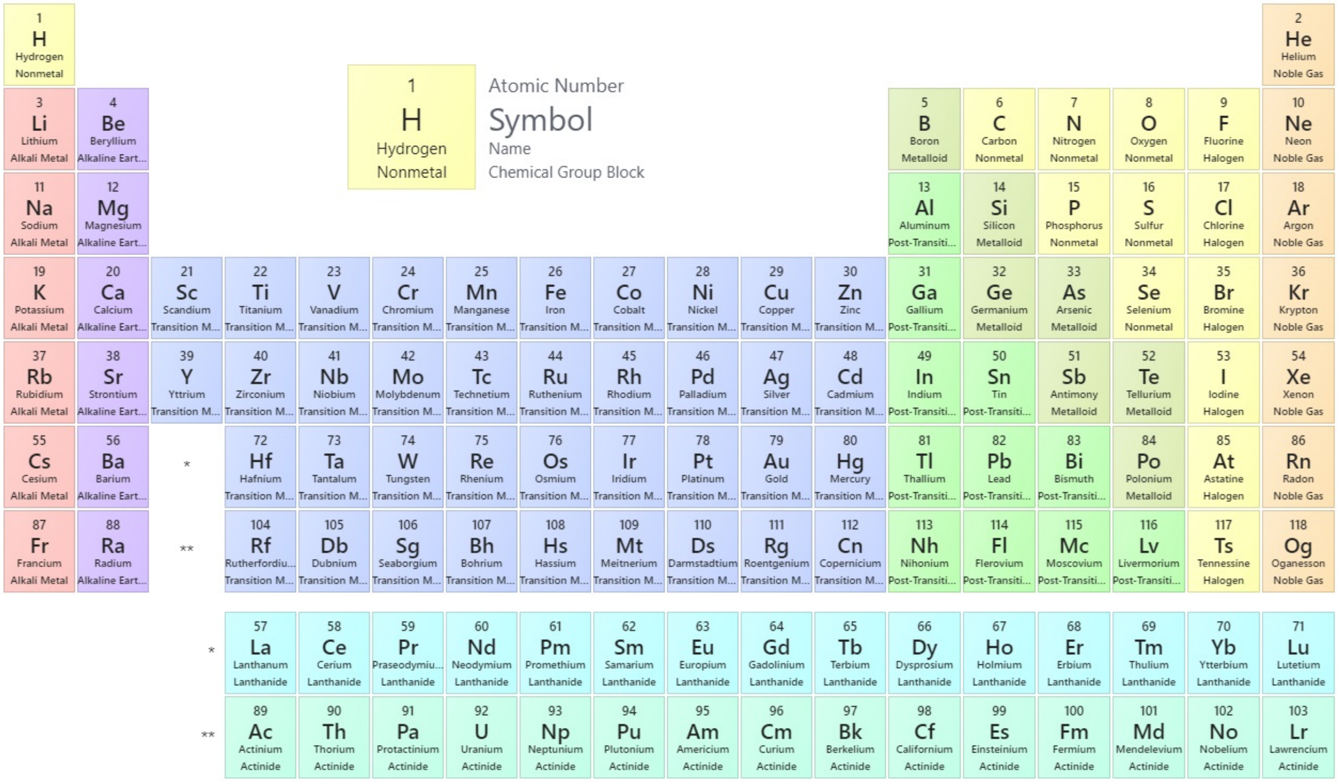
- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

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, 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.

> *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 4.5.1. 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.



1 H Hydrogen Nonmetal	Atomic Number Symbol Name Chemical Group Block

1 H Hydrogen Nonmetal																	2 He Helium Noble Gas
3 Li Lithium Alkali Metal	4 Be Beryllium Alkaline Eart...											5 B Boron Metalloid	6 C Carbon Nonmetal	7 N Nitrogen Nonmetal	8 O Oxygen Nonmetal	9 F Fluorine Halogen	10 Ne Neon Noble Gas
11 Na Sodium Alkali Metal	12 Mg Magnesium Alkaline Eart...											13 Al Aluminum Post-Transiti...	14 Si Silicon Metalloid	15 P Phosphorus Nonmetal	16 S Sulfur Nonmetal	17 Cl Chlorine Halogen	18 Ar Argon Noble Gas
19 K Potassium Alkali Metal	20 Ca Calcium Alkaline Eart...	21 Sc Scandium Transition M...	22 Ti Titanium Transition M...	23 V Vanadium Transition M...	24 Cr Chromium Transition M...	25 Mn Manganese Transition M...	26 Fe Iron Transition M...	27 Co Cobalt Transition M...	28 Ni Nickel Transition M...	29 Cu Copper Transition M...	30 Zn Zinc Transition M...	31 Ga Gallium Post-Transiti...	32 Ge Germanium Metalloid	33 As Arsenic Metalloid	34 Se Selenium Nonmetal	35 Br Bromine Halogen	36 Kr Krypton Noble Gas
37 Rb Rubidium Alkali Metal	38 Sr Strontium Alkaline Eart...	39 Y Yttrium Transition M...	40 Zr Zirconium Transition M...	41 Nb Niobium Transition M...	42 Mo Molybdenum Transition M...	43 Tc Technetium Transition M...	44 Ru Ruthenium Transition M...	45 Rh Rhodium Transition M...	46 Pd Palladium Transition M...	47 Ag Silver Transition M...	48 Cd Cadmium Transition M...	49 In Indium Post-Transiti...	50 Sn Tin Post-Transiti...	51 Sb Antimony Metalloid	52 Te Tellurium Metalloid	53 I Iodine Halogen	54 Xe Xenon Noble Gas
55 Cs Cesium Alkali Metal	56 Ba Barium Alkaline Eart...	*	72 Hf Hafnium Transition M...	73 Ta Tantalum Transition M...	74 W Tungsten Transition M...	75 Re Rhenium Transition M...	76 Os Osmium Transition M...	77 Ir Iridium Transition M...	78 Pt Platinum Transition M...	79 Au Gold Transition M...	80 Hg Mercury Transition M...	81 Tl Thallium Post-Transiti...	82 Pb Lead Post-Transiti...	83 Bi Bismuth Post-Transiti...	84 Po Polonium Metalloid	85 At Astatine Halogen	86 Rn Radon Noble Gas
87 Fr Francium Alkali Metal	88 Ra Radium Alkaline Eart...	**	104 Rf Rutherfordiu... Transition M...	105 Db Dubnium Transition M...	106 Sg Seaborgium Transition M...	107 Bh Bohrium Transition M...	108 Hs Hassium Transition M...	109 Mt Meitnerium Transition M...	110 Ds Darmstadtium Transition M...	111 Rg Roentgenium Transition M...	112 Cn Copernicium Transition M...	113 Nh Nihonium Post-Transiti...	114 Fl Flerovium Post-Transiti...	115 Mc Moscovium Post-Transiti...	116 Lv Livermorium Post-Transiti...	117 Ts Tennessine Halogen	118 Og Oganesson Noble Gas
		*	57 La Lanthanum Lanthanide	58 Ce Cerium Lanthanide	59 Pr Praseodymiu... Lanthanide	60 Nd Neodymium Lanthanide	61 Pm Promethium Lanthanide	62 Sm Samarium Lanthanide	63 Eu Europium Lanthanide	64 Gd Gadolinium Lanthanide	65 Tb Terbium Lanthanide	66 Dy Dysprosium Lanthanide	67 Ho Holmium Lanthanide	68 Er Erbium Lanthanide	69 Tm Thulium Lanthanide	70 Yb Ytterbium Lanthanide	71 Lu Lutetium Lanthanide
		**	89 Ac Actinium Actinide	90 Th Thorium Actinide	91 Pa Protactinium Actinide	92 U Uranium Actinide	93 Np Neptunium Actinide	94 Pu Plutonium Actinide	95 Am Americium Actinide	96 Cm Curium Actinide	97 Bk Berkelium Actinide	98 Cf Californium Actinide	99 Es Einsteinium Actinide	100 Fm Fermium Actinide	101 Md Mendelevium Actinide	102 No Nobelium Actinide	103 Lr Lawrencium Actinide

Figure 4.5.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found Periodic Table of Elements. (Public Domain; PubChem via NIH)

Features of the Periodic Table

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

> *The word halogen comes from the Greek for "salt maker" because these elements combine with other elements to form a group of compounds called salts.*

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.

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

Metals, Nonmetals, and Metalloids

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure $4.5.2$. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure $4.5.2$, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

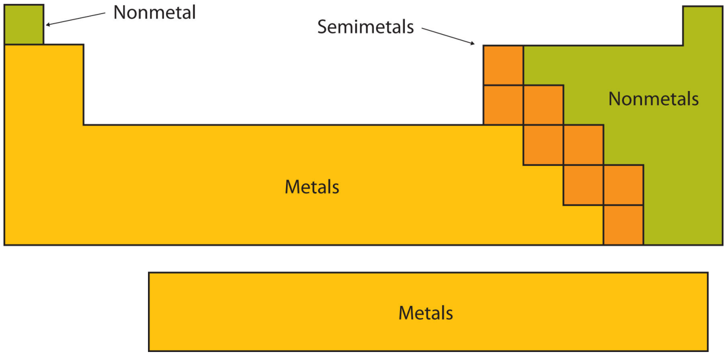


Figure $4.5.2$: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

✓ Example $4.5.1$

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

a. Se
b. Mg
c. Ge

Solution

a. In Figure $4.5.1$, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

? Exercise $4.5.1$

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

Answer

Indium is a metal.

Representative, Transition, and Inner-transition

Another way to categorize the elements of the periodic table is shown in Figure $4.5.3$. 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 between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

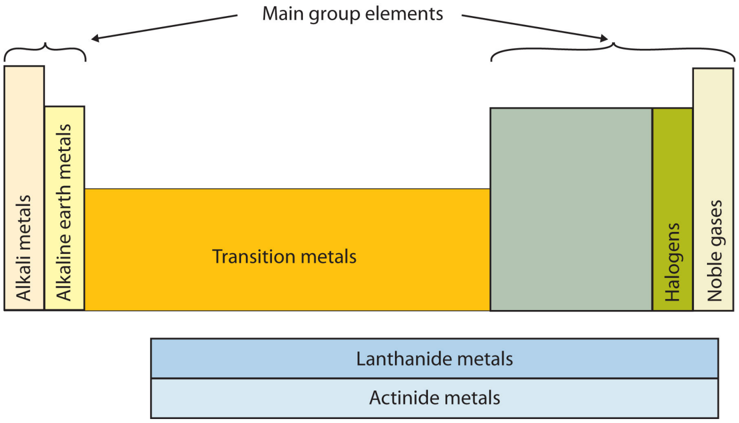


Figure $4.5.3$: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Sections of period table: Alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanide metals, actinide metals.

Descriptive Names

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

Group 1: The Alkali Metals

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

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

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

Group 2: The Alkaline Earth Metals

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

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for "salt forming," which reflects that all 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: 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 4.5.2: 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 4.5.2: 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 4.5.3: Classification of Elements

Classify each element as metal, non metal, 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 4.5.3: Classification of Elements

Classify each element as metal, non metal, 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 a inner transition metal too).

Answer c:

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

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 semi metals. Semi metals exhibit properties intermediate between those of metals and nonmetals. Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of semi metals. Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors. 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.

Contributions & Attributions

4.6: Atomic Mass - The Average Mass of an Element’s Atoms

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.

In chemistry we very rarely deal with only one isotope of an element. We use a mixture of the isotopes of an element in chemical reactions and other aspects of chemistry, because all of the isotopes of an element react in the same manner. That means that we rarely need to worry about the mass of a specific isotope, but instead we need to know the average mass of the atoms of an element. Using the masses of the different isotopes and how abundant each isotope is, we can find the average mass of the atoms of an element. The **atomic mass** of an element is the weighted average mass of the atoms in a naturally occurring sample of the element. Atomic mass is typically reported in atomic mass units.

Calculating Atomic Mass

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

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

Look carefully to see how this equation is used in the following examples.

Example $4.6.1$: Boron Isotopes

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

Solution

Boron has two isotopes. We will use the equation:

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

- Isotope 1: $\%_1 = 0.20$ (Write all percentages as decimals), $\text{mass}_1 = 10$
- Isotope 2: $\%_2 = 0.80$, $\text{mass}_2 = 11$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.20)(10) + (0.80)(11)$$

$$\text{Atomic mass} = 10.8\,\text{amu}$$

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

Example $4.6.2$: Neon Isotopes

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

Solution

Neon has three isotopes. We will use the equation:

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

- Isotope 1: $\%_1 = 0.9092$ (write all percentages as decimals), $\text{mass}_1 = 19.99$

- Isotope 2: $\%_2 = 0.003$, $\text{mass}_2 = 20.99$
- Isotope 3: $\%_3 = 0.0885$, $\text{mass}_3 = 21.99$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99)$$

$$\text{Atomic mass} = 20.17 \text{ amu}$$

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

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that the atomic mass of boron (symbol B) is 10.8, which is what we calculated in Example $4.6.1$, and the atomic mass of neon (symbol Ne) is 20.8, which is what we calculated in Example $4.6.2$. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

Exercise $4.6.1$

Chlorine has two naturally occurring isotopes. In a sample of chlorine, 75.77% of the atoms are Cl-35, with a mass of 34.97 amu Another 24.23% of the atoms are Cl-37, with a mass of 36.97 amu What is the atomic mass of chlorine?

Answer

35.45 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.

4.7: Valence vs Core Electrons

Learning Objectives

- Identify valence electrons using the periodic table and electron configuration.
- Define core and valence electrons.

Electron Shells and the Bohr Model

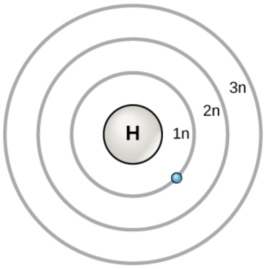


Figure 1. The Bohr model.

It should be stressed that there is a connection between the number of protons in an element, the atomic number that distinguishes one element from another, and the number of electrons it has. In all electrically neutral atoms, the number of electrons is the same as the number of protons. Thus, each element, at least when electrically neutral, has a characteristic number of electrons equal to its atomic number.

An early model of the atom was developed in 1913 by Danish scientist Niels Bohr (1885–1962). In this model, electrons exist within principal shells. An electron normally exists in the lowest energy shell available, which is the one closest to the nucleus. Energy from a photon of light can bump it up to a higher energy shell, but this situation is unstable, and the electron quickly decays back to the ground state. In the process, a photon of light is released.

The Bohr model shows the atom as a central nucleus containing protons and neutrons, with the electrons in circular **orbitals** at specific distances from the nucleus, as illustrated in Figure 1. These orbits form electron shells or energy levels, which are a way of visualizing the number of electrons in the outermost shells. These energy levels are designated by a number and the symbol “n.” For example, 1n represents the first energy level located closest to the nucleus.

Electrons fill orbitals in a consistent order: they first fill the orbitals closest to the nucleus, then they continue to fill orbitals of increasing energy further from the nucleus. If there are multiple orbitals of equal energy, they will be filled with one electron in each energy level before a second electron is added. The electrons of the outermost energy level determine the energetic stability of the atom and its tendency to form chemical bonds with other atoms to form molecules.

Under standard conditions, atoms fill the inner shells first, often resulting in a variable number of electrons in the outermost shell. The innermost shell has a maximum of two electrons but the next two electron shells can each have a maximum of eight electrons. This is known as the **octet rule**, which states, with the exception of the innermost shell, that atoms are more stable energetically when they have eight electrons in their **valence shell**, the outermost electron shell. Examples of some neutral atoms and their electron configurations are shown in Figure 2. Notice that in this figure, helium has a complete outer electron shell, with two electrons filling its first and only shell. Similarly, neon has a complete outer 2n shell containing eight electrons. In contrast, chlorine and sodium have seven and one in their outer shells, respectively, but theoretically they would be more energetically stable if they followed the octet rule and had eight.

Practice Question

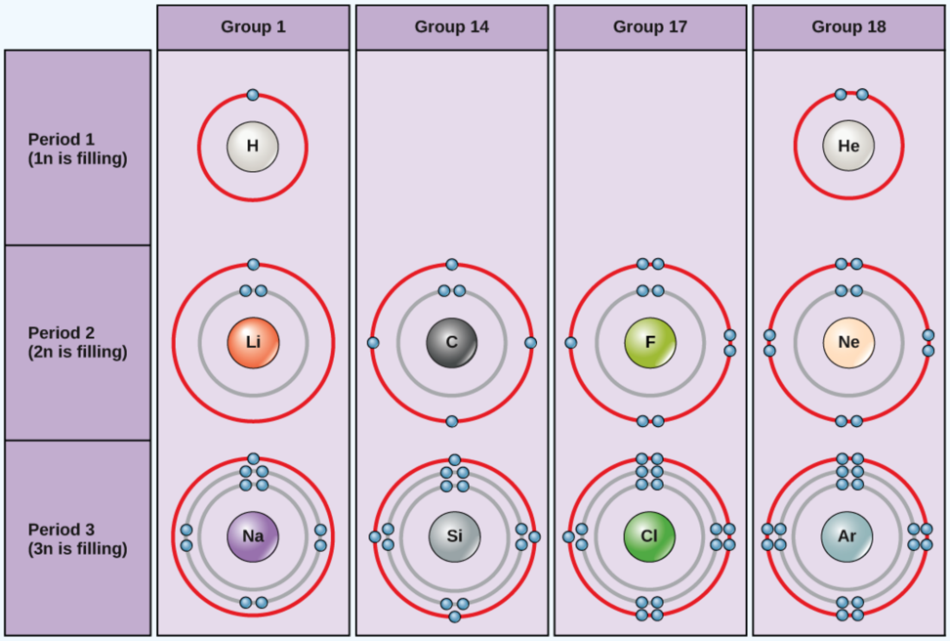


Figure 2. Bohr diagrams for hydrogen, helium, lithium, carbon, fluorine, neon, sodium, silicon, chlorine, and argon.

Bohr diagrams indicate how many electrons fill each principal shell. Group 18 elements (helium, neon, and argon are shown in Figure 2) have a full outer, or valence, shell. A full valence shell is the most stable electron configuration. Elements in other groups have partially filled valence shells and gain or lose electrons to achieve a stable electron configuration.

An atom may give, take, or share electrons with another atom to achieve a full valence shell, the most stable electron configuration. Looking at this figure, how many electrons do elements in group 1 need to lose in order to achieve a stable electron configuration? How many electrons do elements in groups 14 and 17 need to gain to achieve a stable configuration?

[practice-area rows=”2″][/practice-area]
[reveal-answer q=”766153″]**Show Answer**[/reveal-answer]
[hidden-answer a=”766153″]Elements in group 1 need to lose one electron to achieve a stable electron configuration. Elements in groups 14 and 17 need to gain four and one electrons, respectively, to achieve a stable configuration.[/hidden-answer]

Understanding that the organization of the periodic table is based on the total number of protons (and electrons) helps us know how electrons are distributed among the outer shell. The periodic table is arranged in columns and rows based on the number of electrons and where these electrons are located. Take a closer look at the some of the elements in the periodic table’s far right column in Figure 3.

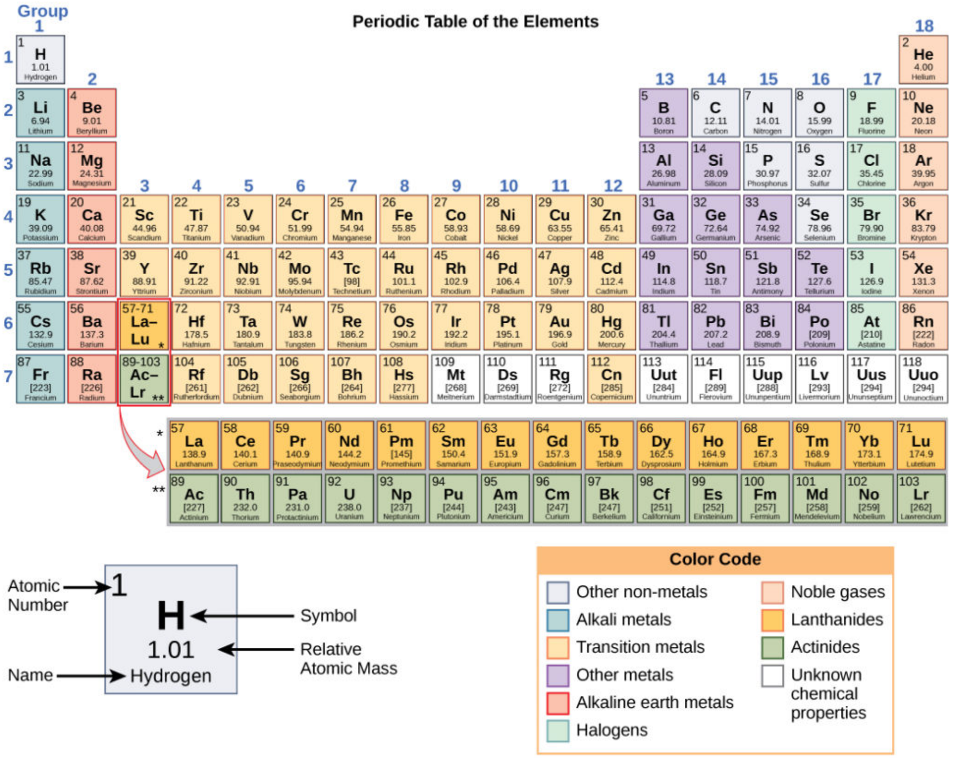


Figure 3. The periodic table shows the atomic mass and atomic number of each element. The atomic number appears above the symbol for the element and the approximate atomic mass appears below it.

The group 18 atoms helium (He), neon (Ne), and argon (Ar) all have filled outer electron shells, making it unnecessary for them to share electrons with other atoms to attain stability; they are highly stable as single atoms. Their non-reactivity has resulted in their being named the **inert gases** (or **noble gases**). Compare this to the group 1 elements in the left-hand column. These elements, including hydrogen (H), lithium (Li), and sodium (Na), all have one electron in their outermost shells. That means that they can achieve a stable configuration and a filled outer shell by donating or sharing one electron with another atom or a molecule such as water. Hydrogen will donate or share its electron to achieve this configuration, while lithium and sodium will donate their electron to become stable. As a result of losing a negatively charged electron, they become positively charged **ions**. Group 17 elements, including fluorine and chlorine, have seven electrons in their outmost shells, so they tend to fill this shell with an electron from other atoms or molecules, making them negatively charged ions. Group 14 elements, of which carbon is the most important to living systems, have four electrons in their outer shell allowing them to make several covalent bonds (discussed below) with other atoms. Thus, the columns of the periodic table represent the potential shared state of these elements’ outer electron shells that is responsible for their similar chemical characteristics.

Electron Orbitals

Although useful to explain the reactivity and chemical bonding of certain elements, the Bohr model of the atom does not accurately reflect how electrons are spatially distributed surrounding the nucleus. They do not circle the nucleus like the earth orbits the sun, but are found in **electron orbitals**. These relatively complex shapes result from the fact that electrons behave not just like particles, but also like waves. Mathematical equations from quantum mechanics known as wave functions can predict within a certain level of probability where an electron might be at any given time. The area where an electron is most likely to be found is called its orbital.

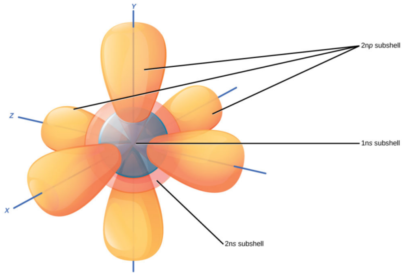


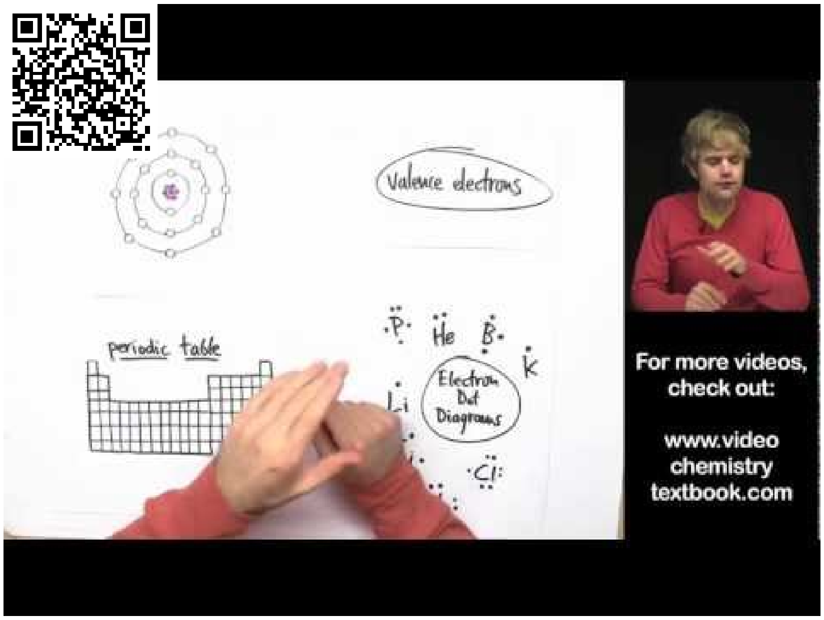
Figure 4. Click for a larger image. The *s* subshells are shaped like spheres. Both the 1n and 2n principal shells have an *s* orbital, but the size of the sphere is larger in the 2n orbital. Each sphere is a single orbital. *p* subshells are made up of three dumbbell-shaped orbitals. Principal shell 2n has a *p* subshell, but shell 1 does not.

Recall that the Bohr model depicts an atom's electron shell configuration. Within each electron shell are subshells, and each subshell has a specified number of orbitals containing electrons. While it is impossible to calculate exactly where an electron is located, scientists know that it is most probably located within its orbital path. Subshells are designated by the letters *s*, *p*, *d*, and *f*. The *s* subshell is spherical in shape and has one orbital. Principal shell 1n has only a single *s* orbital, which can hold two electrons. Principal shell 2n has one *s* and one *p* subshell, and can hold a total of eight electrons. The *p* subshell has three dumbbell-shaped orbitals, as illustrated in Figure 4. Subshells *d* and *f* have more complex shapes and contain five and seven orbitals, respectively. These are not shown in the illustration. Principal shell 3n has *s*, *p*, and *d* subshells and can hold 18 electrons. Principal shell 4n has *s*, *p*, *d* and *f* orbitals and can hold 32 electrons. Moving away from the nucleus, the number of electrons and orbitals found in the energy levels increases. Progressing from one atom to the next in the periodic table, the electron structure can be worked out by fitting an extra electron into the next available orbital.

The closest orbital to the nucleus, called the 1*s* orbital, can hold up to two electrons. This orbital is equivalent to the innermost electron shell of the Bohr model of the atom. It is called the 1*s* orbital because it is spherical around the nucleus. The 1*s* orbital is the closest orbital to the nucleus, and it is always filled first, before any other orbital can be filled. Hydrogen has one electron; therefore, it has only one spot within the 1*s* orbital occupied. This is designated as $1s^1$, where the superscripted 1 refers to the one electron within the 1*s* orbital. Helium has two electrons; therefore, it can completely fill the 1*s* orbital with its two electrons. This is designated as $1s^2$, referring to the two electrons of helium in the 1*s* orbital. On the periodic table Figure 4, hydrogen and helium are the only two elements in the first row (period); this is because they only have electrons in their first shell, the 1*s* orbital. Hydrogen and helium are the only two elements that have the 1*s* and no other electron orbitals in the electrically neutral state.

The second electron shell may contain eight electrons. This shell contains another spherical *s* orbital and three "dumbbell" shaped *p* orbitals, each of which can hold two electrons, as shown in Figure 4. After the 1*s* orbital is filled, the second electron shell is filled, first filling its 2*s* orbital and then its three *p* orbitals. When filling the *p* orbitals, each takes a single electron; once each *p* orbital has an electron, a second may be added. Lithium (Li) contains three electrons that occupy the first and second shells. Two electrons fill the 1*s* orbital, and the third electron then fills the 2*s* orbital. Its **electron configuration** is $1s^2 2s^1$. Neon (Ne), on the other hand, has a total of ten electrons: two are in its innermost 1*s* orbital and eight fill its second shell (two each in the 2*s* and three *p* orbitals); thus, it is an inert gas and energetically stable as a single atom that will rarely form a chemical bond with other atoms. Larger elements have additional orbitals, making up the third electron shell. While the concepts of electron shells and orbitals are closely related, orbitals provide a more accurate depiction of the electron configuration of an atom because the orbital model specifies the different shapes and special orientations of all the places that electrons may occupy.

Video: **Valence Electrons and the Periodic Table by Tyler DeWitt**



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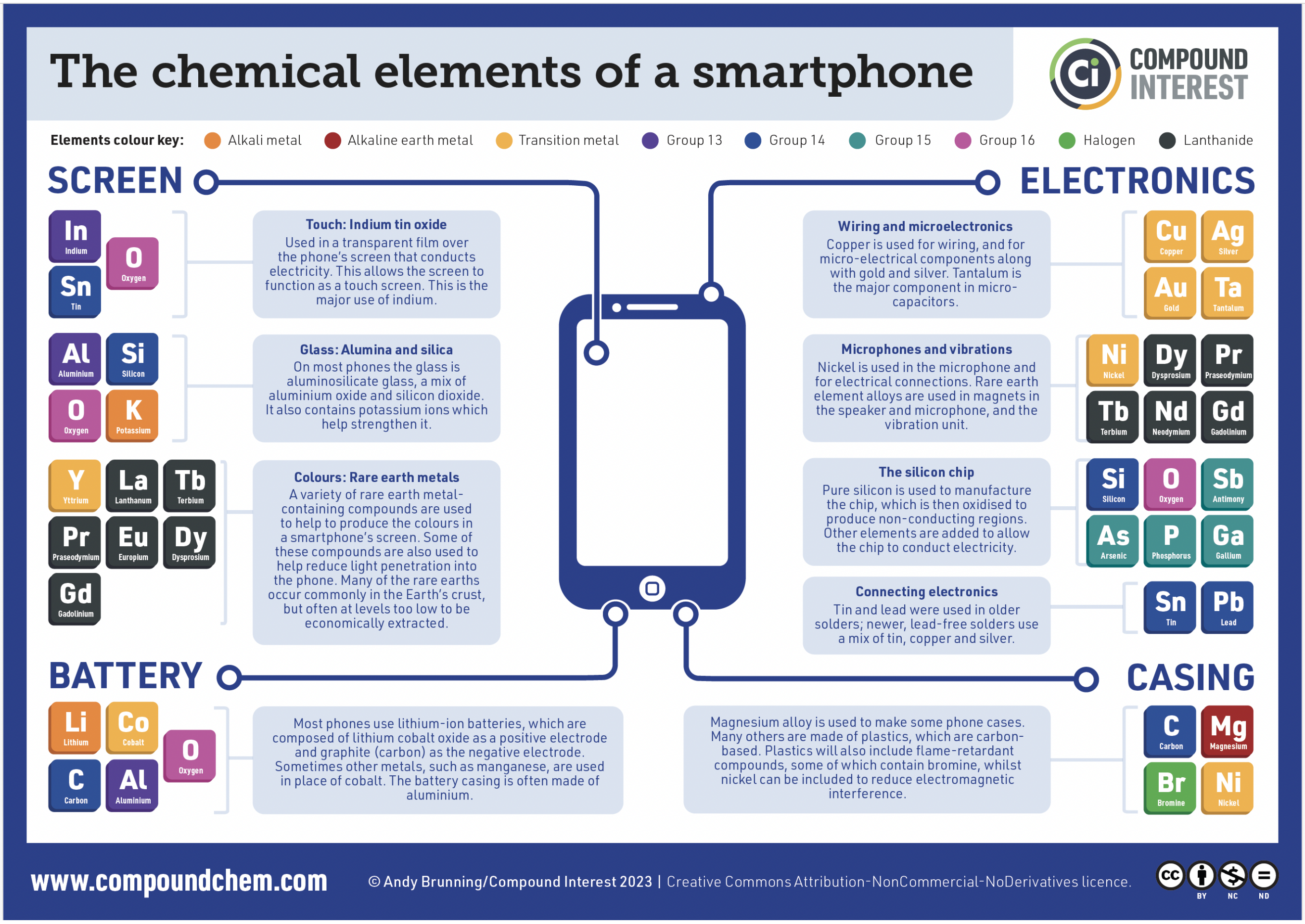
4.8: Content in Context

This section will explore the chemical elements used in smartphone production and highlight those at risk of depletion.

4.8.1: The Chemical Elements of a Smartphones

Video: What's in your Smartphone

The Chemical Elements of a Smartphones:



The chemical elements of a smartphone

COMPOUND INTEREST

Elements colour key: ● Alkali metal ● Alkaline earth metal ● Transition metal ● Group 13 ● Group 14 ● Group 15 ● Group 16 ● Halogen ● Lanthanide

SCREEN

Touch: Indium tin oxide
Used in a transparent film over the phone's screen that conducts electricity. This allows the screen to function as a touch screen. This is the major use of indium.

Glass: Alumina and silica
On most phones the glass is aluminosilicate glass, a mix of aluminium oxide and silicon dioxide. It also contains potassium ions which help strengthen it.

Colours: Rare earth metals
A variety of rare earth metal-containing compounds are used to help to produce the colours in a smartphone's screen. Some of these compounds are also used to help reduce light penetration into the phone. Many of the rare earths occur commonly in the Earth's crust, but often at levels too low to be economically extracted.

ELECTRONICS

Wiring and microelectronics
Copper is used for wiring, and for micro-electrical components along with gold and silver. Tantalum is the major component in micro-capacitors.

Microphones and vibrations
Nickel is used in the microphone and for electrical connections. Rare earth element alloys are used in magnets in the speaker and microphone, and the vibration unit.

The silicon chip
Pure silicon is used to manufacture the chip, which is then oxidised to produce non-conducting regions. Other elements are added to allow the chip to conduct electricity.

Connecting electronics
Tin and lead were used in older solders; newer, lead-free solders use a mix of tin, copper and silver.

BATTERY

Most phones use lithium-ion batteries, which are composed of lithium cobalt oxide as a positive electrode and graphite (carbon) as the negative electrode. Sometimes other metals, such as manganese, are used in place of cobalt. The battery casing is often made of aluminium.

CASING

Magnesium alloy is used to make some phone cases. Many others are made of plastics, which are carbon-based. Plastics will also include flame-retardant compounds, some of which contain bromine, whilst nickel can be included to reduce electromagnetic interference.

www.compoundchem.com

Source: https://www.compoundchem.com/2014/02/19/the-chemical-elements-of-a-smartphone

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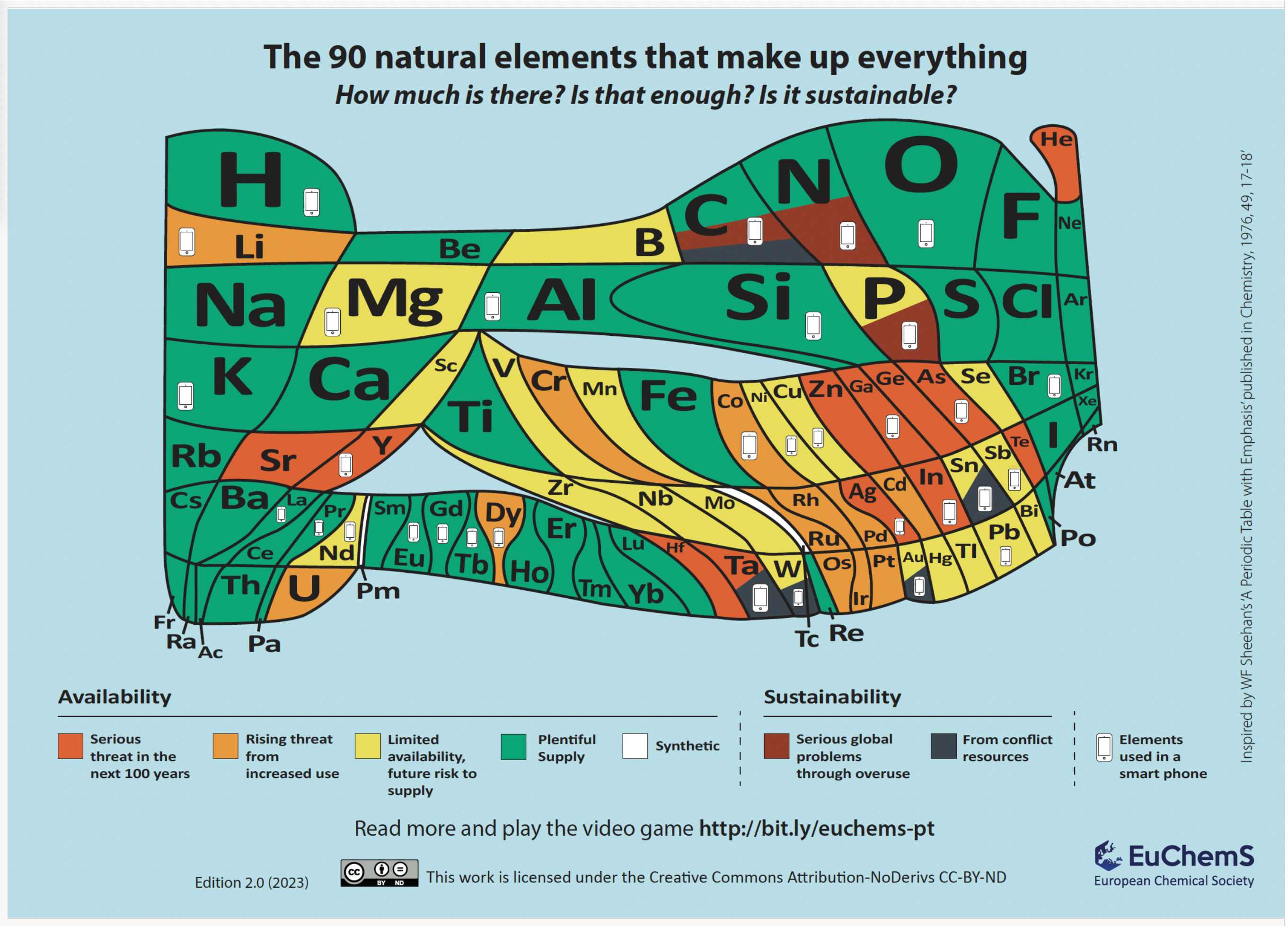
4.8.2: Elements in Danger - Recycle your phones

Video: Elements in danger - Recycle your phones

The 90 natural elements that make up everything:



Source: https://euchems.eu/euchems-periodic-table

4.8.3: Recycled Cell Phones—A Treasure Trove of Valuable Metals

Electronic products are made from valuable resources and materials, including metals, plastics, and glass, all of which require energy to mine and manufacture. Donating or recycling consumer electronics conserves our natural resources and avoids air and water pollution, as well as greenhouse gas emissions that are caused by manufacturing virgin materials.

In 2006, U.S. Geological Survey (USGS) estimated that:

- Recycling one million laptops saves the energy equivalent to the electricity used by more than 3,500 U.S. homes in a year.
- For every million cell phones we recycle, 35 thousand pounds of copper, 772 pounds of silver, 75 pounds of gold and 33 pounds of palladium can be recovered.

Read: U.S. Geological Survey Fact Sheet: Recycled Cell Phones—A Treasure Trove of Valuable Metals

Source: https://www.epa.gov/recycle/electronics-donation-and-recycling

CHAPTER OVERVIEW

5: Chemical Bonds, Ionic Bonds and Ionic Compounds

5.1: Why Atoms Form Bonds

Figure 5.1.1 (Credit: Flickr: Amelia Wells; Source: https://www.flickr.com/photos/speculummundi/4896031682/(opens in new window); License: CC by 2.0(opens in new window))

Have you ever made cupcakes from scratch?

You mix together flour, sugar, eggs, and other ingredients to make the batter, put the batter into cupcake papers, and then put them into the oven to bake. The cupcakes that come out of the oven after baking are different from any of the individual ingredients that went into the batter. Like the ingredients that join together to make cupcakes, atoms of different elements can join together to form entirely different substances called compounds. In cupcakes, the eggs and other wet ingredients cause the dry ingredients to stick together. What causes elements to stick together in compounds? The answer is chemical bonds.

What Is a Chemical Bond?

A chemical bond is a force of attraction between atoms or ions. Bonds form when atoms share or transfer valence electrons. Valence electrons are the electrons in the outer energy level of an atom that may be involved in chemical interactions. Valence electrons are the basis of all chemical bonds.

Q: Why do you think that chemical bonds form?

A: Chemical bonds form because they give atoms a more stable arrangement of electrons.

Why Bonds Form

To understand why chemical bonds form, consider the common compound(opens in new window) known as water, or H_2O. It consists of two hydrogen (H) atoms and one oxygen (O) atom. As you can see in the on the left side of the figure below, each hydrogen atom has just one electron, which is also its sole valence electron. The oxygen atom has six valence electrons. These are the electrons in the outer energy level of the oxygen atom.

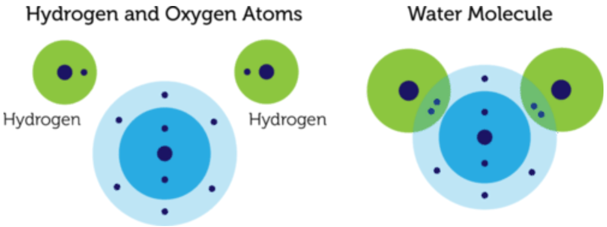


Figure 5.1.2: Copy and Paste Caption here. (Credit: Christopher AuYeung; Source: CK-12 Foundation; License: CC by-NC 3.0(opens in new window))

In the water molecule on the right in the figure above, each hydrogen atom shares a pair of electrons with the oxygen atom. By sharing electrons, each atom has electrons available to fill its sole or outer energy level. The hydrogen atoms each have a pair of shared electrons, so their first and only energy level is full. The oxygen atom has a total of eight valence electrons, so its outer energy level is full. A full outer energy level is the most stable possible arrangement of electrons. It explains why elements form chemical bonds with each other.

Types of Chemical Bonds

Not all chemical bonds form in the same way as the bonds in water. There are actually three different types of chemical bonds, called covalent, ionic, and metallic bonds. Each type of bond is described in more detail within this chapter.

Summary

- A chemical bond is a force of attraction between atoms or ions. Bonds form when atoms share or transfer valence electrons.
- Atoms form chemical bonds to achieve a full outer energy level, which is the most stable arrangement of electrons.
- There are three different types of chemical bonds: covalent, ionic, and metallic bonds.

5.2: Covalent and Ionic Bonds

Learning Objective

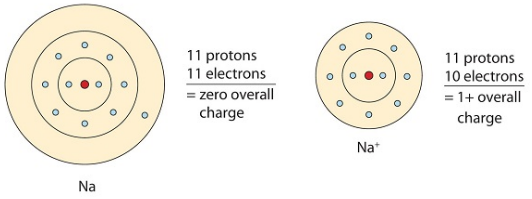
- Identify the difference between covalent and ionic bonds.
- Determine whether a compound is covalent or ionic based on its formula.

There are many types of chemical bonds and forces that bind molecules together. The two most basic types of bonds are characterized as either ionic or covalent. In ionic bonding, atoms transfer electrons to each other. Ionic bonds require at least one electron donor and one electron acceptor. In contrast, atoms with the same electronegativity share electrons in covalent bonds, because neither atom preferentially attracts or repels the shared electrons.

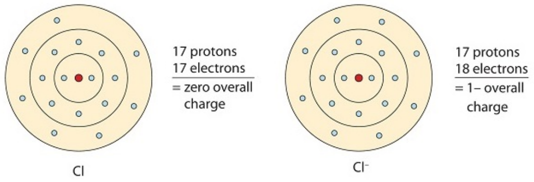
Introduction

Ionic bonding is the complete transfer of valence electron(s) between atoms. It is a type of chemical bond that generates two oppositely charged ions. In ionic bonds, the metal loses electrons to become a positively charged cation, whereas the nonmetal accepts those electrons to become a negatively charged anion. Ionic bonds require an electron donor, often a metal, and an electron acceptor, a nonmetal.

Ionic bonding is observed because metals have few electrons in their outer-most orbitals. By losing those electrons, these metals can achieve noble gas configuration and satisfy the octet rule. Similarly, nonmetals that have close to 8 electrons in their valence shells tend to readily accept electrons to achieve noble gas configuration. In ionic bonding, more than 1 electron can be donated or received to satisfy the octet rule. The charges on the anion and cation correspond to the number of electrons donated or received. In ionic bonds, the net charge of the compound must be zero.



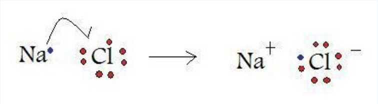
This sodium molecule donates the lone electron in its valence orbital in order to achieve octet configuration. This creates a positively charged cation due to the loss of electron.



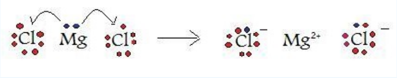
This chlorine atom receives one electron to achieve its octet configuration, which creates a negatively charged anion.

The predicted overall energy of the ionic bonding process, which includes the ionization energy of the metal and electron affinity of the nonmetal, is usually positive, indicating that the reaction is endothermic and unfavorable. However, this reaction is highly favorable because of the electrostatic attraction between the particles. At the ideal interatomic distance, attraction between these particles releases enough energy to facilitate the reaction. Most ionic compounds tend to dissociate in polar solvents because they are often polar. This phenomenon is due to the opposite charges on each ion.

Example 5.2.1: Chloride Salts



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



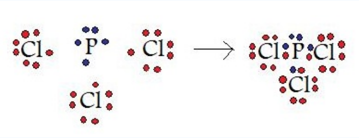
In this example, the magnesium atom is donating both of its valence electrons to chlorine atoms. Each chlorine atom can only accept 1 electron before it can achieve its noble gas configuration; therefore, 2 atoms of chlorine are required to accept the 2 electrons donated by the magnesium. Notice that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

Covalent Bonding

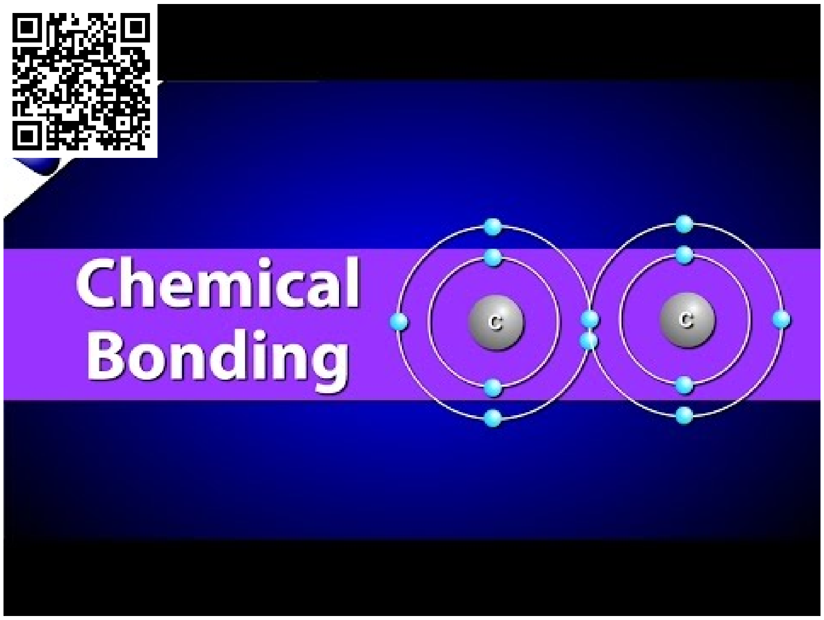
Covalent bonding is the sharing of electrons between atoms. This type of bonding occurs between two atoms of the same element or of elements close to each other in the periodic table. This bonding occurs primarily between nonmetals; however, it can also be observed between nonmetals and metals.

If atoms have similar electronegativities (the same affinity for electrons), covalent bonds are most likely to occur. Because both atoms have the same affinity for electrons and neither has a tendency to donate them, they share electrons in order to achieve octet configuration and become more stable. In addition, the ionization energy of the atom is too large and the electron affinity of the atom is too small for ionic bonding to occur. For example: carbon does not form ionic bonds because it has 4 valence electrons, half of an octet. To form ionic bonds, Carbon molecules must either gain or lose 4 electrons. This is highly unfavorable; therefore, carbon molecules share their 4 valence electrons through single, double, and triple bonds so that each atom can achieve noble gas configurations. Covalent bonds include interactions of the sigma and pi orbitals; therefore, covalent bonds lead to formation of single, double, triple, and quadruple bonds.

✓ Example 5.2.2: PCl_3



In this example, a phosphorous atom is sharing its three unpaired electrons with three chlorine atoms. In the end product, all four of these molecules have 8 valence electrons and satisfy the octet rule.



Problems

1. Do these compounds contain ionic or covalent bonds?

   CH_4 Fe_2O_3 KNO_3 H_2O $BeCl_2$

2. In the following reaction, indicate whether the reactants and products are ionic or covalently bonded.

$$Na_2CO_3 + CaCl_2 \longrightarrow CaCO_3 + 2\,NaCl$$

Solutions

1. From left to right: Covalent, Ionic, Ionic, Covalent, Ionic. In general, a compound containing both metals and nonmetals has at least one ionic bond and compounds containing only nonmetals have covalent bonds.
2. All products and reactants are ionic.

References

1. Vollhardt, K. Peter C., and Neil E. Schore. Organic Chemistry Structure and Function. New York: W. H. Freeman, 2007.
2. Petrucci, Ralph H. General Chemistry: Principles and Modern Applications. Upper Saddle River, NJ: Pearson Education, 2007.
3. Brown, Theodore L., Eugene H. Lemay, and Bruce E. Bursten. Chemistry: The Central Science. 6th ed. Englewood Cliffs, NJ: Prentice Hall, 1994.

5.3: Metallic Bonds

Learning Objective

- Define how metal atoms bond through a sea of delocalized electrons.
- Identify how metallic bonds lead to properties like conductivity and malleability in metals.

Figure 5.3.1: A nineteenth century copper plate. (Credit: John Pearson; Source: http://commons.wikimedia.org/wiki/File:John_Pearson01.png(opens in new window); License: Public Domain)

Why do metals behave the way they do?

The image below is of a copper plate that was made in 1893. The utensil has a great deal of elaborate decoration, and the item is very useful. What would have happened if this plate was made of copper (I) chloride instead? Copper (I) chloride does contain copper, after all. However, the $CuCl$ would end up as a powder when a metalworker pounded on it to shape it. Metals behave in unique ways. The bonding that occurs in a metal is responsible for its distinctive properties: luster, malleability, ductility, and excellent conductivity.

The Metallic Bond

Pure metals are crystalline solids, but unlike ionic compounds, every point in the crystal lattice is occupied by an identical atom. The electrons in the outer energy levels of a metal are mobile and capable of drifting from one metal atom to another. This means that the metal is more properly viewed as an array of positive ions surrounded by a sea of mobile valence electrons. Electrons which are capable of moving freely throughout the empty orbitals of the metallic crystal are called **delocalized electrons** (see below). A **metallic bond** is the attraction of the stationary metal cations to the surrounding mobile electrons.

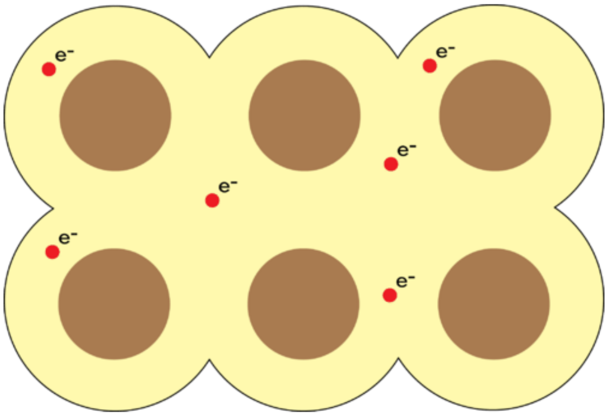


Figure 5.3.2: In a metal, the stationary metal cations are surrounded by a sea of mobile valence electrons that are not associated with any one cation. (Credit: Christopher Auyeung; Source: CK-12 Foundation; License: CC BY-NC 3.0(opens in new window))

Properties of Metals

The metallic bonding model explains the physical properties of metals. Metals conduct electricity and heat very well because of their free-flowing electrons. As electrons enter one end of a piece of metal, an equal number of electrons flow outward from the other end. When light is shone onto the surface of a metal, its electrons absorb small amounts of energy and become excited into one of its many empty orbitals. The electrons immediately fall back down to lower energy levels and emit light. This process is responsible for the high **luster** of metals.

Figure 5.3.3: The American Platinum Eagle is the official platinum bullion coin of the United States and was first minted in 1997. The luster of a metal is due to its metallic bonds. (Credit: Courtesy of the US Mint; Source: Front: http://commons.wikimedia.org/wiki/File:2005_AEPlat_Proof_Obv.png(opens in new window); Back: http://commons.wikimedia.org/wiki/File:American_Platinum_Eagle_2008_Proof_Rev.jpg(opens in new window); License: Public Domain)

Recall that ionic compounds are very brittle. Application of a force results in like-charged ions in the crystal coming too close to one another, causing the crystal to shatter. When a force is applied to a metal, the free-flowing electrons can slip in between the stationary cations and prevent them from coming in contact. Imagine ball bearings that have been coated with oil sliding past one another. As a result, metals are very **malleable** and **ductile**. They can be hammered into shapes, rolled into thin sheets, or pulled into thin wires.

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5.4: Chemical Formula (Molecular and Empirical Formula)

Learning Objectives

- Determine the number of different atoms in a formula.
- Define chemical formula, molecular formula, and empirical formula.

A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a 2:1 ratio. The chemical formula for water is H_2O. Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is H_2SO_4.

Certain groups of atoms are bonded together to form what is called a polyatomic ion that acts as a single unit. Polyatomic ions are discussed in more detail in Section 5.5. Polyatomic ions are enclosed in parenthesis followed by a subscript if more than one of the same ion exist in a chemical formula. The formula $Ca_3(PO_4)_2$ represents a compound with the following:

$$3 \text{ Ca atoms} + 2\ PO_4^{3-} \text{ ions}$$

To count the total number of atoms for formulas with polyatomic ions enclosed in parenthesis, use the subscript as a multiplier for each atom or number of atoms.

$$Ca_3(PO_4)_{\mathbf{2}}$$

$$3 \text{ Ca} + \mathbf{2} \text{ x1 P} + \mathbf{2} \text{ x 4 O} = 3 \text{ Ca atoms} + 2 \text{ P atoms} + 8 \text{ O atoms}$$

Molecular Formula

A **molecular formula** is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

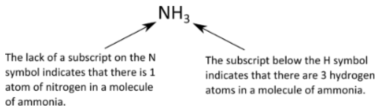


Figure 5.4.1: The molecular formula for ammonia. NH3. There is one atom of nitrogen and 3 atoms of hydrogen in a molecule of ammonia.

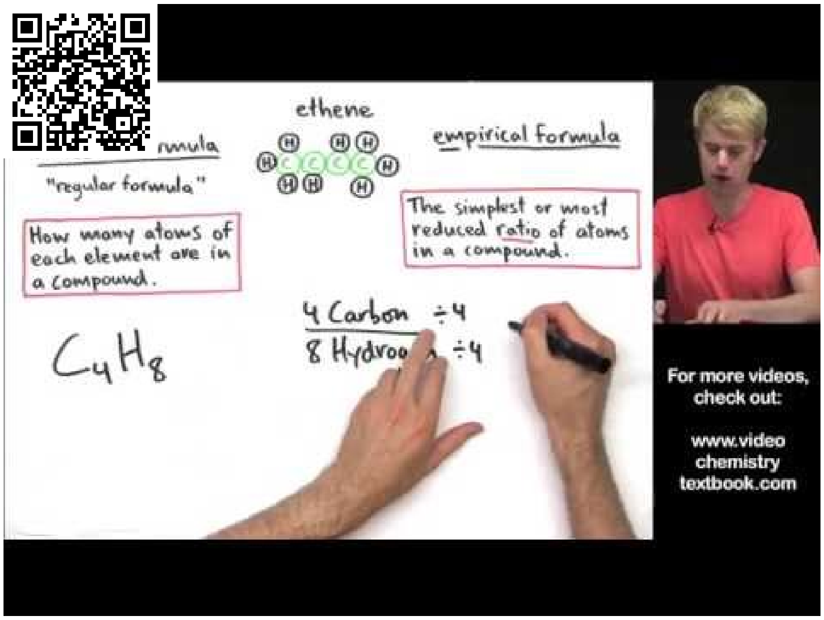
Note from the example that there are some standard rules to follow in writing molecular formulas. The arrangements of the elements depend on the particular structure, which is not of concern at this point. The number of atoms of each kind is indicated by a subscript following the atom. If there is only one atom, no number is written. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. We would not write N_3H for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $C_6H_{12}O_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O. When chemists analyze an unknown compound, often the first step is to determine its empirical formula.

- molecular formula: $C_6H_{12}O_6$
- empirical formula: CH_2O

There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5, then the empirical formula is also the molecular formula.


ethene
rmula
"regular formula"
How many atoms of each element are in a compound.
empirical Formula
The simplest or most reduced ratio of atoms in a compound.
C4H8
4 Carbon ÷4
8 Hydrog ÷4
For more videos, check out:
www.video chemistry textbook.com

5.5: Octet Rule

Learning Objective

- Define how atoms gain, lose, or share electrons to achieve a full outer shell of 8 electrons.

The Octet Rule

In 1904, Richard Abegg formulated what is now known as *Abegg's rule*, which states that the difference between the maximum positive and negative valences of an element is frequently eight. This rule was used later in 1916 when Gilbert N. Lewis formulated the "octet rule" in his cubical atom theory. The **octet rule** refers to the tendency of atoms to prefer to have eight electrons in the *valence shell*. When atoms have fewer than eight electrons, they tend to react and form more stable compounds. Atoms will react to get in the most stable state possible. A complete octet is very stable because all orbitals will be full. Atoms with greater stability have less energy, so a reaction that increases the stability of the atoms will release energy in the form of heat or light ;reactions that decrease stability must absorb energy, getting colder.

> *The Octet Rule: Atoms often gain, lose, or share electrons to achieve the same number of electrons as the noble gas closest to them in the periodic table.*

When discussing the octet rule, we do not consider *d* or *f* electrons. Only the s and p electrons are involved in the octet rule, making it a useful rule for the *main group elements* (elements not in the transition metal or inner-transition metal blocks); an octet in these atoms corresponds to an electron configurations ending with s^2p^6.

Covalent Bonds

Covalent bonds form when atoms share electrons. Hydrogen is a first shell element with only one valence electron, so it can only form one bond creating a duet, an exception to the octet rule. With its four valence electrons, carbon can form four bonds to create an octet.

1. Normally two electrons pairs up and forms a bond, e.g., H_2
2. For most atoms there will be a maximum of eight electrons in the valence shell (octet structure), e.g., CH_4

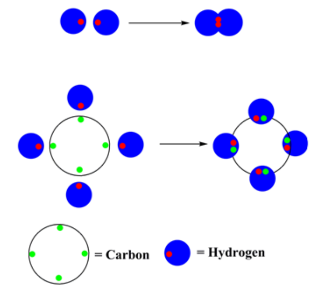


Figure 1: Bonding in H_2 and methane (CH_4)

The other tendency of atoms is to maintain a neutral charge. Only the noble gases (the elements on the right-most column of the periodic table) have zero charge with filled valence octets. All of the other elements have a charge when they have eight electrons all to themselves. The result of these two guiding principles is the explanation for much of the reactivity and bonding that is observed within atoms: atoms seek to share electrons in a way that minimizes charge while fulfilling an octet in the valence shell.

Ionic Bonds

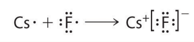
Some atoms do not share electrons. Energetically, it is more favorable to fully gain or lose electrons to form ions. Ionic compounds form through the eletrostatic attraction of the ions to create a crystal lattice.

The formula for table salt is NaCl. It is the result of Na^+ ions and Cl^- ions bonding together. If sodium metal and chlorine gas mix under the right conditions, they will form salt. The sodium loses an electron, and the chlorine gains that electron. In the process, a great amount of light and heat is released. The resulting salt is mostly unreactive — it is stable. It will not undergo any explosive reactions, unlike the sodium and chlorine that it is made of. Why? Referring to the octet rule, atoms attempt to get a noble gas electron configuration, which is eight valence electrons. Sodium has one valence electron, so giving it up would result in the same electron configuration as neon. Chlorine has seven valence electrons, so if it takes one it will have eight (an octet). Chlorine has the electron configuration of argon when it gains an electron.

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

Ionic Bonds Example

Lewis dot symbols can also be used to represent the ions in ionic compounds. The reaction of cesium with fluorine, for example, to produce the ionic compound CsF can be written as follows:

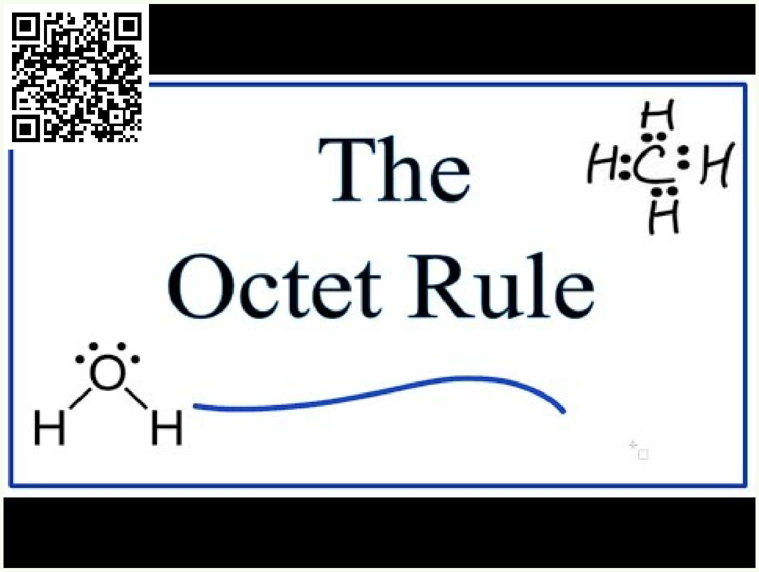


$$\mathrm{Cs}\cdot + :\ddot{\underset{..}{\mathrm{F}}}\cdot \longrightarrow \mathrm{Cs}^+[:\ddot{\underset{..}{\mathrm{F}}}:]^-$$

No dots are shown on Cs$^+$ in the product because cesium has lost its single valence electron to fluorine. The transfer of this electron produces the Cs$^+$ ion, which has the valence electron configuration of Xe, and the F$^-$ ion, which has a total of eight valence electrons (an octet) and the Ne electron configuration. This description is consistent with the statement that among the main group elements, ions in simple binary ionic compounds generally have the electron configurations of the nearest noble gas. The charge of each ion is written in the product, and the anion and its electrons are enclosed in brackets. This notation emphasizes that the ions are associated electrostatically; no electrons are shared between the two elements

Noble Gases

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



Contributors and Attributions

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- National Programme on Technology Enhanced Learning (India)

5.6: Representing Valence Electrons with Dots

Learning Objective

- Draw a Lewis electron dot diagram for an atom or a monatomic ion.

In almost all cases, chemical bonds are formed by interactions of valence electrons in atoms. To facilitate our understanding of how valence electrons interact, a simple way of representing those valence electrons would be useful.

A **Lewis electron dot diagram** (or electron dot diagram, or a Lewis diagram, or a Lewis structure) is a representation of the valence electrons of an atom that uses dots around the symbol of the element. The number of dots equals the number of valence electrons in the atom. These dots are arranged to the right and left and above and below the symbol, with no more than two dots on a side. (The order in which the positions are used does not matter.) For example, the Lewis electron dot diagram for hydrogen is simply

$$\mathbf{H}\cdot$$

Because the side is not important, the Lewis electron dot diagram could also be drawn as follows:

$$\dot{\mathbf{H}} \quad or \quad \cdot\mathbf{H} \quad or \quad \underset{\cdot}{\mathbf{H}}$$

The electron dot diagram for helium, with two valence electrons, is as follows:

$$\mathbf{He:}$$

By putting the two electrons together on the same side, we emphasize the fact that these two electrons are both in the 1*s* subshell; this is the common convention we will adopt, although there will be exceptions later. The next atom, lithium, has an electron configuration of $1s^2 2s^1$, so it has only one electron in its valence shell. Its electron dot diagram resembles that of hydrogen, except the symbol for lithium is used:

$$\mathbf{Li}\cdot$$

Beryllium has two valence electrons in its 2*s* shell, so its electron dot diagram is like that of helium:

$$\mathbf{Be:}$$

The next atom is boron. Its valence electron shell is $2s^2 2p^1$, so it has three valence electrons. The third electron will go on another side of the symbol:

$$\dot{\mathbf{B}}\mathbf{:}$$

Again, it does not matter on which sides of the symbol the electron dots are positioned.

For carbon, there are four valence electrons, two in the 2*s* subshell and two in the 2*p* subshell. As usual, we will draw two dots together on one side, to represent the 2*s* electrons. However, conventionally, we draw the dots for the two *p* electrons on different sides. As such, the electron dot diagram for carbon is as follows:

$$\cdot\dot{\mathbf{C}}\mathbf{:}$$

With N, which has three *p* electrons, we put a single dot on each of the three remaining sides:

$$\cdot\underset{\cdot}{\dot{\mathbf{N}}}\mathbf{:}$$

For oxygen, which has four *p* electrons, we now have to start doubling up on the dots on one other side of the symbol. When doubling up electrons, make sure that each side has no more than two electrons.

$$\cdot\underset{\cdot}{\ddot{\mathbf{O}}}\mathbf{:}$$

Fluorine and neon have seven and eight dots, respectively:

$$\mathbf{:}\underset{\cdot}{\ddot{\mathbf{F}}}\mathbf{:}$$

$$\mathbf{:\overset{..}{\underset{..}{Ne}}:}$$

With the next element, sodium, the process starts over with a single electron because sodium has a single electron in its highest-numbered shell, the n = 3 shell. By going through the periodic table, we see that the Lewis electron dot diagrams of atoms will never have more than eight dots around the atomic symbol.

✓ Example 5.6.1: Lewis Dot Diagrams

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

Solution

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3s electrons:

$$\dot{Al}:$$

2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the n = 4 shell, there are six electrons. Its electron dot diagram is as follows:

$$\mathbf{\cdot\overset{\cdot}{\underset{..}{Se}}:}$$

? Exercise 5.6.1

What is the Lewis electron dot diagram for each element?

a. phosphorus
b. argon

Answer a

$$\mathbf{\cdot\overset{\cdot}{\underset{\cdot}{P}}:}$$

Answer b

$$\mathbf{:\overset{..}{\underset{..}{Ar}}:}$$

5.7: Lewis Structures of Ionic Compounds - Electrons Transferred

Learning Objectives

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

In Section 4.7, we demonstrated that ions are formed by losing electrons to make cations, or by gaining electrons to form anions. The astute reader may have noticed something: many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The **octet rule** explains the favorable trend of atoms having eight electrons in their valence shell. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

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

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

$$\mathbf{Na}\cdot \qquad \cdot\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5$$

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

$$\mathbf{Na}\cdot \curvearrowright \cdot\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}$$

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

$$\mathbf{Na}^+ \qquad \mathbf{:}\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^-$$

$$[Ne] \qquad [Ne]\,3s^2 3p^6$$

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

$$\mathbf{Na}^+ + \mathbf{:}\ddot{\underset{\cdot\cdot}{\mathbf{Cl}}}\mathbf{:}^- \rightarrow Na^+Cl^- \quad or \quad NaCl$$

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

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

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

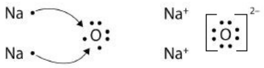
$$\mathbf{Mg}^{2+} + \left[\mathbf{:}\ddot{\underset{\cdot\cdot}{\mathbf{O}}}\mathbf{:}\right]^{2-} \qquad Mg^{2+}O^{2-} \ or \ MgO$$

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

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

$$\mathbf{Na}\cdot \curvearrowright \cdot\ddot{\underset{\cdot}{\mathbf{O}}}\mathbf{:}$$

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



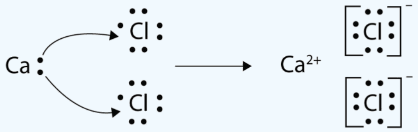
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the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

✓ Example 5.7.1: Synthesis of Calcium Chloride from Elements

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

Solution

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

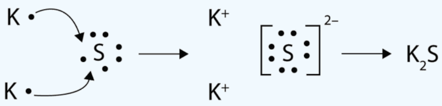


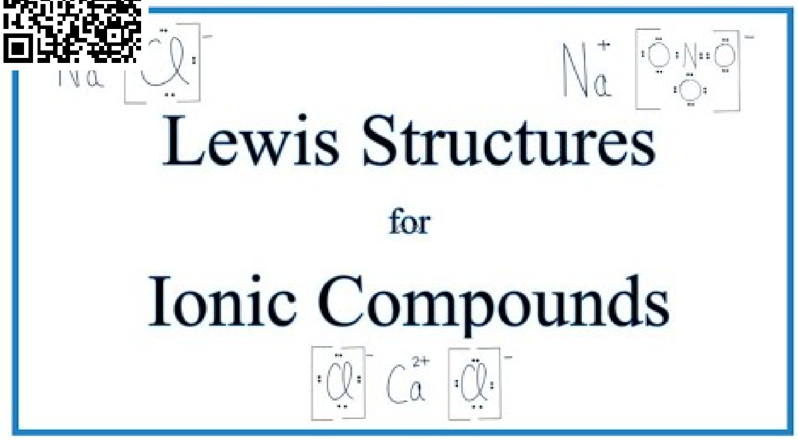
The oppositely charged ions attract each other to make $CaCl_2$.

? Exercise 5.7.1

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

Answer





Lewis Structures

for

Ionic Compounds

- **10.3: Lewis Structures of Ionic Compounds- Electrons Transferred** by Henry Agnew, Marisa Alviar-Agnew is licensed CC BY-NC-SA 3.0.

5.8: Covalent Lewis Structures- Electrons Shared

Learning Objectives

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

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

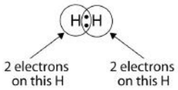
Let us illustrate a covalent bond by using H atoms, with the understanding that H atoms need only two electrons to fill the 1*s* subshell. Each H atom starts with a single electron in its valence shell:

$$\mathbf{H}\cdot \quad \cdot\mathbf{H}$$

The two H atoms can share their electrons:

$$\mathbf{H}:\mathbf{H}$$

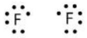
We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



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

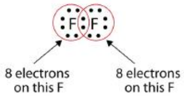
H–H

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

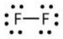


These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.

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



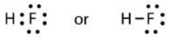
We can also write this using a dash to represent the shared electron pair:



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

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

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



We note that the H atom has a full valence shell with two electrons, while the F atom has a complete octet of electrons.

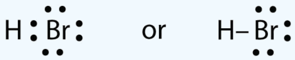
Example 5.8.1:

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

Solution

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

The two atoms can share their unpaired electron:



Exercise 5.8.1

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

Answer

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

Multiple Covalent Bonds

Some molecules are not able to satisfy the octet rule by making only single covalent bonds between the atoms. Consider the compound ethene, which has a molecular formula of C_2H_4. The carbon atoms are bonded together, with each carbon also bonded to two hydrogen atoms.

$$\text{two C atoms} = 2 \times 4 = 8 \text{ valence electrons}$$

$$\text{four H atoms} = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{total of 12 valence electrons in the molecule}$$

If the Lewis electron dot structure was drawn with a single bond between the carbon atoms and with the octet rule followed, it would look like this:

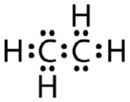


Figure 5.8.1: Incorrect dot structure of ethene. (CK12 License)

This Lewis structure is incorrect because it contains a total of 14 electrons. However, the Lewis structure can be changed by eliminating the lone pairs on the carbon atoms and having to share two pairs instead of only one pair.

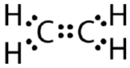


Figure 5.8.2: Correct dot structure for ethene. (CK12 License)

A **double covalent bond** is a covalent bond formed by atoms that share two pairs of electrons. The double covalent bond that occurs between the two carbon atoms in ethane can also be represented by a structural formula and with a molecular model as shown in the figure below.

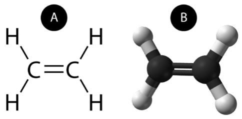


Figure 5.8.3: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4.

A **triple covalent bond** is a covalent bond formed by atoms that share three pairs of electrons. The element nitrogen is a gas that composes the majority of Earth's atmosphere. A nitrogen atom has five valence electrons, which can be shown as one pair and three single electrons. When combining with another nitrogen atom to form a diatomic molecule, the three single electrons on each atom combine to form three shared pairs of electrons.

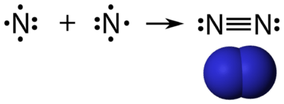
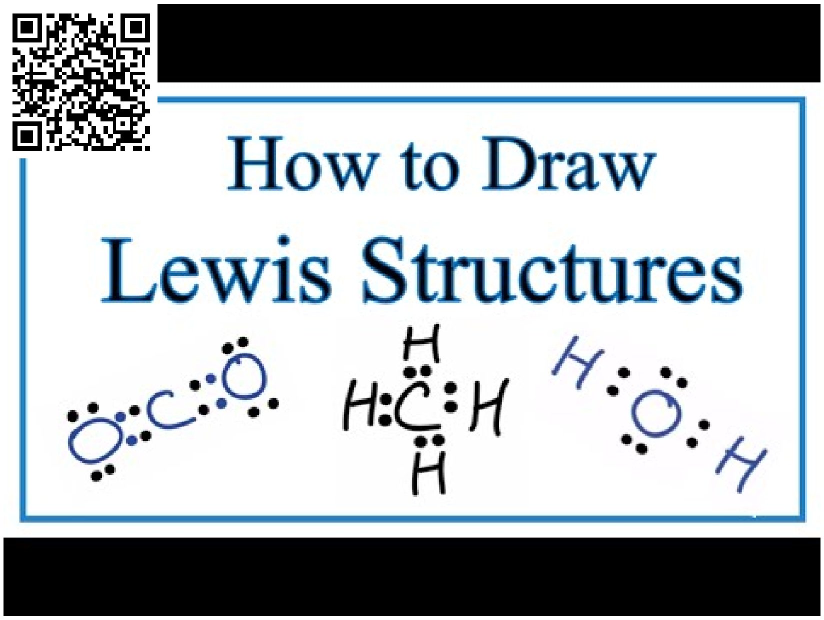


Figure 5.8.4: Triple bond in N_2.

Each nitrogen atom follows the octet rule with one lone pair of electrons, and six electrons that are shared between the atoms.



Contributions & Attributions

-
- Anonymous by request

5.9: Naming Ions and Binary Ionic Compound

Learning Objectives

- To use the rules for naming ionic compounds.

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

Naming Ions

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

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 3.3). Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

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

Table 5.9.1: The Modern and Common System of Cation Names

Element	Stem	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
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 5.9.1 lists the elements that use the common system, along with their respective cation names.

Table 5.9.2: Some Monatomic Anions

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

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 5.9.2 lists the names of some common monatomic ions. The polyatomic ions have their own characteristic names, as discussed earlier.

✓ Example 5.9.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. $SO_3{}^{2-}$
d. $NH_4{}^{+}$
e. Cu^{+}

Solution

a. the calcium ion
b. the sulfide ion
c. the sulfite ion
d. the ammonium ion
e. the copper(I) ion or the cuprous ion

? Exercise 5.9.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. $SO_4{}^{2-}$
d. Ba^{2+}
e. $HCO_3{}^{-}$

Answer a:

iron(II) ion

Answer b:

iron(III) ion

Answer c:

sulfate ion

Answer d:

barium ion

Answer e:

hydrogen carbonate ion or bicarbonate ion

✓ Example 5.9.2

Write the formula for each ion.

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

Solution

a. Br^{-}
b. $PO_4{}^{3-}$
c. Cu^{2+}
d. Mg^{2+}

? Exercise 5.9.2

Write the formula for each ion.

a. the fluoride ion
b. the carbonate ion
c. the stannous ion
d. the potassium ion

Answer a:

F^{-}

Answer b:

$CO_3{}^{2-}$

Answer c:

Sn^{2+}

Answer d:

K^{+}

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

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 5.9.1 for the compound $BaCl_2$. The word *ion* is dropped from both parts.

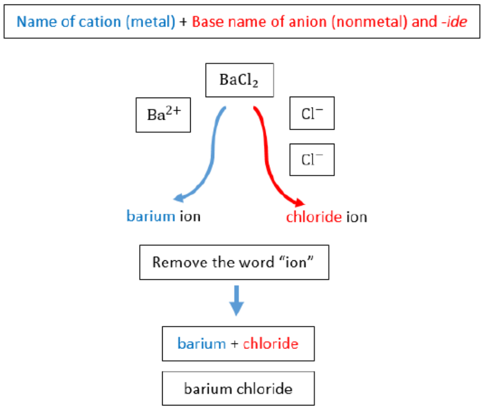


Figure 5.9.1: Naming $BaCl_2$
Naming formula: Name of metal cation + base
name of nonmetal anion + suffix -ide. BaCl2 is named as barium chloride.

Subscripts in the formula do not affect the name.

✓ Example 5.9.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

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

? Exercise 5.9.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

Answer a:

silver iodide

Answer b:

magnesium oxide

Answer c:

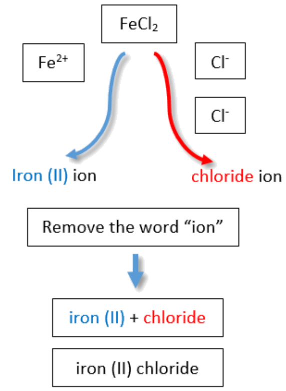
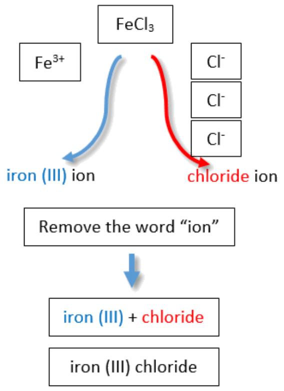
calcium phosphide

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

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1− charge on each chloride ion). In the second compound, the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure 5.9.2).

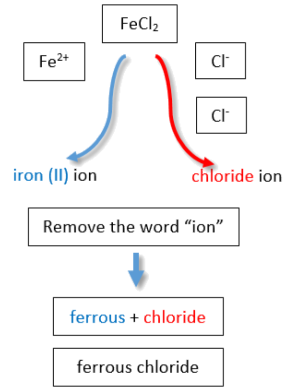
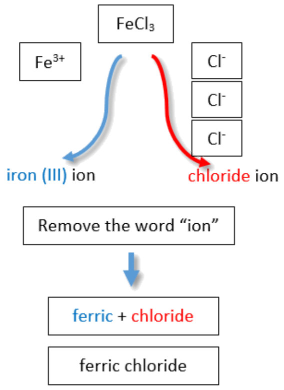
Table 5.9.3: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Modern/Stock System.

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

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

If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure 5.9.3) .

Table 5.9.4: Naming the $FeCl_2$ and $FeCl_3$ Compounds in the Old/Common System.

"Old" base name of cation (metal) and *-ic or -ous* + Base name of anion (nonmetal) and *-ide*	
-ous (for ions with lower charge) 	-ic (for ions with higher charge) 

✓ Example 5.9.4:

Name each ionic compound.

a. Co_2O_3
b. $FeCl_2$

Solution

Solutions to Example 5.7.4

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

Exercise 5.9.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a:

gold(III) chloride

Answer b:

lead(IV) oxide

Answer c:

copper(II) oxide

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium sulfate compound in Figure 5.9.6.

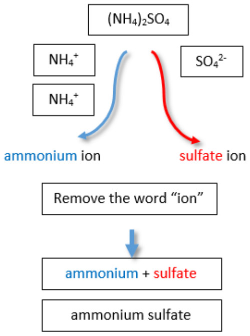


Figure 5.9.2: Naming Ionic Compounds with Polyatomic Ions

Example 5.9.5: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Solutions to Example 5.7.5

Explanation	Answer
a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound's name is ammonium sulfide.	ammonium sulfide
b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound's name is iron(II) phosphate.	iron(II) phosphate

Exercise 5.9.5A

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Figure 5.9.1 is a synopsis of how to name simple ionic compounds.

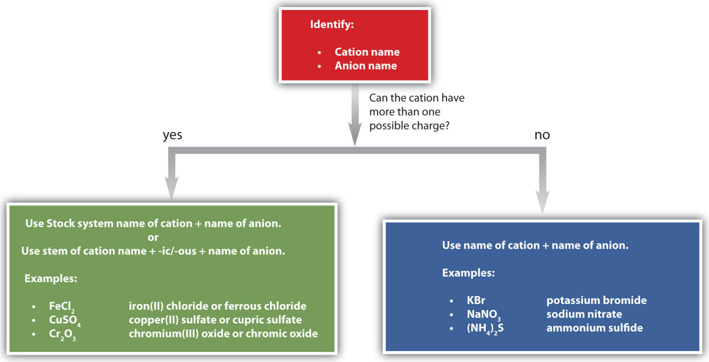
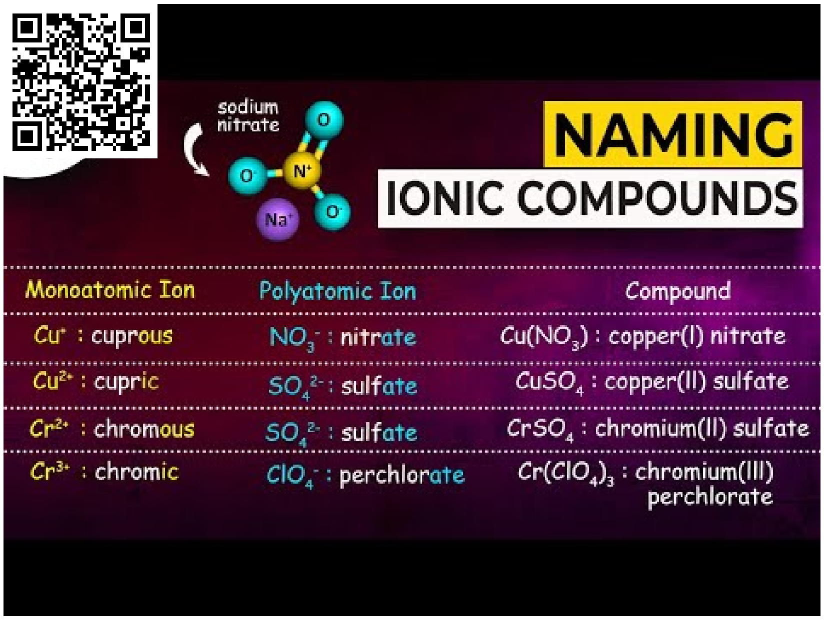


Figure 5.9.3: A Guide to Naming Simple Ionic Compounds.



NAMING IONIC COMPOUNDS

Monoatomic Ion	Polyatomic Ion	Compound
Cu^+ : cuprous	NO_3^- : nitrate	$Cu(NO_3)$: copper(I) nitrate
Cu^{2+} : cupric	SO_4^{2-} : sulfate	$CuSO_4$: copper(II) sulfate
Cr^{2+} : chromous	SO_4^{2-} : sulfate	$CrSO_4$: chromium(II) sulfate
Cr^{3+} : chromic	ClO_4^- : perchlorate	$Cr(ClO_4)_3$: chromium(III) perchlorate

? Exercise $5.9.5B$

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. $(NH_4)_3PO_4$
d. AuF_3
e. AgF

Answer a:

zinc bromide

Answer b:

aluminum oxide

Answer c:

ammonium phosphate

Answer d:

gold(III) fluoride or auric fluoride

Answer e:

silver fluoride

5.10: Naming Polyatomic and Ternary Ionic Compounds

Learning Objectives

- Recognize polyatomic ions in chemical formulas.
- Write the correct formula for ionic compounds with polyatomic ions.
- Use the rules for naming ionic compounds with polyatomic ions.

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Table 5.10.1 lists the most common polyatomic ions.

Table 5.10.1 Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

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

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

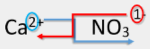
Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 5.10.1: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Example 5.10.1: Steps For Problem Solving, Write the formula for calcium nitrate

Criss Cross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Ca^{2+} NO_3^-
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 5.10.2

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

Solution

Example 5.10.2: Explanation for Writing the Chemical Formula for an Ionic Compound Composed of the Potassium Ion and the Sulfate Ion

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

Exercise 5.10.1

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

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

Answer a:

$MgCO_3$

Answer b:

$Al(CH_3COO)_3$

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium Sulfate compound in Figure 5.10.6.

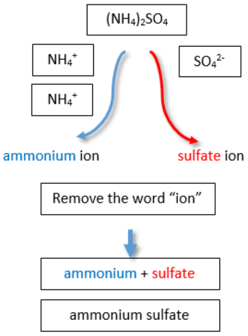
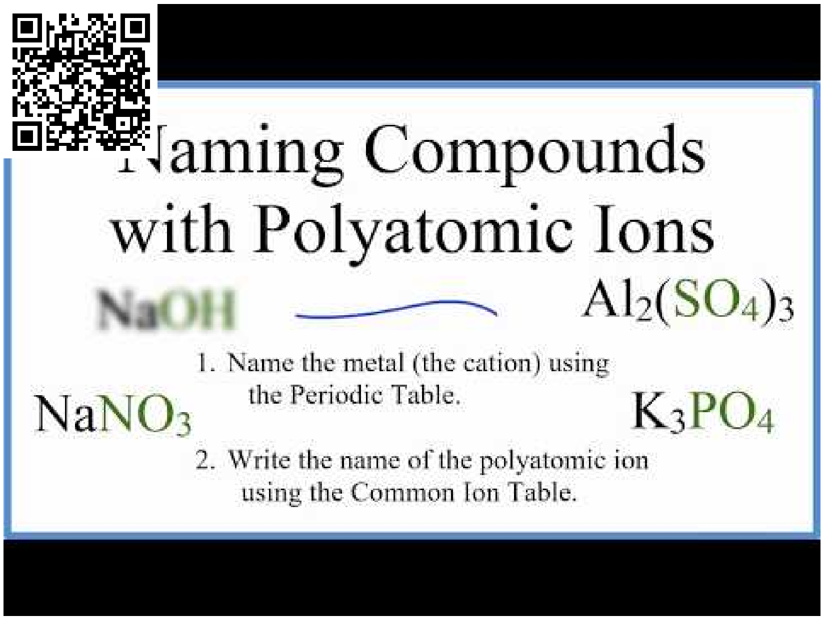


Figure 5.10.2 *Naming Ionic Compounds with Polyatomic Ions*



Naming Compounds with Polyatomic Ions

$NaOH$

$Al_2(SO_4)_3$

1. Name the metal (the cation) using the Periodic Table.

$NaNO_3$

K_3PO_4

2. Write the name of the polyatomic ion using the Common Ion Table.

✓ Example 5.10.3: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. $(NH_4)_2S$
b. $AlPO_4$,
c. $Fe_3(PO_4)_2$

Solution

Example 5.10.3: Explanation for Naming Ionic Compounds

Ionic Compound	Explanation	Answer

Ionic Compound	Explanation	Answer
a. **$(NH_4)_2S$** b.	a. The ammonium ion has a 1+ charge and the sulfide ion has a 2− charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound’s name is ammonium sulfide.	ammonium sulfide
b. $AlPO_4$,	b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate
c. $Fe_3(PO_4)_2$	c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6−, two phosphate ions are needed . The compound’s name is iron(II) phosphate.	iron(II) phosphate

? Exercise 5.10.2

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a:

ammonium phosphate

Answer b:

cobalt(III) nitrite

Contributors and Attributions

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

5.11: Formulas for Ionic Compounds

Learning Objectives

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

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

Ionic compounds exist as alternating positive and negative ions in regular, three-dimensional arrays called crystals (Figure 5.11.1). As you can see, there are no individual NaCl "particles" in the array; instead, there is a continuous lattice of alternating sodium and chloride ions. However, we can use the ratio of sodium ions to chloride ions, expressed in the lowest possible whole numbers, as a way of describing the compound. In the case of sodium chloride, the ratio of sodium ions to chloride ions, expressed in lowest whole numbers, is 1:1, so we use NaCl (one Na symbol and one Cl symbol) to represent the compound. Thus, NaCl is the chemical formula for sodium chloride, which is a concise way of describing the relative number of different ions in the compound. A macroscopic sample is composed of myriads of NaCl pairs; each individual pair called a **formula unit**. Although it is convenient to think that NaCl crystals are composed of individual NaCl units, Figure 5.11.1 shows that no single ion is exclusively associated with any other single ion. Each ion is surrounded by ions of opposite charge.

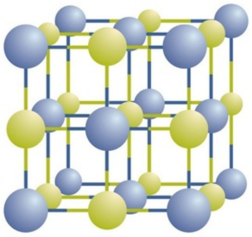


Figure 5.11.1: A Sodium Chloride Crystal. A crystal contains a three-dimensional array of alternating positive and negative ions. The precise pattern depends on the compound. A crystal of sodium chloride, shown here, is a collection of alternating sodium and chlorine ions.

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

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

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

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

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

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

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

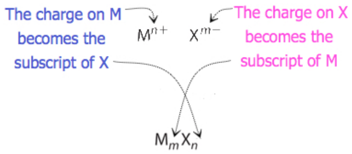
Combining one ion of each does not completely balance the positive and negative charges. The easiest way to balance these charges is to assume the presence of *two* chloride ions for each magnesium ion:

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

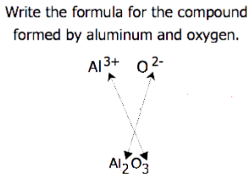
Now the positive and negative charges are balanced. We could write the chemical formula for this ionic compound as MgClCl, but the convention is to use a numerical subscript when there is more than one ion of a given type—$\mathrm{MgCl_2}$. This chemical formula says that there are one magnesium ion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

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

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



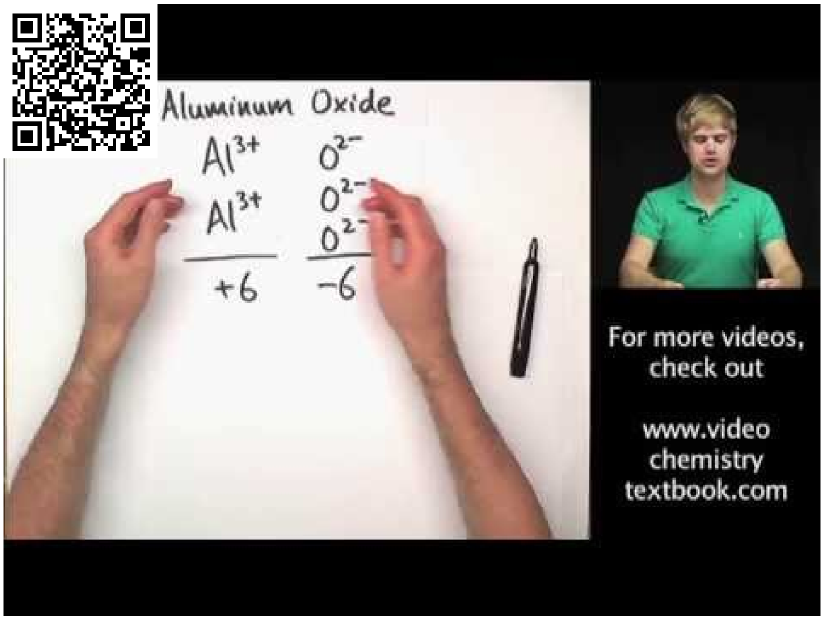
The image shows M with a charge of n+ and X with a charge of m-. The charge on the M becomes the subscript of X, and the charge on X becomes the subscript of M, making the final product M subscript m X subscript n.



The image asks to write the formula for the compound formed by aluminum and oxygen. It shows Al with a charge of 3+ and O with a charge of 2-. By crossing charges, the final result is Al subscript 2 O subscript 3.

Figure 5.11.2: Crossing charges. One method for obtaining subscripts in the empirical formula is by crossing charges.

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


Aluminum Oxide
Al^{3+} O^{2-}
Al^{3+} O^{2-}
O^{2-}
+6 −6
For more videos, check out
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Example 5.11.1

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

a. the sodium ion and the sulfur ion
b. the aluminum ion and the fluoride ion
c. the 3+ iron ion and the oxygen ion

Solution

a. To obtain a valence shell octet, sodium forms an ion with a 1+ charge, while the sulfur ion has a 2− charge. Two sodium 1+ ions are needed to balance the 2− charge on the sulfur ion. Rather than writing the formula as $NaNaS$, we shorten it by convention to Na_2S.
b. The aluminum ion has a 3+ charge, while the fluoride ion formed by fluorine has a 1− charge. Three fluorine 1− ions are needed to balance the 3+ charge on the aluminum ion. This combination is written as AlF_3.

c. Iron can form two possible ions, but the ion with a 3+ charge is specified here. The oxygen atom has a 2− charge as an ion. To balance the positive and negative charges, we look to the least common multiple—6: two iron 3+ ions will give 6+, while three 2− oxygen ions will give 6−, thereby balancing the overall positive and negative charges. Thus, the formula for this ionic compound is Fe_2O_3. Alternatively, use the crossing charges method shown in Figure 3.3.2.

? Exercise 5.11.1

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

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a:

CaO

Answer b:

CuS

Answer c:

Cu_2S

Polyatomic Ions

Some ions consist of groups of atoms ***covalently*** bonded together and have an **overall electric charge**. Because these ions contain more than one atom, they are called polyatomic ions. The Lewis structures, names and formulas of some polyatomic ions are found in Table 3.3.1.

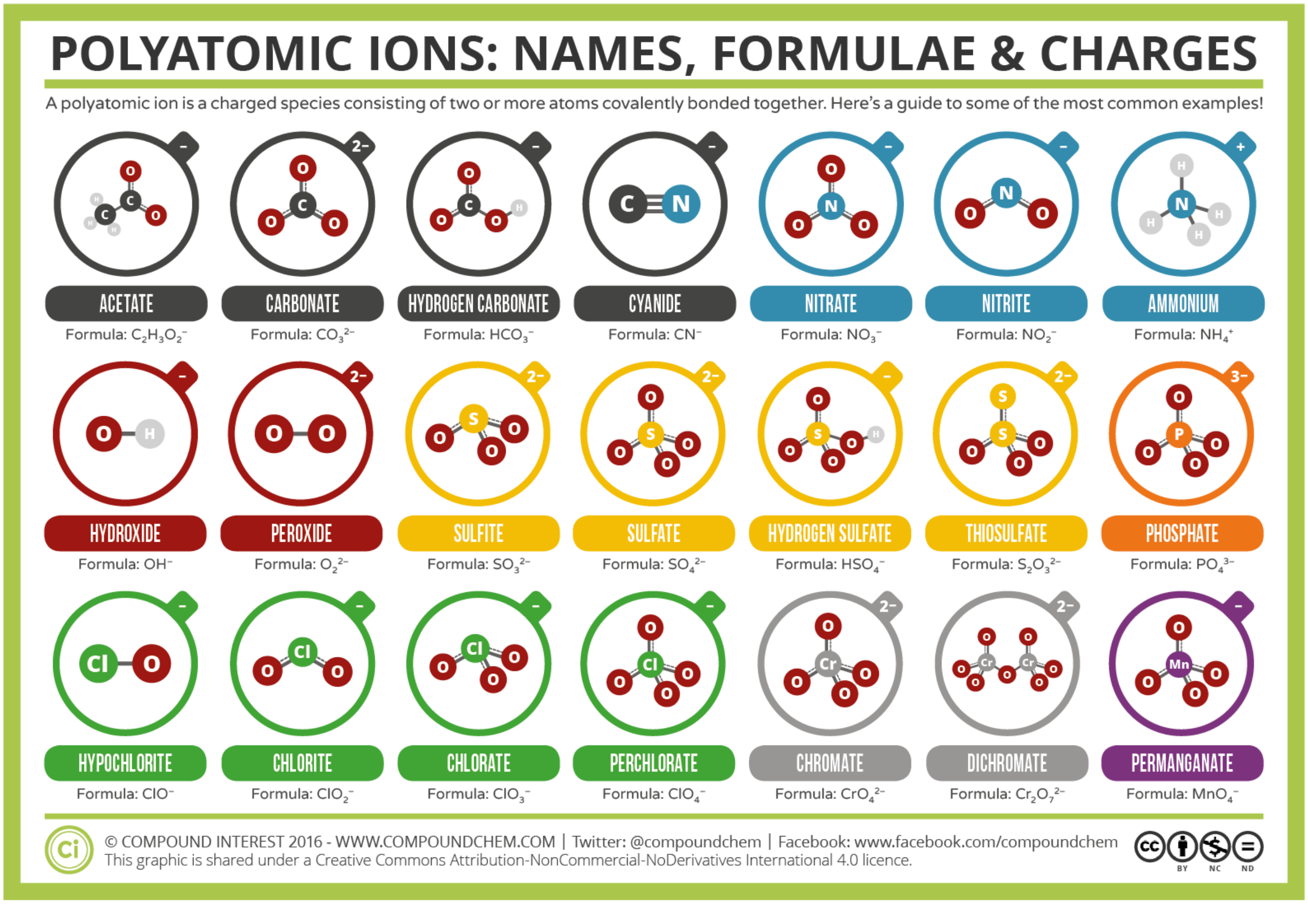


Table 5.11.1: Some Polyatomic Ions

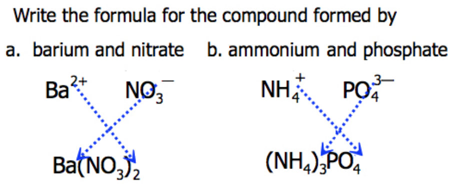
Polyatomic ions have defined formulas, names, and charges that cannot be modified in any way. Table 5.11.2 lists the ion names and ion formulas of the most common polyatomic ions. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1− charge. Figure 5.11.2 lists the most common polyatomic ions.

Table 5.11.2: Ion Names and Ion Formulas of Common Polyatomic Ions

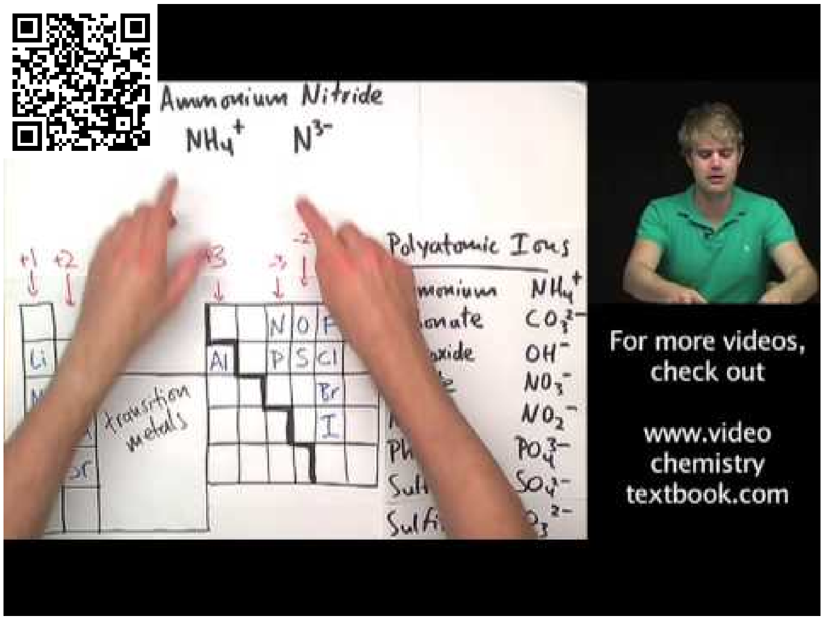
Ion Name	Ion Formula
ammonium ion	NH_4^{+1}
hydroxide ion	OH^{-1}
cyanide ion	CN^{-1}
carbonate ion	CO_3^{-2}
bicarbonate or hydrogen carbonate	HCO_3^{-}
acetate ion	$C_2H_3O_2^{-1}$ **or** $CH_3CO_2^{-1}$
nitrate ion	NO_3^{-1}
nitrite ion	NO_2^{-1}
sulfate ion	SO_4^{-2}
sulfite ion	SO_3^{-2}
phosphate ion	PO_4^{-3}
phosphite ion	PO_3^{-3}

Note that only one polyatomic ion in this Table, the ammonium ion ($NH_4^{+1)}$, is a cation. This polyatomic ion contains one nitrogen and four hydrogens that collectively bear a +1 charge. The remaining polyatomic ions are all negatively-charged and, therefore, are classified as anions. However, only two of these, the hydroxide ion and the cyanide ion, are named using the "-ide" suffix that is typically indicative of negatively-charged particles. The remaining polyatomic anions, which all contain oxygen, in combination with another non-metal, exist as part of a series in which the number of oxygens within the polyatomic unit can vary. As has been repeatedly emphasized in several sections of this text, no two chemical formulas should share a common chemical name. A single suffix, "-ide," is insufficient for distinguishing the names of the anions in a related polyatomic series. Therefore, "-ate" and "-ite" suffixes are employed, in order to denote that the corresponding polyatomic ions are part of a series. Additionally, these suffixes also indicate the relative number of oxygens that are contained within the polyatomic ions. Note that all of the polyatomic ions whose names end in "-ate" contain one more oxygen than those polyatomic anions whose names end in "-ite." Unfortunately, much like the common system for naming transition metals, these suffixes only indicate the *relative* number of oxygens that are contained within the polyatomic ions. For example, the nitrate ion, which is symbolized as NO_3^{-1}, has one more oxygen than the nitrite ion, which is symbolized as NO_2^{-1}. However, the sulfate ion is symbolized as SO_4^{-2}. While both the nitrate ion and the sulfate ion share an "-ate" suffix, the former contains three oxygens, but the latter contains four. Additionally, both the nitrate ion and the sulfite ion contain three oxygens, but these polyatomic ions do not share a common suffix. Unfortunately, the relative nature of these suffixes mandates that the ion formula/ion name combinations of the polyatomic ions must simply be memorized.

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



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



Example 5.11.2

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

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

Solution

a. Potassium ions have a charge of 1+, while sulfate ions have a charge of 2−. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.
b. Calcium ions have a charge of 2+, while nitrate ions have a charge of 1−. We will need two nitrate ions to balance the charge on each calcium ion. The formula for nitrate must be enclosed in parentheses. Thus, we write $Ca(NO_3)_2$ as the formula for this ionic compound.

Exercise 5.11.2

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

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

Answer a:

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

Answer b:

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

Recognizing Ionic Compounds

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

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

Example 5.11.3

Identify each compound as ionic or not ionic.

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

Solution

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

Exercise 5.11.3

Identify each compound as ionic or not ionic.

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

Answer a:

not ionic

Answer b:

ionic

Answer c:

ionic

Answer d:

not ionic

Looking Closer: Blood and Seawater

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

Table showing a comparison of the amounts of ions in blood and seawater.

Ion	Percent in Seawater	Percent in Blood
Na^+	2.36	0.322
Cl^-	1.94	0.366
Mg^{2+}	0.13	0.002

Ion	Percent in Seawater	Percent in Blood
SO_4^{2-}	0.09	—
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
HCO_3^-	0.002	0.165
HPO_4^{2-}, $H_2PO_4^-$	—	0.01

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

Key Takeaways

- Proper chemical formulas for ionic compounds balance the total positive charge with the total negative charge.
- Groups of atoms with an overall charge, called polyatomic ions, also exist.

EXERCISES

1. What information is contained in the formula of an ionic compound?

- Why do the chemical formulas for some ionic compounds contain subscripts, while others do not?

3. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Mg^{2+} and I^-
   b. Na^+ and O^{2-}

4. Write the chemical formula for the ionic compound formed by each pair of ions.

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

5. Write the chemical formula for the ionic compound formed by each pair of ions.

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

6. Write the chemical formula for the ionic compound formed by each pair of ions.

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

7. Write the chemical formula for the ionic compound formed by each pair of ions.

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

8. Write the chemical formula for the ionic compound formed by each pair of ions.

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

9. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Cr^{3+} and O^{2-}
   b. Cr^{2+} and O^{2-}
   c. Pb^{2+} and Cl^-
   d. Pb^{4+} and Cl^-

10. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. Cr^{3+} and NO_3^-
   b. Fe^{2+} and PO_4^{3-}
   c. Ca^{2+} and CrO_4^{2-}
   d. Al^{3+} and OH^-

11. Write the chemical formula for the ionic compound formed by each pair of ions.

   a. NH_4^+ and NO_3^-
   b. H^+ and $Cr_2O_7^{2-}$
   c. Cu^+ and CO_3^{2-}

d. Na^+ and HCO_3^-

12. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

   a. Ba and S
   b. Cs and I

13. For each pair of elements, determine the charge for their ions and write the proper formula for the resulting ionic compound between them.

   a. K and S
   b. Sc and Br

14. Which compounds would you predict to be ionic?

   a. Li_2O
   b. $(NH_4)_2O$
   c. CO_2
   d. $FeSO_3$
   e. C_6H_6
   f. C_2H_6O

15. Which compounds would you predict to be ionic?

   a. $Ba(OH)_2$
   b. CH_2O
   c. NH_2CONH_2
   d. $(NH_4)_2CrO_4$
   e. C_8H_{18}
   f. NH_3

Answers

1. the ratio of each kind of ion in the compound

2. Sometimes more than one ion is needed to balance the charge on the other ion in an ionic compound.

3.

   a. MgI_2
   b. Na_2O

4.

   a. NaBr
   b. $MgBr_2$
   c. MgS

5.

   a. KCl
   b. $MgCl_2$
   c. MgSe

6.

   a. Na_3N
   b. Mg_3N_2
   c. Al_2S_3

7.

   a. Li_3N
   b. Mg_3P_2
   c. Li_3P

8.

   a. $FeBr_3$
   b. $FeBr_2$
   c. Au_2S_3
   d. Au_2S

9.

   a. Cr_2O_3
   b. CrO
   c. $PbCl_2$
   d. $PbCl_4$

10.

   a. $Cr(NO_3)_3$
   b. $Fe_3(PO_4)_2$
   c. $CaCrO_4$

d. $Al(OH)_3$

11.

a. NH_4NO_3
b. $H_2Cr_2O_7$
c. Cu_2CO_3
d. $NaHCO_3$

12.

a. Ba^{2+}, S^{2-}, BaS
b. Cs^{+}, I^{-}, CsI

13.

a. K^{+}, S^{2-}, K_2S
b. Sc^{3+}, Br^{-}, $ScBr_3$

14.

a. ionic
b. ionic
c. not ionic
d. ionic
e. not ionic
f. not ionic

15.

a. ionic
b. not ionic
c. not ionic
d. ionic
e. not ionic
f. not ionic

CHAPTER OVERVIEW

6: Covalent Bonds, Covalent Compounds

6.1: Naming Covalent Compounds

Learning Objectives

- Write the chemical names of covalent molecules.
- Nomenclature of alkanes

In the previous section, the process for using a Lewis structure to determine the chemical formula of a covalent molecule was presented and applied. The chemical name of a molecule, in turn, is derived based on the information included in its chemical formula. As stated previously, a chemical name should uniquely correspond to a single chemical formula. In other words, no two chemical formulas should share a common chemical name.

Recall that covalent bonds are produced when unpaired electrons found within two atoms interact to form a shared pair of electrons. As a Lewis structure is the result of correctly executing this sharing process, the chemical formula and the chemical name of the corresponding covalent molecule must accurately reflect the information shown in that Lewis structure. The procedure for writing the chemical name of a covalent molecule is explained in the following paragraphs.

Naming Covalent Molecules

For a covalent molecule, the information represented in its chemical name must be a direct reflection of its Lewis structure. Therefore, the chemical name of a covalent molecule must not only contain information that indicates the identities of its constituent elements, but also must reflect how many of each of those elements are present within the molecule. Note that, if written properly, the chemical formula for a covalent molecule also contains this information and, therefore, can be used as the basis for developing a chemical name. Elemental names are incorporated into a covalent molecule's chemical name in the order in which their corresponding elemental symbols appear in the chemical formula. The suffix on the second elemental term is replaced with "-ide," in order to indicate its secondary placement within the chemical formula. Finally, since the subscripts in a covalent chemical formula are used to indicate how many times each elemental symbol appears in the molecule's Lewis structure, corresponding numerical prefixes, which are listed below in Table $6.1.1$, must be incorporated into the molecule's chemical name.

Table $6.1.1$: Numerical Prefixes Used When Naming Covalent Molecules

Number of Atoms in the Molecule (Indicated by a Subscript)	Numerical Prefix
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Note that the prefix "mono-" is *never* used to change the first elemental term in a covalent chemical name and should *only* be used as a modifier on the remaining term if the secondary element is oxygen. Finally, an "a" or "o" at the end of a prefix is usually dropped if the name of the element that is being altered begins with a vowel.

These prefix rules are best exemplified when writing the chemical name of "CO," which is the chemical formula of a commonly-known molecule.

This molecule is formed using covalent bonds, as its constituent elements, C (carbon) and O (oxygen), are both non-metals.

Since the elemental symbol "C" appears first in the given chemical formula, "carbon" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." However, this prefix is not used to alter the first elemental term in a covalent chemical name. Therefore, the first word in the chemical name of this molecule is "carbon."

Because the elemental symbol "O" is written second in the given chemical formula, "oxide" becomes the basis of the second word in the molecule's chemical name. The suffix on this elemental term is "-ide," as an indicator of its secondary placement within the chemical formula. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." Because the secondary element is oxygen, this prefix is used as a modifier on the term "oxide." Finally, since "oxide" begins with a vowel, the "o" of "mono-" is dropped. Therefore, the second word in the chemical name of this covalent molecule is "monoxide."

The result of combining these words, "**carbon monoxide**," is the chemically-correct name for CO.

✓ Example 6.1.1

Write the chemical name of SF_2, a covalent molecule that is formed when fluorine and sulfur bond with one another.

Solution

Since the elemental symbol "S" appears first in the given chemical formula, "sulfur" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." However, this prefix is not used to alter the first elemental term in a covalent chemical name. Therefore, the first word in the chemical name of this molecule is "sulfur."

Because the elemental symbol "F" is written second in the given chemical formula, "fluoride" becomes the basis of the second word in the molecule's chemical name. The suffix on this elemental term is "-ide," as an indicator of its secondary placement within the chemical formula. The subscript on this elemental symbol, a "2," corresponds to prefix of "di-." Therefore, the second word in the chemical name of this covalent molecule is "difluoride."

The result of combining these words, "**sulfur difluoride**," is the chemically-correct name for SF_2.

? Exercise 6.1.1

Write the chemical name that corresponds to each of the following chemical formulas.

a. CCl_4, the covalent molecule that is formed when carbon and chlorine bond with one another.
b. PBr_3, a covalent molecule that is formed when bromine and phosphorus bond with one another.
c. N_3P_6, a covalent molecule that is formed when nitrogen and phosphorus bond with one another.
d. I_2O_5, a covalent molecule that is formed when oxygen and iodine bond with one another.
e. HF, the covalent molecule that is formed when hydrogen and fluorine bond with one another.

Answer a

Since the elemental symbol "C" appears first in the given chemical formula, "carbon" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." However, this prefix is not used to alter the first elemental term in a covalent chemical name. Therefore, the first word in the chemical name of this molecule is "carbon."

Because the elemental symbol "Cl" is written second in the given chemical formula, "chloride" becomes the basis of the second word in the molecule's chemical name. The suffix on this elemental term is "-ide," as an indicator of its secondary placement within the chemical formula. The subscript on this elemental symbol, a "4," corresponds to prefix of "tetra-." Therefore, the second word in the chemical name of this covalent molecule is "tetrachloride."

The result of combining these words, "**carbon tetrachloride**," is the chemically-correct name for CCl_4.

Answer b

Since the elemental symbol "P" appears first in the given chemical formula, "phosphorus" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." However, this prefix is not used to alter the first elemental term in a covalent chemical name. Therefore, the first word in the chemical name of this molecule is "phosphorus."

Because the elemental symbol "Br" is written second in the given chemical formula, "bromide" becomes the basis of the second word in the molecule's chemical name. The suffix on this elemental term is "-ide," as an indicator of its secondary placement within the chemical formula. The subscript on this elemental symbol, a "3," corresponds to prefix of "tri-." Therefore, the second word in the chemical name of this covalent molecule is "tribromide."

The result of combining these words, "**phosphorus tribromide**," is the chemically-correct name for PBr_3.

Answer c

Since the elemental symbol "N" appears first in the given chemical formula, "nitrogen" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, a "3," corresponds to prefix of "tri-." Therefore, the first word in the chemical name of this molecule is "trinitrogen."

Because the elemental symbol "P" is written second in the given chemical formula, "phosphide" becomes the basis of the second word in the molecule's chemical name. The suffix on this elemental term is "-ide," as an indicator of its secondary placement within the chemical formula. The subscript on this elemental symbol, a "6," corresponds to prefix of "hexa-." Therefore, the second word in the chemical name of this covalent molecule is "hexaphosphide."

The result of combining these words, "**trinitrogen hexaphosphide**," is the chemically-correct name for N_3P_6.

Answer d

Since the elemental symbol "I" appears first in the given chemical formula, "iodine" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, a "2," corresponds to prefix of "di-." Therefore, the first word in the chemical name of this molecule is "diiodine."

Because the elemental symbol "O" is written second in the given chemical formula, "oxide" becomes the basis of the second word in the molecule's chemical name. The subscript on this elemental symbol, a "5," corresponds to prefix of "penta-." Since "oxide" begins with a vowel, the "a" of "penta-" is dropped. Therefore, the second word in the chemical name of this covalent molecule is "pentoxide."

The result of combining these words, "**diiodine pentoxide**," is the chemically-correct name for I_2O_5.

Answer e

Since the elemental symbol "H" appears first in the given chemical formula, "hydrogen" is the basis of the first word in the molecule's chemical name. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." However, this prefix is not used to alter the first elemental term in a covalent chemical name. Therefore, the first word in the chemical name of this molecule is "hydrogen."

Because the elemental symbol "F" is written second in the given chemical formula, "fluoride" becomes the basis of the second word in the molecule's chemical name. The subscript on this elemental symbol, an unwritten "1," corresponds to prefix of "mono-." However, this prefix is only used as a modifier on the second term in a covalent chemical name if the secondary element is oxygen. Because the secondary element is fluorine, the "mono-" prefix is not applied. Therefore, the second word in the chemical name of this molecule is "fluoride."

The result of combining these words, "**hydrogen fluoride**," is the chemically-correct name for HF.

Alkanes

Most of the hydrocarbons in petroleum belong to a family of compounds called the **alkanes**, in which all carbon atoms are linked by single bonds. The names of all alkanes end with *-ane*. Whether or not the carbons are linked together end-to-end in a ring (called *cyclic alkanes* or *cycloalkanes*) or whether they contain side chains and branches, the name of every carbon-hydrogen chain that lacks any double bonds or functional groups will end with the suffix *-ane*.

Alkanes with unbranched carbon chains are simply named by the number of carbons in the chain. The first four members of the series (in terms of number of carbon atoms) are named as follows:

1. CH_4 = **methane** = one hydrogen-saturated carbon
2. C_2H_6 = **ethane** = two hydrogen-saturated carbons
3. C_3H_8 = **propane** = three hydrogen-saturated carbons
4. C_4H_{10} = **butane** = four hydrogen-saturated carbons

Alkanes with five or more carbon atoms are named by adding the suffix *-ane* to the appropriate numerical multiplier, except the terminal *-a* is removed from the basic numerical term. Hence, C_5H_{12} is called *pentane*, C_6H_{14} is called *hexane*, C_7H_{16} is called *heptane* and so forth.

Straight-chain alkanes are sometimes indicated by the prefix *n-* (for normal) to distinguish them from branched-chain alkanes having the same number of carbon atoms. Although this is not strictly necessary, the usage is still common in cases where there is an important difference in properties between the straight-chain and branched-chain isomers: e.g. *n-hexane* is a neurotoxin while its branched-chain isomers are not.

IUPAC nomenclature

The IUPAC nomenclature is a system on which most organic chemists have agreed to provide guidelines to allow them to learn from each others' works. Nomenclature, in other words, provides a foundation of language for organic chemistry.

Number of Hydrogen to Carbons

This equation describes the relationship between the number of hydrogen and carbon atoms in alkanes:

$$\mathbf{H = 2C + 2}$$

where "C" and "H" are used to represent the number of carbon and hydrogen atoms present in one molecule. If C = 2, then H = 6.

Many textbooks put this in the following format:

$$\mathbf{C_nH_{2n+2}}$$

where "C_n" and "H_{2n+2}" represent the number of carbon and hydrogen atoms present in one molecule. If $C_n = 3$, then $H_{2n+2} = 2(3) + 2 = 8$. (For this formula look to the "n" for the number, the "C" and the "H" letters themselves do not change.)

Progressively longer hydrocarbon chains can be made and are named systematically, depending on the number of carbons in the longest chain.

The following table contains the systematic names for the first ten straight chain alkanes.

Name	Molecular Formula	Condensed Structural Formula
Methane	CH_4	CH_4
Ethane	C_2H_6	CH_3CH_3
Propane	C_3H_8	$CH_3CH_2CH_3$
Butane	C_4H_{10}	$CH_3(CH_2)_2CH_3$
Pentane	C_5H_{12}	$CH_3(CH_2)_3CH_3$
Hexane	C_6H_{14}	$CH_3(CH_2)_4CH_3$
Heptane	C_7H_{16}	$CH_3(CH_2)_5CH_3$
Octane	C_8H_{18}	$CH_3(CH_2)_6CH_3$

Name	Molecular Formula	Condensed Structural Formula
Nonane	C_9H_{20}	$CH_3(CH_2)_7CH_3$
Decane	$C_{10}H_{22}$	$CH_3(CH_2)_8CH_3$

Contributors

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6.2: Writing Lewis Structures for Covalent Compounds

Learning Objectives

- Draw Lewis structures for covalent compounds.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions.

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

For CO_3^{2-}, for example, we add two electrons to the total because of the −2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons).
- The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 6.2.1: Water

Write the Lewis Structure for H_2O.

Solution

Solutions to Example 10.4.1

Steps for Writing Lewis Structures	Example 6.2.1
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.

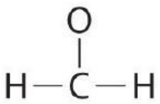
Steps for Writing Lewis Structures	Example 6.2.1
2. Arrange the atoms to show specific connections.	H O H Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives H -O- H with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary.
7. Final check.	The Lewis structure gives oxygen an octet and each hydrogen 2 electrons.

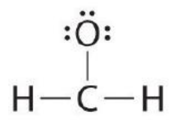
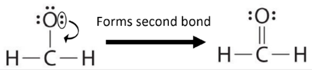
✓ Example 6.2.2

Write the Lewis structure for the CH_2O molecule

Solution

Solutions to Example 10.4.2

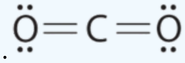
Steps for Writing Lewis Structures	Example 6.2.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following:  6 electrons are used, and 6 are left over.

Steps for Writing Lewis Structures	Example 6.2.2
4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:  Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.
5. If any electrons are left over, place them on the central atom.	Not necessary. There are no electrons left to place on the central atom.
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond: 
7. Final check	Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 6.2.1

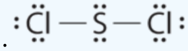
Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

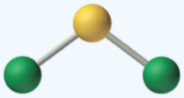
Answer CO_2



Carbon dioxide

Answer SCl_2





Sulfur dichloride

The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Writing Lewis Structures for Polyatomic Ions (CK-12)

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia (NH_3) molecule in a coordinate covalent bond.

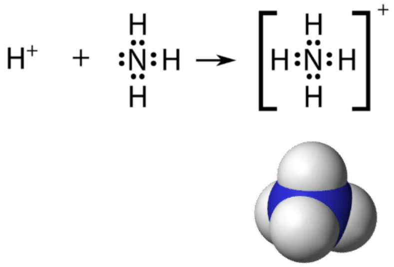


Figure 6.2.3: The ammonium ion. (CK12 License)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$$1 \text{ N atom} = 5 \text{ valence electrons}$$

$$4 \text{ H atoms} = 4 \times 1 = 4 \text{ valence electrons}$$

$$\text{subtract 1 electron for the } 1+ \text{ charge of the ion}$$

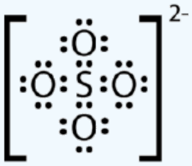
$$\text{total of 8 valence electrons in the ion}$$

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

✓ Exercise 6.2.2

Draw the Lewis electron dot structure for the sulfate ion.

Answer (CK12 License)



6.3: Resonance - Equivalent Lewis Structures for The Same Molecule

Learning Objectives

- Explain the concept of resonance and how it works with within molecules.

Resonance

There are some cases in which more than one viable Lewis structure can be drawn for a molecule. An example is the ozone (O_3) molecule in Figure 6.3.1. There are a total of 18 electrons in the structure and so the following two structures are possible.

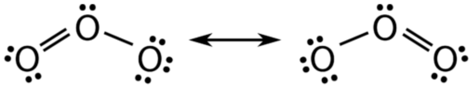


Figure 6.3.1: Resonance forms of ozone. Note the use of the double-headed arrow.

The structure on the left (6.3.1) can be converted to the structure on the right by a shifting of electrons without altering the positions of the atoms.

It was once thought that the structure of a molecule such as O_3 consisted of one single bond and one double bond which then shifted back and forth as shown above. However, further studies showed that the two bonds are identical. Any double covalent bond between two given atoms is typically shorter than a single covalent bond. Studies of the O_3 and other similar molecules showed that the bonds were identical in length. Interestingly, the length of the bond is in between the lengths expected for an $O-O$ single bond and a double bond.

Resonance is the use of two or more Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms are best thought of as being "one and a half" bonds, as opposed to either a pure single bond or a pure double bond. This "half-bond" can be shown as a dotted line in both the Lewis structure and the molecular model (Figure 6.3.2).

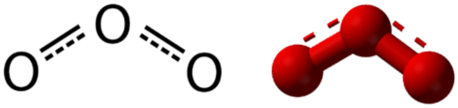


Figure 6.3.2: "Half-bond" model of ozone molecule. This is a better description of the electronic structure of ozone than either of the resonance structures in Figure 6.3.1.

Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, NO_3^- in Figure 6.3.3.

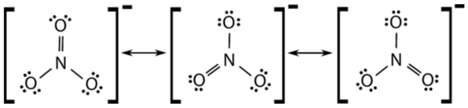


Figure 6.3.3: Resonance structure of nitrate anion.

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

6.4: Molecular Geometry

Learning Objective

- Determine the shape of simple molecules.

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

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

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

Figure 6.4.1: Beryllium hydride and carbon dioxide bonding.

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

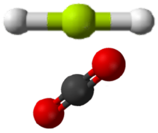


Figure 6.4.2: Beryllium hydride and carbon dioxide models. (CK12 Licence)

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

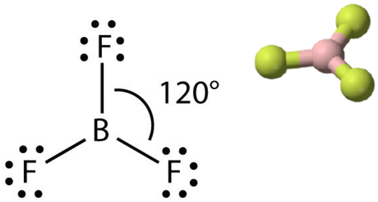


Figure 6.4.3: Boron trifluoride bonding. (CK12 Licence)

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

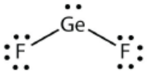


Figure 6.4.4: Germanium difluoride bonding.

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

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 6.4.1 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

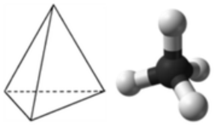


Figure 6.4.5: Tetrahedral structure of methane. (CK12 Licence)

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

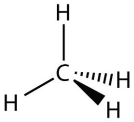


Figure 6.4.6: Methane bonding. (CK12 Licence)

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

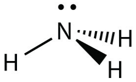


Figure 6.4.7: Ammonia bonding. (CK12 Licence)

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

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

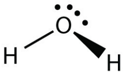


Figure 6.4.8: Water bonding.

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

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 6.4.9.

Figure 6.4.9: Lewis Electron Dot Diagram of Formaldehyde.

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

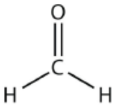


Figure 6.4.10: Formaldehyde bonding.

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

Table 6.4.1 summarizes the shapes of molecules based on the number of electron groups and surrounding atoms.

Table 6.4.1: Summary of Molecular Shapes

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

✓ Example 6.4.1

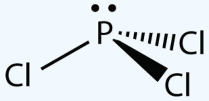
What is the approximate shape of each molecule?

a. PCl3
b. NOF

Solution

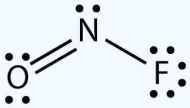
The first step is to draw the Lewis structure of the molecule.

For PCl_3, the electron dot diagram is as follows:



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

The electron dot diagram for NOF is as follows:



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

? Exercise 6.4.1

What is the approximate molecular shape of CH_2Cl_2?

Answer

Tetrahedral

? Exercise 6.4.2

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

Answer

Trigonal planar about both central C atoms.

Summary

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

6.5: Exception for Octet Rule

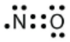
Learning Objectives

- Identify common exceptions to the octet rule

Exceptions to the Octet Rule (BC Campus)

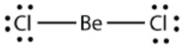
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

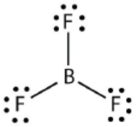


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

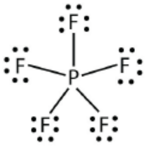
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

Example 6.5.3: Octet Violations

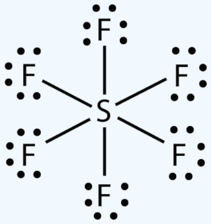
Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:

b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



Exercise 6.5.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer

The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

The octet rule states that atoms tend to achieve eight electrons in their valence shell for stability, but there are key exceptions. These include incomplete octets (e.g., boron), expanded octets (e.g., phosphorus), and odd-electron species (e.g., nitrogen oxide). Understanding these exceptions is essential for accurately predicting molecular behavior and bonding properties.

6.6: Content in Context

In this unit, we'll explore how chemistry helps us understand the composition of the air around us and the effects of pollutants on our health and environment. Now that you've learned to name chemical compounds, we'll apply that knowledge by identifying common air pollutants such as carbon monoxide, sulfur dioxide, and nitrogen oxides. Understanding their chemical properties will help us investigate where these pollutants come from and how they impact the world we live in.

6.6.1: Earth's Atmosphere- Divisions and Composition

Learning Objectives

- Describe the different layers of the atmosphere.
- Know the composition of air.

The Earth's atmosphere is composed of several layers (Figure 6.6.1.1). The lowest layer, the troposphere, extends from the Earth's surface up to about 6 miles or 10 kilometers (km) in altitude. Virtually all human activities occur in the troposphere. Mt. Everest, the tallest mountain on the planet, is only about 5.6 miles (9 km) high. The next layer, the stratosphere, continues from 6 miles (10 km) to about 31 miles (50 km). Most commercial airplanes fly in the lower part of the stratosphere.

Most atmospheric ozone is concentrated in a layer in the stratosphere, about 9 to 18 miles (15 to 30 km) above the Earth's surface (see the figure below). Ozone is a molecule that contains three oxygen atoms. At any given time, ozone molecules are constantly formed and destroyed in the stratosphere. The total amount has remained relatively stable during the decades that it has been measured.

The ozone layer in the stratosphere absorbs a portion of the radiation from the sun, preventing it from reaching the planet's surface. Most importantly, it absorbs the portion of UV light called UVB. UVB has been linked to many harmful effects, including skin cancers, cataracts, xand harm to some crops and marine life.

The mesosphere starts just above the stratosphere and extends to 85 kilometers (53 miles) high. Meteors burn up in this layer

The thermosphere starts just above the mesosphere and extends to 600 kilometers (372 miles) high. Aurora and satellites occur in this layer. The ionosphere is an abundant layer of electrons and ionized atoms and molecules that stretches from about 48 kilometers (30 miles) above the surface to the edge of space at about 965 km (600 mi), overlapping into the mesosphere and thermosphere. This dynamic region grows and shrinks based on solar conditions and divides further into the sub-regions: D, E and F; based on what wavelength of solar vvradiation is absorbed. The ionosphere is a critical link in the chain of Sun-Earth interactions. This region is what makes radio communications possible. The exosphere is is the upper limit of the atmosphere. It extends from the top of the thermosphere up to 10,000 km (6,200 mi).

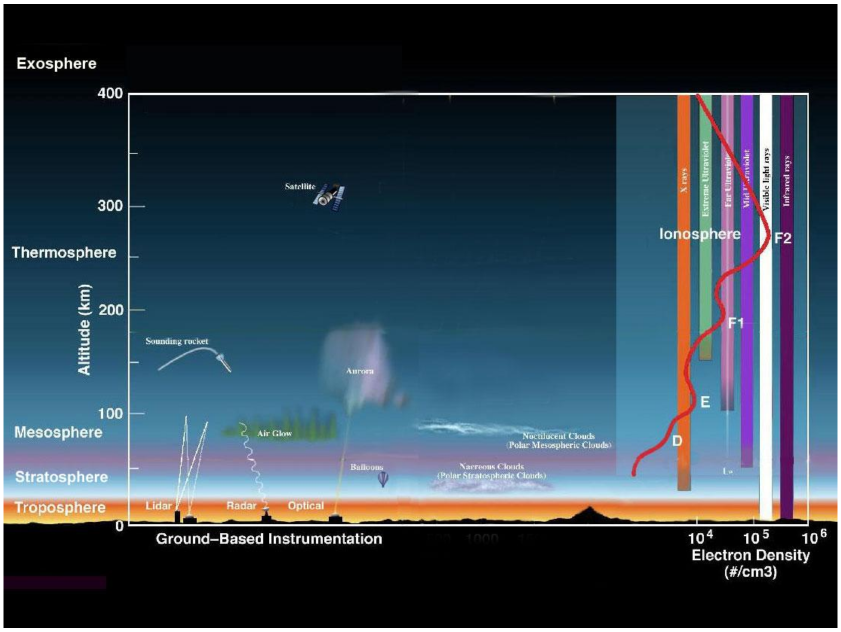


***Figure 6.6.1.1** The Earth's atmosphere*
***Credit:** NASA/Goddard*

Composition of the Atmosphere

Except for water vapor, whose atmospheric abundance varies from practically zero up to 4%, the fractions of the major atmospheric components N_2, O_2, and Ar are remarkably uniform below about 100 km (Table 6.6.1.1). At greater heights, diffusion becomes the principal transport process, and the lighter gases become relatively more abundant. In addition, photochemical processes result in the formation of new species whose high reactivities would preclude their existence in significant concentrations at the higher pressures found at lower elevations.

The atmospheric gases fall into three abundance categories: major, minor, and trace. Nitrogen, the most abundant component, has accumulated over time as a result of its geochemical inertness; a very small fraction of it passes into the other phases as a result of biological activity and natural fixation by lightning. It is believed that denitrifying bacteria in marine sediments may provide the major route for the return of N_2 to the atmosphere. Oxygen is almost entirely of biological origin, and cycles through the hydrosphere, the biosphere, and sedimentary rocks. Argon consists mainly of Ar^{40} which is a decay product of K^{40} in the mantle and crust.

Table 6.6.1.1 *The Major Components of Clean Dry Air in the Lower Atmosphere*

nitrogen	N_2	78.08 %
oxygen	O_2	20.95 %
argon	Ar	0.93 %

The most abundant of the minor gases aside from water vapor is carbon dioxide (Table 6.6.1.2). Next in abundance are neon and helium. Helium is a decay product of radioactive elements in the earth, but neon and the other inert gases are primordial, and have probably been present in their present relative abundances since the earth's formation. Two of the minor gases, ozone and carbon monoxide, have abundances that vary with time and location. A variable abundance implies an imbalance between the rates of formation and removal. In the case of carbon monoxide, whose major source is anthropogenic (a small amount is produced by biological action), the variance is probably due largely to localized differences in fuel consumption, particularly in internal combustion engines. The nature of the carbon monoxide sink (removal mechanism) is not entirely clear; it may be partly microbial.

Table 6.6.1.2 *The Minor Components of Air in the Lower Atmosphere.*

water	H_2O	0-4 %
carbon dioxide	CO_2	325 ppm
neon	Ne	18 ppm
helium	He	5 ppm
methane	CH_4	2 ppm
krypton	Kr	1 ppm
hydrogen	H_2	0.5 ppm
nitrous oxide	N_2O	0.3 ppm
carbon monoxide	CO	0.05-0.2 ppm
ozone	O_3	0.02 - 10 ppm
xenon	Xe	0.08 ppm

Summary

- The layers of the atmosphere include the trophosphere, stratosphere, mesosphere, thermosphere, and exosphere.
- The major components of dry clean air are N_2, O_2, and Ar.
- The most abundant of the minor gases are water vapor and CO_2

Contributors and Attributions

- US Environmental Protection Agency (EPA)
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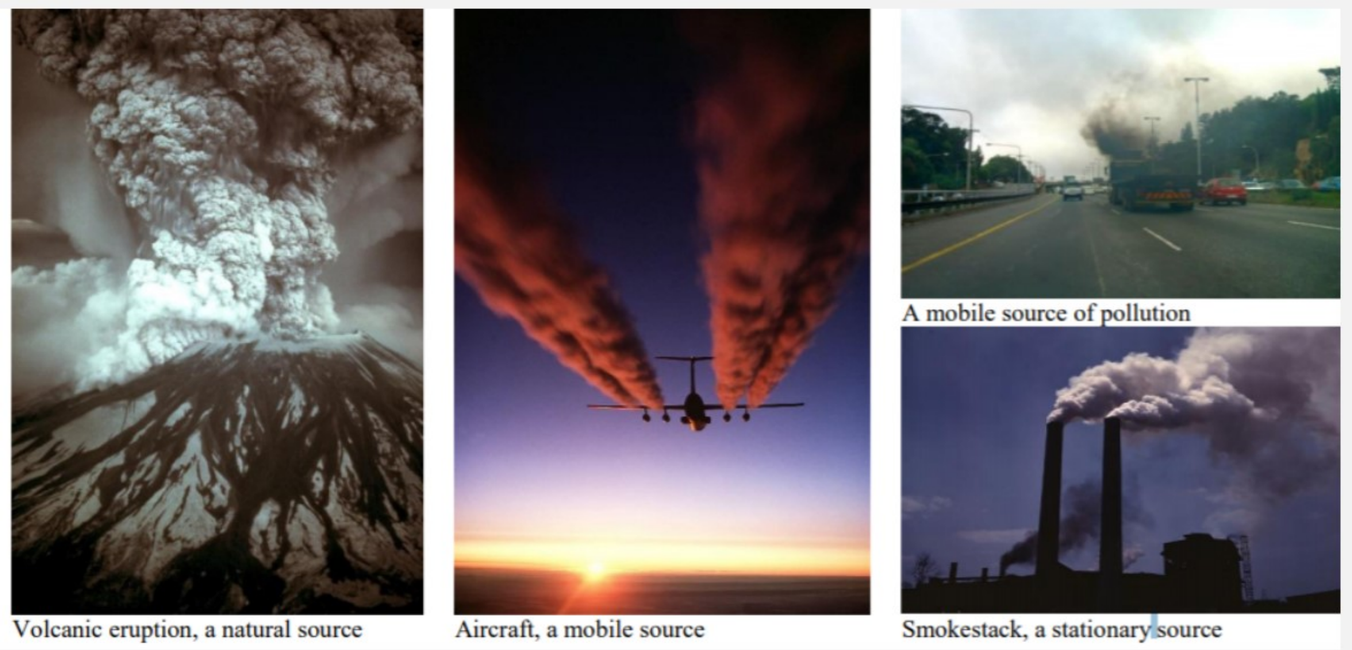
6.6.2: Outdoor Air Pollution and Sources

Learning Objectives

- Define air pollution.
- Know the sources and effects of major outdoor pollutants

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. 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.

Outdoor pollutants can come from **stationary (point)** sources or **mobile (nonpoint)** sources (Figure 6.6.2.1). Stationary sources have a fixed location, for example power plant smokestacks, burning, construction sites, farmlands and surface mines among others. Mobile sources of air pollutants move from place to place while emitting pollutants. Examples of mobile sources include vehicles, aircrafts, ships, and trains.



Volcanic eruption, a natural source

Aircraft, a mobile source

A mobile source of pollution

Smokestack, a stationary source

Figure 6.6.2.1: Images showing various sources of air pollution, including natural and anthropogenic, stationary and mobile. Source: All images obtained from Wiki Commons commons.wikimedia.org/wiki/Air_pollution

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 Clean Air Act, the Environmental Protection Agency (EPA) establishes air quality standards to protect public health and the environment. EPA has set national air quality standards for six common air pollutants namely: 1) carbon monoxide; 2) ground-level ozone; 3) nitrogen dioxide; 4) Sulfur dioxide; 5) lead; and 6) particulate matter (also known as particle pollution). Of the six pollutants, particle pollution and ground-level ozone are the most widespread health threats. 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.

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 come from mobile sources. CO can cause harmful health effects by reducing oxygen delivery to the body's organs (like the heart and brain) and tissues. At extremely high levels, CO can cause death.
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.
3. **Nitrogen dioxide (NO_2)**: is one of a group of highly reactive gasses known as "oxides of nitrogen," or "nitrogen oxides (NO_x)." Other nitrogen oxides include nitrous acid 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 utility, and industrial combustion.
4. **Sulfur dioxide (SO_2)**: Sulfur dioxide is one of a group of highly reactive gasses 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.
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 in 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 larger than 2.5 micrometers and smaller than 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 6.6.2.2). 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.

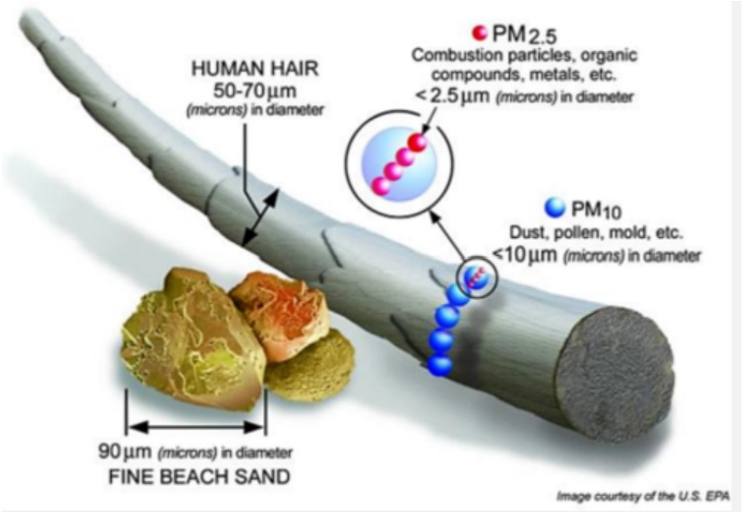


Figure 6.6.2.2 Graphic showing size comparison of particulate material (PM_{10} and $PM_{2.5}$) compared to fine beach sand and human hair. Source: US EPA http://www.epa.gov/air/particlepollution/basic.html

Volatile Organic Compounds

Volatile organic compounds (VOCs) are carbon-containing chemicals emitted as gases from natural and human-made sources. Natural sources of VOCs include plants, the largest source, and 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, 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 VOC. Also, 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 mobile sources (e.g., cars, trucks, buses) and stationary sources (e.g., factories, refineries, power plants), as well as indoor sources (e.g., 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

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 6.6.2.1) 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 6.6.2.1: Sources and health effects of criteria pollutants

Pollutant	Sources	Health Effects
Ground-level Ozone (O_3)	Secondary pollutant typically formed by chemical reaction of volatile organic compounds (VOCs) and NO_x in the presence of sunlight.	Decreases 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, emergency department (ED) visits, and premature mortality.
Particulate Matter (PM)	Emitted or formed through chemical reactions; fuel combustion (e.g., burning coal, wood, diesel); industrial processes; agriculture (plowing, field burning); and unpaved roads.	Short-term exposures can aggravate heart or lung diseases leading to respiratory symptoms, increased medication use, hospital admissions, ED visits, and premature mortality; long-term exposures can lead to the development of heart or lung disease and premature mortality.
Lead	Smelters (metal refineries) and other metal industries; combustion of leaded gasoline in piston engine aircraft; waste incinerators; and battery manufacturing.	Damages the developing nervous system, resulting in IQ loss and impacts on learning, memory, and behavior in children. Cardiovascular and renal effects in adults and early effects related to anemia.
Oxides of Nitrogen (NO_x)	Fuel combustion (e.g., electric utilities, industrial boilers, and vehicles) and wood burning.	Aggravate lung diseases leading to respiratory symptoms, hospital admissions, and ED visits; increased susceptibility to respiratory infection.
Carbon Monoxide (CO)	Fuel combustion (especially vehicles), industrial processes, fires, waste combustion, and residential wood burning.	Reduces the amount of oxygen reaching the body’s organs and tissues; aggravates heart disease, resulting in chest pain and other symptoms leading to hospital admissions and ED visits.
Sulfur Dioxide (SO_2)	Fuel combustion (especially high-sulfur coal); electric utilities and industrial processes; and natural sources such as volcanoes.	Aggravates asthma and increased respiratory symptoms. Contributes to particle formation with associated health effects.

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 (e.g., 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.

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 their reaction products (CO, NO_x, Ozone) as well as particulate matter, SO_2, and VOCs.
- 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.)
- TextMap: General CHemistry (Averill and Eldredge)
- Marisa Alviar-Agnew (Sacramento City College)
- Wikipedia

6.6.2.1: Automobile Emissions

Learning Objective

- Identify the different pollutant gases in automobile emissions.

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 motor vehicle emissions (and from stationary internal combustion engines), which can also include:

- Crankcase blow-by
- Evaporation of unused gasoline

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.

Carbon monoxide and nitrogen oxides are the two main pollutant gases from automobile emissions. Ozone is a result of the reaction between nitrogen oxides and volatile organic compounds (VOCs).

Carbon Monoxide

Carbon Monoxide (CO) is a colorless, odorless gas that can be harmful when inhaled in large amounts. CO is released 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.

Harmful effects

Breathing air with a high concentration of CO reduces the amount of oxygen that can be transported in the blood stream to critical organs like the heart and brain.

At very high levels, which are possible indoors or in other enclosed environments, CO can cause dizziness, confusion, unconsciousness and death.

Very high levels of CO are not likely to occur outdoors. However, when CO levels are elevated outdoors, they can be of particular concern for people with some types of heart disease. These people already have a reduced ability for getting oxygenated blood to their hearts in situations where the heart needs more oxygen than usual. They are especially vulnerable to the effects of CO when exercising or under increased stress. In these situations, short-term exposure to elevated CO may result in reduced oxygen to the heart accompanied by chest pain also known as angina.

Nitrogen Oxides

Nitrogen dioxide (NO_2) is one of a group of highly reactive gasses known as "oxides of nitrogen," or "nitrogen oxides (NO_x)." Other nitrogen oxides include nitrous acid 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 utility, and industrial combustion.

NO_2 primarily gets in the air from the burning of fuel. NO_2 forms from emissions from cars, trucks and buses, power plants, and off-road equipment.

Health effects

Breathing air with a high concentration of NO_2 can irritate airways in the human respiratory system. Such exposures over short periods can aggravate respiratory diseases, particularly asthma, leading to respiratory symptoms (such as coughing, wheezing or difficulty breathing), hospital admissions and visits to emergency rooms. Longer exposures to elevated concentrations of NO_2 may contribute to the development of asthma and potentially increase susceptibility to respiratory infections. People with asthma, as well as children and the elderly are generally at greater risk for the health effects of NO_2.

NO_2 along with other NO_x reacts with other chemicals in the air to form both particulate matter and ozone. Both of these are also harmful when inhaled due to effects on the respiratory system.

Environmental effects

NO_2 and other NO_x interact with water, oxygen and other chemicals in the atmosphere to form acid rain. Acid rain harms sensitive ecosystems such as lakes and forests.

The nitrate particles that result from NO_x make the air hazy and difficult to see though. This affects the many national parks that we visit for the view. NO_x in the atmosphere contributes to nutrient pollution in coastal waters.

Ozone

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 is a gas composed of three atoms of oxygen (O_3). Ozone occurs both in the Earth's upper atmosphere and at ground level. Ozone can be good or bad, depending on where it is found.

Called stratospheric ozone, good ozone occurs naturally in the upper atmosphere, where it forms a protective layer that shields us from the sun's harmful ultraviolet rays. This beneficial ozone has been partially destroyed by manmade chemicals, causing what is sometimes called a "hole in the ozone." The good news is, this hole is diminishing.

Ozone at ground level is a harmful air pollutant, because of its effects on people and the environment, and it is the main ingredient in “smog."

How does ground-level ozone form?

Tropospheric, or ground level ozone, is not emitted directly into the air, but is created by chemical reactions between oxides of nitrogen (NOx) and volatile organic compounds (VOC). This happens when pollutants emitted by cars, power plants, industrial boilers, refineries, chemical plants, and other sources chemically react in the presence of sunlight.

Ozone is most likely to reach unhealthy levels on hot sunny days in urban environments, but can still reach high levels during colder months. Ozone can also be transported long distances by wind, so even rural areas can experience high ozone levels.

This graphic depicts the formation of ozone from car and factory emissions

Health effects

Ozone in the air we breathe can harm our health especially on hot sunny days when ozone can reach unhealthy levels. People most at risk from breathing air containing ozone include people with asthma, children, older adults, and people who are active outdoors, especially outdoor workers.

Environmental effects

Ozone affects sensitive vegetation and ecosystems, including forests, parks, wildlife refuges and wilderness areas. In particular, ozone harms sensitive vegetation during the growing season.

Summary

- Carbon monoxide and nitrogen oxides are the two main compounds from automobile emissions.
- Ozone but is created by chemical reactions between oxides of nitrogen (NOx) and volatile organic compounds (VOC).
- Different levels of exposure to automobile emissions and ozone can lead to various health problems.

Contributor

- US Environmental Protection Agency
- Marisa Alviar-Agnew (Sacramento City College)

6.6.3: Indoor air pollution

Learning Objectives

- Name the different indoor pollutants and their sources.
- Know the harmful effects of indoor pollutants.

In both developed and developing nations, indoor air pollution poses a greater health risk than outdoor air pollution. According to the World Health Organization (WHO) and other agencies such as the Environmental Protection Agency (EPA), indoor air generally contains higher concentrations of toxic pollutants than outdoor air. Additionally, people generally spend more time indoors than outdoors, hence, the health effects from indoor air pollution in workplaces, schools, and homes are far greater than outdoor. Indoor pollution sources that release gases or particles into the air are the primary cause of indoor air quality problems in homes. Inadequate ventilation can increase indoor pollutant levels by not bringing in enough outdoor air to dilute emissions from indoor sources and by not carrying indoor air pollutants out of the home.

Outdoor air enters and leaves a building by infiltration, natural ventilation, and mechanical ventilation. In infiltration, outdoor air flows into the house through openings, joints, and cracks in walls, floors, and ceilings, and around windows and doors. In natural ventilation, air moves through opened windows and doors. Air movement associated with infiltration and natural ventilation is caused by air temperature differences between indoors and outdoors and by wind. Finally, there are a number of mechanical ventilation devices, from outdoor-vented fans that intermittently remove air from a single room, such as bathrooms and kitchen, to air handling systems that use fans and duct work to continuously remove indoor air and distribute filtered and conditioned outdoor air to strategic points throughout the house. The rate at which outdoor air replaces indoor air is described as the air exchange rate. When there is little infiltration, natural ventilation, or mechanical ventilation, the air exchange rate is low and pollutant levels can increase. High temperature and humidity levels can also increase concentrations of some pollutants.

There are many sources of indoor air pollution in any home (Figure $6.6.3.1$). These include combustion sources such as oil, gas, kerosene, coal, wood, and tobacco products; building materials and furnishings as diverse as deteriorated, asbestos-containing insulation, wet or damp carpet, and cabinetry or furniture made of certain pressed wood products; products for household cleaning and maintenance, personal care, or hobbies; central heating and cooling systems and humidification devices. Pollutants causing indoor air pollution can also originate from outside sources such as radon, pesticides, and outdoor air pollution. Radon is a naturally occurring radioactive gas produced from the decay of uranium in rock. If a building/home is constructed in an area with uranium containing rock, the gas can seep through the foundations and accumulate in basements. Exposure to radon can cause lung cancer.

The relative importance of any single source depends on how much of a given pollutant it emits and how hazardous those emissions are. In some cases, factors such as how old the source is and whether it is properly maintained are significant. For example, an improperly adjusted gas stove can emit significantly more carbon monoxide than one that is properly adjusted. Some sources, such as building materials, furnishings, and household products like air fresheners, release pollutants more or less continuously. Other sources, related to activities carried out in the home, release pollutants intermittently. These include smoking, the use of unvented or malfunctioning stoves, furnaces, or space heaters, the use of solvents in cleaning and hobby activities, the use of paint strippers in redecorating activities, and the use of cleaning products and pesticides in house-keeping. High pollutant concentrations can remain in the air for long periods after some of these activities.

Risks from indoor air pollution differ between less industrialized and industrialized nations. Indoor pollution has a greater impact in less industrialized nations where many people use cheaper sources of fuel such as wood, charcoal, and crop waste among others for cooking and heating, often with little or no ventilation. The most significant indoor pollutant, therefore, is soot and carbon monoxide. In industrialized nations, the primary indoor air health risks are cigarette smoke and radon.

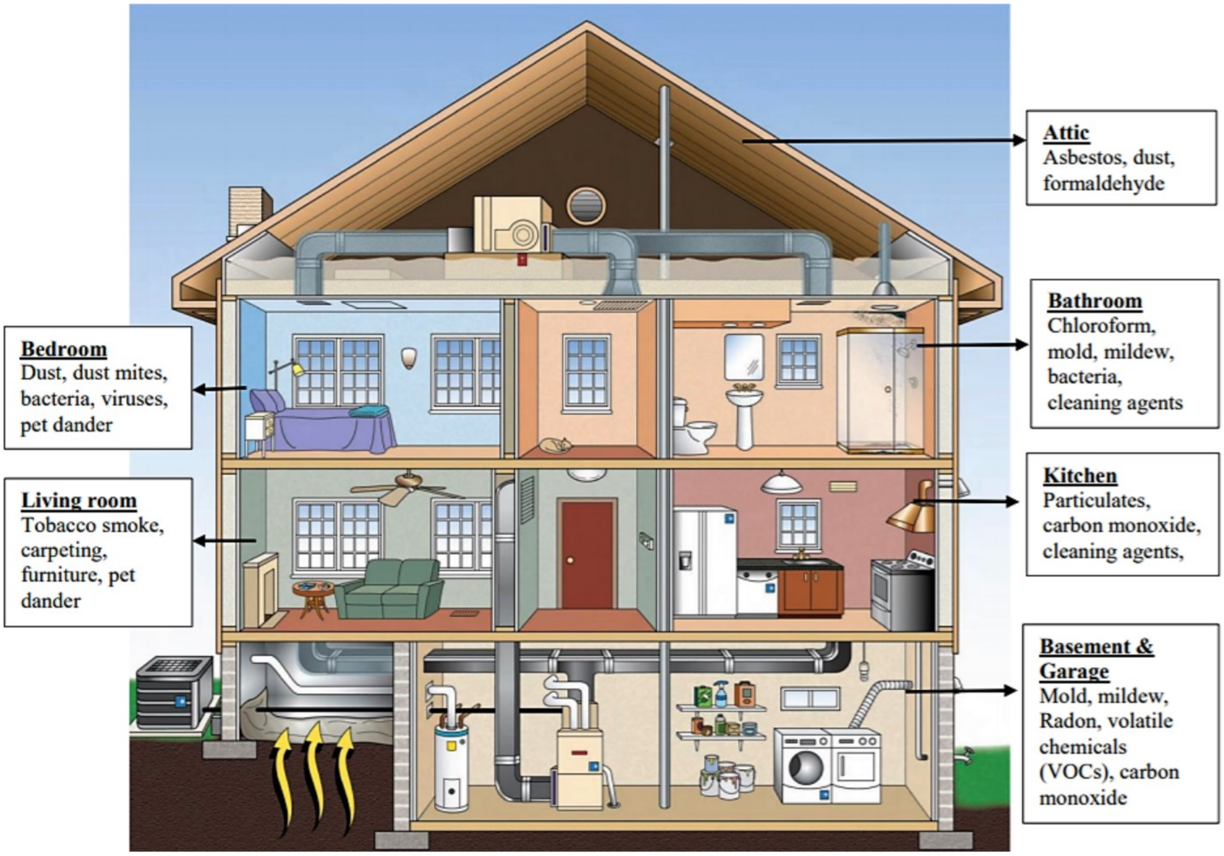


Figure 6.6.3.1: Sources of indoor air pollution. Image source: EPA http://www.epa.gov/iaq/pdfs/careforyourair.pdf

Combustion Pollutants

Combustion pollutants are gases or particles that come from burning materials. In homes, the major source of combustion pollutants are improperly vented or unvented fuel-burning appliances such as space heaters, wood stoves, gas stoves, water heaters, dryers, and fireplaces. The types and amounts of pollutants produced depends on the type of appliance, how well the appliance is installed, maintained and vented and the kind of fuel it uses. Common combustion pollutants include:

Carbon monoxide (CO) which is a colorless, odorless gas that interferes with the delivery of oxygen throughout the body. Carbon monoxide causes headaches, dizziness, weakness, nausea and even death. Average levels in homes without gas stoves vary from 0.5 to 5 parts per million (ppm). Levels near properly adjusted gas stoves are often 5 to 15 ppm and those near poorly adjusted stoves may be 30 ppm or higher.

Nitrogen dioxide (NO_2) which is a colorless, odorless gas that causes eye, nose and throat irritation, shortness of breath, and an increased risk of respiratory infection. Average level in homes without combustion appliances is about half that of outdoors. In homes with gas stoves, kerosene heaters or un-vented gas space heaters, indoor levels often exceed outdoor levels.

Particulate Matter (PM) which refers to fine particles that forms in smoke when wood or other organic matter burns.

Other pollutants from wood smoke In addition to particle pollution, wood smoke contains several toxic harmful air pollutants including benzene, formaldehyde, acrolein, and polycyclic aromatic hydrocarbons (PAHs).

Secondhand Smoke

Secondhand smoke is a mixture of the smoke given off by the burning of tobacco products, such as cigarettes, cigars or pipes and the smoke exhaled by smokers. Secondhand smoke is also called environmental tobacco smoke (ETS). Exposure to secondhand smoke is sometimes called involuntary or passive smoking. Secondhand smoke, classified by EPA as a Group A carcinogen, contains more than 7,000 substances. Secondhand smoke exposure commonly occurs indoors, particularly in homes and cars. Secondhand smoke can move between rooms of a home and between apartment units. Opening a window or increasing ventilation in a home or car is not protective from secondhand smoke.

Radon

Radon is a radioactive gas that results from the natural decay of uranium and radiumfound in nearly all rocks and soils. Elevated radon levels have been found in every state. Radon is in the atmosphere and can also be found in ground water. The national average for radon in outdoor air is 0.4 picocuries per liter (pCi/L), while the average for indoor air is 1.3 pCi/L.

Any building can have high levels of radon, including new and old homes, well-sealed and drafty homes, office buildings and schools, and homes with or without basements. Radon gas can get into buildings through cracks in solid floors and walls, construction joints, gaps in suspended floors and around service pipes, cavities inside walls, the water supply and building materials. Testing is the only way to know if your home has elevated radon levels.EPA recommends fixing your home when the radon level is at or above 4 pCi/L.

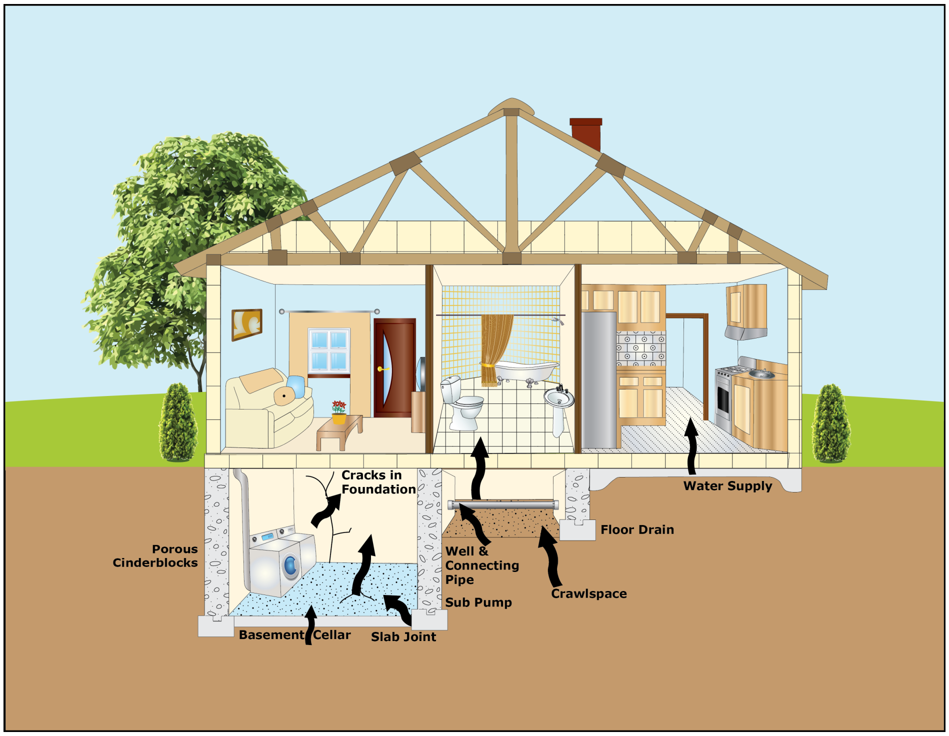


***Figure* $6.6.3.2$*: Sources of Radon. Image source: EPA*https://www.epa.gov/radiation/radionuclide-basics-radon

EPA estimates that about 21,000 lung cancer deaths each year in the U.S. are radon-related. Exposure to radon is the second leading cause of lung cancer after smoking. For most people, radon is the single greatest environmental source of radiation exposure. EPA recommends that all homes and schools be tested for radon. For smokers, the risk of lung cancer is heightened due to the combined effects of radon and smoking.

Other Indoor Pollutants

Molds are usually not a problem indoors, unless mold spores land on a wet or damp spot and begin growing. Molds have the potential to cause health problems. Molds produce allergens (substances that can cause allergic reactions) and irritants. Inhaling or touching mold or mold spores may cause allergic reactions in sensitive individuals. Allergic responses include hay fever-type symptoms, such as sneezing, runny nose, red eyes, and skin rash.

Allergic reactions to mold are common. They can be immediate or delayed. Molds can also cause asthma attacks in people with asthma who are allergic to mold. In addition, mold exposure can irritate the eyes, skin, nose, throat, and lungs of both mold-allergic and non-allergic people. Symptoms other than the allergic and irritant types are not commonly reported as a result of inhaling mold.

Volatile organic compounds (VOCs) are emitted as gases from certain solids or liquids. VOCs include a variety of chemicals, some of which may have short- and long-term adverse health effects. Concentrations of many VOCs are consistently higher indoors (up to ten times higher) than outdoors. VOCs are emitted by a wide array of products numbering in the thousands.

Organic chemicals are widely used as ingredients in household products. Paints, varnishes and wax all contain organic solvents, as do many cleaning, disinfecting, cosmetic, degreasing and hobby products. Fuels are 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.

EPA's Office of Research and Development's "Total Exposure Assessment Methodology (TEAM) Study" (Volumes I through IV, completed in 1985) found levels of about a dozen common organic pollutants to be 2 to 5 times higher inside homes than outside, regardless of whether the homes were located in rural or highly industrial areas. TEAM studies indicated that while people are using products containing organic chemicals, they can expose themselves and others to very high pollutant levels, and elevated concentrations can persist in the air long after the activity is completed.

VOCs are emitted by a wide array of products used in homes including paints and lacquers, paint strippers, cleaning supplies, varnishes and waxes, pesticides, building materials and furnishings, office equipment, moth repellents, air fresheners, and dry-cleaned clothing. VOCs evaporate into the air when these products are used or sometimes even when they are stored.

Volatile organic compounds irritate the eyes, nose and throat, and cause headaches, nausea, and damage to the liver, kidneys and central nervous system. Some of them can cause cancer.

Asbestos is a mineral fiber that occurs in rock and soil. Because of its fiber strength and heat resistance asbestos has been used in a variety of building construction materials for insulation and as a fire retardant. Asbestos has also been used in a wide range of manufactured goods, mostly in building materials (roofing shingles, ceiling and floor tiles, paper products, and asbestos cement products), friction products (automobile clutch, brake, and transmission parts), heat-resistant fabrics, packaging, gaskets, and coatings.

Asbestos fibers may be released into the air by the disturbance of asbestos-containing material during product use, demolition work, building or home maintenance, repair, and remodeling. In general, exposure may occur only when the asbestos-containing material is disturbed or damaged in some way to release particles and fibers into the air.

Three of the major health effects associated with asbestos exposure are:

· lung cancer

· mesothelioma, a rare form of cancer that is found in the thin lining of the lung, chest and the abdomen and heart

· asbestosis, a serious progressive, long-term, non-cancer disease of the lungs

Improving Your Indoor Air Quality

Take steps to help improve your air quality and reduce your IAQ-related health risks at little or no cost by:

Controlling the sources of pollution: Usually the most effective way to improve indoor air is to eliminate individual sources or reduce their emissions.

Ventilating: Increasing the amount of fresh air brought indoors helps reduce pollutants inside. When weather permits, open windows and doors, or run an air conditioner with the vent control open. Bathroom and kitchen fans that exhaust to the outdoors also increase ventilation and help remove pollutants.

Always ventilate and follow manufacturers' instructions when you use products or appliances that may release pollutants into the indoor air.

Changing filters regularly: Central heaters and air conditioners have filters to trap dust and other pollutants in the air. Make sure to change or clean the filters regularly, following the instructions on the package.

Adjusting humidity: The humidity inside can affect the concentrations of some indoor air pollutants. For example, high humidity keeps the air moist and increases the likelihood of mold.

Keep indoor humidity between 30 and 50 percent. Use a moisture or humidity gauge, available at most hardware stores, to see if the humidity in your home is at a good level. To increase humidity, use a vaporizer or humidifier. To decrease humidity, open the windows if it is not humid outdoors. If it is warm, turn on the air conditioner or adjust the humidity setting on the humid

Get a quick glimpse of some of the most important ways to protect the air in your home by touring the Indoor Air Quality (IAQ) House. Room-by-room, you'll learn about the key pollutants and how to address them.

Interactive Version

https://www.epa.gov/indoor-air-quality-iaq/interactive-tour-indoor-air-quality-demo-house#mainHouse

Text Version

https://www.epa.gov/indoor-air-quality-iaq/text-version-indoor-air-quality-house-tour

Summary

- Indoor pollutants include combustion pollutants (from burning of oil, gas, kerosene, coal, wood, and tobacco products); building materials and furnishings products for household cleaning and maintenance, personal care, or hobbies; central heating and cooling systems and humidification devices.

- Pollutants causing indoor air pollution can also originate from outside sources such as radon, pesticides, and outdoor air pollution.
- Indoor pollutants can cause serious health problems to sensitive groups.
- Several suggestions were provided to improve indoor air quality.

Contributors and Attributions

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US Environmental Protection Agency

CHAPTER OVERVIEW

7: Chemical Reactions

7.1: The Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and $\rightarrow$ appropriately when writing a chemical reaction.

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

Reactants and Products

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

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

$$\underbrace{\text{Flour} + \text{Butter} + \text{Sugar} + \text{Vanilla} + \text{Baking Soda} + \text{Eggs} + \text{Chocolate Chips}}_{\text{Ingredients} = \text{Reactants}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}}$$

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\mathrm{SO_2 + O_2}$, are reactants and sulfur trioxide, $\mathrm{SO_3}$, is the product.

$$\underbrace{\mathrm{2\,SO_2(g) + O_2(g)}}_{\text{Reactants}} \rightarrow \underbrace{\mathrm{2\,SO_3(g)}}_{\text{Products}}$$

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

$$\text{Reactants } \rightarrow \text{Products}$$

There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table $7.1.1$ shows a listing of symbols used in chemical equations.

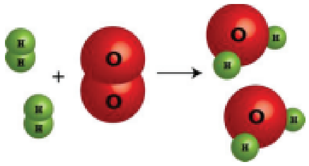
Table $7.1.1$: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Description
$+$	used to separate multiple reactants or products	(s)	reactant or product in the solid state
$\rightarrow$	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
$\rightleftharpoons$	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state

Symbol	Description	Symbol	Description
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

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

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



2. They could write a word equation for the chemical reaction:
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. They could write the equation in chemical shorthand.

$$2\text{H}_2\,(g)+\text{O}_2\,(g)\rightarrow 2\text{H}_2\text{O}\,(g)$$

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

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

$$\text{Ca(NO}_3)_2\,(aq)+2\text{NaOH}\,(aq)\rightarrow\text{Ca(OH)}_2\,(s)+2\text{NaNO}_3\,(aq)$$

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

$$\text{Cu}\,(s)+\text{AgNO}_3\,(aq)\rightarrow\text{Cu(NO}_3)_2\,(aq)+\text{Ag}\,(s)$$

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

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

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).
3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table 7.1.1: *Diatomic Elements*

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

✓ Example 7.1.1

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

a. $\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$
b. Gaseous propane, $\mathrm{C_3H_8}$, 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 $(\mathrm{C_3H_8})$ and oxygen $(\mathrm{O_2})$

Product: carbon dioxide $(\mathrm{CO_2})$ and water $(\mathrm{H_2O})$

$$\mathrm{C_3H_8}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathrm{HF}\,(g) + \mathrm{K_2CO_3}\,(aq) \rightarrow \mathrm{KF}\,(aq) + \mathrm{H_2O}\,(l) + \mathrm{CO_2}\,(g)$$

? Exercise 7.1.1

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

a. Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.
b. $\mathrm{HCl}\,(aq) + \mathrm{LiOH}\,(aq) \rightarrow \mathrm{LiCl}\,(aq) + \mathrm{H_2O}\,(l)$
c. Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer a

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

Answer b

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

Answer c

$Cu(s) + O_2(g) \rightarrow CuO(s)$

Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

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7.2: 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 Mass 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 two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

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

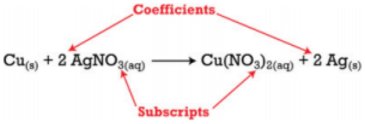
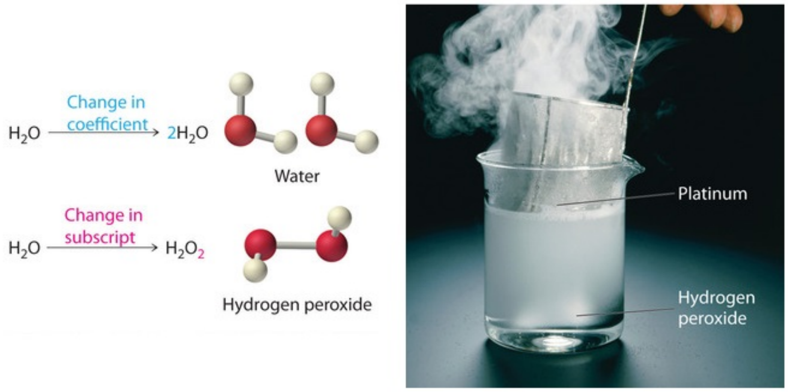


Figure 7.2.1: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.

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

Balancing a Chemical Equation

Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 7.2.1.



Original molecule H_2O: if the coefficient 2 is added in front, that makes 2 water molecules; but if the subscript 2 is added to make H_2O_2, that's hydrogen peroxide.

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

Steps in Balancing a Chemical Equation

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

Example 7.2.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane (C_7H_{16}).

$$\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)}$$

Solution

Solutions to Example 7.2.1

Steps	**Example**
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_7H_{16}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side: $\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow \underline{7}CO_2(g) + H_2O(g)}$ • 7 carbon atoms on both reactant and product sides b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side: $\mathrm{C_7H_{16}(l) + O_2(g) \rightarrow 7\,CO_2(g) + \underline{8}H_2O(g)}$ • 16 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2, on the reactant side: $\mathrm{C_7H_{16}(l) + \underline{11}O_2(g) \rightarrow 7\,CO_2(g) + 8\,H_2O(g)}$ • 22 oxygen atoms on both reactant and product sides
5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

Example 7.2.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

Solution

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

Solutions to Example 7.4.2

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18}. We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.**	a. The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products: $C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$ • 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$ • 18 hydrogen atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.**	The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $C_8H_{18}(l) + \underline{\frac{25}{2}}O_2(g) \rightarrow 8\,CO_2(g) + 9\,H_2O(g)$ • 25 oxygen atoms on both reactant and product sides The equation is now balanced, but we usually write equations with whole number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$

Steps	Example
5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

Example 7.2.3: Precipitation of Lead (II) Chloride

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

Solution

Solutions to Example 7.4.3

Steps	Example
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $\mathrm{Pb(NO_3)_2(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)}$
2. **Adjust the coefficients.**	There are twice as many chloride ions in the reactants as there are in the products. Place a 2 in front of the NaCl in order to balance the chloride ions. $\mathrm{Pb(NO_3)_2(aq) + \underline{2}NaCl(aq) \rightarrow NaNO_3(aq) + PbCl_2(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on reactant side, 1 Na atom on product side • 2 Cl atoms on both reactant and product sides
3. **Balance polyatomic ions as a unit.**	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $\mathrm{Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow \underline{2}NaNO_3(aq) + PbCl_2(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides
4. **Balance the remaining atoms.**	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$\mathrm{Pb(NO_3)_2(aq) + 2\,NaCl(aq) \rightarrow 2\,NaNO_3(aq) + PbCl_2(s)}$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO_3^- atoms on both reactant and product sides

? Exercise 7.2.1

Is each chemical equation balanced?

a. $2\,Hg(\ell)^{+}O_2(g) \rightarrow Hg_2O_2(s)$
b. $C_2H_4(g) + 2\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(g)$
c. $Mg(NO_3)_2(s) + 2\,Li(s) \rightarrow Mg(s)^{+}2\,LiNO_3(s)$

Answer a

yes

Answer b

no

Answer c

yes

? Exercise 7.2.2

Balance the following chemical equations.

a. $N_2(g) + O_2(g) \rightarrow NO_2(g)$
b. $Pb(NO_3)_2(aq) + FeCl_3(aq) \rightarrow Fe(NO_3)_3(aq) + PbCl_2(s)$
c. $C_6H_{14}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

Answer a

N_2 (g) + $2O_2$ (g) → $2NO_2$ (g)

Answer b

$3Pb(NO_3)_2(aq) + 2FeCl_3(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3PbCl_2(s)$

Answer c

$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(g)$

Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

7.3: Acid Base Reaction

Learning Objectives

- Identify common acids and bases
- Define acid-base reactions
- Recognize and identify examples of acid-base reactions

Acid-Base Reactions

***Video 7.3.1**: An introduction to acids and bases.*

An acid-base reaction is one in which a hydrogen ion, H^+, is transferred from one chemical species to another. Such reactions are of central importance to numerous natural and technological processes, ranging from the chemical transformations that take place within cells and the lakes and oceans, to the industrial-scale production of fertilizers, pharmaceuticals, and other substances essential to society. The subject of acid-base chemistry, therefore, is worthy of thorough discussion, which is unfortunately beyond the scope of this course.

For purposes of this brief introduction, we will consider only the more common types of acid-base reactions that take place in aqueous solutions. In this context, an acid is a substance that will dissolve in water to yield hydronium ions, H_3O^+. As an example, consider the equation shown here:

$$\mathrm{HCl}(aq) + \mathrm{H_2O}(aq) \rightarrow \mathrm{Cl^-}(aq) + \mathrm{H_3O^+}(aq) \tag{7.3.1}$$

The process represented by this equation confirms that hydrogen chloride is an acid. When dissolved in water, H_3O^+ ions are produced by a chemical reaction in which H^+ ions are transferred from HCl molecules to H_2O molecules (Figure 7.3.1).

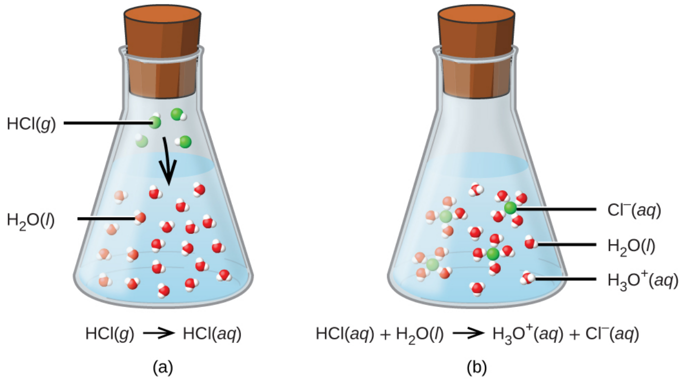


***Figure 7.3.1**: When hydrogen chloride gas dissolves in water, (a) it reacts as an acid, transferring protons to water molecules to yield (b) hydronium ions (and solvated chloride ions)*

The nature of HCl is such that its reaction with water as just described is essentially 100% efficient: Virtually every HCl molecule that dissolves in water will undergo this reaction. Acids that completely react in this fashion are called strong acids, and HCl is one among just a handful of common acid compounds that are classified as strong (Table 7.3.1). A far greater number of compounds behave as weak acids and only partially react with water, leaving a large majority of dissolved molecules in their original form and generating a relatively small amount of hydronium ions. Weak acids are commonly encountered in nature, being the substances partly responsible for the tangy taste of citrus fruits, the stinging sensation of insect bites, and the unpleasant smells associated with body odor. A familiar example of a weak acid is acetic acid, the main ingredient in food vinegars:

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons CH_3CO_2^-(aq) + H_3O^+(aq) \tag{7.3.2}$$

When dissolved in water under typical conditions, only about 1% of acetic acid molecules are present in the ionized form, $CH_3CO_2^-$ (Figure 7.3.2). (The use of a double-arrow in the equation above denotes the partial reaction aspect of this process, a concept addressed fully in the chapters on chemical equilibrium.)

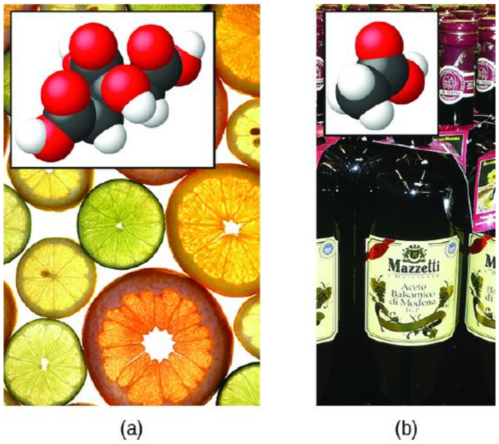


Figure 7.3.2: (a) Fruits such as oranges, lemons, and grapefruit contain the weak acid citric acid. (b) Vinegars contain the weak acid acetic acid. (credit a: modification of work by Scott Bauer; credit b: modification of work by Brücke-Osteuropa/Wikimedia Commons)

Table 7.3.1: Common Strong Acids

Compound Formula	Name in Aqueous Solution
HBr	hydrobromic acid
HCl	hydrochloric acid
HI	hydroiodic acid
HNO_3	nitric acid
$HClO_4$	perchloric acid
H_2SO_4	sulfuric acid

A base is a substance that will dissolve in water to yield hydroxide ions, OH^-. The most common bases are ionic compounds composed of alkali or alkaline earth metal cations (groups 1 and 2) combined with the hydroxide ion—for example, NaOH and $Ca(OH)_2$. When these compounds dissolve in water, hydroxide ions are released directly into the solution. For example, KOH and $Ba(OH)_2$ dissolve in water and dissociate completely to produce cations (K^+ and Ba^{2+}, respectively) and hydroxide ions, OH^-. These bases, along with other hydroxides that completely dissociate in water, are considered strong bases.

Consider as an example the dissolution of lye (sodium hydroxide) in water:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq) \tag{7.3.3}$$

This equation confirms that sodium hydroxide is a base. When dissolved in water, NaOH dissociates to yield Na^+ and OH^- ions. This is also true for any other ionic compound containing hydroxide ions. Since the dissociation process is essentially complete when ionic compounds dissolve in water under typical conditions, NaOH and other ionic hydroxides are all classified as strong bases.

Unlike ionic hydroxides, some compounds produce hydroxide ions when dissolved by chemically reacting with water molecules. In all cases, these compounds react only partially and so are classified as weak bases. These types of compounds are also abundant in nature and important commodities in various technologies. For example, global production of the weak base ammonia is typically well over 100 metric tons annually, being widely used as an agricultural fertilizer, a raw material for chemical synthesis of other compounds, and an active ingredient in household cleaners (Figure $7.3.3$). When dissolved in water, ammonia reacts partially to yield hydroxide ions, as shown here:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \tag{7.3.4}$$

This is, by definition, an acid-base reaction, in this case involving the transfer of H^+ ions from water molecules to ammonia molecules. Under typical conditions, only about 1% of the dissolved ammonia is present as NH_4^+ ions.


(a)

(b)

Figure $7.3.3$: *Ammonia is a weak base used in a variety of applications. (a) Pure ammonia is commonly applied as an agricultural fertilizer. (b) Dilute solutions of ammonia are effective household cleansers. (credit a: modification of work by National Resources Conservation Service; credit b: modification of work by pat00139)*

The chemical reactions described in which acids and bases dissolved in water produce hydronium and hydroxide ions, respectively, are, by definition, acid-base reactions. In these reactions, water serves as both a solvent *and* a reactant. A neutralization reaction is a specific type of acid-base reaction in which the reactants are an acid and a base, the products are often a salt and water, and neither reactant is the water itself:

$$\text{acid} + \text{base} \rightarrow \text{salt} + \text{water} \tag{7.3.5}$$

To illustrate a neutralization reaction, consider what happens when a typical antacid such as milk of magnesia (an aqueous suspension of solid $Mg(OH)_2$) is ingested to ease symptoms associated with excess stomach acid (HCl):

$$Mg(OH)_2(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + 2\,H_2O(l). \tag{7.3.6}$$

Note that in addition to water, this reaction produces a salt, magnesium chloride.

Example 7.3.1: Writing Equations for Acid-Base Reactions

Write balanced chemical equations for the acid-base reactions described here:

a. the weak acid hydrogen hypochlorite reacts with water
b. a solution of barium hydroxide is neutralized with a solution of nitric acid

Solution

(a) The two reactants are provided, HOCl and H_2O. Since the substance is reported to be an acid, its reaction with water will involve the transfer of H^+ from HOCl to H_2O to generate hydronium ions, H_3O^+ and hypochlorite ions, OCl^-.

$$HOCl(aq) + H_2O(l) \rightleftharpoons OCl^-(aq) + H_3O^+(aq)$$

A double-arrow is appropriate in this equation because it indicates the HOCl is a weak acid that has not reacted completely.

(b) The two reactants are provided, $Ba(OH)_2$ and HNO_3. Since this is a neutralization reaction, the two products will be water and a salt composed of the cation of the ionic hydroxide (Ba^{2+}) and the anion generated when the acid transfers its hydrogen ion (NO_3^-).

$$\mathrm{Ba(OH)_2}(aq) + 2\,\mathrm{HNO_3}(aq) \rightarrow \mathrm{Ba(NO_3)_2}(aq) + 2\,\mathrm{H_2O}(l)$$

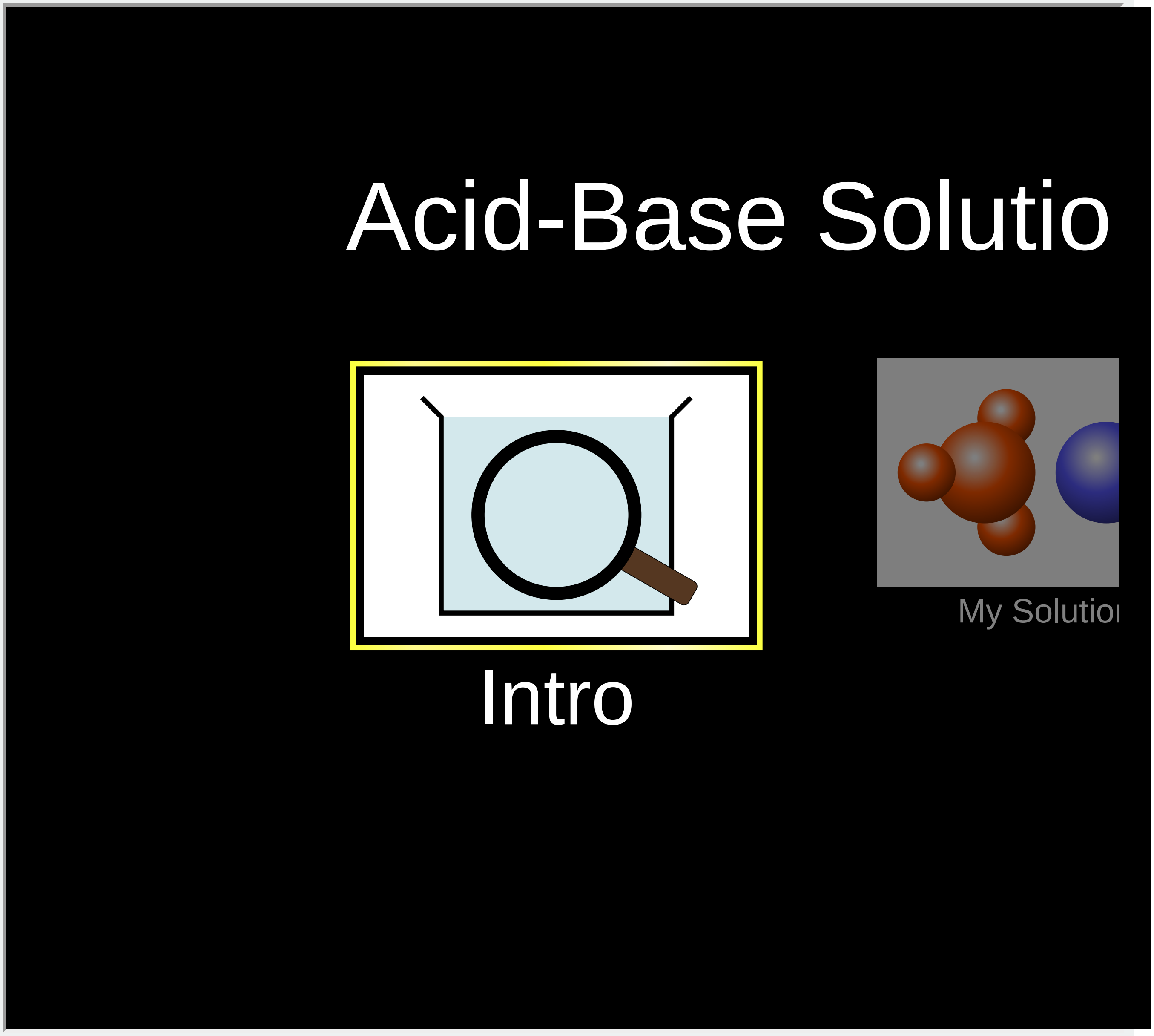
? Exercise 7.3.1

Write the net ionic equation representing the neutralization of any strong acid with an ionic hydroxide. (Hint: Consider the ions produced when a strong acid is dissolved in water.)

Answer

$$\mathrm{H_3O^+}(aq) + \mathrm{OH^-}(aq) \rightarrow 2\,\mathrm{H_2O}(l)$$

Phet Simulation



Chemistry in Everyday Life

Stomach Antacids

Our stomachs contain a solution of roughly 0.03 M HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

\ce{CaCO3}(s)+\ce{2HCl}(aq)⇌\ce{CaCl2}(aq)+\ce{H2O}(l)+\ce{CO2}(g)

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction:

\ce{Mg(OH)2}(s)⇌$Mg^{2+}(aq)+2OH^-(aq)$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that:

$$H_3O^+ + OH^- \rightleftharpoons 2H_2O(l)$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect. Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in concert with magnesium hydroxide to balance the side effects of the two substances.

Culinary Aspects of Chemistry

Examples of acid-base chemistry are abundant in the culinary world. One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter "rises." Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure 7.3.4). It turns out that fish have volatile amines (bases) in their systems, which are neutralized by the acids to yield involatile ammonium salts. This reduces the odor of the fish, and also adds a "sour" taste that we seem to enjoy.

$$CH_3COOH + NH_2CH_2CH_2CH_2CH_2NH_2 \longrightarrow CH_3COO^- + NH_3^+CH_2CH_2CH_2CH_2NH_2$$

$$\text{Acetic acid} + \text{Putrescine} \longrightarrow \text{Acetate ion} + \text{Putrescinium ion}$$

Figure* 7.3.4*: *A neutralization reaction takes place between citric acid in lemons or acetic acid in vinegar, and the bases in the flesh of fish.*

Pickling is a method used to preserve vegetables using a naturally produced acidic environment. The vegetable, such as a cucumber, is placed in a sealed jar submerged in a brine solution. The brine solution favors the growth of beneficial bacteria and suppresses the growth of harmful bacteria. The beneficial bacteria feed on starches in the cucumber and produce lactic acid as a waste product in a process called fermentation. The lactic acid eventually increases the acidity of the brine to a level that kills any harmful bacteria, which require a basic environment. Without the harmful bacteria consuming the cucumbers they are able to last much longer than if they were unprotected. A byproduct of the pickling process changes the flavor of the vegetables with the acid making them taste sour.

Summary

Video* 7.3.2*: *A summary of acid/base reactions.*

Chemical reactions are classified according to similar patterns of behavior. A large number of important reactions are included in three categories: precipitation, acid-base, and oxidation-reduction (redox). Precipitation reactions involve the formation of one or more insoluble products. Acid-base reactions involve the transfer of hydrogen ions between reactants. Redox reactions involve a change in oxidation number for one or more reactant elements. Writing balanced equations for some redox reactions that occur in aqueous solutions is simplified by using a systematic approach called the half-reaction method.

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7.4: Redox Reaction

Learning Objectives

- Define *oxidation* and *reduction*.
- Assign oxidation numbers to atoms in simple compounds.
- Recognize a reaction as an oxidation-reduction reaction.

Consider this chemical reaction:

$$\mathrm{Mg(s)+Cl_2(g)\rightarrow MgCl_2}$$

The reactants are two electrically neutral elements; they have the same number of electrons as protons. The product, however, is ionic; it is composed of Mg^{2+} and Cl^{-} ions. Somehow, the individual Mg atoms lose two electrons to make the Mg^{2+} ion, while the Cl atoms gain an electron to become Cl^{-} ions. This reaction involves the *transfer of electrons* between atoms.

The process of losing and gaining electrons occurs simultaneously. However, mentally we can separate the two processes. **Oxidation** is defined as the loss of one or more electrons by an atom. **Reduction** is defined as the gain of one or more electrons by an atom. So oxidation and reduction always occur together; it is only mentally that we can separate them. Chemical reactions that involve the transfer of electrons are called **oxidation-reduction (or redox) reactions**.

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use **oxidation numbers** to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on four rules. Oxidation numbers are not necessarily equal to the charge on the atom (although sometimes they can be); we must keep the concepts of charge and oxidation numbers separate.

Assigning Oxidation Numbers

The rules for assigning oxidation numbers to atoms are as follows:

1. Atoms in their elemental state are assigned an oxidation number of 0.
2. Atoms in monatomic (i.e., one-atom) ions are assigned an oxidation number equal to their charge. Oxidation numbers are usually written with the sign first, then the magnitude, to differentiate them from charges.
3. In compounds, fluorine is assigned a −1 oxidation number; oxygen is usually assigned a −2 oxidation number (except in peroxide compounds [where it is −1] and in binary compounds with fluorine [where it is positive]); and hydrogen is usually assigned a +1 oxidation number (except when it exists as the hydride ion [H^{-}], in which case rule 2 prevails).
4. In compounds, all other atoms are assigned an oxidation number so that the sum of the oxidation numbers on all the atoms in the species equals the charge on the species (which is zero if the species is neutral).

Here are some examples for practice. In H_2, both H atoms have an oxidation number of 0 by rule 1. In $MgCl_2$, magnesium has an oxidation number of +2, while chlorine has an oxidation number of −1 by rule 2. In H_2O, the H atoms each have an oxidation number of +1, while the O atom has an oxidation number of −2, even though hydrogen and oxygen do not exist as ions in this compound (rule 3). By contrast, by rule 3, each H atom in hydrogen peroxide (H_2O_2) has an oxidation number of +1, while each O atom has an oxidation number of −1. We can use rule 4 to determine oxidation numbers for the atoms in SO_2. Each O atom has an oxidation number of −2; for the sum of the oxidation numbers to equal the charge on the species (which is zero), the S atom is assigned an oxidation number of +4. Does this mean that the sulfur atom has a 4+ charge on it? No, it means only that the S atom is assigned a +4 oxidation number by our rules of apportioning electrons among the atoms in a compound.

Remembering Oxidation and Reduction

It is common to remember the difference between oxidation and reduction using one of two pneumonic devices:

1. "LeO goes GeR"
   In Greek Mythology, Leo was a lion (like the zodiac sign). And what do lion's do? Well, they roar but that doesn't really fit, so we go with "Ger".
   Loses **e**lectrons = **O**xidation
   Gains **e**lectrons = **R**eduction

Figure 7.4.1: *Lion image by Tambako, reused under the Creative Commons Attribution 2.0 Generic license.*

✓ Example 7.4.1

Assign oxidation numbers to the atoms in each substance.

a. Cl_2
b. GeO_2
c. $Ca(NO_3)_2$

Solution

1. Cl_2 is the elemental form of chlorine. Rule 1 states each atom has an oxidation number of 0.
2. By rule 3, oxygen is normally assigned an oxidation number of −2. For the sum of the oxidation numbers to equal the charge on the species (zero), the Ge atom is assigned an oxidation number of +4.
3. $Ca(NO_3)_2$ can be separated into two parts: the Ca^{2+} ion and the NO_3^- ion. Considering these separately, the Ca^{2+} ion has an oxidation number of +2 by rule 2. Now consider the NO_3^- ion. Oxygen is assigned an oxidation number of −2, and there are three of them. According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation $x + 3(-2) = -1$

   where x is the oxidation number of the N atom and the −1 represents the charge on the species. Evaluating for x,

   $x + (-6) = -1 x = +5$

   Thus the oxidation number on the N atom in the NO_3^- ion is +5.

Exercise 7.4.1: Phosphoric Acid

Assign oxidation numbers to the atoms in H_3PO_4.

Answer

H: +1; O: −2; P: +5

All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When an oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidized*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction can also be defined in terms of increasing or decreasing oxidation numbers, respectively.

Example 7.4.2

Identify what is being oxidized and reduced in this redox reaction.

$$2\,\mathrm{Na} + \mathrm{Br_2} \rightarrow 2\,\mathrm{NaBr}$$

Solution

Both reactants are the elemental forms of their atoms, so the Na and Br atoms have oxidation numbers of 0. In the ionic product, the Na^+ ions have an oxidation number of +1, while the Br^- ions have an oxidation number of −1.

$$2\underset{0}{Na} + \underset{0}{Br_2} \rightarrow 2\underset{+1-1}{NaBr}$$

Sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; bromine is decreasing its oxidation number from 0 to −1, so it is being reduced:

Because oxidation numbers are changing, this is a redox reaction. The total number of electrons being lost by sodium (two, one lost from each Na atom) is gained by bromine (two, one gained for each Br atom).

Exercise 7.4.2

Identify what is being oxidized and reduced in this redox reaction.

$$\mathrm{C} + \mathrm{O_2} \rightarrow \mathrm{CO_2}$$

Answer

C is being oxidized from 0 to +4; O is being reduced from 0 to −2.

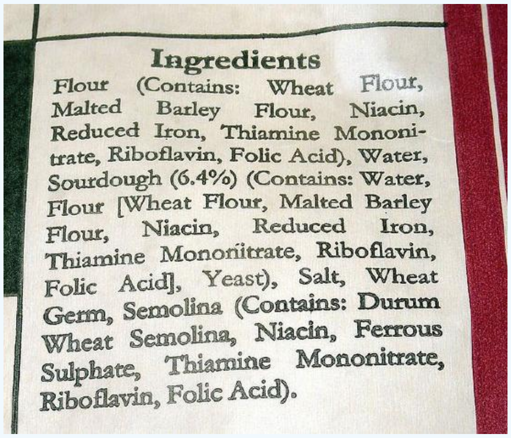
Oxidation reactions can become quite complex, as attested by the following redox reaction:

$$6H^+(aq) + 2\underset{+7}{MnO_4^-}(aq) + 5\underset{-1}{H_2O_2}(l) \rightarrow 2\underset{+2}{Mn^{2+}}(aq) + 5\underset{0}{O_2}(g) + 8H_2O(l)$$

To demonstrate that this is a redox reaction, the oxidation numbers of the species being oxidized and reduced are listed; can you determine what is being oxidized and what is being reduced? This is also an example of a net ionic reaction; spectator ions that do not change oxidation numbers are not displayed in the equation. Eventually, we will need to learn techniques for writing correct (i.e., balanced) redox reactions.

Food and Drink Application: Fortifying Food with Iron

Iron is an essential mineral in our diet; iron-containing compounds like the heme protein in hemoglobin could not function without it. Most biological iron has the form of the Fe^{2+} ion; iron with other oxidation numbers is almost inconsequential in human biology (although the body does contain an enzyme to reduce Fe^{3+} to Fe^{2+}, so Fe^{3+} must have some biological significance, albeit minor). To ensure that we ingest enough iron, many foods are enriched with iron. Although Fe^{2+} compounds are the most logical substances to use, some foods use "reduced iron" as an ingredient (bread and breakfast cereals are the most well-known examples). Reduced iron is simply iron metal; iron is added as a fine metallic powder. The metallic iron is oxidized to Fe^{2+} in the digestive system and then absorbed by the body, but the question remains: Why are we ingesting metallic iron? Why not just use Fe^{2+} salts as an additive?



Ingredients

Flour (Contains: Wheat Flour, Malted Barley Flour, Niacin, Reduced Iron, Thiamine Mononitrate, Riboflavin, Folic Acid), Water, Sourdough (6.4%) (Contains: Water, Flour [Wheat Flour, Malted Barley Flour, Niacin, Reduced Iron, Thiamine Mononitrate, Riboflavin, Folic Acid], Yeast), Salt, Wheat Germ, Semolina (Contains: Durum Wheat Semolina, Niacin, Ferrous Sulphate, Thiamine Mononitrate, Riboflavin, Folic Acid).

Figure 7.4.1 Ingredients. Many prepared foods list reduced iron in their ingredients list.

Although it is difficult to establish conclusive reasons, a search of scientific and medical literature suggests a few reasons. One reason is that fine iron filings do not affect the taste of the product. The size of the iron powder (several dozen micrometers) is not noticeable when chewing iron-supplemented foods, and the tongue does not detect any changes in flavor that can be detected when using Fe^{2+} salts. Fe^{2+} compounds can affect other properties of foodstuffs during preparation and cooking, like dough pliability, yeast growth, and color. Finally, of the common iron substances that might be used, metallic iron is the least expensive. These factors appear to be among the reasons why metallic iron is the supplement of choice in some foods.

Summary

- Oxidation-reduction (redox) reactions involve the transfer of electrons from one atom to another.
- Oxidation numbers are used to keep track of electrons in atoms.
- There are rules for assigning oxidation numbers to atoms.
- Oxidation is an increase in oxidation number (loss of electrons); reduction is a decrease in oxidation number (gain of electrons).

7.5: Combustion Reactions

Learning Objectives

- Recognize a reaction as a combustion reaction.

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.

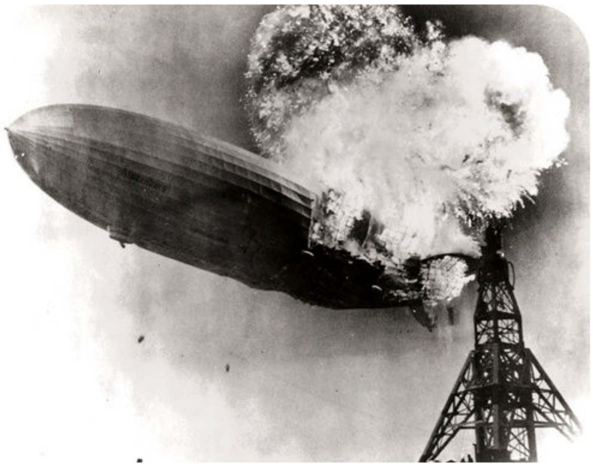


Figure 7.5.1: Explosion of the Hindenberg.

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 7.5.1: 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 7.5.1: Combustion of Hexane

Write the balanced equation for the combustion of hexane, C_6H_{14}, which is a liquid at room temperature.

Answer

$2\ C_6H_{14}\ (l) + 19\ O_2\ (g) \rightarrow 12\ CO_2\ (g) + 14\ H_2O\ (g)$

7.6: Content in Context

In this unit, we'll explore how air pollutants interact with other chemicals in the atmosphere to form new pollutants. Now that you're familiar with chemical reactions, we'll apply that knowledge to understand processes like how nitrogen oxides react with volatile organic compounds to create ground-level ozone, or how sulfur dioxide can transform into harmful particles. By investigating these reactions, you'll gain a deeper understanding of how pollutants evolve in the environment and their wider impact on air quality and health.

7.6.1: Smog

Learning Objectives

- Define smog.
- Describe industrial smog.
- Describe photochemical smog.

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. 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 $7.6.1.3$ and $7.6.1.4$). There are two main types of smog: industrial and photochemical smog. **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. **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.

Figure $7.6.1.3$: Smog over Almaty city, Kazakhstan. Photo by Igors Jefimovs. Source: Wikicommons commons.wikimedia.org/wiki/C...dia/File:Smog_ over_Almaty.jpg

Figure 7.6.1.4: Smog over Santiago in Chile. Source: German Wikipedia commons.wikimedia.org/wiki/C...#/media/File:Santiago30std.jpg

Industrial Smog

Industrial smog or London-type smog is mainly a product of burning large amounts of high sulfur coal. 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.

$$SO_2 + OH \rightarrow HOSO_2 \tag{7.6.1.1}$$

$$HOSO_2 + O_2 \rightarrow HO_2 + SO_3 \tag{7.6.1.2}$$

$$SO_3 + H_2O \rightarrow H_2SO_4 \tag{7.6.1.3}$$

Concerns about the harmful effects of acid rain have led to strong pressure on industry to minimize the release of SO_2 and NO. 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 ($CaSO_3 \cdot 2H_2O$; Figure 7.6.1.5).

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 (e.g., 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.

Photochemical Smog

Photochemical smog is a type of air pollution due to the reaction of solar radiation with airborne pollutant mixtures of nitrogen oxides (NOx) and volatile organic compounds (hydrocarbons). Smog is a byproduct of modern industrialization. Due to industry and the number of motor vehicles, this is more of a problem in large cities that have a warm, sunny and dry climate.

- **Oxidation**: Photochemical smog is also referred to as oxidizing smog. Oxidation reactions have been defined several ways. In terms of oxygen transfer, oxidation is a gain of oxygen. Oxidation can also be defined as a loss of hydrogen. The most important use of oxidation is described in terms of electron transfer. Oxidation can be described as an increase in oxidation number or loss of electrons. Oxidation numbers represents a distribution of charge. In other words, oxidation numbers represent the charge of the atom if the compound was composed of ions.

- **Reduction**: Reduction can involve the gain of hydrogen or loss of oxygen. Reduction can refer to the gain of electrons, which results in a decrease in oxidation number.

Formation of Photochemical Smog

The different reactions involved in the formation of photochemical smog are given below.

Step 1: People begin driving in the morning, nitrogen is burned or oxidized

$$N_2 + O_2 \rightarrow 2NO$$

- Oxidation number of N_2 is 0. The nitrogen in NO has acquired an oxidation number of +2.

Step 2: After a few hours, NO combines with O_2, in another oxidation reaction

$$2NO + O_2 \rightarrow 2NO_2$$

- The nitrogen in NO has an oxidation number of +2. The nitrogen in NO_2 has an oxidation number of +4.

Step 3: Nitrogen dioxide absorbs light energy, resulting in a reduction reaction

$$NO_2 \rightarrow NO + O$$

- The nitrogen in NO_2 has an oxidation number of +4 and the nitrogen in NO is +2.

Step 4: In sunlight, atomic oxygen combines with oxygen gas to form ozone

$$O + O_2 \rightarrow O_3$$

Step 5: Reaction is temperature and sunlight dependent

$$O_3 + NO \rightleftharpoons NO_2 + O_2$$

Alternative Reactions

NO and NO_2 can also react with the hydrocarbons instead of ozone to form other volatile compounds known as PAN (peroxyacetyl nitrate) as shown in Figure . 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. Panoramic view of Santiago covered by a layer of smog on May 10, 2006. The Metropolitan Region of Santiago facing the driest autumn last 28 years due to lack of rainfall, which coupled with poor air circulation, causes an increase in smog.

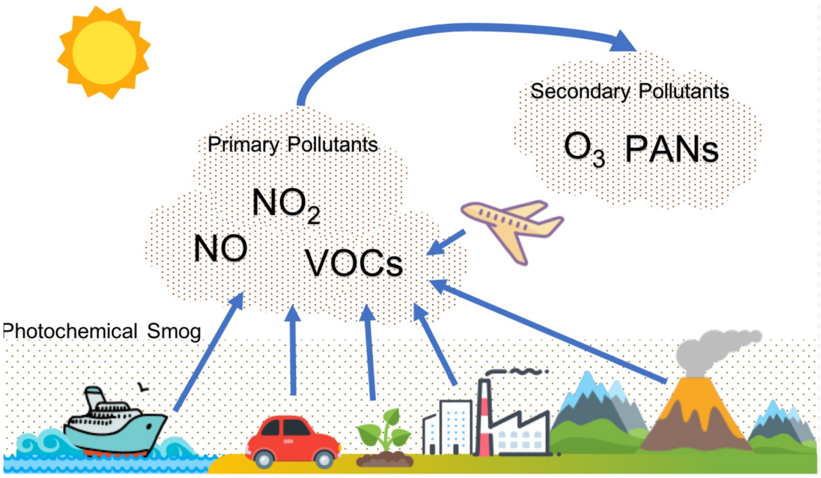


Figure 7.6.1.1 The photochemical smog formation diagram. Source: Wikipedia

Health Hazards

Because ozone is highly reactive, it has the ability to oxidize and destroy lung tissue. Short term exposures to elevated levels of ozone (above .75 ppm) have been linked to a host of respiratory irritations including coughing, wheezing, substernal soreness, pharyngitis, and dyspnea. Prolonged exposure to the molecule has been proven to cause a permanent reduction in lung function, as well as elevate the risk of developing asthma. Sulfur dioxide is a common component of London smog. Epidemiological studies have linked short term sulfur dioxide exposure to respiratory irritations including coughing, wheezing, and pharyngitis.

Other Harmful Effects of Smog

Plants are harmed by exposure to nitrogen oxides, ozone, and peroxyacetyl nitrate (PAN, see above), all oxidants present in a smoggy atmosphere. PAN is the most harmful of these constituents, damaging younger plant leaves, especially. Ozone exposure causes formation of yellow spots on leaves, a condition called chlorotic stippling. Some plant species, including sword-leaf lettuce, black nightshade, quickweed, and double-fortune tomato, are extremely susceptible to damage by oxidant species 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.

Controlling Photochemical Smog

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. Many scientists have suggested that pumping gas at night could reduce photochemical ozone formation by limiting the amount of exposure VOCs have with sunlight.

Preventing Smog with Green Chemistry

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 can help 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.

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 and “end-of-pipe” measures. These primitive measures implemented in the early 1970s 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. (And it would be much more efficient if those drivers who feel that they must drive “sport utility” behemoths would switch to vehicles of a more sensible size.) 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.

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 probably unsolvable problem of a very limited range between charges and the need for relatively heavy batteries. 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.

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 process 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.

Green chemistry can be applied to devices and processes other than automobiles to reduce smog-forming emissions. This is especially true in the area of organic solvents used for parts cleaning and other industrial operations, vapors of which are often released to the atmosphere. The substitution of water with proper additives or even the use of supercritical carbon dioxide fluid can eliminate such emissions.

Summary

- 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. 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.
- 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 to sunlight, creating a brown haze above cities.
- Photochemical smog is formed from the reactions of natural and man-made emissions of nitrogen oxides and VOCs.
- 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.
- 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.
- Catalytic converters in gas powered vehicles help reduce photochemical emissions.
- The practice of green chemistry and the application of the principles of industrial ecology can help to reduce smog.

Contributors and Attributions

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- Marisa Alviar-Agnew (Sacramento City College)

- Wikipedia
- www.epa.gov

7.6.2: Acid Rain

Learning Objective

- Describe acid rain and its effects.

Acid rain is a term referring to a mixture of wet and dry deposition (deposited material) from the atmosphere (Figure $7.6.2.1$) 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 mild solution of sulfuric acid and nitric acid.

$$SO_2 + HOH \rightarrow H_2SO_3 \tag{7.6.2.1}$$

$$2NO_2 + HOH \rightarrow HNO_2 + HNO_3 \tag{7.6.2.2}$$

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.

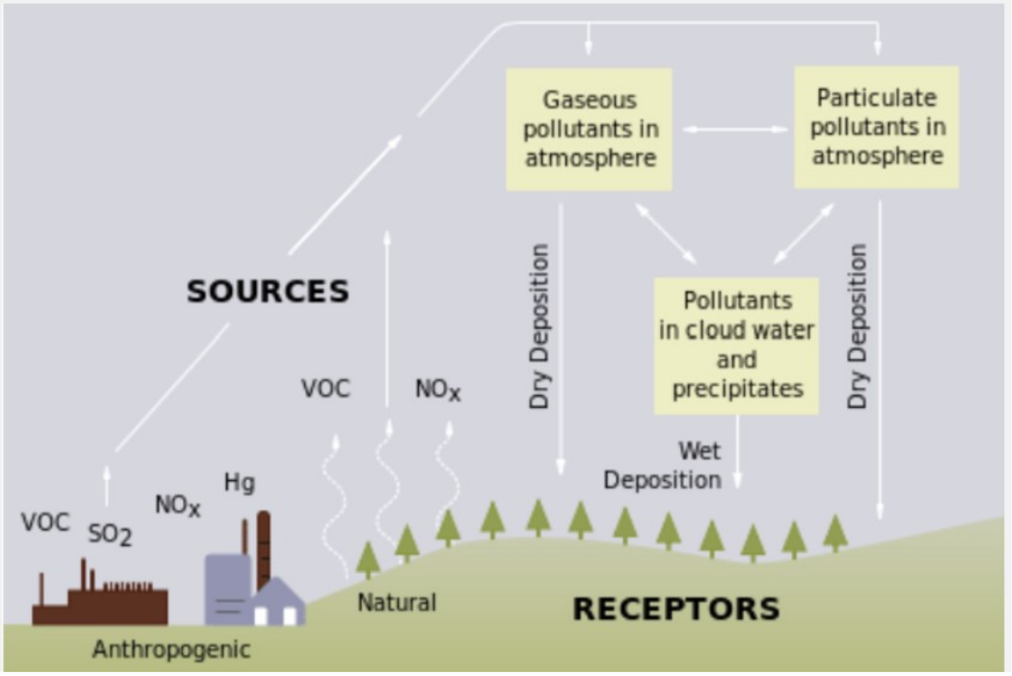

Figure 7.6.2.1 Processes involved in acid deposition.

Acid rain is measured using a scale called “pH.” 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 weak carbonic acid, giving the resulting mixture a pH of approximately 5.6 at typical atmospheric concentrations of CO_2. As of 2000, the most acidic rain falling in the U.S. has a pH of about 4.3.

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 nation’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 (i.e., the pH value goes down) 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. Acid rain causes slower growth, injury, or death of forests as shown in Figure $7.6.2.2$ 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 $7.6.2.2$. Acid rain damage to a forest in the Czech Republic.

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:

$$CaCO_{3(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + H_2O_{(l)} + CO_{2(g)} \tag{7.6.2.3}$$

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, cultural objects (such as statues, monuments, and tombstones), and cars (Figure $7.6.2.3$).

Figure $7.6.2.3$. A gargoyle that has been damaged by acid rain.

Sulfates and nitrates that form in the atmosphere from sulfur dioxide (SO_2) and nitrogen oxides (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—sulfur dioxide (SO_2) and nitrogen oxides (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.

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 rain water in the soil, streams, lakes, and marshes (and other bodies of water) can be harmful to trees, plants, 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).

Contributors and Attributions

- Essentials of Environmental Science by Kamala Doršner is licensed under CC BY 4.0. Modified from the original by Matthew R. Fisher.
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7.6.3: National Ambient Air Quality Standards

Learning Objectives

- Know the EPA standards for criteria air pollutants.

The Clean Air Act, which was last amended in 1990, requires EPA to set National Ambient Air Quality Standards (40 CFR part 50) for pollutants considered harmful to public health and the environment. The Clean Air Act identifies two types of national ambient air quality standards. ***Primary standards*** provide public health protection, including protecting the health of "sensitive" populations such as asthmatics, children, and the elderly. ***Secondary standards*** provide public welfare protection, including protection against decreased visibility and damage to animals, crops, vegetation, and buildings.

The EPA has set National Ambient Air Quality Standards for six principal pollutants, which are called "criteria" air pollutants. Periodically, the standards are reviewed and may be revised. The current standards are listed in Table 7.6.3.1. Units of measure for the standards are parts per million (ppm) by volume, parts per billion (ppb) by volume, and micrograms per cubic meter of air (μg/m^3).

Table 7.6.3.1 National Ambient Air Quality Standards for Criteria Air Pollutants.

Pollutant [links to historical tables of NAAQS reviews]		Primary/ Secondary	Averaging Time	Level	Form
Carbon Monoxide (CO)		primary	8 hours	9 ppm	Not to be exceeded more than once per year
			1 hour	35 ppm	
Lead (Pb)		primary and secondary	Rolling 3 month average	0.15 μg/m^3 (1)	Not to be exceeded
Nitrogen Dioxide (NO_2)		primary	1 hour	100 ppb	98th percentile of 1-hour daily maximum concentrations, averaged over 3 years
		primary and secondary	1 year	53 ppb (2)	Annual Mean
Ozone (O_3)		primary and secondary	8 hours	0.070 ppm (3)	Annual fourth-highest daily maximum 8-hour concentration, averaged over 3 years
Particle Pollution (PM)	$PM_{2.5}$	primary	1 year	12.0 μg/m^3	annual mean, averaged over 3 years
		secondary	1 year	15.0 μg/m^3	annual mean, averaged over 3 years
		primary and secondary	24 hours	35 μg/m^3	98th percentile, averaged over 3 years

Pollutant [links to historical tables of NAAQS reviews]		Primary/ Secondary	Averaging Time	Level	Form
	PM$_{10}$	primary and secondary	24 hours	150 µg/m^3	Not to be exceeded more than once per year on average over 3 years
Sulfur Dioxide (SO$_2$)		primary	1 hour	75 ppb (4)	99th percentile of 1-hour daily maximum concentrations, averaged over 3 years
		secondary	3 hours	0.5 ppm	Not to be exceeded more than once per year

(1) In areas designated nonattainment for the Pb standards prior to the promulgation of the current (2008) standards, and for which implementation plans to attain or maintain the current (2008) standards have not been submitted and approved, the previous standards (1.5 µg/m3 as a calendar quarter average) also remain in effect.

(2) The level of the annual NO2 standard is 0.053 ppm. It is shown here in terms of ppb for the purposes of clearer comparison to the 1-hour standard level.

(3) Final rule signed October 1, 2015, and effective December 28, 2015. The previous (2008) O3 standards additionally remain in effect in some areas. Revocation of the previous (2008) O3 standards and transitioning to the current (2015) standards will be addressed in the implementation rule for the current standards.

(4) The previous SO2 standards (0.14 ppm 24-hour and 0.03 ppm annual) will additionally remain in effect in certain areas: (1) any area for which it is not yet 1 year since the effective date of designation under the current (2010) standards, and (2)any area for which an implementation plan providing for attainment of the current (2010) standard has not been submitted and approved and which is designated nonattainment under the previous SO2 standards or is not meeting the requirements of a SIP call under the previous SO2 standards (40 CFR 50.4(3)). A SIP call is an EPA action requiring a state to resubmit all or part of its State Implementation Plan to demonstrate attainment of the required NAAQS.

Air Quality by Location

You can get live air quality data by location at airnow.gov. Here is the link: https://www.airnow.gov

Paying the Price

Pollution has a cost. Manufacturing activities that cause air pollution impose health and clean-up costs on the whole of society, whereas the neighbors of an individual who chooses to fire-proof his home may benefit from a reduced risk of a fire spreading to their own homes. A manufacturing activity that causes air pollution is an example of a negative externality in production. A negative externality in production occurs “when a firm’s production reduces the well-being of others who are not compensated by the firm." For example, if a laundry firm exists near a polluting steel manufacturing firm, there will be increased costs for the laundry firm because of the dirt and smoke produced by the steel manufacturing firm. If external costs exist, such as those created by pollution, the manufacturer will choose to produce more of the product than would be produced if the manufacturer were required to pay all associated environmental costs. Because responsibility or consequence for self-directed action lies partly outside the self, an element of externalization is involved. If there are external benefits, such as in public safety, less of the good may be produced than would be the case if the producer were to receive payment for the external benefits to others. However, goods and services that involve negative externalities in production, such as those that produce pollution, tend to be over-produced and underpriced since the externality is not being priced into the market.

Pollution can also create costs for the firms producing the pollution. Sometimes firms choose, or are forced by regulation, to reduce the amount of pollution that they are producing. The associated costs of doing this are called abatement costs, or marginal abatement costs if measured by each additional unit. In 2005 pollution abatement capital expenditures and operating costs in the US amounted to nearly $27 billion.

Summary

- The EPA has set National Ambient Air Quality Standards for six principal pollutants, which are called "criteria" air pollutants.
- The six criteria pollutants are carbon monoxide, lead, nitrogen dioxide, ozone, particulate matter, and sulfur dioxide.
- The key air pollutants can cause various environmental and health problems that could affect the respiratory, nervous, and cardiovascular systems.

Contributors and Attributions

- US EPA
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- Marisa Alviar-Agnew (Sacramento City College)

7.6.4: 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 7.6.4.1). 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.

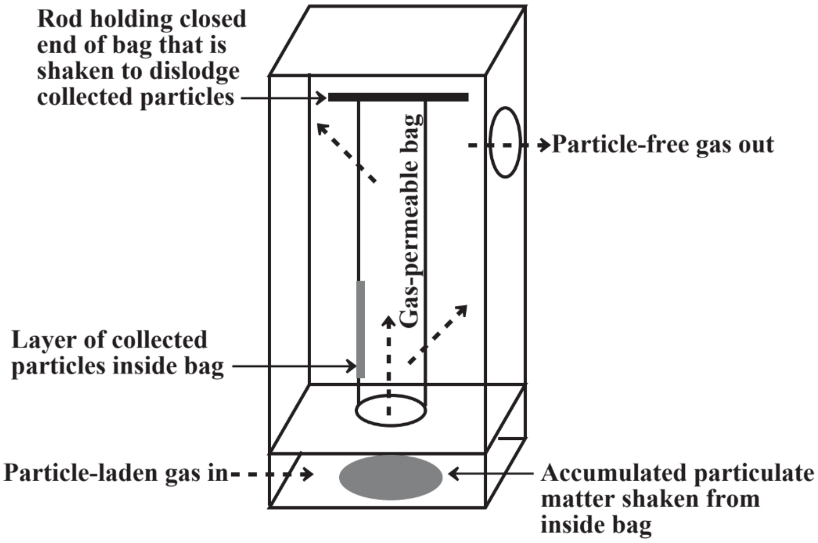

Figure 7.6.4.1. Illustration of a unit of a single bag in a baghouse for removal of particulate matter from a gas stream. The flow of gas is illustrated by the dashed lines. A bag is mounted on an opening in a plate such that particle-laden gas is forced up through the inside of the cylindrical bag and 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

$$\mathrm{CaSO_3} + \frac{1}{2}\,\mathrm{O_2} + 2\,\mathrm{H_2O} \rightarrow \mathrm{CaSO_4 \cdot 2H_2O}$$

to generate gypsum, $\mathrm{CaSO_4 \cdot 2H_2O}$. 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 $7.6.4.2$). 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.

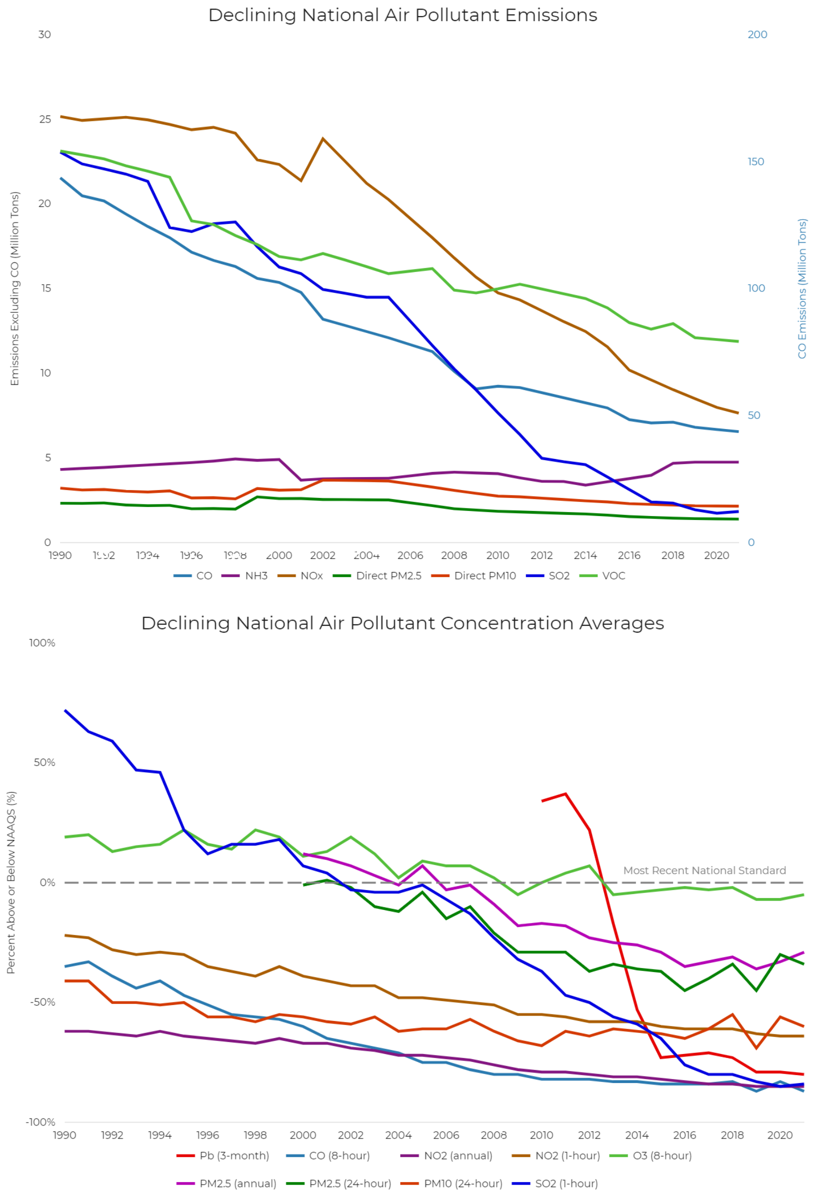

Figure 7.6.4.2: Declining air pollutant emissions and the resultant air pollutant concentrations since the passage of the 1990 Amendments to the Clean Air Act of 1970 and the 1990 Pollution Prevention Act. NAAQS are National Ambient Air Quality Standards. (US Environmental Protection Agency, www.epa.gov)

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 7.6.4.3). Studies have estimated that the benefits of reducing pollution exceed the costs by a factor of more than 30-to-1.

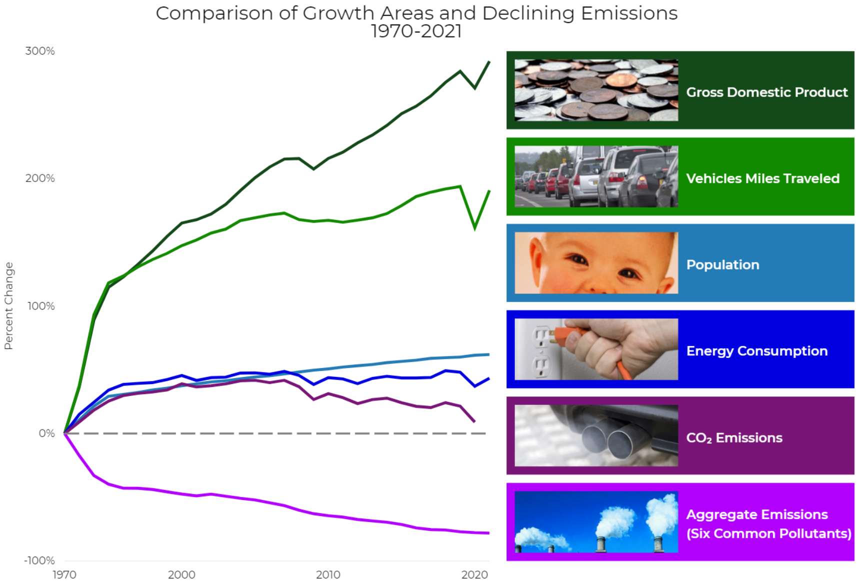

Figure 7.6.4.3: Some of the benefits that have increased as a result of pollution reduction since the implementation of Clean Air Acts. (US Environmental Protection Agency, www.epa.gov)

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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CHAPTER OVERVIEW

8: Atomic Mass, Reaction Stoichiometry

8.1: Isotopes and Atomic Weights

Learning Objectives

- Explain how isotopes differ from one another.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Isotopes

As introduced previously, atoms of a specific element are distinguished from other elements by their atomic number, (the number of protons). Atoms of the same element always have the same number of protons, however, the number of neutrons can vary. **Isotopes** are atoms of the same element that contain *different* numbers of *neutrons*. This difference in neutron amount affects the mass number (A) but not the atomic number (Z). In a chemical laboratory, isotopes of an element appear and react the same. For this reason, it is difficult to distinguish between different isotopes. In contrast, nuclear scientists can identify and separate different types of atomic nuclei. The technology required for this process is more sophisticated that what could be found in a typical chemical laboratory.

Figure 8.1.1 shows an easy way to represent isotopes with a **nuclear symbol**, which includes the atomic or element symbol (represented by X), the mass number, A, and the atomic number, Z. Thus, for the isotope of carbon that has 6 protons and 6 neutrons, the symbol is:

$$^{12}_{6}\mathrm{C}$$

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

It is also common to state the mass number after the name of an element to indicate a particular isotope. *Carbon-12* represents an isotope of carbon with 6 protons and 6 neutrons, while *uranium-238* is an isotope of uranium that has 92 protons and 146 neutrons.

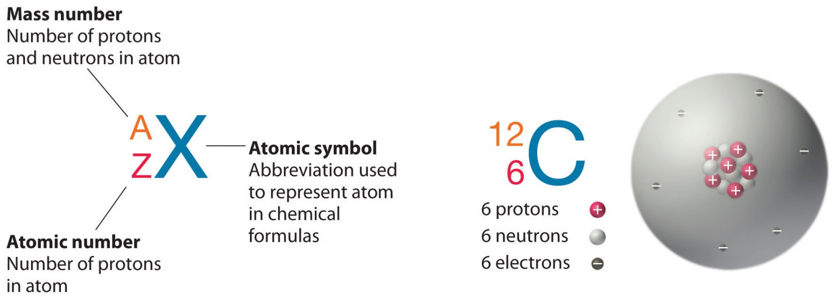


Figure 8.1.1: Nuclear Symbol. Unlike protons, the number of neutrons is not absolutely fixed for most elements. Atoms that have the same number of protons, and hence the same atomic number, but different numbers of neutrons are called isotopes. All isotopes of an element have the same number of protons and electrons, which means they exhibit the same chemistry. The isotopes of an element differ only in their atomic mass, which is given by the mass number (A), the sum of the numbers of protons and neutrons. (CC BY-NC-SA 4.0; anonymous by request)

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

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium*. An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. Figure 8.1.2 compares the three isotopes of hydrogen.

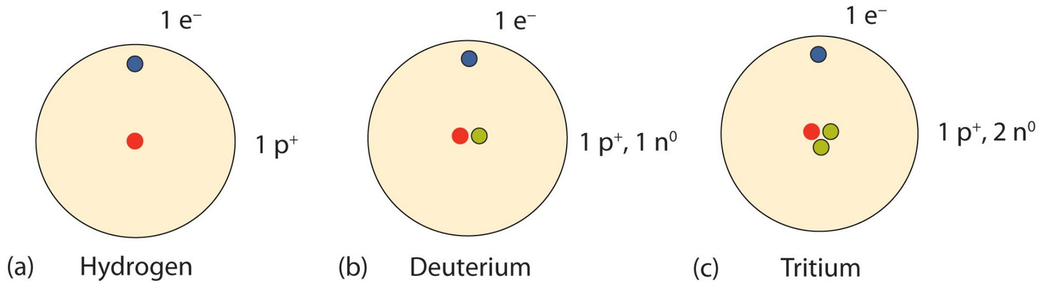


Figure 8.1.2: Isotopes of Hydrogen. Most hydrogen atoms have only a proton in the nucleus (a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

There are currently over 3,500 isotopes known for all the elements. When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. A/Z and symbol-mass formats can be used to display periodic table information. When viewing either of these two notations, isotopic differences can be obtained.

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

Look at the **A/Z formats** for the three isotopes of hydrogen in Table 8.1.1. Note how the atomic number (bottom value) remains the same while the atomic masses (top number) are varied. All isotopes of a particular element will vary in neutrons and mass. This variance in mass will be visible in the **symbol-mass format** of same isotopes as well.

Table 8.1.1

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

Both A/Z or symbol-mass formats can be utilized to determine the amount of subatomic particles (protons, neutrons, and electrons) contained inside an isotope. When given either format, these mass values should be used to calculate the number of neutrons in the nucleus.

Atomic Weight

Since most naturally occurring elements samples are mixtures of isotopes, it is useful to use an average weight of an element. The **atomic mass** of an element is the weighted mass of all the naturally presented isotopes (on earth). To determine the most abundant isotopic form of an element, compare given isotopes to the weighted average on the periodic table. For example, the three hydrogen isotopes in Figure 8.1.2 are H-1, H-2, and H-3. The atomic mass or weighted average of hydrogen is around 1.008 amu (look again to the periodic table). Of the three hydrogen isotopes, H-1 is closest in mass to the weighted average; therefore, it is the most abundant. The other two isotopes of hydrogen are quite rare, but are very exciting in the world of nuclear science.

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

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \cdots \tag{8.1.1}$$

Averages like Equation 1 are known as *weighted averages*. An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all the abundances are not provided, it is safe to assume that all numbers should add up to 100%.

For example, Boron has two naturally occurring isotopes. In a sample of boron, 20% of the atoms are B-10, which is an isotope of boron with 5 neutrons and mass of 10 amu. The other 80% of the atoms are B-11, which is an isotope of boron with 6 neutrons and a mass of 11 amu. How do we calculate the atomic mass of boron?

Boron has two isotopes so we will use the Equation 8.1.1 and substitute the relative abundances and atomic masses of Boron into Equation 8.1.1:

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

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

✓ Example 8.1.1: Atomic Weight of Neon

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

Solution

Neon has three isotopes. We will use the equation:

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

Substitute these into the equation, and we get:

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

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

? Exercise 8.1.1

Magnesium has the three isotopes listed in the following table:

Table showing the 3 isotopes of magnesium, the exact mass of each, and the percent abundance of each.

Isotope	Exact Mass (amu)	Percent Abundance (%)
^{24}Mg	23.98504	78.70
^{25}Mg	24.98584	10.13
^{26}Mg	25.98259	11.17

Use these data to calculate the atomic mass of magnesium.

Answer

24.31 amu

Applications of Isotopes

During the Manhattan project, the majority of federal funding was dedicated the separation of uranium isotopes. The two most common isotopes of uranium are U-238 and U-235. About 99.3% of uranium is of the U-238 variety, this form is not fissionable and will not work in a nuclear weapon or reaction. The remaining 0.7% is U-235 which is fissionable, but first had to be separated from U-238. This separation process is called *enrichment*. During World War II, a nuclear facility was built in Oak Ridge, Tennessee to accomplish this project. At the time, the enrichment process only produced enough U-235 for one nuclear weapon. This fuel was placed inside the smaller of the two atomic bombs (Little Boy) dropped over Japan.

Figure 8.1.3: A billet of highly enriched uranium that was recovered from scrap processed at the Y-12 National Security Complex Plant. Original and unrotated.

A billet of highly enriched uranium that was recovered from scrap processed at the Y-12 National Security Complex Plant.

Uranium is a natural element that can be found in several different countries. Countries that do not have natural uranium supplies would need to obtain it from one of the countries below. Most nuclear reactors that provide energy rely on U-235 as a source of fuel. Fortunately, reactors only need 2-5% U-235 for the production of megawatts or even gigawatts of power. If the purification process exceeds this level, than it is likely a country is focusing on making nuclear weapons. For example, Manhattan Project scientists enriched U-235 up to 90% in order to produce the Little Boy weapon.

Abbreviations like HEU (highly enriched uranium) and LEU (low-enriched uranium) are used frequently by nuclear scientists and groups. HEU is defined as being over 20% pure U-235 and would not be used in most commercial nuclear reactors. This type of material is used to fuel larger submarines and aircraft carriers. If the purification of U-235 reaches 90%, then the HEU is further classified as being weapons grade material. This type of U-235 could be used to make a nuclear weapon (fission or even fusion based). As for LEU, its U-235 level would be below this 20% mark. LEU would be used for commercial nuclear reactors and smaller, nuclear powered submarines. LEU is not pure enough to be used in a conventional nuclear weapon, but could be used in a dirty bomb. This type of weapon uses conventional explosives like dynamite to spread nuclear material. Unlike a nuclear weapon, dirty bombs are not powerful enough to affect large groups of buildings or people. Unfortunately, the spread of nuclear material would cause massive chaos for a community and would result in casualties.

Summary

- The isotopes of an element have different masses and are identified by their mass numbers.
- An element's atomic mass is the weighted average of the masses of the isotopes of an element
- An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all the abundances are not provided, it is safe to assume that all numbers should add up to 100%.

Concept Review Exercises

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

Answers

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

Contributors and Attributions

-
- Elizabeth R. Gordon (Furman University)

8.2: The Mole

Learning Objectives

- Describe the unit *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 8.2.1: 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 ($\$6.022 \times 10^{23}$) 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 $6.02214076 \times 10^{23}$ 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 ($^{12}_{6}C$). 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.

8.3: 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.

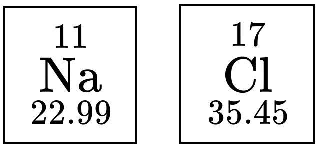
Figure 8.3.1: How much sodium is in this spoonful of salt? (via Flickr; Gianni pepenera).

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 8.3.1

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 8.3.1

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.

8.4: 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 8.4.1). However, atoms and molecules may be counted indirectly by using a common trick of "counting by weighing."

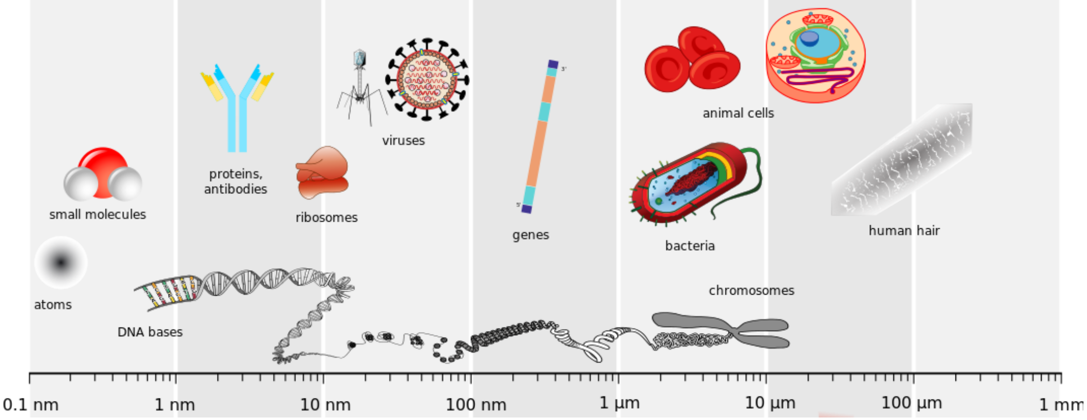


Figure 8.4.1: A comparison of scale. (CC BY-SA 3.0; Wikipedia)

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

$$\text{mass of nails} = 1340\,\text{g} - 113\,\text{g} = 1227\,\text{g}$$

Therefore

$$\text{nails in box} = \frac{1227\,\text{g}}{0.450\ \text{g/nail}} = 2{,}726.7\,\text{nails} = \boxed{2{,}730\,\text{nails}}$$

You have just counted the number of nails in the box by weighing them (rather than by counting them individually).

? ✏️ Exercise 8.4.1

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

8.5: 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**.

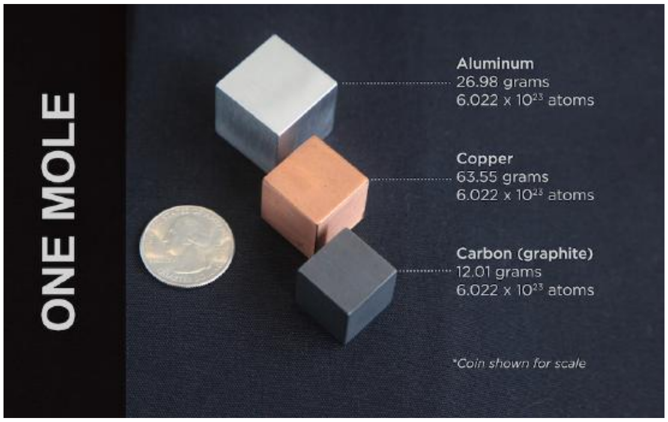


Figure 8.5.1: One mole of aluminum, copper, and carbon. (R. Press/N. Hanacek/NIST)

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 8.5.2: 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 8.5.1) is essential to understanding the mole concept and correctly performing mole calculations.

Table 8.5.1: 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 8.5.1

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 8.5.2

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	$Ca(NO_3)_2$

	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 8.5.1

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.

8.6: Grams-->Moles-->Moles-->Grams

Learning Objectives

- To convert between mass units and mole units.

As we just discussed, **molar mass** is defined as the mass (in grams) of 1 mole of substance (or Avogadro's number of molecules or formula units).

The simplest type of manipulation using molar mass as a *conversion factor* is a mole-gram conversion (or its reverse, a gram-mole conversion).

We also established that 1 mol of Al has a mass of 26.98 g (Example 8.6.1). Stated mathematically,

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

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

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

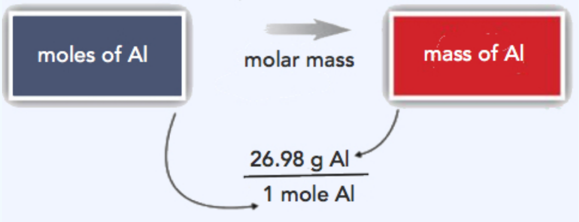
The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do "by eye."

Example 8.6.1

What is the mass of 3.987 mol of Al?

Solution

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



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

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

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

Our final answer is expressed to four significant figures.

Exercise 8.6.1

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

Answer

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

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 8.6.1 is a chart for determining what conversion factor is needed, and Figure 8.6.2 is a flow diagram for the steps needed to perform a conversion.

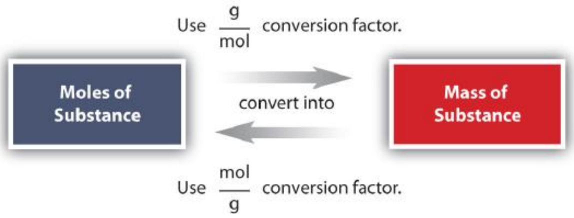


Figure 8.6.1 A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles.

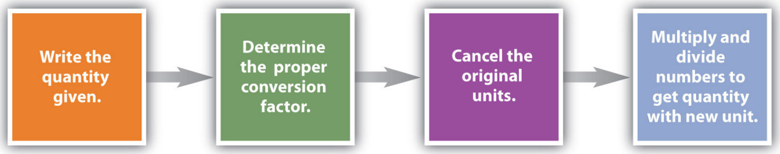


Figure 8.6.2 A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

✓ Example 8.6.2

A biochemist needs 0.00655 mol of bilirubin ($C_{33}H_{36}N_4O_6$) for an experiment. How many grams of bilirubin will that be?

Solution

To convert from moles to mass, we need the molar mass of bilirubin, which we can determine from its chemical formula:

33 C molar mass:	33 × 12.01 g =	396.33 g
36 H molar mass:	36 × 1.01 g =	36.36 g
4 N molar mass:	4 × 14.01 g =	56.04 g
6 O molar mass:	6 × 16.00 g =	96.00 g
Total:		584.73 g

The molar mass of bilirubin is 584.73 g. Using the relationship

$$1 \text{ mol bilirubin} = 584.73 \text{ g bilirubin}$$

we can construct the appropriate conversion factor for determining how many grams there are in 0.00655 mol. Following the steps from Figure 8.6.2:

$$0.00655 \text{ mol bilirubin} \times \frac{584.73 \text{ g bilirubin}}{\text{mol bilirubin}} = 3.83 \text{ g bilirubin}$$

The mol bilirubin unit cancels. The biochemist needs 3.83 g of bilirubin.

? Exercise 8.6.2

A chemist needs 457.8 g of $KMnO_4$ to make a solution. How many moles of $KMnO_4$ is that?

Answer

$$457.8 \text{ g KMnO}_4 \times \frac{1 \text{ mol KMnO}_4}{158.04 \text{ g KMnO}_4} = 2.897 \text{ mol KMnO}_4$$

To Your Health: Minerals

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

Table 8.6.1: Essential Minerals and their Composition in Humans

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

Table 8.6.1 illustrates several things. First, the needs of men and women for some minerals are different. The extreme case is for iron; women need over twice as much as men do. In all other cases where there is a different RDI, men need more than women.

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

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

CHAPTER OVERVIEW

9: Molecular Polarity, Intermolecular Forces and Solubility

9.1: Electronegativity and Polarity

Learning Objectives

- Explain how polar compounds differ from nonpolar compounds.
- Determine if a molecule is polar or nonpolar.
- Given a pair of compounds, predict which would have a higher melting or boiling point.

Bond Polarity

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

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure 9.1.1.

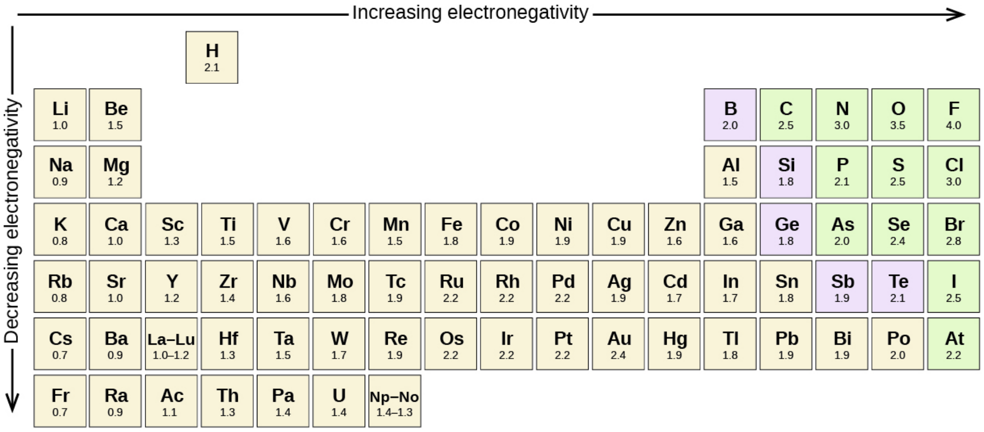


Figure 9.1.1: Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

difference of the electronegativities of the two atoms involved in the covalent bond

Electronegativity Difference	Bond Type
0–0.4	pure covalent
0.5–2.0	polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

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

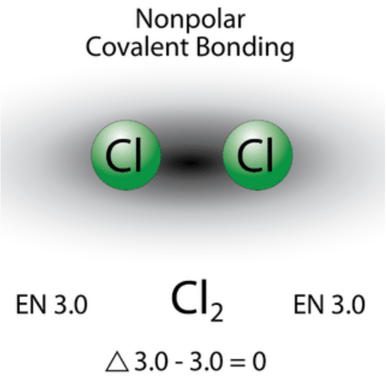


Figure 9.1.2: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

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

Polar Covalent Bonds

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

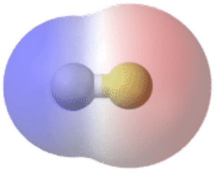


Figure 9.1.3: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

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

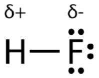


Figure 9.1.4: Use of δ to indicate partial charge.

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

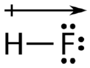


Figure 9.1.5: Use of crossed arrow to indicate polarity.

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

✓ Example 9.1.1: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure 9.1.1, we can calculate the difference of the electronegativities of the atoms involved in the bond.

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

? Exercise 9.1.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a

likely ionic

Answer b

polar covalent

Molecular Polarity

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

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.

Figure 9.1.6: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide $(\mathrm{CO_2})$ is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of $\mathrm{CO_2}$ is zero.

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

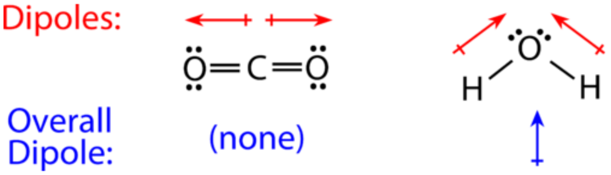


Figure 9.1.7: The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as $\mathrm{CH_4}$ is nonpolar. However, if one of the peripheral H atoms is replaced with another atom that has a different electronegativity, the molecule becomes polar. A trigonal planar molecule $(\mathrm{BF_3})$ may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule $(\mathrm{NH_3})$ is polar.

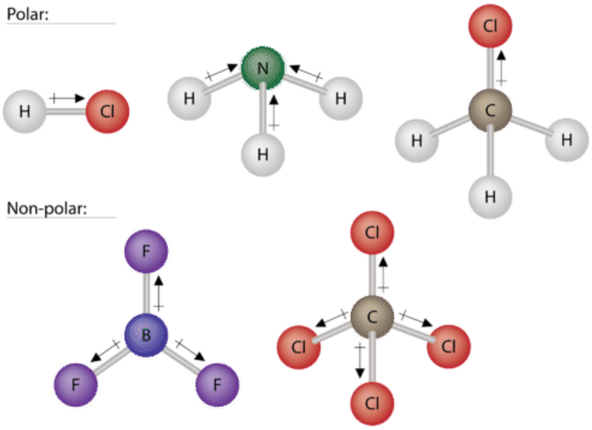


Figure 9.1.8: Some examples of polar and nonpolar molecules based on molecular geometry.

To summarize, to be polar, a molecule must:

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

Steps to Identify Polar Molecules

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

Properties of Polar Molecules

Polar molecules tend to align when placed in an electric field with the positive end of the molecule oriented toward the negative plate and the negative end toward the positive plate (Figure $9.1.14$). We can use an electrically charged object to attract polar molecules, but nonpolar molecules are not attracted. Also, polar solvents are better at dissolving polar substances, and nonpolar solvents are better at dissolving nonpolar substances.

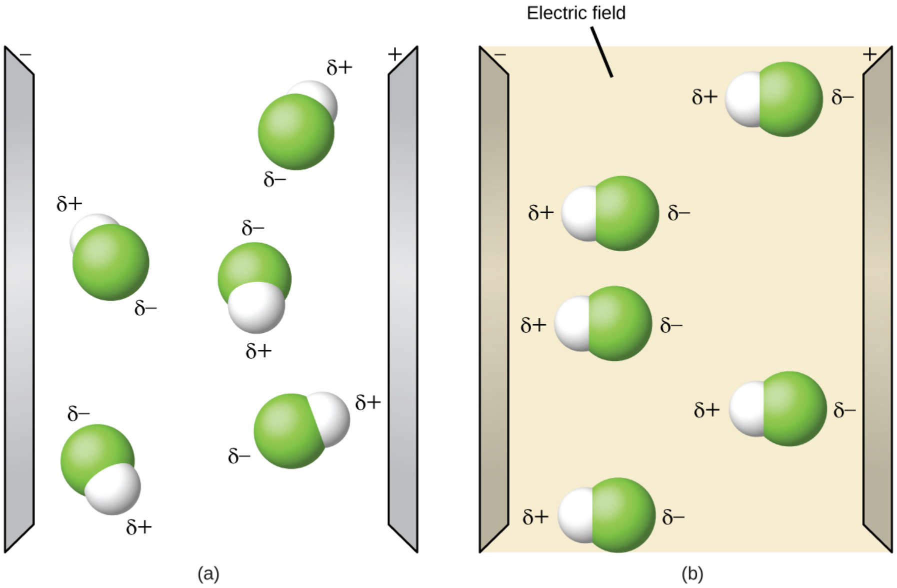


Figure $9.1.9$: (a) Molecules are always randomly distributed in the liquid state in the absence of an electric field. (b) When an electric field is applied, polar molecules like HF will align to the dipoles with the field direction. (OpenStax CC-BY-SA);

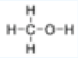
While molecules can be described as "polar covalent" or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

✓ Example $9.1.2$:

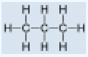
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: 

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

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

? Exercise $9.1.2$

Label each of the following as polar or nonpolar.

a. SO_3
b. NH_3

Answer a

nonpolar

Answer b

polar

Contributions & Attributions

- StackExchange (thomij).

- **10.8: Electronegativity and Polarity - Why Oil and Water Do not Mix** by Henry Agnew, Marisa Alviar-Agnew is licensed CK-12. Original source: https://www.ck12.org/c/chemistry/.

9.2: Intermolecular Forces

Learning Objectives

- Describe the types of intermolecular forces possible between atoms or molecules in condensed phases (dispersion forces, dipole-dipole attractions, and hydrogen bonding)
- Identify the types of intermolecular forces experienced by specific molecules based on their structures
- Explain the relation between the intermolecular forces present within a substance and the temperatures associated with changes in its physical state

As was the case for gaseous substances, the kinetic molecular theory may be used to explain the behavior of solids and liquids. In the following description, the term *particle* will be used to refer to an atom, molecule, or ion. Note that we will use the popular phrase "intermolecular attraction" to refer to attractive forces between the particles of a substance, regardless of whether these particles are molecules, atoms, or ions.

Consider these two aspects of the molecular-level environments in solid, liquid, and gaseous matter:

- Particles in a solid are tightly packed together and often arranged in a regular pattern; in a liquid, they are close together with no regular arrangement; in a gas, they are far apart with no regular arrangement.
- Particles in a solid vibrate about fixed positions and do not generally move in relation to one another; in a liquid, they move past each other but remain in essentially constant contact; in a gas, they move independently of one another except when they collide.

The differences in the properties of a solid, liquid, or gas reflect the strengths of the attractive forces between the atoms, molecules, or ions that make up each phase. The phase in which a substance exists depends on the relative extents of its intermolecular forces (IMFs) and the kinetic energies (KE) of its molecules. IMFs are the various forces of attraction that may exist between the atoms and molecules of a substance due to electrostatic phenomena, as will be detailed in this module. These forces serve to hold particles close together, whereas the particles' KE provides the energy required to overcome the attractive forces and thus increase the distance between particles. Figure 9.2.1 illustrates how changes in physical state may be induced by changing the temperature, hence, the average KE, of a given substance.

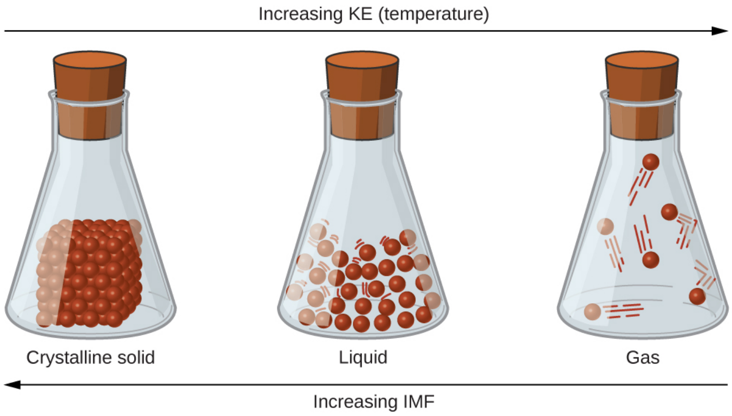


Figure 9.2.1: Transitions between solid, liquid, and gaseous states of a substance occur when conditions of temperature or pressure favor the associated changes in intermolecular forces. (Note: The space between particles in the gas phase is much greater than shown.)

Three sealed flasks are labeled, "Crystalline solid," "Liquid," and "Gas," from left to right. The first flask holds a cube composed of small spheres sitting on the bottom while the second flask shows a lot of small spheres in the bottom that are spaced a small distance apart from one another and have lines around them to indicate motion. The third flask shows a few spheres spread far from one another with larger lines to indicate motion. There is a right-facing arrow that spans the top of all three flasks. The arrow is labeled, "Increasing K E (temperature)." There is a left-facing arrow that spans the bottom of all three flasks. The arrow is labeled, "Increasing I M F."

As an example of the processes depicted in this figure, consider a sample of water. When gaseous water is cooled sufficiently, the attractions between H_2O molecules will be capable of holding them together when they come into contact with each other; the gas condenses, forming liquid H_2O. For example, liquid water forms on the outside of a cold glass as the water vapor in the air is cooled by the cold glass, as seen in Figure 9.2.2.

Figure 9.2.2: Condensation forms when water vapor in the air is cooled enough to form liquid water, such as (a) on the outside of a cold beverage glass or (b) in the form of fog. (credit a: modification of work by Jenny Downing; credit b: modification of work by Cory Zanker)

Image a shows a brown colored beverage in a glass with condensation on the outside. Image b shows a body of water with fog hovering above the surface of the water.

We can also liquefy many gases by compressing them, if the temperature is not too high. The increased pressure brings the molecules of a gas closer together, such that the attractions between the molecules become strong relative to their KE. Consequently, they form liquids. Butane, C_4H_{10}, is the fuel used in disposable lighters and is a gas at standard temperature and pressure. Inside the lighter's fuel compartment, the butane is compressed to a pressure that results in its condensation to the liquid state, as shown in Figure 9.2.3.

Figure 9.2.3: Gaseous butane is compressed within the storage compartment of a disposable lighter, resulting in its condensation to the liquid state. (credit: modification of work by "Sam-Cat"/Flickr)

Finally, if the temperature of a liquid becomes sufficiently low, or the pressure on the liquid becomes sufficiently high, the molecules of the liquid no longer have enough KE to overcome the IMF between them, and a solid forms. A more thorough discussion of these and other changes of state, or phase transitions, is provided in a later module of this chapter.

Forces between Molecules

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not *intra*molecular forces. *Intra*molecular forces are those *within* the molecule that keep the molecule together, for example, the bonds between the atoms. *Inter*molecular forces are the attractions *between* molecules, which determine many of the physical properties of a substance. Figure 9.2.4 illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

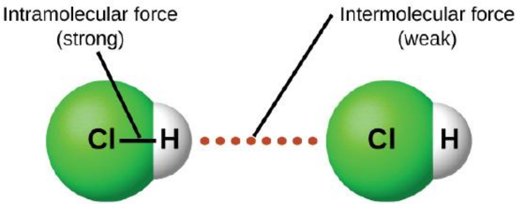


Figure 9.2.4: Intramolecular forces keep a molecule intact. Intermolecular forces hold multiple molecules together and determine many of a substance's properties.

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

All of the attractive forces between neutral atoms and molecules are known as van der Waals forces, although they are usually referred to more informally as intermolecular attraction. We will consider the various types of IMFs in the next three sections of this module.

Dispersion Forces

One of the three van der Waals forces is present in all condensed phases, regardless of the nature of the atoms or molecules composing the substance. This attractive force is called the London dispersion force in honor of German-born American physicist Fritz London who, in 1928, first explained it. This force is often referred to as simply the dispersion force. Because the electrons of an atom or molecule are in constant motion (or, alternatively, the electron's location is subject to quantum-mechanical variability), at any moment in time, an atom or molecule can develop a temporary, instantaneous dipole if its electrons are distributed asymmetrically. The presence of this dipole can, in turn, distort the electrons of a neighboring atom or molecule, producing an induced dipole. These two rapidly fluctuating, temporary dipoles thus result in a relatively weak electrostatic attraction between the species—a so-called dispersion force like that illustrated in Figure 9.2.5.

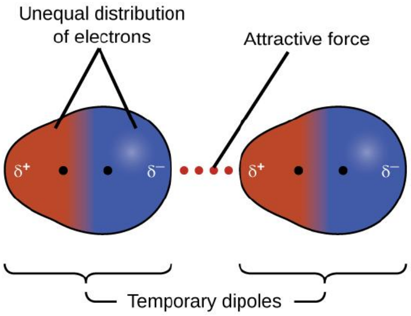


Figure 9.2.5: Dispersion forces result from the formation of temporary dipoles, as illustrated here for two nonpolar diatomic molecules.

Two pairs of molecules are shown where each molecule has one larger blue side labeled "delta sign, negative sign" and a smaller red side labeled "delta sign, positive sign." Toward the middle of the both molecules, but still on each distinct side, is a black dot. Between the two images is a dotted line labeled, "Attractive force." In the first image, the red and blue sides are labeled, "Unequal distribution of electrons." Below both images are brackets. The brackets are labeled, "Temporary dipoles."

Dispersion forces that develop between atoms in different molecules can attract the two molecules to each other. The forces are relatively weak, however, and become significant only when the molecules are very close. Larger and heavier atoms and molecules exhibit stronger dispersion forces than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in Table 9.2.1.

Table 9.2.1: Melting and Boiling Points of the Halogens

Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K

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

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

✓ Example 9.2.1: London Forces and Their Effects

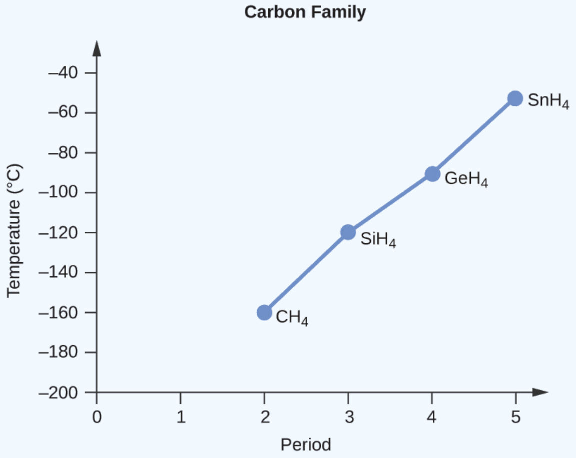
Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

Solution

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

$$CH_4 < SiH_4 < GeH_4 < SnH_4$$

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



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

Exercise 9.2.1

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

Answer

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

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

The shapes of molecules also affect the magnitudes of the dispersion forces between them. For example, boiling points for the isomers *n*-pentane, isopentane, and neopentane (shown in Figure 9.2.6) are 36 °C, 27 °C, and 9.5 °C, respectively. Even though these compounds are composed of molecules with the same chemical formula, C_5H_{12}, the difference in boiling points suggests that dispersion forces in the liquid phase are different, being greatest for *n*-pentane and least for neopentane. The elongated shape of *n*-pentane provides a greater surface area available for contact between molecules, resulting in correspondingly stronger dispersion forces. The more compact shape of isopentane offers a smaller surface area available for intermolecular contact and, therefore, weaker dispersion forces. Neopentane molecules are the most compact of the three, offering the least available surface area for intermolecular contact and, hence, the weakest dispersion forces. This behavior is analogous to the connections that may be formed between strips of VELCRO brand fasteners: the greater the area of the strip's contact, the stronger the connection.

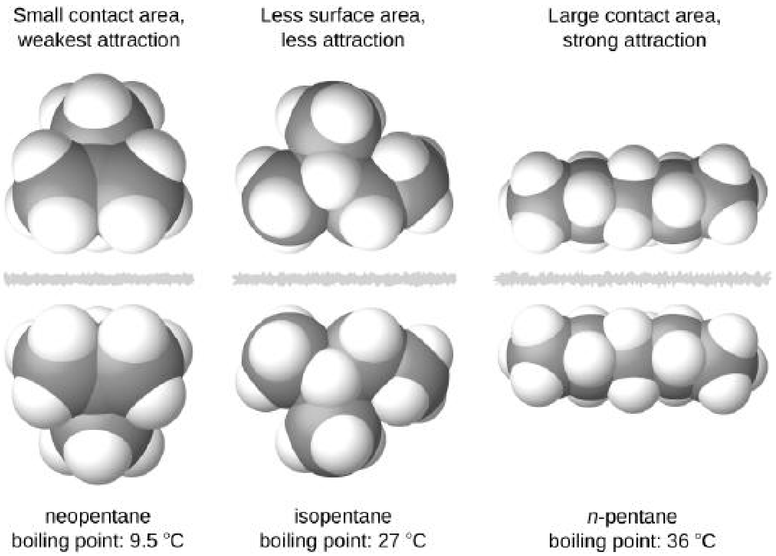


Figure 9.2.6: The strength of the dispersion forces increases with the contact area between molecules, as demonstrated by the boiling points of these pentane isomers.

Three images of molecules are shown. The first shows a cluster of large, gray spheres each bonded together and to several smaller, white spheres. There is a gray, jagged line and then the mirror image of the first cluster of spheres is shown. Above these two clusters is the label, "Small contact area, weakest attraction," and below is the label, "neopentane boiling point: 9.5 degrees C." The second shows a chain of three gray spheres bonded by the middle sphere to a fourth gray sphere. Each gray sphere is bonded to several smaller, white spheres. There is a jagged, gray line and then the mirror image of the first chain appears. Above these two chains is the label, "Less surface area, less attraction," and below is the label, "isopentane boiling point: 27 degrees C." The third image shows a chain of five gray spheres bonded together and to several smaller, white spheres. There is a jagged gray line and then the mirror image of the first chain appears. Above these chains is the label, "Large contact area, strong attraction," and below is the label, "n-pentane boiling point 36 degrees C."

Applications: Geckos and Intermolecular Forces

Geckos have an amazing ability to adhere to most surfaces. They can quickly run up smooth walls and across ceilings that have no toe-holds, and they do this without having suction cups or a sticky substance on their toes. And while a gecko can lift its feet easily as it walks along a surface, if you attempt to pick it up, it sticks to the surface. How are geckos (as well as spiders and some other insects) able to do this? Although this phenomenon has been investigated for hundreds of years, scientists only recently uncovered the details of the process that allows geckos' feet to behave this way.

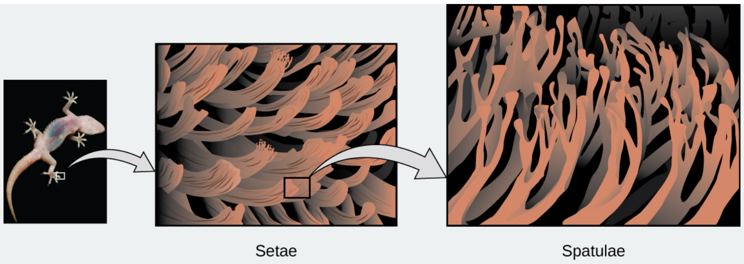


Figure 9.2.7: Geckos' toes contain large numbers of tiny hairs (setae), which branch into many triangular tips (spatulae). Geckos adhere to surfaces because of van der Waals attractions between the surface and a gecko's millions of spatulae. By changing how the spatulae contact the surface, geckos can turn their stickiness "on" and "off." (credit photo: modification of work by "JC*+A!"/Flickr)

Three figures are shown. The first is a photo of the bottom of a gecko's foot. The second is bigger version which shows the setae. The third is a bigger version of the setae and shows the spatulae.

Geckos' toes are covered with hundreds of thousands of tiny hairs known as *setae*, with each seta, in turn, branching into hundreds of tiny, flat, triangular tips called *spatulae*. The huge numbers of spatulae on its setae provide a gecko, shown in Figure 9.2.7, with a large total surface area for sticking to a surface. In 2000, Kellar Autumn, who leads a multi-institutional gecko research team, found that geckos adhered equally well to both polar silicon dioxide and nonpolar gallium arsenide. This proved that geckos stick to surfaces because of dispersion forces—weak intermolecular attractions arising from temporary, synchronized charge distributions between adjacent molecules. Although dispersion forces are very weak, the total attraction over millions of spatulae is large enough to support many times the gecko's weight.

In 2014, two scientists developed a model to explain how geckos can rapidly transition from "sticky" to "non-sticky." Alex Greaney and Congcong Hu at Oregon State University described how geckos can achieve this by changing the angle between their spatulae and the surface. Geckos' feet, which are normally nonsticky, become sticky when a small shear force is applied. By curling and uncurling their toes, geckos can alternate between sticking and unsticking from a surface, and thus easily move across it. Further investigations may eventually lead to the development of better adhesives and other applications.

Watch this video to learn more about Kellar Autumn's research that determined that van der Waals forces are responsible for a gecko's ability to cling and climb.

Dipole-Dipole Attractions

Recall from the chapter on chemical bonding and molecular geometry that *polar* molecules have a partial positive charge on one side and a partial negative charge on the other side of the molecule—a separation of charge called a *dipole*. Consider a polar molecule such as hydrogen chloride, HCl. In the HCl molecule, the more electronegative Cl atom bears the partial negative charge, whereas the less electronegative H atom bears the partial positive charge. An attractive force between HCl molecules results from the attraction between the positive end of one HCl molecule and the negative end of another. This attractive force is called a dipole-dipole attraction—the electrostatic force between the partially positive end of one polar molecule and the partially negative end of another, as illustrated in Figure 9.2.8.

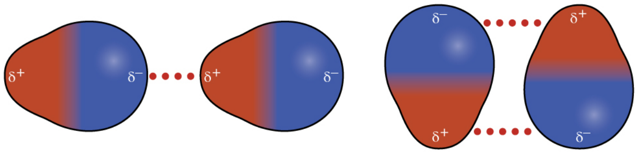


Figure 9.2.8: This image shows two arrangements of polar molecules, such as HCl, that allow an attraction between the partial negative end of one molecule and the partial positive end of another.

Two pairs of molecules are shown where each molecule has one larger blue side labeled “delta sign, negative sign” and a smaller red side labeled “delta sign, positive sign. In the first pair, the red sides of the two molecules both face to the left and the blue side to the right. A horizontal dotted line lies in between the two. In the second pair, the molecules face up and down, with the red and blue ends aligning. A horizontal dotted line lies between the red and blue ends facing upward and another lies between the red and blue ends facing downward.

The effect of a dipole-dipole attraction is apparent when we compare the properties of HCl molecules to nonpolar F_2 molecules. Both HCl and F_2 consist of the same number of atoms and have approximately the same molecular mass. At a temperature of 150 K, molecules of both substances would have the same average KE. However, the dipole-dipole attractions between HCl molecules are sufficient to cause them to “stick together” to form a liquid, whereas the relatively weaker dispersion forces between nonpolar F_2 molecules are not, and so this substance is gaseous at this temperature. The higher normal boiling point of HCl (188 K) compared to F_2 (85 K) is a reflection of the greater strength of dipole-dipole attractions between HCl molecules, compared to the attractions between nonpolar F_2 molecules. We will often use values such as boiling or freezing points, or enthalpies of vaporization or fusion, as indicators of the relative strengths of IMFs of attraction present within different substances.

✓ Example 9.2.2: Dipole-Dipole Forces and Their Effects

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

Solution

CO and N_2 are both diatomic molecules with masses of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N_2 is nonpolar, its molecules cannot exhibit dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N_2 molecules, so CO is expected to have the higher boiling point.

A common method for preparing oxygen is the decomposition

? Exercise 9.2.2

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

Answer

ICl. ICl and Br_2 have similar masses (~160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br_2 is nonpolar and does not. The relatively stronger dipole-dipole attractions require more energy to overcome, so ICl will have the higher boiling point.

Hydrogen Bonding

Nitrosyl fluoride (ONF, molecular mass 49 amu) is a gas at room temperature. Water (H_2O, molecular mass 18 amu) is a liquid, even though it has a lower molecular mass. We clearly cannot attribute this difference between the two compounds to dispersion forces. Both molecules have about the same shape and ONF is the heavier and larger molecule. It is, therefore, expected to experience more significant dispersion forces. Additionally, we cannot attribute this difference in boiling points to differences in the dipole moments of the molecules. Both molecules are polar and exhibit comparable dipole moments. The large difference between the boiling points is due to a particularly strong dipole-dipole attraction that may occur when a molecule contains a hydrogen atom bonded to a fluorine, oxygen, or nitrogen atom (the three most electronegative elements). The very large difference in electronegativity between the H atom (2.1) and the atom to which it is bonded (4.0 for an F atom, 3.5 for an O atom, or 3.0 for a N atom), combined with the very small size of a H atom and the relatively small sizes of F, O, or N atoms, leads to *highly concentrated partial charges* with these atoms. Molecules with F-H, O-H, or N-H moieties are very strongly attracted to similar moieties in nearby molecules, a particularly strong type of dipole-dipole attraction called hydrogen bonding. Examples of hydrogen bonds include HF···HF, H_2O···HOH, and H_3N···HNH_2, in which the hydrogen bonds are denoted by dots. Figure 9.2.9 illustrates hydrogen bonding between water molecules.

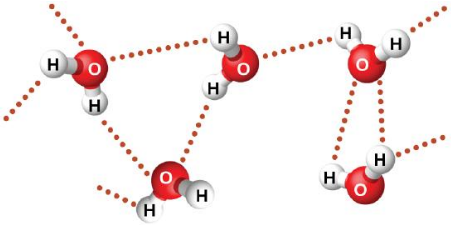


Figure 9.2.9: Water molecules participate in multiple hydrogen-bonding interactions with nearby water molecules.

Five water molecules are shown near one another, but not touching. A dotted line lies between many of the hydrogen atoms on one molecule and the oxygen atom on another molecule.

Despite use of the word "bond," keep in mind that hydrogen bonds are *intermolecular* attractive forces, not *intramolecular* attractive forces (covalent bonds). Hydrogen bonds are much weaker than covalent bonds, only about 5 to 10% as strong, but are generally much stronger than other dipole-dipole attractions and dispersion forces.

Hydrogen bonds have a pronounced effect on the properties of condensed phases (liquids and solids). For example, consider the trends in boiling points for the binary hydrides of group 15 (NH_3, PH_3, AsH_3, and SbH_3), group 16 hydrides (H_2O, H_2S, H_2Se, and H_2Te), and group 17 hydrides (HF, HCl, HBr, and HI). The boiling points of the heaviest three hydrides for each group are plotted in Figure 9.2.10. As we progress down any of these groups, the polarities of the molecules decrease slightly, whereas the sizes of the molecules increase substantially. The effect of increasingly stronger dispersion forces dominates that of increasingly weaker dipole-dipole attractions, and the boiling points are observed to increase steadily. For the group 15, 16, and 17 hydrides, the boiling points for each class of compounds increase with increasing molecular mass for elements in periods 3, 4, and 5.

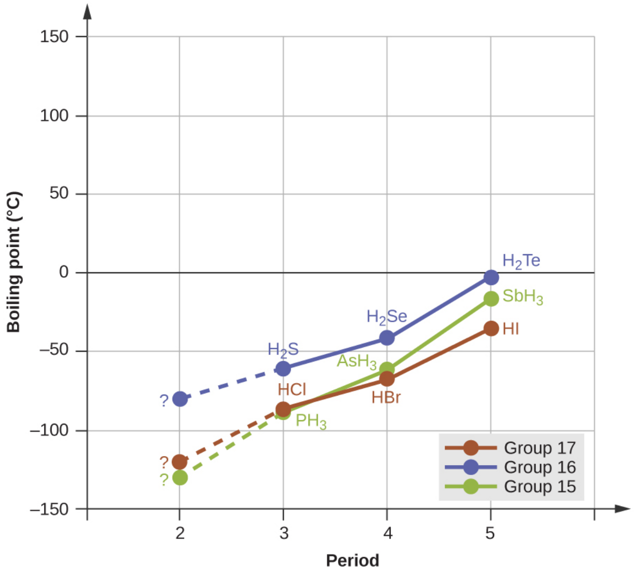


Figure 9.2.10: In comparison to periods 3−5, the binary hydrides of period 2 elements in groups 17, 16 and 15 (F, O and N, respectively) exhibit anomalously high boiling points due to hydrogen bonding.

A line graph is shown where the y-axis is labeled "Boiling point (, degree sign, C)" and has values of " negative 150" to "150" from bottom to top in increments of 50. The x-axis is labeled "Period" and has values of "0" to "5" in increments of 1. Three lines are shown on the graph and are labeled in the legend. The red line is labeled as "halogen family," the blue is "oxygen family" and the green is "nitrogen family." The first point on the red line is labeled "question mark" and is at point "2, negative 120". The second point on the line is labeled "H C l" and is at point "3, negative 80" while the third point on the line is labeled "H B r" and is at point "4, negative 60". The fourth point on the line is labeled "H I" and is at point "5, negative 40." The first point on the green line is labeled "question mark" and is at point "2, negative 125." The second point on the line is labeled "P H, subscript 3" and is at point "3, negative 80" while the third point on the line is labeled "A s H, subscript 3" and is at point "4, negative 55." The fourth point on the line is labeled "S b H, subscript 3" and is at point "5, negative 10." The first point on the blue line is labeled "question mark" and is at point "2, negative 80." The second point on the line is labeled "H, subscript 2, S" and is at point "3, negative 55" while the third point on the line is labeled "H, subscript 2, S e" and is at point "4, negative 45." The fourth point on the line is labeled "H, subscript 2, T e" and is at point "5, negative 3."

If we use this trend to predict the boiling points for the lightest hydride for each group, we would expect NH_3 to boil at about −120 °C, H_2O to boil at about −80 °C, and HF to boil at about −110 °C. However, when we measure the boiling points for these compounds, we find that they are dramatically higher than the trends would predict, as shown in Figure $9.2.10$. The stark contrast between our naïve predictions and reality provides compelling evidence for the strength of hydrogen bonding.

✓ Example $9.2.3$: Effect of Hydrogen Bonding on Boiling Points

Consider the compounds dimethylether (CH_3OCH_3), ethanol (CH_3CH_2OH), and propane ($CH_3CH_2CH_3$). Their boiling points, not necessarily in order, are −42.1 °C, −24.8 °C, and 78.4 °C. Match each compound with its boiling point. Explain your reasoning.

Solution

The VSEPR-predicted shapes of CH_3OCH_3, CH_3CH_2OH, and $CH_3CH_2CH_3$ are similar, as are their molar masses (46 g/mol, 46 g/mol, and 44 g/mol, respectively), so they will exhibit similar dispersion forces. Since $CH_3CH_2CH_3$ is nonpolar, it may exhibit *only* dispersion forces. Because CH_3OCH_3 is polar, it will also experience dipole-dipole attractions. Finally, CH_3CH_2OH has an −OH group, and so it will experience the uniquely strong dipole-dipole attraction known as hydrogen bonding. So the ordering in terms of strength of IMFs, and thus boiling points, is $CH_3CH_2CH_3 < CH_3OCH_3 < CH_3CH_2OH$. The boiling point of propane is −42.1 °C, the boiling point of dimethylether is −24.8 °C, and the boiling point of ethanol is 78.5 °C.

? Exercise $9.2.3$

Ethane (CH_3CH_3) has a melting point of −183 °C and a boiling point of −89 °C. Predict the melting and boiling points for methylamine (CH_3NH_2). Explain your reasoning.

Answer

> The melting point and boiling point for methylamine are predicted to be significantly greater than those of ethane. CH_3CH_3 and CH_3NH_2 are similar in size and mass, but methylamine possesses an −NH group and therefore may exhibit hydrogen bonding. This greatly increases its IMFs, and therefore its melting and boiling points. It is difficult to predict values, but the known values are a melting point of −93 °C and a boiling point of −6 °C.

Hydrogen Bonding and DNA

Deoxyribonucleic acid (DNA) is found in every living organism and contains the genetic information that determines the organism’s characteristics, provides the blueprint for making the proteins necessary for life, and serves as a template to pass this information on to the organism’s offspring. A DNA molecule consists of two (anti-)parallel chains of repeating nucleotides, which form its well-known double helical structure, as shown in Figure $9.2.10$.

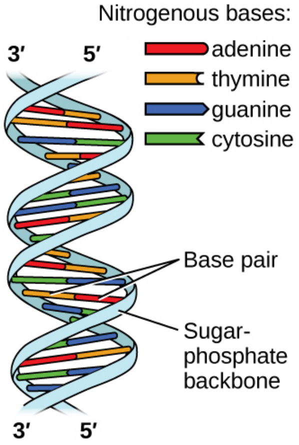


Figure 9.2.11: Two separate DNA molecules form a double-stranded helix in which the molecules are held together via hydrogen bonding. (credit: modification of work by Jerome Walker, Dennis Myts)

Two images are shown. The first lies on the left side of the page and shows a helical structure like a twisted ladder where the rungs of the ladder, labeled “Base pair” are red, yellow, green and blue paired bars. The red and yellow bars, which are always paired together, are labeled in the legend, which is titled “Nitrogenous bases” as “adenine” and “thymine,” respectively. The blue and green bars, which are always paired together, are labeled in the legend as “guanine” and “cytosine,” respectively. At the top of the helical structure, the left-hand side rail, or “Sugar, dash, phosphate backbone,” is labeled as “3, prime” while the right is labeled as “5, prime.” These labels are reversed at the bottom of the helix. To the right of the page is a large Lewis structure. The top left corner of this structure, labeled “5, prime,” shows a phosphorus atom single bonded to three oxygen atoms, one of which has a superscripted negative charge, and double bonded to a fourth oxygen atom. One of the single bonded oxygen atoms is single bonded to the left corner of a five-membered ring with an oxygen atom at its top point and which is single bonded to an oxygen atom on the bottom left. This oxygen atom is single bonded to a phosphorus atom that is single bonded to two other hydrogen atoms and double bonded to a fourth oxygen atom. The lower left of these oxygen atoms is single bonded to another oxygen atom that is single bonded to a five-membered ring with an oxygen in the upper bonding site. The bottom left of this ring has a hydroxyl group attached to it while the upper right carbon is single bonded to a nitrogen atom that is part of a five-membered ring bonded to a six-membered ring. Both of these rings have points of unsaturation and nitrogen atoms bonded into their structures. On the right side of the six-membered ring are two single bonded amine groups and a double bonded oxygen. Three separate dotted lines extend from these sites to corresponding sites on a second six-membered ring. This ring has points of unsaturation and a nitrogen atom in the bottom right bonding position that is single bonded to a five-membered ring on the right side of the image. This ring is single bonded to a carbon that is single bonded to an oxygen that is single bonded to a phosphorus. The phosphorus is single bonded to two other oxygen atoms and double bonded to a fourth oxygen atom. This group is labeled “5, prime.” The five-membered ring is also bonded on the top side to an oxygen that is bonded to a phosphorus single bonded to two other oxygen atoms and double bonded to a fourth oxygen atom. The upper left oxygen of this group is single bonded to a carbon that is single bonded to a five-membered ring with an oxygen in the bottom bonding position. This ring has a hydroxyl group on its upper right side that is labeled “3, prime” and is bonded on the left side to a nitrogen that is a member of a five-membered ring. This ring is bonded to a six-membered ring and both have points of unsaturation. This ring has a nitrogen on the left side, as well as an amine group, that have two dotted lines leading from them to an oxygen and amine group on a six membered ring. These dotted lines are labeled “Hydrogen bonds.” The six membered ring also has a double bonded oxygen on its lower side and a nitrogen atom on its left side that is single bonded to a five-membered ring. This ring connects to the two phosphate groups mentioned at the start of this to form a large circle. The name “guanine” is written below the lower left side of this image while the name “cytosine” is written on the lower right. The name “thymine” is written above the right side of the image and “adenine” is written on the top right. Three sections are indicated below the images where the left is labeled “Sugar, dash, phosphate backbone,” the middle is labeled “Bases” and the right is labeled “Sugar, dash, phosphate backbone.”

Each nucleotide contains a (deoxyribose) sugar bound to a phosphate group on one side, and one of four nitrogenous bases on the other. Two of the bases, cytosine (C) and thymine (T), are single-ringed structures known as pyrimidines. The other two, adenine (A) and guanine (G), are double-ringed structures called purines. These bases form complementary base pairs consisting of one purine and one pyrimidine, with adenine pairing with thymine, and cytosine with guanine. Each base pair is held together by hydrogen bonding. A and T share two hydrogen bonds, C and G share three, and both pairings have a similar shape and structure Figure 9.2.12

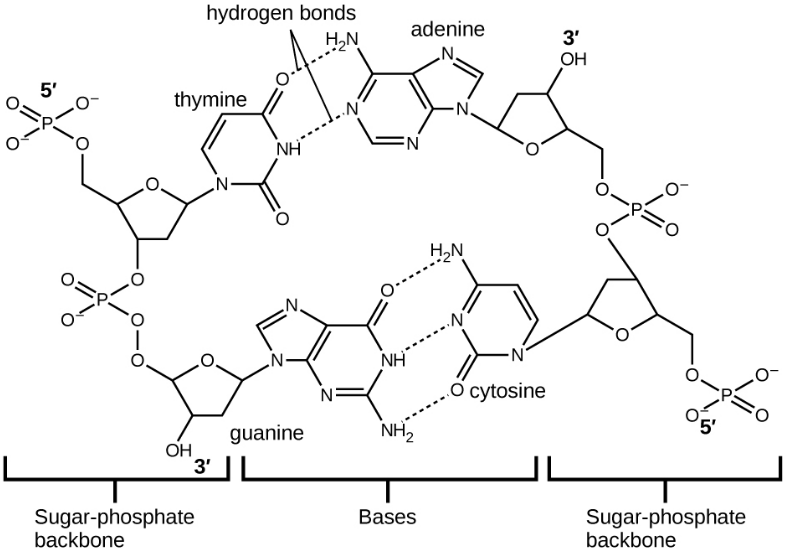


Figure 9.2.12: The geometries of the base molecules result in maximum hydrogen bonding between adenine and thymine (AT) and between guanine and cytosine (GC), so-called “complementary base pairs.”

A large Lewis structure is shown. The top left corner of this structure, labeled “5, prime,” shows a phosphorus atom single bonded to three oxygen atoms, one of which has a superscripted negative charge, and double bonded to a fourth oxygen atom. One of the single bonded oxygen atoms is single bonded to the left corner of a five-membered ring with an oxygen atom at its top point and which is single bonded to an oxygen atom on the bottom left. This oxygen atom is single bonded to a phosphorus atom that is single bonded to two other hydrogen atoms and double bonded to a fourth oxygen atom. The lower left of these oxygen atoms is single bonded to another oxygen atom that is single bonded to a five-membered ring with an oxygen in the upper bonding site. The bottom left of this ring has a hydroxyl group attached to it while the upper right carbon is single bonded to a nitrogen atom that is part of a five-membered ring bonded to a six-membered ring. Both of these rings have points of unsaturation and nitrogen atoms bonded into their structures. On the right side of the six-membered ring are two single bonded amine groups and a double bonded oxygen. Three separate dotted lines extend from these sites to corresponding sites on a second six-membered ring. This ring has points of unsaturation and a nitrogen atom in the bottom right bonding position that is single bonded to a five-membered ring on the right side of the image. This ring is single bonded to a carbon that is single bonded to an oxygen that is single bonded to a phosphorus. The phosphorus is single bonded to two other oxygen atoms and double bonded to a fourth oxygen atom. This group is labeled “5, prime.” The five-membered ring is also bonded on the top side to an oxygen that is bonded to a phosphorus single bonded to two other oxygen atoms and double bonded to a fourth oxygen atom. The upper left oxygen of this group is single bonded to a carbon that is single bonded to a five-membered ring with an oxygen in the bottom bonding position. This ring has a hydroxyl group on its upper right side that is labeled “3, prime” and is bonded on the left side to a nitrogen that is a member of a five-membered ring. This ring is bonded to a six-membered ring and both have points of unsaturation. This ring has a nitrogen on the left side, as well as an amine group, that have two dotted lines leading from them to an oxygen and amine group on a six membered ring. These dotted lines are labeled “Hydrogen bonds.” The six membered ring also has a double bonded oxygen on its lower side and a nitrogen atom on its left side that is single bonded to a five-membered ring. This ring connects to the two phosphate groups mentioned at the start of this to form a large circle. The name “guanine” is written below the lower left side of this image while the name “cytosine” is written on the lower right. The name “thymine” is written above the right side of the image and “adenine” is written on the top right. Three sections are indicated below the images where the left is labeled “Sugar, dash, phosphate backbone,” the middle is labeled “Bases” and the right is labeled “Sugar, dash, phosphate backbone.”

The cumulative effect of millions of hydrogen bonds effectively holds the two strands of DNA together. Importantly, the two strands of DNA can relatively easily “unzip” down the middle since hydrogen bonds are relatively weak compared to the covalent bonds that hold the atoms of the individual DNA molecules together. This allows both strands to function as a template for replication.

Summary

The physical properties of condensed matter (liquids and solids) can be explained in terms of the kinetic molecular theory. In a liquid, intermolecular attractive forces hold the molecules in contact, although they still have sufficient kinetic energy to move past each other. Intermolecular attractive forces, collectively referred to as van der Waals forces, are responsible for the behavior of liquids and solids and are electrostatic in nature. Dipole-dipole attractions result from the electrostatic attraction of the partial negative end of one dipolar molecule for the partial positive end of another. The temporary dipole that results from the motion of the electrons in an atom can induce a dipole in an adjacent atom and give rise to the London dispersion force. London forces increase with increasing molecular size. Hydrogen bonds are a special type of dipole-dipole attraction that results when hydrogen is bonded to one of the three most electronegative elements: F, O, or N.

9.3: Solubility

Learning Objectives

- Describe the effects of temperature and pressure on solubility
- State Henry's law and use it in calculations involving the solubility of a gas in a liquid
- Explain the degrees of solubility possible for liquid-liquid solutions

Imagine adding a small amount of salt to a glass of water, stirring until all the salt has dissolved, and then adding a bit more. You can repeat this process until the salt concentration of the solution reaches its natural limit, a limit determined primarily by the relative strengths of the solute-solute, solute-solvent, and solvent-solvent attractive forces discussed in the previous two modules of this chapter. You can be certain that you have reached this limit because, no matter how long you stir the solution, undissolved salt remains. The concentration of salt in the solution at this point is known as its solubility.

The solubility of a solute in a particular solvent is the maximum concentration that may be achieved under given conditions when the dissolution process is at equilibrium. Referring to the example of salt in water:

$$\mathrm{NaCl}(s) \rightleftharpoons \mathrm{Na}^+(aq) + \mathrm{Cl}^-(aq) \tag{9.3.1}$$

When a solute's concentration is equal to its solubility, the solution is said to be saturated with that solute. If the solute's concentration is less than its solubility, the solution is said to be unsaturated. A solution that contains a relatively low concentration of solute is called dilute, and one with a relatively high concentration is called concentrated.

If we add more salt to a saturated solution of salt, we see it fall to the bottom and no more seems to dissolve. In fact, the added salt does dissolve, as represented by the forward direction of the dissolution equation. Accompanying this process, dissolved salt will precipitate, as depicted by the reverse direction of the equation. The system is said to be at equilibrium when these two reciprocal processes are occurring at equal rates, and so the amount of undissolved and dissolved salt remains constant. Support for the simultaneous occurrence of the dissolution and precipitation processes is provided by noting that the number and sizes of the undissolved salt crystals will change over time, though their combined mass will remain the same.

Video 9.3.1: Watch this impressive video showing the precipitation of sodium acetate from a supersaturated solution.

Solutions may be prepared in which a solute concentration exceeds its solubility. Such solutions are said to be supersaturated, and they are interesting examples of nonequilibrium states. For example, the carbonated beverage in an open container that has not yet "gone flat" is supersaturated with carbon dioxide gas; given time, the CO_2 concentration will decrease until it reaches its equilibrium value.

Solutions of Gases in Liquids

In an earlier module of this chapter, the effect of intermolecular attractive forces on solution formation was discussed. 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, but 100 times less than the solubility of chloromethane, $CHCl_3$. Considering the role of the solvent's chemical structure, note that the solubility of oxygen in the 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 9.3.1). This is one of the major impacts resulting from the thermal pollution of natural bodies of water.

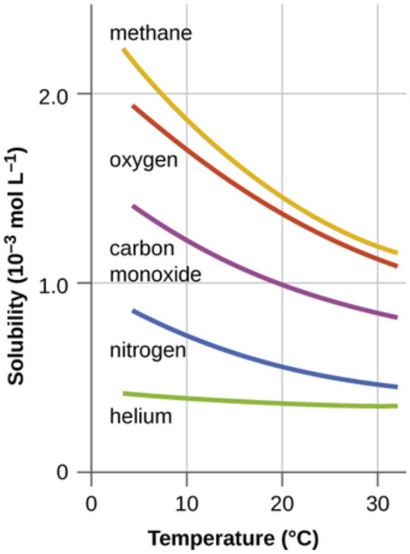


Figure 9.3.1: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

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 may have serious consequences for the health of the water's ecosystems and, in severe cases, can result in large-scale fish kills (Figure 9.3.2).

Figure 9.3.2: (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 solute in the gas to which the solution is exposed. Gas solubility increases as the pressure of the gas increases. Carbonated beverages provide a nice illustration of this relationship. The carbonation process involves exposing the beverage to a relatively high pressure of carbon dioxide gas and then sealing the beverage container, thus saturating the beverage with CO_2 at 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 9.3.3). 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."

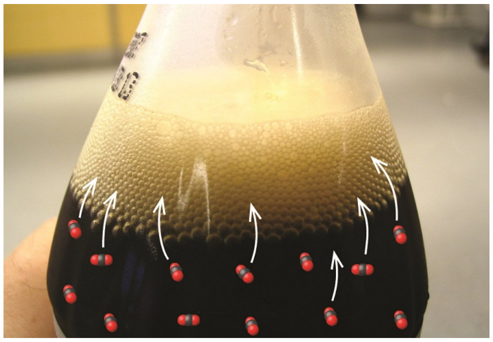


Figure 9.3.3: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (credit: modification of work by Derrick Coetzee)

A dark brown liquid is shown in a clear, colorless container. A thick layer of beige bubbles appear at the surface of the liquid. In the liquid, thirteen small clusters of single black spheres with two red spheres attached to the left and right are shown. Red spheres represent oxygen atoms and black represent carbon atoms. Seven white arrows point upward in the container from these clusters to the bubble layer at the top of the liquid.

For many gaseous solutes, the relation between solubility, C_g, and partial pressure, P_g, is a proportional one:

$$C_{\mathrm{g}} = kP_{\mathrm{g}}$$

where k is a proportionality constant that depends on the identities of the gaseous solute and solvent, and on the solution temperature. This is a mathematical statement of Henry's law: The quantity of an ideal gas that dissolves in a definite volume of liquid is directly proportional to the pressure of the gas.

✓ Example 9.3.1: Application of Henry's Law

At 20 °C, the concentration of dissolved oxygen in water exposed to gaseous oxygen at a partial pressure of 101.3 kPa (760 torr) is 1.38×10^{-3} mol L^{-1}. Use Henry's law to determine the solubility of oxygen when its partial pressure is 20.7 kPa (155 torr), the approximate pressure of oxygen in earth's atmosphere.

Solution

According to Henry's law, for an ideal solution the solubility, C_g, of a gas (1.38×10^{-3} mol L^{-1}, in this case) is directly proportional to the pressure, P_g, of the undissolved gas above the solution (101.3 kPa, or 760 torr, in this case). Because we know both C_g and P_g, we can rearrange this expression to solve for k.

$$\begin{aligned} C_{\mathrm{g}} &= kP_{\mathrm{g}} \\ k &= \frac{C_{\mathrm{g}}}{P_{\mathrm{g}}} \\ &= \frac{1.38 \times 10^{-3}\ \mathrm{mol\ L^{-1}}}{101.3\ \mathrm{kPa}} \\ &= 1.36 \times 10^{-5}\ \mathrm{mol\ L^{-1}\ kPa^{-1}} \\ &\quad (1.82 \times 10^{-6}\ \mathrm{mol\ L^{-1}\ torr^{-1}}) \end{aligned}$$

Now we can use k to find the solubility at the lower pressure.

$$\begin{aligned} &C_{\mathrm{g}} = kP_{\mathrm{g}} \\ &1.36 \times 10^{-5}\ \mathrm{mol\ L^{-1}\ kPa^{-1}} \times 20.7\ \mathrm{kPa} \\ &(\text{or } 1.82 \times 10^{-6}\ \mathrm{mol\ L^{-1}\ torr^{-1}} \times 155\ \mathrm{torr}) \\ &= 2.82 \times 10^{-4}\ \mathrm{mol\ L^{-1}} \end{aligned}$$

Note that various units may be used to express the quantities involved in these sorts of computations. Any combination of units that yield to the constraints of dimensional analysis are acceptable.

Exercise 9.3.1

Exposing a 100.0 mL sample of water at 0 °C to an atmosphere containing a gaseous solute at 20.26 kPa (152 torr) resulted in the dissolution of 1.45×10^{-3} g of the solute. Use Henry's law to determine the solubility of this gaseous solute when its pressure is 101.3 kPa (760 torr).

Answer

7.25×10^{-3} g in 100.0 mL or 0.0725 g/L

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.

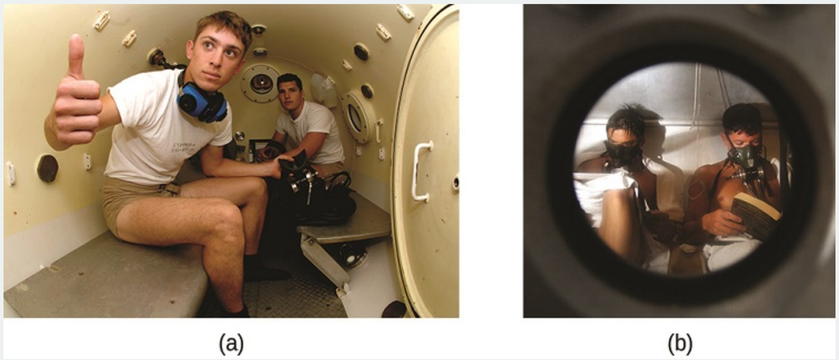
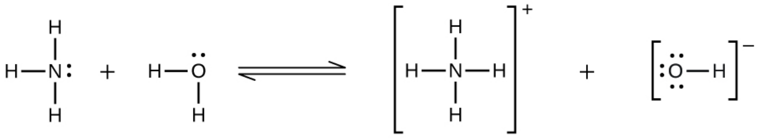

(a)
(b)

Figure 9.3.4: (a) US Navy divers undergo training in a recompression chamber. (b) Divers receive hyperbaric oxygen therapy.

Two photos are shown. The first shows two people seated in a steel chamber on benches that run length of the chamber on each side. The chamber has a couple of small circular windows and an open hatch-type door. One of the two people is giving a thumbs up gesture. The second image provides a view through a small, circular window. Inside the two people can be seen with masks over their mouths and noses. The people appear to be reading.

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 preventive measures are unsuccessful, divers with DCS are often provided hyperbaric oxygen therapy in pressurized vessels called decompression (or recompression) chambers (Figure 9.3.4).

Deviations from Henry's law are observed when a chemical reaction takes place between the gaseous solute and the solvent. 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. In 1986, more than 1700 people in Cameroon were killed when a cloud of gas, almost certainly carbon dioxide, bubbled from Lake Nyos (Figure 9.3.5), a deep lake in a volcanic crater. The water at the bottom of Lake Nyos is saturated with carbon dioxide by volcanic activity beneath the lake. 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 9.3.5: (a) It is believed that the 1986 disaster that killed more than 1700 people near Lake Nyos in Cameroon resulted when a large volume of carbon dioxide gas was released from the lake. (b) A CO_2 vent has since been installed to help outgas the lake in a slow, controlled fashion and prevent a similar catastrophe from happening in the future. (credit a: modification of work by Jack Lockwood; credit b: modification of work by Bill Evans)

Two photos are shown. The first is an aerial view of a lake surrounded by green hills. The second shows a large body of water with a fountain sending liquid up into the air several yards or meters above the surface of the water.

Solutions of Liquids in Liquids

We know that some liquids mix with each other in all proportions; in other words, they have infinite mutual solubility and are said to be miscible. Ethanol, sulfuric acid, and ethylene glycol (popular for use as antifreeze, pictured in Figure 9.3.6) are examples of liquids that are completely miscible with water. Two-cycle motor oil is miscible with gasoline.

Figure 9.3.6: Water and antifreeze are miscible; mixtures of the two are homogeneous in all proportions. (credit: “dno1967”/Wikimedia commons)

Liquids that mix with water in all proportions are usually polar substances or substances that form hydrogen bonds. For such liquids, the dipole-dipole attractions (or hydrogen bonding) of the solute molecules with the solvent molecules are at least as strong as those between molecules in the pure solute or in the pure solvent. Hence, the two kinds of molecules mix easily. Likewise, nonpolar liquids are miscible with each other because there is no appreciable difference in the strengths of solute-solute, solvent-solvent, and solute-solvent intermolecular attractions. The solubility of polar molecules in polar solvents and of nonpolar molecules in nonpolar solvents is, again, an illustration of the chemical axiom “like dissolves like.”

Two liquids that do not mix to an appreciable extent are called immiscible. Layers are formed when we pour immiscible liquids into the same container. Gasoline, oil (Figure 9.3.7), benzene, carbon tetrachloride, some paints, and many other nonpolar liquids are immiscible with water. The attraction between the molecules of such nonpolar liquids and polar water molecules is ineffectively weak. The only strong attractions in such a mixture are between the water molecules, so they effectively squeeze out the molecules of the nonpolar liquid. The distinction between immiscibility and miscibility is really one of degrees, so that miscible liquids are of infinite mutual solubility, while liquids said to be immiscible are of very low (though not zero) mutual solubility.

Figure 9.3.7: Water and oil are immiscible. Mixtures of these two substances will form two separate layers with the less dense oil floating on top of the water. (credit: “Yortw”/Flickr)

This is a photo of a clear, colorless martini glass containing a golden colored liquid layer resting on top of a clear, colorless liquid.

Two liquids, such as bromine and water, that are of moderate mutual solubility are said to be partially miscible. Two partially miscible liquids usually form two layers when mixed. In the case of the bromine and water mixture, the upper layer is water, saturated with bromine, and the lower layer is bromine saturated with water. Since bromine is nonpolar, and, thus, not very soluble in water, the water layer is only slightly discolored by the bright orange bromine dissolved in it. Since the solubility of water in bromine is very low, there is no noticeable effect on the dark color of the bromine layer (Figure 9.3.8).

Figure 9.3.8: Bromine (the deep orange liquid on the left) and water (the clear liquid in the middle) are partially miscible. The top layer in the mixture on the right is a saturated solution of bromine in water; the bottom layer is a saturated solution of water in bromine. (credit: Paul Flowers)

This figure shows three test tubes. The first test tube holds a dark orange-brown substance. The second test tube holds a clear substance. The amount of substance in both test tubes is the same. The third test tube holds a dark orange-brown substance on the bottom with a lighter orange substance on top. The amount of substance in the third test tube is almost double of the first two.

Solutions of Solids in Liquids

The dependence of solubility on temperature for a number of inorganic solids in water is shown by the solubility curves in Figure 9.3.9. Reviewing these data indicate a general trend of increasing solubility with temperature, although there are exceptions, as illustrated by the ionic compound cerium sulfate.

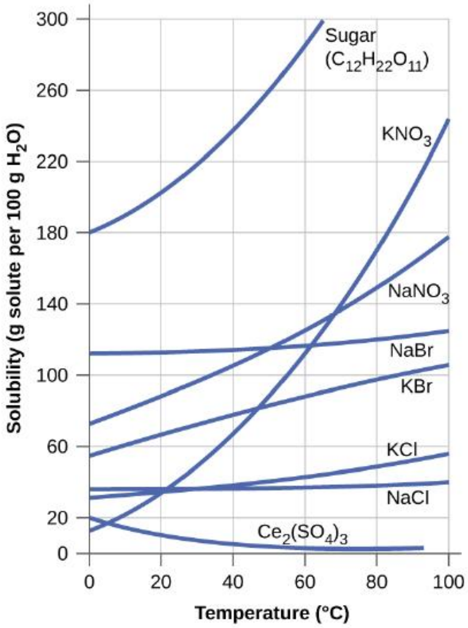


Figure 9.3.9: This graph shows how the solubility of several solids changes with temperature.

This shows a graph of the solubility of sugar C subscript 12 H subscript 22 O subscript 11, K N O subscript 3, N a N O subscript 3, N a B r, K B r, N a subscript 2 S O subscript 4, K C l, and C e subscript 2 left parenthesis S O subscript 4 right parenthesis subscript 3 in g solute per 100 g H subscript 2 O at temperatures ranging from 0 degrees Celsius to 100 degrees Celsius. At 0 degrees Celsius, solubilities are approximately 180 for sugar C subscript 12 H subscript 22 O subscript 11, 115 for K N O subscript 3, 75 for N a N O subscript 3, 115 for N a B r, 55 for K B r, 7 for N a subscript 2 S O subscript 4, 25 for K C l, and 20 for C e subscript 2 left parenthesis S O subscript 4 right parenthesis subscript 3. At 0 degrees Celsius, solubilities are approximately 180 for sugar C subscript 12 H subscript 22 O subscript 11, 115 for K N O subscript 3, 75 for N a N O subscript 3, 115 for N a B r, 55 for K B r, 7 for N a subscript 2 S O subscript 4, 25 for K C l, and 20 for C e subscript 2 left parenthesis S O subscript 4 right parenthesis subscript 3. At 100 degrees Celsius, sugar C subscript 12 H subscript 22 O subscript 11 has exceeded the upper limit of solubility indicated on the graph, 240 for K N O subscript 3, 178 for N a N O subscript 3, 123 for N a B r, 105 for K B r, 52 for N a subscript 2 S O subscript 4, 58 for K C l, and the graph for C e subscript 2 left parenthesis S O subscript 4 right parenthesis subscript 3 stops at about 92 degrees Celsius where the solubility is nearly zero. The graph for N a subscript 2 S O subscript 4 is shown in red. All others substances are shown in blue. The solubility of this substance increases until about 30 degrees Celsius and declines beyond that point with increasing temperature.

The temperature dependence of solubility can be exploited to prepare supersaturated solutions of certain compounds. A solution may be saturated with the compound at an elevated temperature (where the solute is more soluble) and subsequently cooled to a lower temperature without precipitating the solute. The resultant solution contains solute at a concentration greater than its equilibrium solubility at the lower temperature (i.e., it is supersaturated) and is relatively stable. Precipitation of the excess solute can be initiated by adding a seed crystal (see the video in the Link to Learning earlier in this module) or by mechanically agitating the solution. Some hand warmers, such as the one pictured in Figure 9.3.10, take advantage of this behavior.

Figure 9.3.10: This hand warmer produces heat when the sodium acetate in a supersaturated solution precipitates. Precipitation of the solute is initiated by a mechanical shockwave generated when the flexible metal disk within the solution is “clicked.” (credit: modification of work by “Velela”/Wikimedia Commons)

Three photos of hand warmers are shown side by side with an arrow pointing from the first photo to the second, and another arrow pointing from the second photo to the third. The first packet contains a clear colorless liquid and a small metal disc can be seen. In the second packet, the disc can’t be seen and a dispersion of white liquid is beginning. In the third packet, all of the liquid is white.

Video 9.3.2: This video shows the crystallization process occurring in a hand warmer.

Summary

The extent to which one substance will dissolve in another is determined by several factors, including the types and relative strengths of intermolecular attractive forces that may exist between the substances’ atoms, ions, or molecules. This tendency to dissolve is quantified as substance’s solubility, its maximum concentration in a solution at equilibrium under specified conditions. A saturated solution contains solute at a concentration equal to its solubility. A supersaturated solution is one in which a solute’s concentration exceeds its solubility—a nonequilibrium (unstable) condition that will result in solute precipitation when the solution is appropriately perturbed. Miscible liquids are soluble in all proportions, and immiscible liquids exhibit very low mutual solubility. Solubilities for gaseous solutes decrease with increasing temperature, while those for most, but not all, solid solutes increase with temperature. The concentration of a gaseous solute in a solution is proportional to the partial pressure of the gas to which the solution is exposed, a relation known as Henry’s law.

9.4: Content in Context

This section applies chemical principles—molecular polarity, intermolecular forces, and solubility—to real-world scenarios, exploring topics like oil slicks, miscibility, and hydrogen bonding to understand their impact on everyday life and the environment.

9.4.1: Oil Slicks and Miscibility

The central principle governing solubility is the concept of "like dissolves like." In other words, molecules with similar intermolecular forces are able to dissolve in each other, while molecules with dramatically different intermolecular forces generally cannot. A classic example is the fact that "oil and water don't mix." Oil is composed largely of nonpolar hydrocarbons, which exert only relatively weak London dispersion forces. Water molecules, in contrast, are held together by relatively strong hydrogen bonding between a hydrogen atom on one molecule (which has a partial positive charge) and the oxygen atom on another molecule (which has a partial negative charge).

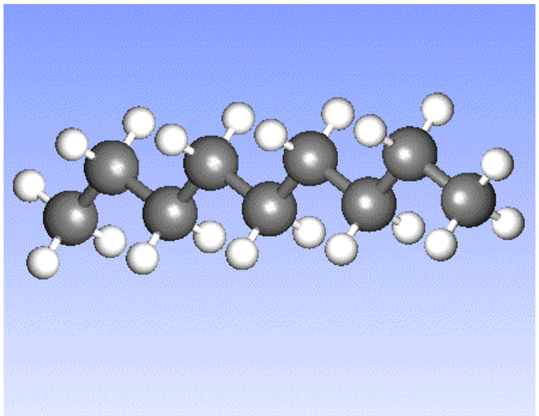


Figure 9.4.1.1 *A nonpolar hydrocarbon molecule is held together by London dispersion forces*

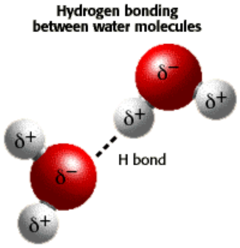


Figure 9.4.1.2 *A polar water molecule is held together by hydrogen bonding*

IWC replica Réplica de reloj The strong intermolecular forces between the water molecules make them tend to cluster together, excluding the more weakly attractive oil molecules, which then group together on their own. On a macro level, this causes water and oil to separate into two different layers. Since oil is less dense than water, the oil layer floats above the water layer, creating an oil "slick" on the surface of the water.

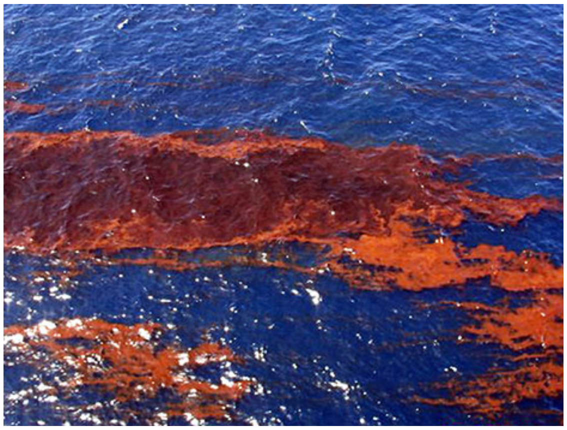


Figure 9.4.1.3 *Oil slick on the surface of the Gulf of Mexico*

This phenomenon takes on serious environmental consequences when crude oil is spilled into natural bodies of water, whether through an oil tanker accident such as the Exxon Valdez oil spill in 1989, or through an oil well accident such as the BP Deepwater Horizon oil spill in 2010. Overall, the National Research Council estimates that an average of three million gallons of oil and other petroleum products are spilled into U.S. waters every year (1). The sticky, viscous oil slick that forms on the water's surface can trap and damage seabirds, turtles, and marine mammals, and if washed ashore also threatens biologically important shorelines and beaches (2). While this surface oil can sometimes be removed by skimming or burning, these techniques are not always possible or practical, and have their own potential side effects. (For example, the burning of oil generates carbon dioxide and other airborne pollutants.) An alternative strategy is to try to disperse the oil into the water in small droplets, a process called emulsification. While this does not physically remove the oil from the water, it does break up the potentially dangerous oil slick and prevents it from reaching the shore. Emulsification may also speed up biodegradation of the oil by naturally occurring bacteria by as much as 50% (3).

While the action of wind and waves may naturally disperse some of the oil into small droplets, another approach is to use a chemical dispersant to emulsify the oil. Dispersants tend to be mixtures of three types of chemicals: solvents, additives, and surface-active agents, commonly called surfactants. Surfactants, which are the key to the dispersant's effectiveness, are compounds containing two different kinds of chemical groups: one that is oil-compatible (lipophilic or hydrophobic) and one that is water-compatible (hydrophilic). The lipophilic end of the surfactant molecule has London dispersion forces similar to those of the oil, allowing this portion of the molecule to "dissolve" in the oil layer. In contrast, the hydrophilic end has dipole-dipole forces and/or hydrogen bonding similar to that of water, allowing this portion of the molecule to "dissolve" in the water. This dual action reduces what is called the interfacial tension between the oil and the water, allowing the oil to break up into small droplets surrounded by surfactant molecules.

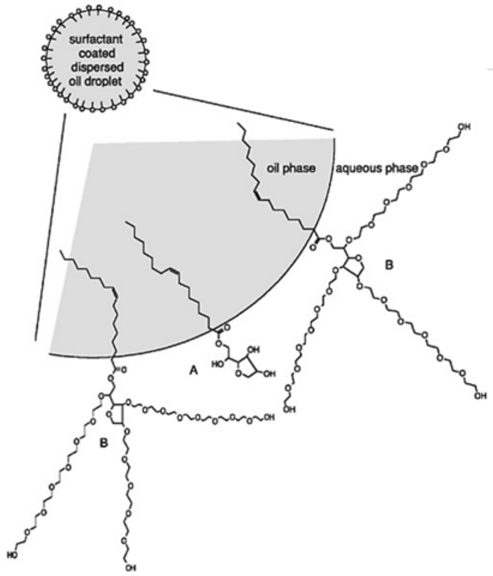


Figure 9.4.1.4 *Interaction of surfactants with an oil-water interface. Surfactant A is sorbitan monooleate, while surfactant B is ethoxylated sorbitan monooleate*

Natural water turbulence then pulls these droplets apart and down into the water column (a process called entrainment), allowing the oil to disperse through the water. Note that this does NOT mean that the oil has actually dissolved in the water! The oil and water are still insoluble in each other; however, the dispersant enables the oil to exist in small, stable droplets within the water column instead of in a single layer on top of the water(1).

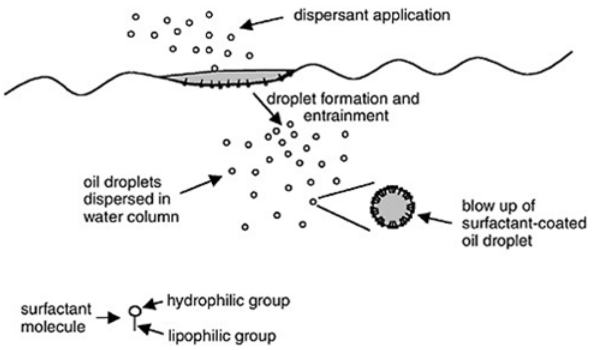


Figure 9.4.1.5 *Diagram showing the action of a dispersant on an oil slick*

Video showing the action of a dispersant on an oil slick: www.fox10tv.com/dpp/news/gulf...ispersant-work

The predominant dispersant used in the BP Deepwater Horizon oil spill was Corexit EC9500A, produced by the Nalco company. The primary surfactants found in this dispersant are sorbitan monooleate and ethoxylated sorbitan monooleates (i.e., surfactants A and B in the diagram above) along with dioctyl sodium sulfosuccinate, which is a detergent that is also commonly found in laxatives (1, 4).

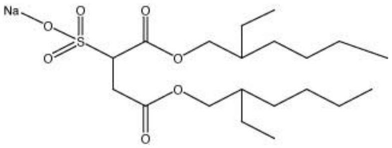


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Figure 9.4.1.6 *Chemical structure of dioctyl sodium sulfosuccinate*

Corexit EC9500A has a reasonable dispersant effectiveness of about 55% for South Louisiana crude oil, and the EPA has deemed it no more toxic than other dispersant chemicals currently in use (5, 3). Nevertheless, the application of this dispersant, both in the Deepwater Horizon incident and for other oil spills, remains controversial. As the National Research Council points out, adding a dispersant to an oil spill increases the hydrocarbon load on the ecosystem and has the potential to harm sensitive marine organisms. Thus, the decision to use a dispersant is always a trade-off: it increases the risk to organisms in the water column and the seafloor while decreasing risks to organisms on the water surface and shoreline (1).

From ChemPRIME: 10.18: Solubility and Molecular Structure

References:

1) National Research Council. Oil Spill Dispersants: Efficacy and Effects. Washington, D.C.: National Academies Press, 2005.
2) Graham, P. (2010). Deep Sea Oil Spill Cleanup Techniques: Applicability, Trade-offs and Advantages. ProQuest Discovery Guides, www.csa.com/discoveryguides/oil/review.pdf
3) www.epa.gov/bpspill/dispersants-qanda.html
4) http://www.nytimes.com/gwire/2010/06/09/09greenwire-ingredients-of-controversial-dispersants-used-42891.html
5) www.epa.gov/osweroe1/content/...s/corex950.htm

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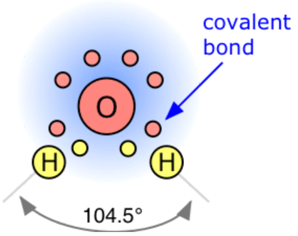
9.4.2: Hydrogen Bonding Allows Life to Exist

Learning Objectives

- Identify three special properties of water that make it unusual for a molecule of its size, and explain how these result from hydrogen bonding.
- Explain what is meant by hydrogen bonding and the molecular structural features that bring it about.
- Describe the "structure", such as it is, of liquid water.
- Sketch out structural examples of hydrogen bonding in three small molecules other than H_2O.
- Describe the roles of hydrogen bonding in proteins and in DNA.

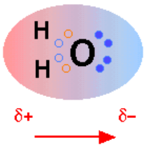
Most students of chemistry quickly learn to relate the structure of a molecule to its general properties. Thus we generally expect small molecules to form gases or liquids, and large ones to exist as solids under ordinary conditions. And then we come to H_2O, and are shocked to find that many of the predictions are way off, and that water (and by implication, life itself) should not even exist on our planet! In this section we will learn why this tiny combination of three nuclei and ten electrons possesses special properties that make it unique among the more than 15 million chemical species we presently know.

In water, each hydrogen nucleus is covalently bound to the central oxygen atom by a pair of electrons that are shared between them. In H_2O, only two of the six outer-shell electrons of oxygen are used for this purpose, leaving four electrons which are organized into two non-bonding pairs. The four electron pairs surrounding the oxygen tend to arrange themselves as far from each other as possible in order to minimize repulsions between these clouds of negative charge. This would ordinarily result in a tetrahedral geometry in which the angle between electron pairs (and therefore the H-O-H *bond angle*) is 109.5°. However, because the two non-bonding pairs remain closer to the oxygen atom, these exert a stronger repulsion against the two covalent bonding pairs, effectively pushing the two hydrogen atoms closer together. The result is a distorted tetrahedral arrangement in which the H—O—H angle is 104.5°.

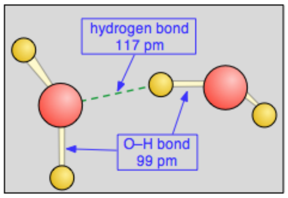


Water's large dipole moment leads to hydrogen bonding

The H_2O molecule is electrically neutral, but the positive and negative charges are not distributed uniformly. This is illustrated by the gradation in color in the schematic diagram here. The electronic (negative) charge is concentrated at the oxygen end of the molecule, owing partly to the nonbonding electrons (solid blue circles), and to oxygen's high nuclear charge which exerts stronger attractions on the electrons. This charge displacement constitutes an *electric dipole*, represented by the arrow at the bottom; you can think of this dipole as the electrical "image" of a water molecule.

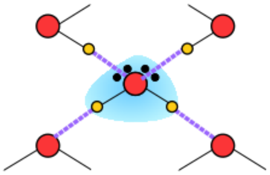


Opposite charges attract, so it is not surprising that the negative end of one water molecule will tend to orient itself so as to be close to the positive end of another molecule that happens to be nearby. The strength of this *dipole-dipole attraction* is less than that of a normal chemical bond, and so it is completely overwhelmed by ordinary thermal motions in the gas phase. However, when the H_2O molecules are crowded together in the liquid, these attractive forces exert a very noticeable effect, which we call (somewhat misleadingly) *hydrogen bonding*. And at temperatures low enough to turn off the disruptive effects of thermal motions, water freezes into ice in which the hydrogen bonds form a rigid and stable network.



Notice that the hydrogen bond (shown by the dashed green line) is somewhat longer than the covalent O—H bond. It is also *much weaker*, about 23 kJ mol^{-1} compared to the O–H covalent bond strength of 492 kJ mol^{-1}.

Water has long been known to exhibit many physical properties that distinguish it from other small molecules of comparable mass. Although chemists refer to these as the "anomalous" properties of water, they are by no means mysterious; all are entirely predictable consequences of the way the size and nuclear charge of the oxygen atom conspire to distort the electronic charge clouds of the atoms of other elements when these are chemically bonded to the oxygen.



The combination of large bond dipoles and short dipole–dipole distances results in very strong dipole–dipole interactions called hydrogen bonds, as shown for ice in Figure $9.4.2.6$. A hydrogen bond is usually indicated by a dotted line between the hydrogen atom attached to O, N, or F (the *hydrogen bond donor*) and the atom that has the lone pair of electrons (the *hydrogen bond acceptor*). Because each water molecule contains two hydrogen atoms and two lone pairs, a tetrahedral arrangement maximizes the number of hydrogen bonds that can be formed. In the structure of ice, each oxygen atom is surrounded by a distorted tetrahedron of hydrogen atoms that form bridges to the oxygen atoms of adjacent water molecules. The bridging hydrogen atoms are *not* equidistant from the two oxygen atoms they connect, however. Instead, each hydrogen atom is 101 pm from one oxygen and 174 pm from the other. In contrast, each oxygen atom is bonded to two H atoms at the shorter distance and two at the longer distance, corresponding to two O–H covalent bonds and two O⋯H hydrogen bonds from adjacent water molecules, respectively. The resulting open, cagelike 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.

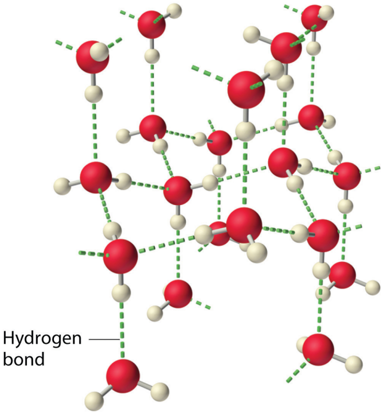


Figure $\boldsymbol{9.4.2.6}$***:*** *The Hydrogen-Bonded Structure of Ice.*

Each water molecule accepts two hydrogen bonds from two other water molecules and donates two hydrogen atoms to form hydrogen bonds with two more water molecules, producing an open, cagelike structure. 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.

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

Boiling Point

Molecules with hydrogen atoms bonded to electronegative atoms such as O, N, and F (and to a much lesser extent Cl and S) tend to exhibit unusually strong intermolecular interactions. These result in much higher boiling points than are observed for substances in which London dispersion forces dominate, as illustrated for the covalent hydrides of elements of groups 14–17 in Figure 9.4.2.5. Methane and its heavier congeners in group 14 form a series whose boiling points increase smoothly with increasing molar mass. This is the expected trend in nonpolar molecules, for which London dispersion forces are the exclusive intermolecular forces. In contrast, the hydrides of the lightest members of groups 15–17 have boiling points that are more than 100°C greater than predicted on the basis of their molar masses. The effect is most dramatic for water: if we extend the straight line connecting the points for H_2Te and H_2Se to the line for period 2, we obtain an estimated boiling point of −130°C for water! Imagine the implications for life on Earth if water boiled at −130°C rather than 100°C.

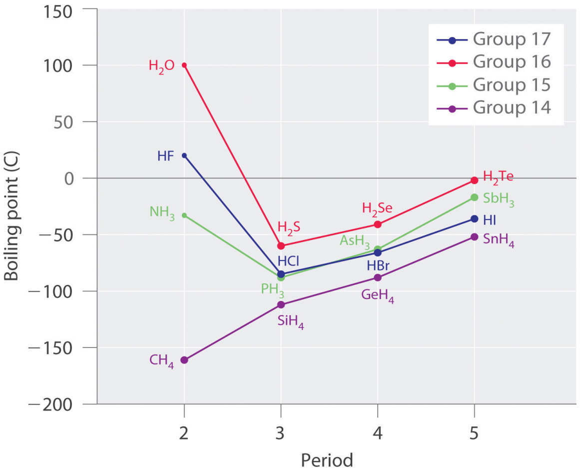
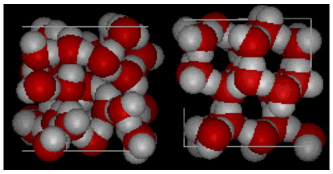


Figure 9.4.2.5: The Effects of Hydrogen Bonding on Boiling Points. These plots of the boiling points of the covalent hydrides of the elements of groups 14–17 show that the boiling points of the lightest members of each series for which hydrogen bonding is possible (HF, NH_3, and H_2O) are anomalously high for compounds with such low molecular masses.

Why do strong intermolecular forces produce such anomalously high boiling points and other unusual properties, such as high enthalpies of vaporization and high melting points? The answer lies in the highly polar nature of the bonds between hydrogen and very electronegative elements such as O, N, and F. The large difference in electronegativity results in a large partial positive charge on hydrogen and a correspondingly large partial negative charge on the O, N, or F atom. Consequently, H–O, H–N, and H–F bonds have very large bond dipoles that can interact strongly with one another. Because a hydrogen atom is so small, these dipoles can also approach one another more closely than most other dipoles.

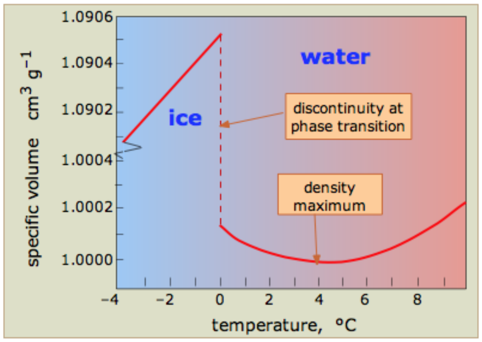
Ice Floats on Water

The most energetically favorable configuration of H_2O molecules is one in which each molecule is hydrogen-bonded to four neighboring molecules. Owing to the thermal motions described above, this ideal is never achieved in the liquid, but when water freezes to ice, the molecules settle into exactly this kind of an arrangement in the ice crystal. This arrangement requires that the molecules be somewhat farther apart then would otherwise be the case; as a consequence, ice, in which hydrogen bonding is at its maximum, has a more open structure, and thus a lower density than water.



Here are three-dimensional views of a typical local structure of water (left) and ice (right.) Notice the greater openness of the ice structure which is necessary to ensure the strongest degree of hydrogen bonding in a uniform, extended crystal lattice. The more crowded and jumbled arrangement in liquid water can be sustained only by the greater amount of thermal energy available above the freezing point.

When ice melts, the more vigorous thermal motion disrupts much of the hydrogen-bonded structure, allowing the molecules to pack more closely. Water is thus one of the very few substances whose solid form has a lower density than the liquid at the freezing point. Localized clusters of hydrogen bonds still remain, however; these are continually breaking and reforming as the thermal motions jiggle and shove the individual molecules. As the temperature of the water is raised above freezing, the extent and lifetimes of these clusters diminish, so the density of the water increases.



At higher temperatures, another effect, common to all substances, begins to dominate: as the temperature increases, so does the amplitude of thermal motions. This more vigorous jostling causes the average distance between the molecules to increase, reducing the density of the liquid; this is ordinary thermal expansion. Because the two competing effects (hydrogen bonding at low temperatures and thermal expansion at higher temperatures) both lead to a decrease in density, it follows that there must be some temperature at which the density of water passes through a maximum. This temperature is 4° C; this is the temperature of the water you will find at the bottom of an ice-covered lake in which this most dense of all water has displaced the colder water and pushed it nearer to the surface.

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.

Contributors

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CHAPTER OVERVIEW

10: Radiation, Electromagnetic Waves, Biological Effects of UV Radiation

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10.1: Radiation and Electromagnetic waves

Learning Objectives

- Explain the basic behavior of waves.
- Identify the forms of radiation present in the electromagnetic spectrum.
- Distinguish forms of radiation based on wavelength, frequency, and energy
- Describe the wave nature of light
- Use appropriate equations to calculate related light-wave properties such as period, frequency, wavelength, and energy
- Describe the particle nature of light

The nature of light has been a subject of inquiry since antiquity. In the seventeenth century, Isaac Newton performed experiments with lenses and prisms and was able to demonstrate that white light consists of the individual colors of the rainbow combined together. Newton explained his optics findings in terms of a "corpuscular" view of light, in which light was composed of streams of extremely tiny particles travelling at high speeds according to Newton's laws of motion. Others in the seventeenth century, such as Christiaan Huygens, had shown that optical phenomena such as reflection and refraction could be equally well explained in terms of light as waves travelling at high speed through a medium called "luminiferous aether" that was thought to permeate all space. Early in the nineteenth century, Thomas Young demonstrated that light passing through narrow, closely spaced slits produced interference patterns that could not be explained in terms of Newtonian particles but could be easily explained in terms of waves. Later in the nineteenth century, after James Clerk Maxwell developed his theory of electromagnetic radiation and showed that light was the visible part of a vast spectrum of electromagnetic waves, the particle view of light became thoroughly discredited. By the end of the nineteenth century, scientists viewed the physical universe as roughly comprising two separate domains: matter composed of particles moving according to Newton's laws of motion, and electromagnetic radiation consisting of waves governed by Maxwell's equations. Today, these domains are referred to as classical mechanics and classical electrodynamics (or classical electromagnetism). Although there were a few physical phenomena that could not be explained within this framework, scientists at that time were so confident of the overall soundness of this framework that they viewed these aberrations as puzzling paradoxes that would ultimately be resolved somehow within this framework. As we shall see, these paradoxes led to a contemporary framework that intimately connects particles and waves at a fundamental level called wave-particle duality, which has superseded the classical view.

Visible light and other forms of electromagnetic radiation play important roles in chemistry, since they can be used to infer the energies of electrons within atoms and molecules. Much of modern technology is based on electromagnetic radiation. For example, radio waves from a mobile phone, X-rays used by dentists, the energy used to cook food in your microwave, the radiant heat from red-hot objects, and the light from your television screen are forms of electromagnetic radiation that all exhibit wavelike behavior.

Waves

A wave is an oscillation or periodic movement that can transport energy from one point in space to another. Common examples of waves are all around us. Shaking the end of a rope transfers energy from your hand to the other end of the rope, dropping a pebble into a pond causes waves to ripple outward along the water's surface, and the expansion of air that accompanies a lightning strike generates sound waves (thunder) that can travel outward for several miles. In each of these cases, kinetic energy is transferred through matter (the rope, water, or air) while the matter remains essentially in place. An insightful example of a wave occurs in sports stadiums when fans in a narrow region of seats rise simultaneously and stand with their arms raised up for a few seconds before sitting down again while the fans in neighboring sections likewise stand up and sit down in sequence. While this wave can quickly encircle a large stadium in a few seconds, none of the fans actually travel with the wave-they all stay in or above their seats.

Waves need not be restricted to travel through matter. As Maxwell showed, electromagnetic waves consist of an electric field oscillating in step with a perpendicular magnetic field, both of which are perpendicular to the direction of travel. These waves can travel through a vacuum at a constant speed of 2.998×10^8 m/s, the speed of light (denoted by c).

All waves, including forms of electromagnetic radiation, are characterized by, a wavelength (denoted by λ, the lowercase Greek letter lambda), a frequency (denoted by ν, the lowercase Greek letter nu), and an amplitude. As can be seen in Figure 10.1.1, the wavelength is the distance between two consecutive peaks or troughs in a wave (measured in meters in the SI system). Electromagnetic waves have wavelengths that fall within an enormous range-wavelengths of kilometers (10^3 m) to picometers

(10^{-12} m) have been observed. The frequency is the number of wave cycles that pass a specified point in space in a specified amount of time (in the SI system, this is measured in seconds). A cycle corresponds to one complete wavelength. The unit for frequency, expressed as cycles per second [s^{-1}], is the hertz (Hz). Common multiples of this unit are megahertz, (1 MHz = 1 × 10^6 Hz) and gigahertz (1 GHz = 1 × 10^9 Hz). The amplitude corresponds to the magnitude of the wave's displacement and so, in Figure, this corresponds to one-half the height between the peaks and troughs. The amplitude is related to the intensity of the wave, which for light is the brightness, and for sound is the loudness.

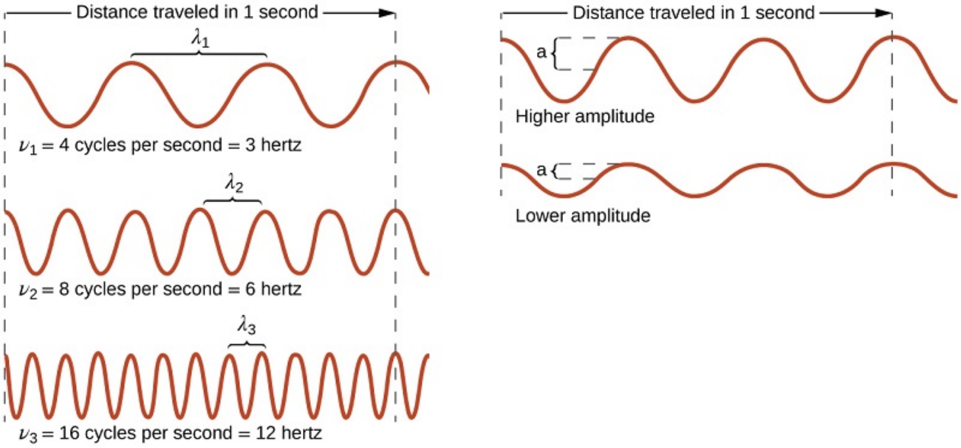


Figure 10.1.1: One-dimensional sinusoidal waves show the relationship among wavelength, frequency, and speed. The wave with the shortest wavelength has the highest frequency. Amplitude is one-half the height of the wave from peak to trough. (CC by 4.0; OpenStax)

This figure includes 5 one-dimensional sinusoidal waves in two columns. The column on the left includes three waves, and the column on the right includes two waves. In each column, dashed vertical line segments extend down the left and right sides of the column. A right pointing arrow extends from the left dashed line to the right dashed line in both columns and is labeled, “Distance traveled in 1 second.” The waves all begin on the left side at a crest. The wave at the upper left shows 3 peaks to the right of the starting point. A bracket labeled, “lambda subscript 1,” extends upward from the second and third peaks. Beneath this wave is the label, “nu subscript 1 equals 4 cycles per second equals 3 hertz.” The wave below has six peaks to the right of the starting point with a bracket similarly connecting the third and fourth peaks which is labeled, “lambda subscript 2.” Beneath this wave is the label, “nu subscript 2 equals 8 cycles per second equals 6 hertz” The third wave in the column has twelve peaks to the right of the starting point with a bracket similarly connecting the seventh and eighth peaks which is labeled, “lambda subscript 3.” Beneath this wave is the label, “nu subscript 3 equals 12 cycles per second equals 12 hertz.” All waves in this column appear to have the same vertical distance from peak to trough. In the second column, the two waves are similarly shown, but lack the lambda labels. The top wave in this column has a greater vertical distance between the peaks and troughs and is labeled, “Higher amplitude.” The wave beneath it has a lesser distance between the peaks and troughs and is labeled, “Lower amplitude.”

The product of a wave's wavelength (λ) and its frequency (ν), $\lambda\nu$, is the speed of the wave. Thus, for electromagnetic radiation in a vacuum:

$$c = 2.998 \times 10^8 \, \text{ms}^{-1} = \lambda\nu \tag{10.1.1}$$

Wavelength and frequency are inversely proportional: As the wavelength increases, the frequency decreases. The inverse proportionality is illustrated in Figure $10.1.2$ This figure also shows the electromagnetic spectrum, the range of all types of electromagnetic radiation. Each of the various colors of visible light has specific frequencies and wavelengths associated with them, and you can see that visible light makes up only a small portion of the electromagnetic spectrum. Because the technologies developed to work in various parts of the electromagnetic spectrum are different, for reasons of convenience and historical legacies, different units are typically used for different parts of the spectrum. For example, radio waves are usually specified as frequencies (typically in units of MHz), while the visible region is usually specified in wavelengths (typically in units of nm or angstroms).

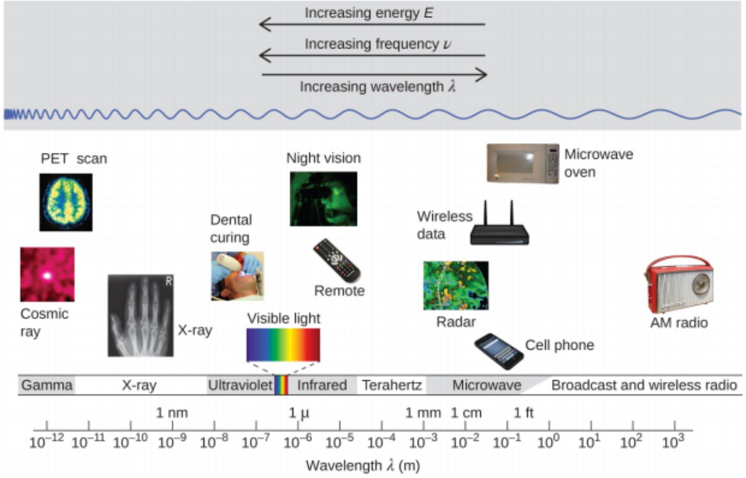


Figure 10.1.2: Portions of the electromagnetic spectrum are shown in order of decreasing frequency and increasing wavelength. Examples of some applications for various wavelengths include positron emission tomography (PET) scans, X-ray imaging, remote controls, wireless Internet, cellular telephones, and radios. (credit “Cosmic ray": modification of work by NASA; credit “PET scan": modification of work by the National Institute of Health; credit “X-ray": modification of work by Dr. Jochen Lengerke; credit “Dental curing": modification of work by the Department of the Navy; credit “Night vision": modification of work by the Department of the Army; credit “Remote": modification of work by Emilian Robert Vicol; credit “Cell phone": modification of work by Brett Jordan; credit “Microwave oven": modification of work by Billy Mabray; credit “Ultrasound": modification of work by Jane Whitney; credit “AM radio": modification of work by Dave Clausen)

✓ Example 10.1.1: Determining the Frequency and Wavelength of Radiation

A sodium streetlight gives off yellow light that has a wavelength of 589 nm (1 nm = 1 × 10^{-9} m). What is the frequency of this light?

Solution

We can rearrange the Equation 10.1.1 to solve for the frequency:

$$\nu = \frac{c}{\lambda}$$

Since *c* is expressed in meters per second, we must also convert 589 nm to meters.

$$\nu = \left(\frac{2.998 \times 10^{8} \ \cancel{\text{m}}\, \text{s}^{-1}}{589 \ \cancel{\text{nm}}}\right)\left(\frac{1 \times 10^{9} \ \cancel{\text{nm}}}{1 \ \cancel{\text{m}}}\right) = 5.09 \times 10^{14} \, \text{s}^{-1}$$

? Exercise 10.1.1

One of the frequencies used to transmit and receive cellular telephone signals in the United States is 850 MHz. What is the wavelength in meters of these radio waves?

Answer

0.353 m = 35.3 cm

Wireless Communication

Many valuable technologies operate in the radio (3 kHz-300 GHz) frequency region of the electromagnetic spectrum. At the low frequency (low energy, long wavelength) end of this region are AM (amplitude modulation) radio signals (540-2830 kHz) that can travel long distances. FM (frequency modulation) radio signals are used at higher frequencies (87.5-108.0 MHz). In AM radio, the information is transmitted by varying the amplitude of the wave (Figure 10.1.5). In FM radio, by contrast, the amplitude is constant and the instantaneous frequency varies.

Figure 10.1.3: Radio and cell towers are typically used to transmit long-wavelength electromagnetic radiation. Increasingly, cell towers are designed to blend in with the landscape, as with the Tucson, Arizona, cell tower (right) disguised as a palm tree. (credit left: modification of work by Sir Mildred Pierce; credit middle: modification of work by M.O. Stevens)

Other technologies also operate in the radio-wave portion of the electromagnetic spectrum. For example, 4G cellular telephone signals are approximately 880 MHz, while Global Positioning System (GPS) signals operate at 1.228 and 1.575 GHz, local area wireless technology (Wi-Fi) networks operate at 2.4 to 5 GHz, and highway toll sensors operate at 5.8 GHz. The frequencies associated with these applications are convenient because such waves tend not to be absorbed much by common building materials.

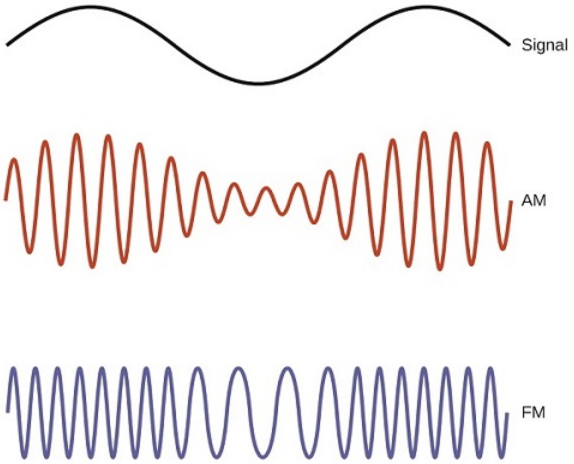


Figure 10.1.4: This schematic depicts how amplitude modulation (AM) and frequency modulation (FM) can be used to transmit a radio wave.

One particularly characteristic phenomenon of waves results when two or more waves come into contact: They interfere with each other. Figure 10.1.5 shows the interference patterns that arise when light passes through narrow slits closely spaced about a wavelength apart. The fringe patterns produced depend on the wavelength, with the fringes being more closely spaced for shorter wavelength light passing through a given set of slits. When the light passes through the two slits, each slit effectively acts as a new source, resulting in two closely spaced waves coming into contact at the detector (the camera in this case). The dark regions in Figure 10.1.5 correspond to regions where the peaks for the wave from one slit happen to coincide with the troughs for the wave from the other slit (destructive interference), while the brightest regions correspond to the regions where the peaks for the two waves (or their two troughs) happen to coincide (constructive interference). Likewise, when two stones are tossed close together into a pond, interference patterns are visible in the interactions between the waves produced by the stones. Such interference patterns cannot be explained by particles moving according to the laws of classical mechanics.

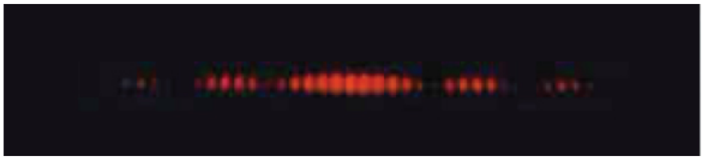


Figure 10.1.5: Interference fringe patterns are shown for light passing through two closely spaced, narrow slits. The spacing of the fringes depends on the wavelength, with the fringes being more closely spaced for the shorter-wavelength blue light. (credit: PASCO)

Dorothy Hodgkin

Because the wavelengths of X-rays (10-10,000 picometers [pm]) are comparable to the size of atoms, X-rays can be used to determine the structure of molecules. When a beam of X-rays is passed through molecules packed together in a crystal, the X-rays collide with the electrons and scatter. Constructive and destructive interference of these scattered X-rays creates a specific diffraction pattern. Calculating backward from this pattern, the positions of each of the atoms in the molecule can be determined very precisely. One of the pioneers who helped create this technology was Dorothy Crowfoot Hodgkin.

She was born in Cairo, Egypt, in 1910, where her British parents were studying archeology. Even as a young girl, she was fascinated with minerals and crystals. When she was a student at Oxford University, she began researching how X-ray crystallography could be used to determine the structure of biomolecules. She invented new techniques that allowed her and her students to determine the structures of vitamin B_{12}, penicillin, and many other important molecules. Diabetes, a disease that affects 382 million people worldwide, involves the hormone insulin. Hodgkin began studying the structure of insulin in 1934, but it required several decades of advances in the field before she finally reported the structure in 1969. Understanding the structure has led to better understanding of the disease and treatment options.

Not all waves are travelling waves. Standing waves (also known as stationary waves) remain constrained within some region of space. As we shall see, standing waves play an important role in our understanding of the electronic structure of atoms and molecules. The simplest example of a standing wave is a one-dimensional wave associated with a vibrating string that is held fixed at its two end points. Figure $10.1.6$ shows the four lowest-energy standing waves (the fundamental wave and the lowest three harmonics) for a vibrating string at a particular amplitude. Although the string's motion lies mostly within a plane, the wave itself is considered to be one dimensional, since it lies along the length of the string. The motion of string segments in a direction perpendicular to the string length generates the waves and so the amplitude of the waves is visible as the maximum displacement of the curves seen in Figure $10.1.6$. The key observation from the figure is that only those waves having an integer number, n, of half-wavelengths between the end points can form. A system with fixed end points such as this restricts the number and type of the possible waveforms. This is an example of quantization, in which only discrete values from a more general set of continuous values of some property are observed. Another important observation is that the harmonic waves (those waves displaying more than one-half wavelength) all have one or more points between the two end points that are not in motion. These special points are nodes. The energies of the standing waves with a given amplitude in a vibrating string increase with the number of half-wavelengths n. Since the number of nodes is $n – 1$, the energy can also be said to depend on the number of nodes, generally increasing as the number of nodes increases.

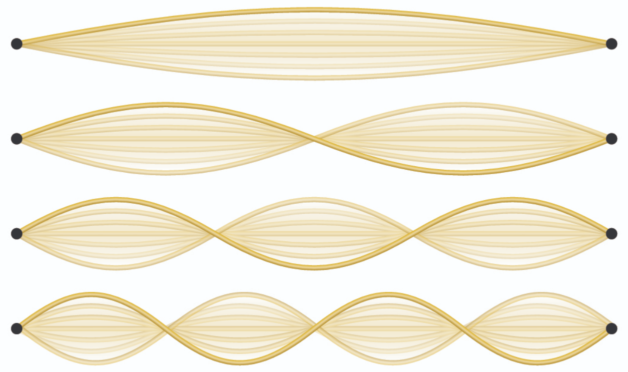


Figure 10.1.6: A vibrating string shows some one-dimensional standing waves. Since the two end points of the string are held fixed, only waves having an integer number of half-wavelengths can form. The points on the string between the end points that are not moving are called the nodes.

An example of two-dimensional standing waves is shown in Figure $10.1.7$ which shows the vibrational patterns on a flat surface. Although the vibrational amplitudes cannot be seen like they could in the vibrating string, the nodes have been made visible by sprinkling the drum surface with a powder that collects on the areas of the surface that have minimal displacement. For one-dimensional standing waves, the nodes were points on the line, but for two-dimensional standing waves, the nodes are lines on the surface (for three-dimensional standing waves, the nodes are two-dimensional surfaces within the three-dimensional volume). Because of the circular symmetry of the drum surface, its boundary conditions (the drum surface being tightly constrained to the circumference of the drum) result in two types of nodes: radial nodes that sweep out all angles at constant radii and, thus, are seen as circles about the center, and angular nodes that sweep out all radii at constant angles and, thus, are seen as lines passing through the center. The upper left image in Figure $10.1.7$ shows two radial nodes, while the image in the lower right shows the vibrational pattern associated with three radial nodes and two angular nodes.

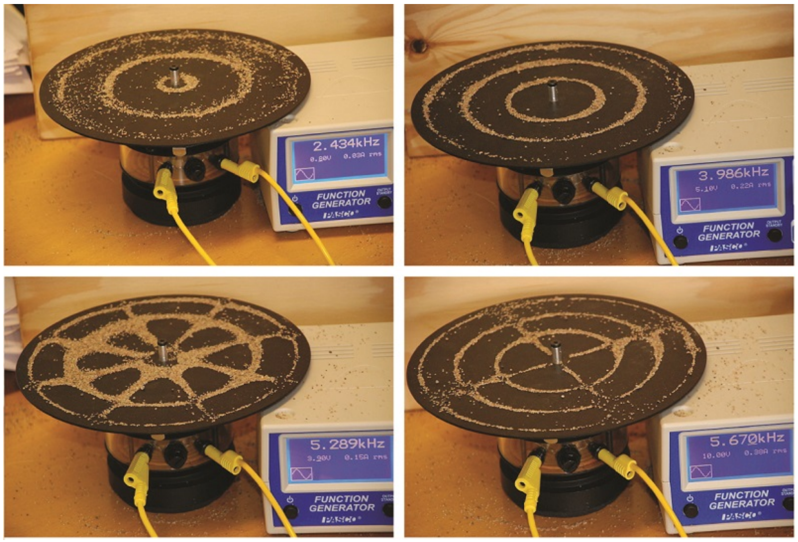


Figure $10.1.7$: *Two-dimensional standing waves can be visualized on a vibrating surface. The surface has been sprinkled with a powder that collects near the nodal lines. There are two types of nodes visible: radial nodes (circles) and angular nodes (radii). For a more animated video, check this link out.*

Blackbody Radiation and the Ultraviolet Catastrophe

The last few decades of the nineteenth century witnessed intense research activity in commercializing newly discovered electric lighting. This required obtaining a better understanding of the distributions of light emitted from various sources being considered. Artificial lighting is usually designed to mimic natural sunlight within the limitations of the underlying technology. Such lighting consists of a range of broadly distributed frequencies that form a continuous spectrum. Figure $10.1.8$ shows the wavelength distribution for sunlight. The most intense radiation is in the visible region, with the intensity dropping off rapidly for shorter wavelength ultraviolet (UV) light, and more slowly for longer wavelength infrared (IR) light.

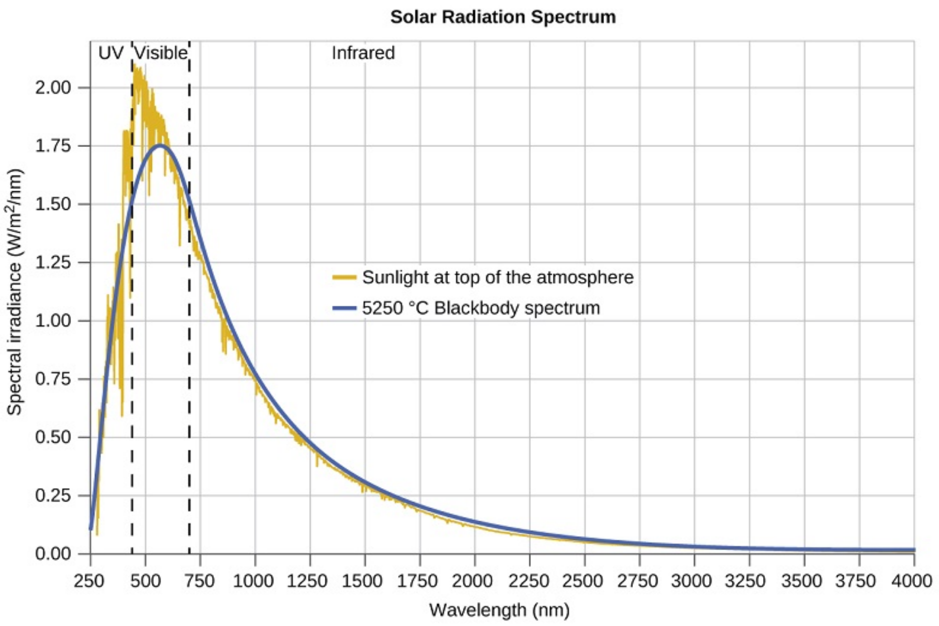


Figure 10.1.8: The spectral distribution (light intensity vs. wavelength) of sunlight reaches the Earth's atmosphere as UV light, visible light, and IR light. The unabsorbed sunlight at the top of the atmosphere has a distribution that approximately matches the theoretical distribution of a blackbody at 5250 °C, represented by the blue curve. (credit: modification of work by American Society for Testing and Materials (ASTM) Terrestrial Reference Spectra for Photovoltaic Performance Evaluation)

In Figure 10.1.8, the solar distribution is compared to a representative distribution, called a blackbody spectrum, that corresponds to a temperature of 5250 °C. The blackbody spectrum matches the solar spectrum quite well. A blackbody is a convenient, ideal emitter that approximates the behavior of many materials when heated. It is “ideal” in the same sense that an ideal gas is a convenient, simple representation of real gases that works well, provided that the pressure is not too high nor the temperature too low. A good approximation of a blackbody that can be used to observe blackbody radiation is a metal oven that can be heated to very high temperatures. The oven has a small hole allowing for the light being emitted within the oven to be observed with a spectrometer so that the wavelengths and their intensities can be measured. Figure 10.1.8 shows the resulting curves for some representative temperatures. Each distribution depends only on a single parameter: the temperature. The maxima in the blackbody curves, λ_{max}, shift to shorter wavelengths as the temperature increases, reflecting the observation that metals being heated to high temperatures begin to glow a darker red that becomes brighter as the temperature increases, eventually becoming white hot at very high temperatures as the intensities of all of the visible wavelengths become appreciable. This common observation was at the heart of the first paradox that showed the fundamental limitations of classical physics that we will examine.

Physicists derived mathematical expressions for the blackbody curves using well-accepted concepts from the theories of classical mechanics and classical electromagnetism. The theoretical expressions as functions of temperature fit the observed experimental blackbody curves well at longer wavelengths, but showed significant discrepancies at shorter wavelengths. Not only did the theoretical curves not show a peak, they absurdly showed the intensity becoming infinitely large as the wavelength became smaller, which would imply that everyday objects at room temperature should be emitting large amounts of UV light. This became known as the “ultraviolet catastrophe” because no one could find any problems with the theoretical treatment that could lead to such unrealistic short-wavelength behavior. Finally, around 1900, Max Planck derived a theoretical expression for blackbody radiation that fit the experimental observations exactly (within experimental error). Planck developed his theoretical treatment by extending the earlier work that had been based on the premise that the atoms composing the oven vibrated at increasing frequencies (or decreasing wavelengths) as the temperature increased, with these vibrations being the source of the emitted electromagnetic radiation. But where the earlier treatments had allowed the vibrating atoms to have any energy values obtained from a continuous set of energies (perfectly reasonable, according to classical physics), Planck found that by restricting the vibrational energies to discrete values for each frequency, he could derive an expression for blackbody radiation that correctly had the intensity dropping rapidly for the short wavelengths in the UV region.

$$E = nh\nu,\ n = 1, 2, 3,\ \ldots$$

The quantity h is a constant now known as Planck's constant, in his honor. Although Planck was pleased he had resolved the blackbody radiation paradox, he was disturbed that to do so, he needed to assume the vibrating atoms required quantized energies, which he was unable to explain. The value of Planck's constant is very small, 6.626×10^{-34} joule seconds (J s), which helps explain why energy quantization had not been observed previously in macroscopic phenomena.

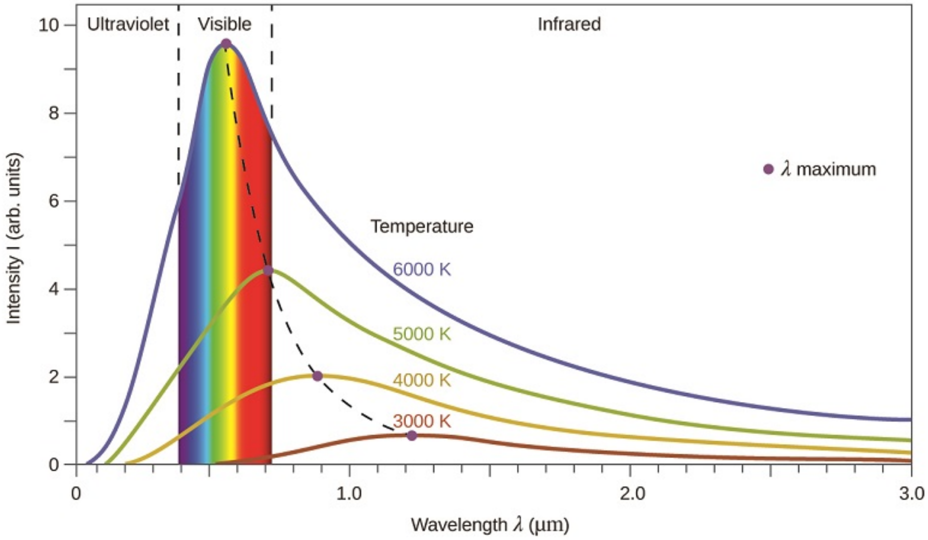


Figure 10.1.9: Blackbody spectral distribution curves are shown for some representative temperatures.

A graph is shown with a horizontal axis labeled, “Wavelength lambda (micrometers)” and a vertical axis labeled, “Intensity I (a r b. units).” The horizontal axis begins at 0 and extends to 3.0 with markings provided every 0.1 micrometer. Similarly, the vertical axis begins at 0 and extends to 10 with markings every 1 unit. Two vertical dashed lines are drawn. The first appears at about 0.39 micrometers and the second at about 0.72 micrometers. To the left of the first of these lines, the label, “Ultraviolet,” appears at the top of the graph. Between these lines, the label, “Visible,” appears at the top of the graph. To the right of the second of these lines, the label, “Infrared,” appears at the top of the graph. To the far right of the graph in open space a purple dot is placed which is labeled, “lambda maximum.” A “Temperature” label is located in a central region of the graph. A blue curve begins on the horizontal axis at about 0.05 micrometers. This curve increases steeply to a maximum value between the two vertical line segments of approximately 9.5 at about 0.55 micrometers. This curve decreases rapidly at first, then tapers off to reach a value of about 1.5 at the far right end of the graph. This blue curve is labeled 6000 K beneath the “Temperature” label. Curves are similarly drawn in green for 5000 K, orange for 4000 K, and red for 3000 K. As the temperature decreases, the height of the peak is lower and shifted right on the graph. The maximum value for the green curve is around 4.5 at 7.2 micrometers. This curve tapers at the right end of the graph to a value around 0.6. The maximum for the orange curve is around 2 at about 0.9 micrometers. This curve tapers at the right end of the graph to a value around 0.2. The maximum for the red curve is around 0.7 at about 1.2 micrometers. This curve tapers at the right end of the graph to a value around 0.1. The entire region under the blue curve that is between the two dashed lines, indicating the visible region, is shaded with vertical bands of color. The colors extending left to right across this region are violet, indigo, blue, green, yellow, orange, and red. A purple dot is placed at the peak of each of the four colored curves. These peaks are connected with a dashed curve.

The Photoelectric Effect

The next paradox in the classical theory to be resolved concerned the photoelectric effect (Figure 10.1.10). It had been observed that electrons could be ejected from the clean surface of a metal when light having a frequency greater than some threshold frequency was shone on it. Surprisingly, the kinetic energy of the ejected electrons did not depend on the brightness of the light, but increased with increasing frequency of the light. Since the electrons in the metal had a certain amount of binding energy keeping them there, the incident light needed to have more energy to free the electrons. According to classical wave theory, a wave's energy depends on its intensity (which depends on its amplitude), not its frequency. One part of these observations was that the number of electrons ejected within in a given time period was seen to increase as the brightness increased. In 1905, Albert Einstein was able to resolve the paradox by incorporating Planck's quantization findings into the discredited particle view of light (Einstein actually won his Nobel prize for this work, and not for his theories of relativity for which he is most famous).

Einstein argued that the quantized energies that Planck had postulated in his treatment of blackbody radiation could be applied to the light in the photoelectric effect so that the light striking the metal surface should not be viewed as a wave, but instead as a stream of particles (later called photons) whose energy depended on their frequency, according to Planck's formula, $E = h\nu$ (or, in terms of wavelength using $c = \nu\lambda$, $E = \frac{hc}{\lambda}$). Electrons were ejected when hit by photons having sufficient energy (a frequency greater than the threshold). The greater the frequency, the greater the kinetic energy imparted to the escaping electrons by the collisions. Einstein also argued that the light intensity did not depend on the amplitude of the incoming wave, but instead corresponded to the number of photons striking the surface within a given time period. This explains why the number of ejected electrons increased with increasing brightness, since the greater the number of incoming photons, the greater the likelihood that they would collide with some of the electrons.

With Einstein's findings, the nature of light took on a new air of mystery. Although many light phenomena could be explained either in terms of waves or particles, certain phenomena, such as the interference patterns obtained when light passed through a double slit, were completely contrary to a particle view of light, while other phenomena, such as the photoelectric effect, were completely contrary to a wave view of light. Somehow, at a deep fundamental level still not fully understood, light is both wavelike and particle-like. This is known as wave-particle duality.

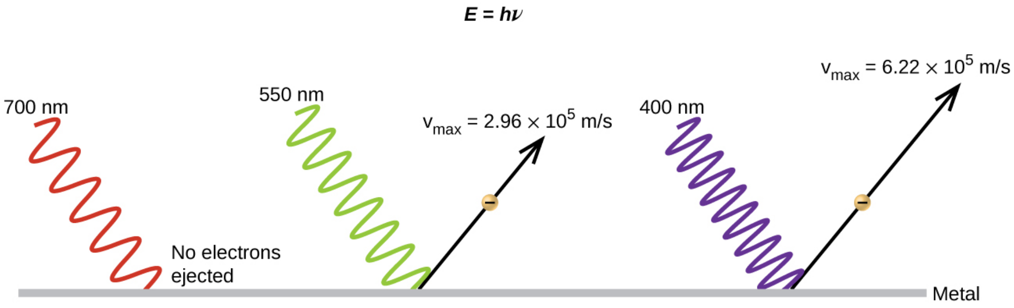


Figure 10.1.10: Photons with low frequencies do not have enough energy to cause electrons to be ejected via the photoelectric effect. For any frequency of light above the threshold frequency, the kinetic energy of an ejected electron will increase linearly with the energy of the incoming photon.

The figure includes three diagrams of waves approaching a flat, horizontal surface that is labeled, “Metal,” from an angle around 45 degrees above and to the left relative to the surface. At the top of the diagram at the center is the label, “E equals h nu.” At the left, a sinusoidal wave reaches the surface and stops. The portion of the diagram near the flat metal surface is labeled, “No electrons ejected,” and the wave is labeled, “700 n m.” To the right, a second similar, more compressed wave, which is labeled, “550 n m,” reaches the flat surface. This time, an arrow extends up and to the right at an angle of approximately 45 degrees. A tiny yellow circle with a negative sign in it is at the center of the arrow shaft. Above this arrow is the equation, “v subscript max equals 2.96 times 10 superscript 5 m per s.” To the far right, a third similar, even more compressed wave, which is labeled “400 n m” reaches the flat surface. This time, an arrow extends up and to the right at an angle of approximately 45 degrees. A tiny yellow circle with a negative sign in it is at the center of the arrow shaft. Above this arrow is the equation “v subscript max equals 6.22 times 10 superscript 5 m per s.”

✓ Example 10.1.2: Calculating the Energy of Radiation

When we see light from a neon sign, we are observing radiation from excited neon atoms. If this radiation has a wavelength of 640 nm, what is the energy of the photon being emitted?

Solution

We use the part of Planck's equation that includes the wavelength, λ, and convert units of nanometers to meters so that the units of λ and c are the same.

$$\begin{aligned} E &= \frac{hc}{\lambda} \\ &= \frac{(6.626 \times 10^{-34}\ \mathrm{J}\ \cancel{\mathrm{s}})(2.998 \times 10^{8}\ \mathrm{m}\ \cancel{\mathrm{s}}^{-1})}{(640\ \cancel{\mathrm{nm}})\left(\frac{1\ \mathrm{m}}{10^{9}\ \cancel{\mathrm{nm}}}\right)} \\ &= 3.10 \times 10^{-19}\ \mathrm{J} \end{aligned}$$

? Exercise 10.1.2

The microwaves in an oven are of a specific frequency that will heat the water molecules contained in food. (This is why most plastics and glass do not become hot in a microwave oven-they do not contain water molecules.) This frequency is about 3 × 10^9 Hz. What is the energy of one photon in these microwaves?

Answer

2×10^{-24} J

✓ Example 10.1.3: Photoelectric Effect

Identify which of the following statements are false and, where necessary, change the italicized word or phrase to make them true, consistent with Einstein's explanation of the photoelectric effect.

a. Increasing the brightness of incoming light *increases* the kinetic energy of the ejected electrons.
b. Increasing the *wavelength* of incoming light increases the kinetic energy of the ejected electrons.
c. Increasing the brightness of incoming light *increases* the number of ejected electrons.
d. Increasing the *frequency* of incoming light can increase the number of ejected electrons.

Solution

a. False. Increasing the brightness of incoming light *has no effect* on the kinetic energy of the ejected electrons. Only energy, not the number or amplitude, of the photons influences the kinetic energy of the electrons.
b. False. Increasing the *frequency* of incoming light increases the kinetic energy of the ejected electrons. Frequency is proportional to energy and inversely proportional to wavelength. Frequencies above the threshold value transfer the excess energy into the kinetic energy of the electrons.
c. True. Because the number of collisions with photons increases with brighter light, the number of ejected electrons increases.
d. True with regard to the threshold energy binding the electrons to the metal. Below this threshold, electrons are not emitted and above it they are. Once over the threshold value, further increasing the frequency does not increase the number of ejected electrons

? Exercise 10.1.3

Calculate the threshold energy in kJ/mol of electrons in aluminum, given that the lowest frequency photon for which the photoelectric effect is observed is $9.87 \times 10^{14}\ Hz$.

Answer

$3.94\ kJ/mol$

Summary

Light and other forms of electromagnetic radiation move through a vacuum with a constant speed, *c*, of 2.998×10^{8} m s^{-1}. This radiation shows wavelike behavior, which can be characterized by a frequency, *ν*, and a wavelength, *λ*, such that $c = \lambda\nu$. Light is an example of a travelling wave. Other important wave phenomena include standing waves, periodic oscillations, and vibrations. Standing waves exhibit quantization, since their wavelengths are limited to discrete integer multiples of some characteristic lengths. Electromagnetic radiation that passes through two closely spaced narrow slits having dimensions roughly similar to the wavelength will show an interference pattern that is a result of constructive and destructive interference of the waves. Electromagnetic radiation also demonstrates properties of particles called photons. The energy of a photon is related to the frequency (or alternatively, the wavelength) of the radiation as $E = h\nu$ (or $E = \frac{hc}{\lambda}$), where *h* is Planck's constant. That light demonstrates both wavelike and particle-like behavior is known as wave-particle duality. All forms of electromagnetic radiation share these properties, although various forms including X-rays, visible light, microwaves, and radio waves interact differently with matter and have very different practical applications. Electromagnetic radiation can be generated by exciting matter to higher energies, such as by heating it. The emitted light can be either continuous (incandescent sources like the sun) or discrete (from specific types of excited atoms). Continuous spectra often have distributions that can be approximated as blackbody radiation at some appropriate temperature.

10.2: Content in Context

You learned about different types of radiation, electromagnetic waves, frequency, and energy. In this section, we will focus on UV radiation, its harmful effects, the stratospheric ozone layer, and how to protect ourselves from the dangers of UV rays.

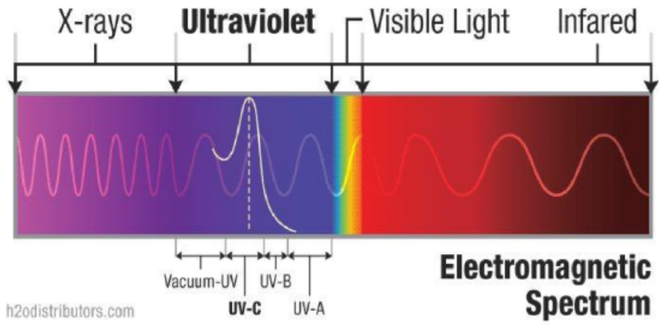
10.2.1: UVA, UVB, UVC and Its' Effect

Learning Objectives

- Differentiate the three regions of UV radiation based on their energy, impact on human molecules, and wavelength.
- Explain how UV radiation impacts the skin leading to sunburn and skin cancer
- Understand the link between UV exposure and eye damage like cataracts

UV Radiation

Solar energy (sunlight) contains light we can see, and some we cannot. *Visible light* has wavelengths of 750 to 400 nm. *Ultraviolet (UV) light* has shorter wavelengths, cannot be seen, and has higher energy. Infrared (IR) radiation is the major source of heat for Earth. Though UV is a fraction of sunlight, it can be damaging to living organisms. All of these are forms of energy in the *electromagnetic spectrum*.



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Just as visible light components have names (red, orange, yellow, green, blue, indigo, violet), so do the types of UV light: UV-A, UV-B, UV-C and vacuum-UV. UV-A has lowest energy and is least damaging; *UV-A* is also called "black light." *UV-B* and *UV-C* have higher energies and can cause break bonds of molecules, causing changes in DNA and thus skin cancers.

The majority of UV-B is absorbed by ozone in the stratosphere. Though UV-C is most damaging, it is totally absorbed by oxygen and ozone. In recent years, depletion of the ozone layer has allowed more UV light to reach us, resulting in more cases of skin cancers. Consequently, we have become aware of the need to protect ourselves from UV light.

What protects us from UV light? One strategy would be to avoid exposure to any type of sunlight. Since we cannot avoid sunlight while outdoors, we can physically or chemically block the sun. A wide variety of commercial sunscreens are available with *sun protection factors (SPF)* ranging from SPF 2 to SPF 100. These lotions contain organic molecules that absorb UV light. Some materials, such as glass and plastic also absorb UV light, while still allowing visible light through.

UV light type	Wavelength	Relative Energy	Comments
UV-A	320 – 400 nm	lowest energy	reaches Earth in greatest amount
UV-B	280 – 320 nm	higher energy than UV-A, but less than UV-C	most is absorbed by ozone
UV-C	200 – 280 nm	highest energy	absorbed by ozone and oxygen

Health Effects of UV Radiation

Here is an overview of the major health problems linked to overexposure to UV radiation. Understanding these risks and taking sensible precautions will help you enjoy the sun while reducing your chances of sun-related health problems.

Skin Cancer

Each year, more new cases of skin cancer are diagnosed in the U.S. than new cases of breast, prostate, lung, and colon cancer combined. One in five Americans will develop skin cancer in their lifetime. One American dies from skin cancer every hour.

Unprotected exposure to UV radiation is the most preventable risk factor for skin cancer.

Melanoma

Melanoma, the most serious form of skin cancer, is now one of the most common cancers among adolescents and young adults ages 15-29. While melanoma accounts for about three percent of skin cancer cases, it causes more than 75 percent of skin cancer deaths. UV exposure and sunburns, particularly during childhood, are risk factors for the disease. Not all melanomas are exclusively sun-related—other possible influences include genetic factors and immune system deficiencies.

Nonmelanoma Skin Cancers

Non-melanoma skin cancers are less deadly than melanomas. Nevertheless, they can spread if left untreated, causing disfigurement and more serious health problems. There are two primary types of non-melanoma skin cancers: basal cell and squamous cell carcinomas. If caught and treated early, these two cancers are rarely fatal.

- **Basal cell carcinomas** are the most common type of skin cancer tumors. They usually appear as small, fleshy bumps or nodules on the head and neck, but can occur on other skin areas. Basal cell carcinoma grows slowly, and it rarely spreads to other parts of the body. It can, however, penetrate to the bone and cause considerable damage.
- **Squamous cell carcinomas** are tumors that may appear as nodules or as red, scaly patches. This cancer can develop into large masses, and unlike basal cell carcinoma, it can spread to other parts of the body.

Premature Aging and Other Skin Damage

Other UV-related skin disorders include actinic keratoses and premature aging of the skin. Actinic keratoses are skin growths that occur on body areas exposed to the sun. The face, hands, forearms, and the “V” of the neck are especially susceptible to this type of lesion. Although premalignant, actinic keratoses are a risk factor for squamous cell carcinoma. Look for raised, reddish, rough-textured growths and seek prompt medical attention if you discover them.

Chronic exposure to the sun also causes premature aging, which over time can make the skin become thick, wrinkled, and leathery. Since it occurs gradually, often manifesting itself many years after the majority of a person’s sun exposure, premature aging is often regarded as an unavoidable, normal part of growing older. However, up to 90 percent of the visible skin changes commonly attributed to aging are caused by the sun. With proper protection from UV radiation, most premature aging of the skin can be avoided.

Cataracts and Other Eye Damage

Cataracts are a form of eye damage in which a loss of transparency in the lens of the eye clouds vision. If left untreated, cataracts can lead to blindness. Research has shown that UV radiation increases the likelihood of certain cataracts. Although curable with modern eye surgery, cataracts diminish the eyesight of millions of Americans and cost billions of dollars in medical care each year.

Other kinds of eye damage include pterygium (tissue growth that can block vision), skin cancer around the eyes, and degeneration of the macula (the part of the retina where visual perception is most acute). All of these problems can be lessened with proper eye protection. Look for sunglasses, glasses or contact lenses if you wear them, that offer 99 to 100 percent UV protection.

Immune System Suppression

Scientists have found that overexposure to UV radiation may suppress proper functioning of the body’s immune system and the skin’s natural defenses. For example, the skin normally mounts a defense against foreign invaders such as cancers and infections. But overexposure to UV radiation can weaken the immune system, reducing the skin’s ability to protect against these invaders.

Original Source for 'Health Effects of UV Radiation': https://www.epa.gov/sunsafety/health-effects-uv-radiation

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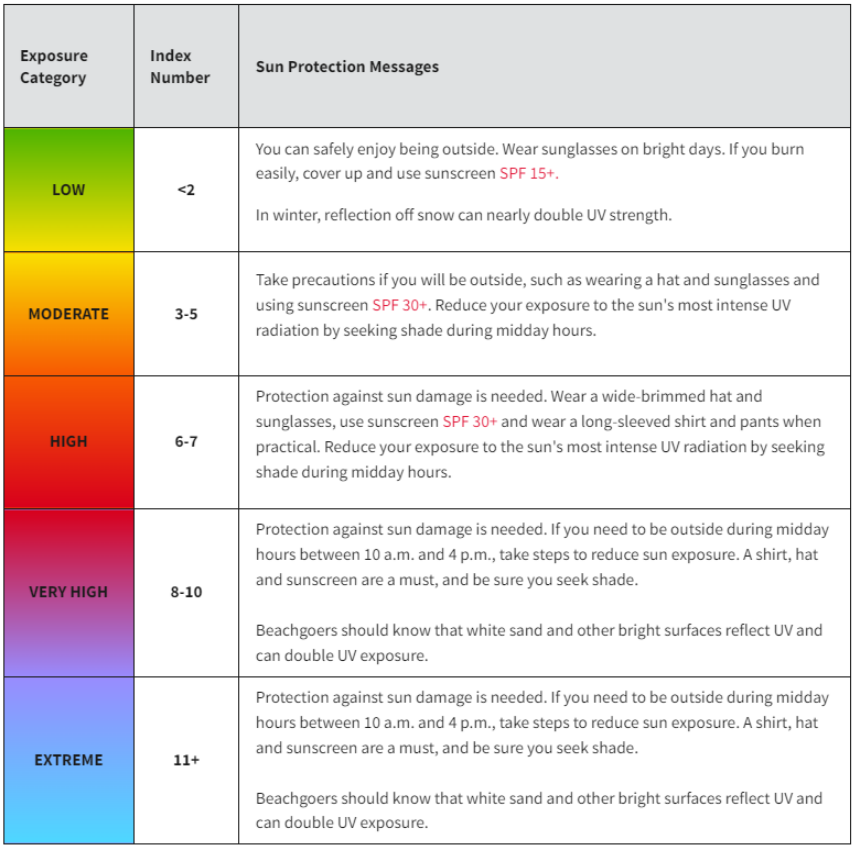
10.2.2: UV Index

Learning Objectives

- Identify the different UV Index levels and the associated risk of sunburn.
- Demonstrate how to use the UV Index to plan outdoor activities and choose appropriate sun protection measures.
- Discuss how factors like time of day, location, cloud cover, and elevation can impact the UV Index.

The UV Index provides a daily forecast of the expected intensity of ultraviolet (UV) radiation from the sun. Some exposure to sunlight is enjoyable. However, too much sun can be dangerous. Overexposure to the sun's ultraviolet radiation can cause immediate damage, such as sunburn, and long-term problems, such as skin cancer and cataracts.

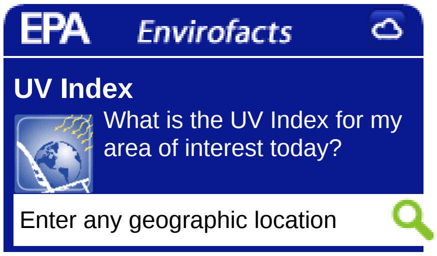
It is important to remember that people of all skin types need to be protected from overexposure to the sun. Overexposure to UV radiation poses the risk of serious health effects for everyone, but not everyone is equally at risk. For example, you may be at greater risk of contracting skin cancer if your skin always burns; if you have blond or red hair; or blue, green, or gray eyes. Other factors indicating an increased risk of skin cancer include: a history of blistering sunburns in early childhood, the presence of many moles, or a family history of skin cancer. However, it is a good idea to remember that all people, no matter what skin type, are equally at risk of eye damage.



Exposure Category	Index Number	Sun Protection Messages
LOW	<2	You can safely enjoy being outside. Wear sunglasses on bright days. If you burn easily, cover up and use sunscreen SPF 15+. In winter, reflection off snow can nearly double UV strength.
MODERATE	3-5	Take precautions if you will be outside, such as wearing a hat and sunglasses and using sunscreen SPF 30+. Reduce your exposure to the sun's most intense UV radiation by seeking shade during midday hours.
HIGH	6-7	Protection against sun damage is needed. Wear a wide-brimmed hat and sunglasses, use sunscreen SPF 30+ and wear a long-sleeved shirt and pants when practical. Reduce your exposure to the sun's most intense UV radiation by seeking shade during midday hours.
VERY HIGH	8-10	Protection against sun damage is needed. If you need to be outside during midday hours between 10 a.m. and 4 p.m., take steps to reduce sun exposure. A shirt, hat and sunscreen are a must, and be sure you seek shade. Beachgoers should know that white sand and other bright surfaces reflect UV and can double UV exposure.
EXTREME	11+	Protection against sun damage is needed. If you need to be outside during midday hours between 10 a.m. and 4 p.m., take steps to reduce sun exposure. A shirt, hat and sunscreen are a must, and be sure you seek shade. Beachgoers should know that white sand and other bright surfaces reflect UV and can double UV exposure.

For more information on UV Index, click on the **A Guide To The UV Index**.

UV Index Widget:



EPA Envirofacts

UV Index

What is the UV Index for my area of interest today?

Enter any geographic location

Original Source: https://www.epa.gov/enviro/uv-index-description

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10.2.3: Stratospheric Ozone and Ozone Depletion

Learning Objectives

- Explain where the ozone layer is located in the atmosphere (stratosphere).
- Describe the natural formation of ozone through UV radiation interaction with oxygen molecules.
- Describe the depletion of the ozone layer.
- Explain how chlorine and bromine atoms react with ozone that leads to the depletion of the ozone layer.
- Identify chlorofluorocarbons (CFCs) as the primary ozone-depleting substance.
- Policies to Reduce Ozone Destruction

The earth's stratospheric ozone layer plays a critical role in absorbing ultraviolet radiation emitted by the sun. In the last thirty years, it has been discovered that stratospheric ozone is depleting as a result of anthropogenic pollutants. There are a number of chemical reactions that can deplete stratospheric ozone; however, some of the most significant of these involves the catalytic destruction of ozone by halogen radicals such as chlorine and bromine.

Introduction

The atmosphere of the Earth is divided into five layers. In order of closest and thickest to farthest and thinnest the layers are listed as follows: troposphere, stratosphere, mesosphere, thermosphere and exosphere. The majority of the ozone in the atmosphere resides in the stratosphere, which extends from six miles above the Earth’s surface to 31 miles. Humans rely heavily on the absorption of ultraviolet B rays by the ozone layer because UV-B radiation causes skin cancer and can lead to genetic damage. The ozone layer has historically protected the Earth from the harmful UV rays, although in recent decades this protection has diminished due to stratospheric ozone depletion.

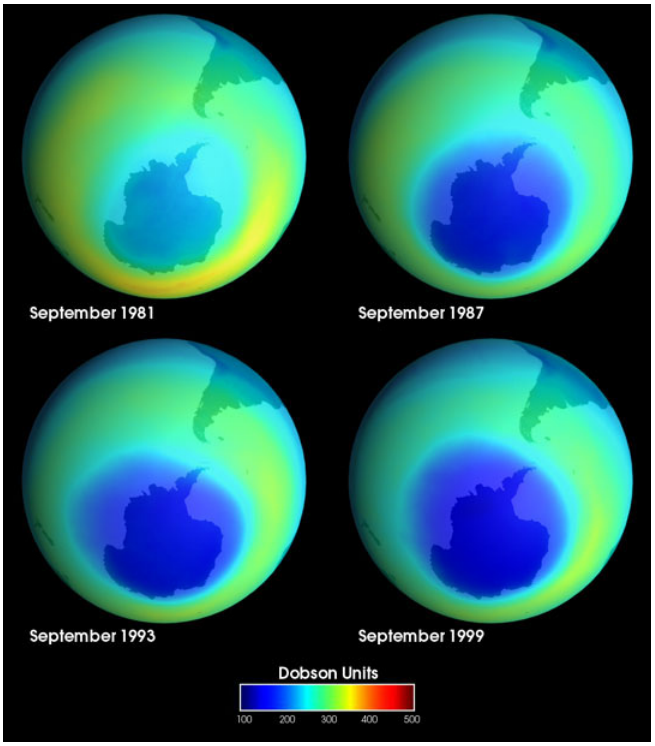

Figure courtesy of NASA.

Ozone depletion is largely a result of man-made substances. Humans have introduced gases and chemicals into the atmosphere that have rapidly depleted the ozone layer in the last century. This depletion makes humans more vulnerable to the UV-B rays which are known to cause skin cancer as well as other genetic deformities. The possibility of ozone depletion was first introduced by scientists in the late 1960's as dreams of super sonic transport began to become a reality. Scientists had long been aware that nitric oxide (NO) can catalytically react with ozone (O_3) to produce O_2 molecules; however, NO molecules produced at ground level have a half life far too short to make it into the stratosphere. It was not until the advent of commercial super sonic jets (which fly in the stratosphere and at an altitude much higher then conventional jets) that the potential for NO to react with stratospheric ozone became a possibility. The threat of ozone depletion from commercial super sonic transport was so great that it is often cited as the main reason why the US federal government pulled support for its development in 1971. Fear of ozone depletion was abated until 1974 when Sherwood Rowland and Mario Molina discovered that chlorofluorocarbons could be photolyzed by high energy photons in the stratosphere. They discovered that this process could releasing chlorine radicals that would catalytically react with O_3 and destroy the molecule. This process is called the Rowland-Molina theory of O_3 depletion.

The Chapman Cycle

The stratosphere is in a constant cycle with oxygen molecules and their interaction with ultraviolet rays. This process is considered a cycle because of its constant conversion between different molecules of oxygen. The ozone layer is created when ultraviolet rays react with oxygen molecules (O_2) to create ozone (O_3) and atomic oxygen (O). This process is called the *Chapman cycle*.

Step 1: An oxygen molecules is photolyzed by solar radiation, creating two oxygen radicals:

$$h\nu + O_2 \rightarrow 2O^{\cdot}$$

Step 2: Oxygen radicals then react with molecular oxygen to produce ozone:

$$O_2 + O^{\cdot} \rightarrow O_3$$

Step 3: Ozone then reacts with an additional oxygen radical to form molecular oxygen:

$$O_3 + O^{\cdot} \rightarrow 2O_2$$

Step 4: Ozone can also be recycled into molecular oxygen by reacting with a photon:

$$O_3 + h\nu \rightarrow O_2 + O^{\cdot}$$

It is important to keep in mind that ozone is constantly being created and destroyed by the Chapman cycle and that these reactions are natural processes, which have been taking place for millions of years. Because of this, the thickness the ozone layer at any particular time can vary greatly. It is also important to know that O_2 is constantly being introduced into the atmosphere through photosynthesis, so the ozone layer has the capability of regenerating itself.

Chemistry of Ozone Depletion

CFC molecules are made up of chlorine, fluorine and carbon atoms and are extremely stable. This extreme stability allows CFC's to slowly make their way into the stratosphere (most molecules decompose before they can cross into the stratosphere from the troposphere). This prolonged life in the atmosphere allows them to reach great altitudes where photons are more energetic. When the CFC's come into contact with these high energy photons, their individual components are freed from the whole. The following reaction displays how Cl atoms have an ozone destroying cycle:

$$Cl + O_3 \rightarrow ClO + O_2 \tag{step 1}$$

$$ClO + O^{\cdot} \rightarrow Cl + O_2 \tag{step 2}$$

$$O_3 + O^{\cdot} \rightarrow 2O_2 \tag{Overall reaction}$$

Chlorine is able to destroy so much of the ozone because it acts as a catalyst. Chlorine initiates the breakdown of ozone and combines with a freed oxygen to create two oxygen molecules. After each reaction, chlorine begins the destructive cycle again with another ozone molecule. One chlorine atom can thereby destroy thousands of ozone molecules. Because ozone molecules are being broken down they are unable to absorb any ultraviolet light so we experience more intense UV radiation at the earths surface.

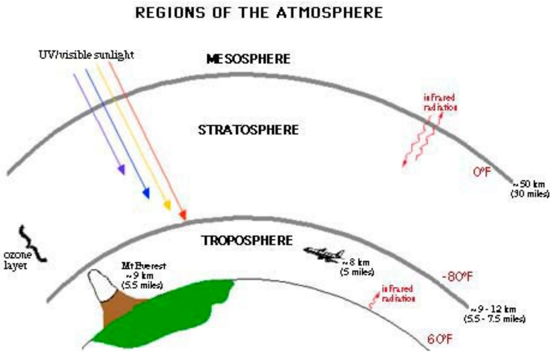

Figure courtesy of NOAA.

From 1985 to 1988, researchers studying atmospheric properties over the south pole continually noticed significantly reduced concentrations of ozone directly over the continent of Antarctica. For three years it was assumed that the ozone data was incorrect and was due to some type of instrument malfunction. In 1988, researchers finally realized their error and concluded that an enormous hole in the ozone layer had indeed developed over Antarctica. Examination of NASA satellite data later showed that the hole had begun to develop in the mid 1970's.

The ozone hole over Antarctica is formed by a slew of unique atmospheric conditions over the continent that combine to create an ideal environment for ozone destruction.

- Because Antarctica is surrounded by water, winds over the continent blow in a unique clockwise direction creating a so called "polar vortex" that effectively contains a single static air mass over the continent. As a result, air over Antarctica does not mix with air in the rest of the earth's atmosphere.
- Antarctica has the coldest winter temperatures on earth, often reaching -110 F. These chilling temperatures result in the formation of polar stratospheric clouds (PSC's) which are a conglomeration of frozen H_2O and HNO_3. Due to their extremely cold temperatures, PSC's form an electrostatic attraction with CFC molecules as well as other halogenated compounds

As spring comes to Antarctica, the PSC's melt in the stratosphere and release all of the halogenated compounds that were previously absorbed to the cloud. In the antarctic summer, high energy photons are able to photolyze the halogenated compounds, freeing halogen radicals that then catalytically destroy O_3. Because Antarctica is constantly surrounded by a polar vortex, radical halogens are not able to be diluted over the entire globe. The ozone hole develops as result of this process.

Resent research suggests that the strength of the polar vortex from any given year is directly correlated to the size of the ozone hole. In years with a strong polar vortex, the ozone hole is seen to expand in diameter, whereas in years with a weaker polar vortex, the ozone hole is noted to shrink

Policies to Reduce Ozone Destruction

One success story in reducing pollutants that harm the atmosphere concerns ozone-destroying chemicals. In 1973, scientists calculated that CFCs could reach the stratosphere and break apart. This would release chlorine atoms, which would then destroy ozone. Based only on their calculations, the United States and most Scandinavian countries banned CFCs in spray cans in 1978. More confirmation that CFCs break down ozone was needed before more was done to reduce production of ozone-destroying chemicals. In 1985, members of the British Antarctic Survey reported that a 50% reduction in the ozone layer had been found over Antarctica in the previous three springs.

Two years after the British Antarctic Survey report, the “Montreal Protocol on Substances that Deplete the Ozone Layer” was ratified by nations all over the world. The **Montreal Protocol**controls the production and consumption of 96 chemicals that damage the ozone layer (Figure 10.2.310.2.3). CFCs have been mostly phased out since 1995, although they were used in developing nations until 2010. Some of the less hazardous substances will not be phased out until 2030. The Protocol also requires that wealthier nations donate money to develop technologies that will replace these chemicals.

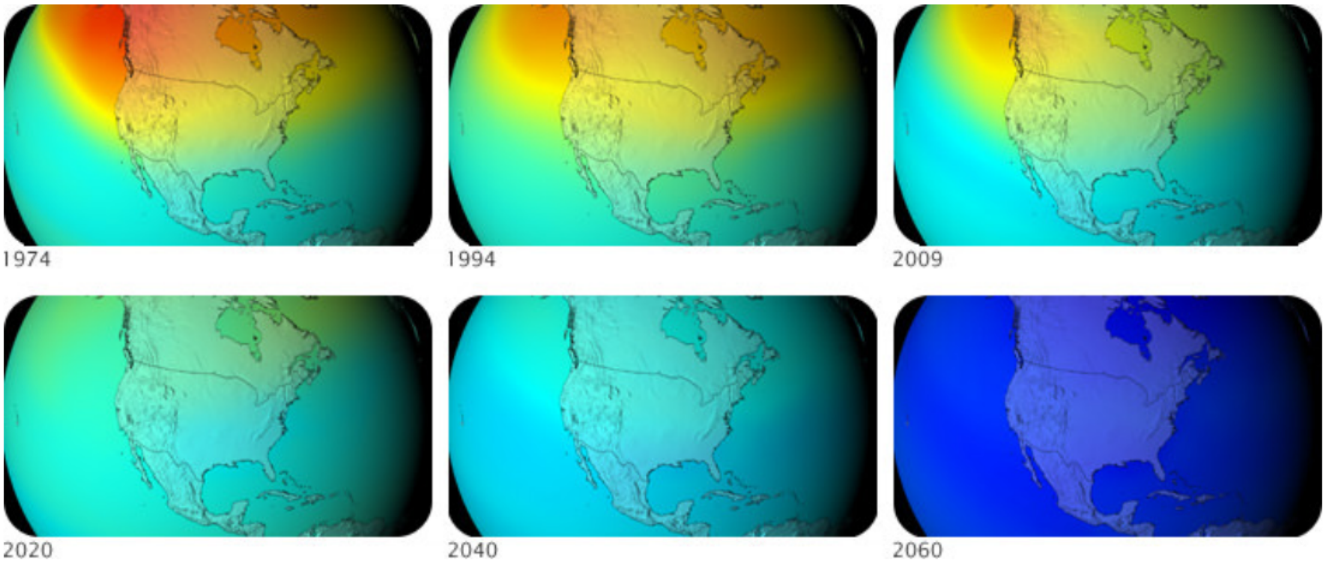

Figure 10.2.310.2.3. Ozone levels over North America decreased between 1974 and 2009. Models of the future predict what ozone levels would have been if CFCs were not being phased out. Warmer colors indicate more ozone.

Because CFCs take many years to reach the stratosphere and can survive there a long time before they break down, the ozone hole will probably continue to grow for some time before it begins to shrink. The ozone layer will reach the same levels it had before 1980 around 2068 and 1950 levels in one or two centuries.

Ozone Depleting Substances

The following substances are listed as ozone depleting substances under Title VI of the United State Clean Air Act:

Table 10.2.3.1: Ozone Depleting Substances And Their Ozone-Depletion Potential. Taken directly from the Clean Air Act, as of June 2010.

Substance	Ozone- depletion potential
chlorofluorocarbon-11 (CFC–11)	1.0
chlorofluorocarbon-12 (CFC–12)	1.0
chlorofluorocarbon-13 (CFC–13)	1.0
chlorofluorocarbon-111 (CFC–111)	1.0
chlorofluorocarbon-112 (CFC–112)	1.0
chlorofluorocarbon-113 (CFC–113)	0.8
chlorofluorocarbon-114 (CFC–114)	1.0
chlorofluorocarbon-115 (CFC–115)	0.6
chlorofluorocarbon-211 (CFC–211)	1.0
chlorofluorocarbon-212 (CFC–212)	1.0
chlorofluorocarbon-213 (CFC–213)	1.0
chlorofluorocarbon-214 (CFC–214)	1.0
chlorofluorocarbon-215 (CFC–215)	1.0
chlorofluorocarbon-216 (CFC–216)	1.0
chlorofluorocarbon-217 (CFC–217)	1.0
halon-1211	3.0

Substance	Ozone- depletion potential
halon-1301	10.0
halon-2402	6.0
carbon tetrachloride	1.1
methyl chloroform	0.1
hydrochlorofluorocarbon-22 (HCFC–22)	0.05
hydrochlorofluorocarbon-123 (HCFC–123)	0.02
hydrochlorofluorocarbon-124 (HCFC–124)	0.02
hydrochlorofluorocarbon-141(b) (HCFC–141(b))	0.1
hydrochlorofluorocarbon-142(b) (HCFC–142(b))	0.06

Summary

- Any disruption of the balance that results in a higher rate of ozone destruction than ozone creation would result in depletion of ozone.
- Some compounds that release chlorine or bromine when they are exposed to intense UV light in the stratosphere contribute to ozone depletion, and are called ozone-depleting substances (ODS)
- The Montreal Protocol is an international agreement that committed all parties (signatory nations) to a schedule for phasing out the production and use of CFCs and other substances known to be harmful to the ozone layer.

Contributors and Attributions

- Libretext: Supplemental Module, Physical and Theoretical Chemistry
- Marisa Alviar-Agnew (Sacramento City College)
- 'Policies to Reduce Ozone Destruction' is from Bio LibreText

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10.2.4: Sunscreen

Learning Objectives

- Learn the importance of using sunscreen to protect skin from the sun's rays. .
- Understand how physical and chemical sunscreen ingredients protect the skin from UV rays.

Exposure to Ultraviolet Radiation

Of all the types of ionizing radiation, people come into more contact with ultraviolet radiation. Sources of this radiation would include the sun, ultraviolet lights, and tanning beds. The sun produces three different forms of ultraviolet rays (UVA, UVB and UVC). Fortunately, the deadliest of these three rays (UVC) never reaches the earth's surface and is absorbed by Earths' ozone layer. However, both UVA and UVB reach the Earth's surface to affect living systems.

UVA is produced naturally by the sun and artificially by lamps and tanning beds. For years, sunbathers believed that UVA provided only a youthful glow. Now, the American Academy of Dermatologists has condemned the use of tanning beds. Scientific research has correlated certain types of skin cancers (mainly basal and squamous cell) with the use of these devices. Besides skin cancer, tanning beds cause premature aging and damage to your eyes. If a tanning bed is not properly sanitized between uses, skin diseases can be transmitted from one consumer to another.

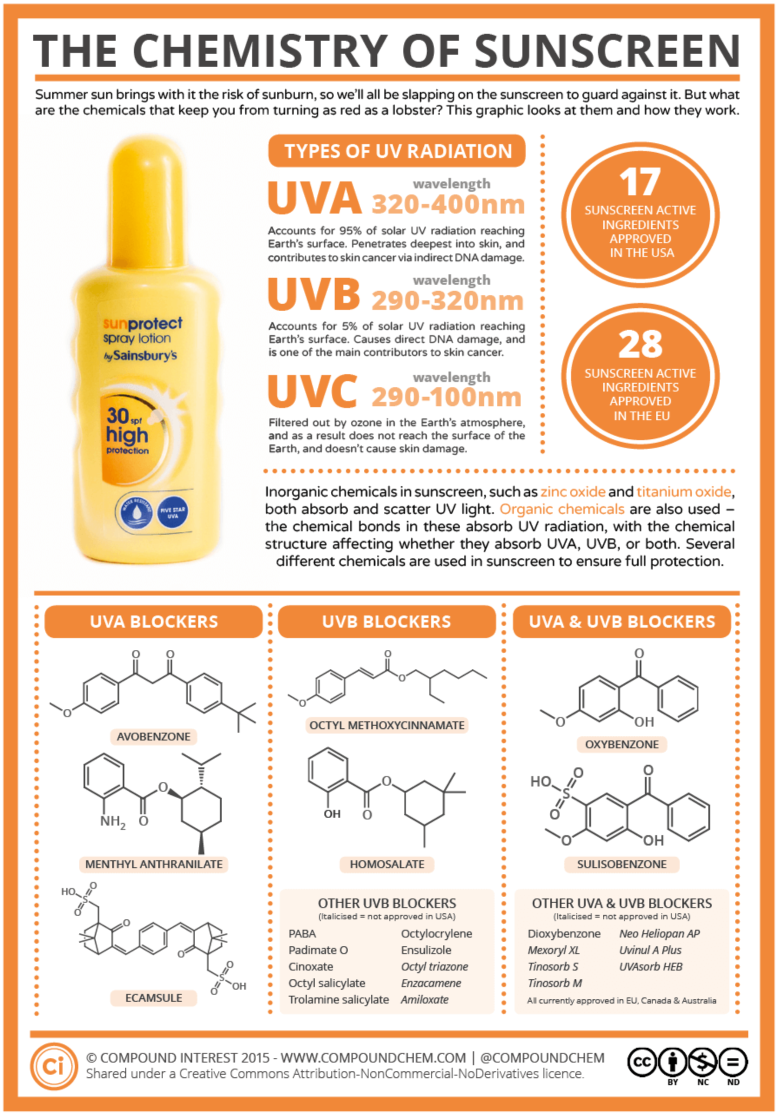
Sunblocks and Sunscreens

UVB radiation in sunlight allows the skin to produce vitamin D. This vitamin prevents bone disorders like rickets and osteoporosis (brittle bone disease). The American Academy of Dermatology suggests vitamin D be obtained through foods or nutritional supplements. Excessive exposure to UV can be damaging and the pigment **melanin**, deposited in cells at the base of the epidermis, helps to protect the underlying layers of the skin from this damage. Melanin also colors the skin and variations in the amount of melanin produces colors from pale yellow to black. The darker the skin tone, the more melanin one has, and the less likely skin cancer will occur.

Figure 10.2.4.3: The author's great, great- grandmother was diagnosed with skin cancer on her nose during the 1950's.

Excess exposure to the sun can cause sunburn. This is common in humans, but light skinned animals like cats and pigs can also be sunburned, especially on the ears. Skin cancer can also result from excessive exposure to the sun. As holes in the ozone layer increase exposure to the sun’s UV rays, so too does the rate of skin cancer in humans and animals.

Sunscreens and sunblocks are designed to protect skin from ultraviolet rays. *Sunblocks* contain inorganic ingredients like zinc oxide or titanium dioxide. These chemicals act as UV filters by reflecting the sun's UV rays. Sunblocks can have grainy textures due to the inorganic components. The thick nature of a sunblock can make it difficult to spread evenly on the skin. *Sunscreens* contain organic compounds like oxybenzone, avobenzone, homosalate, and octinoxate (Figure 10.2.4.4). By absorbing ultraviolet rays, these compounds decompose and give off heat. Sunscreens apply smoother than sunblocks. Often, manufacturers will combine sunscreen and sunblock ingredients to make their products.



THE CHEMISTRY OF SUNSCREEN

Summer sun brings with it the risk of sunburn, so we'll all be slapping on the sunscreen to guard against it. But what are the chemicals that keep you from turning as red as a lobster? This graphic looks at them and how they work.

TYPES OF UV RADIATION

UVA wavelength 320-400nm

Accounts for 95% of solar UV radiation reaching Earth's surface. Penetrates deepest into skin, and contributes to skin cancer via indirect DNA damage.

UVB wavelength 290-320nm

Accounts for 5% of solar UV radiation reaching Earth's surface. Causes direct DNA damage, and is one of the main contributors to skin cancer.

UVC wavelength 290-100nm

Filtered out by ozone in the Earth's atmosphere, and as a result does not reach the surface of the Earth, and doesn't cause skin damage.

17 SUNSCREEN ACTIVE INGREDIENTS APPROVED IN THE USA

28 SUNSCREEN ACTIVE INGREDIENTS APPROVED IN THE EU

Inorganic chemicals in sunscreen, such as zinc oxide and titanium oxide, both absorb and scatter UV light. Organic chemicals are also used – the chemical bonds in these absorb UV radiation, with the chemical structure affecting whether they absorb UVA, UVB, or both. Several different chemicals are used in sunscreen to ensure full protection.

UVA BLOCKERS

UVB BLOCKERS

UVA & UVB BLOCKERS

OTHER UVB BLOCKERS

(Italicised = not approved in USA)

PABA
Padimate O
Cinoxate
Octyl salicylate
Trolamine salicylate
Octylocrylene
Ensulizole
Octyl triazone
Enzacamene
Amiloxate

OTHER UVA & UVB BLOCKERS

(Italicised = not approved in USA)

Dioxybenzone
Mexoryl XL
Tinosorb S
Tinosorb M
Neo Heliopan AP
Uvinul A Plus
UVAsorb HEB

All currently approved in EU, Canada & Australia

Figure 10.2.4.4: Effect of Light on skin

Sun protection factor (SPF) measures a product's protection from UVB rays. SPF does not quantify protection from UVA radiation. The American Academy of Dermatology recommends you select a sunscreen or sunblock with a minimum SPF factor of 30. This SPF value means a lotion can filter out 97% of UVB rays. Moving to a SPF of 50 will only filter out 1 more percentage of UVB rays. Increasing a SPF factor past 30 really does very little in shielding skin from UVB radiation.

Products that protect skin from UVA must be labeled as being broad spectrum. Sunblocks provide UVA and UVB protection, but sunscreens can vary on what they can screen. Every two hours, sunscreens/sunblocks should be reapplied to the skin. Sweating and swimming can remove sunscreen/sunblock products. No sun products are waterproof, but some are labeled as being water resistant. Consumers are encouraged to reapply these products every two hours as well.

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CHAPTER OVERVIEW

11: Climate and The Environment

11.1: Carbon Cycle

Learning Objectives

- Define the carbon cycle and explain its importance for life on Earth.
- Identify the different carbon reservoirs (atmosphere, oceans, land, living organisms)
- Identify human activities that contribute to increased carbon dioxide levels in the atmosphere (burning fossil fuels, deforestation).

Carbon, just like all other elements, cycles through the environment and is constantly in the process of changing forms and locations. In this section, as in many other pieces of scientific literature, we will periodically refer to carbon by its chemical symbol, C. There is no new carbon in the world, rather, all carbon is continuously recycled from one form to another. All plants, animals (including humans!), fungi, bacteria, and archaea are made of mostly carbon-based molecules such as lipids, carbohydrates, proteins, and nucleic acids. Carbon is also prevalent in soils, rocks and sediments, water bodies (dissolved), and the atmosphere. These locations where carbon resides are known as pools or reservoirs, and the processes that move carbon from one location to another are called fluxes. Figure 7.1.1 shows a simplified version of the global carbon cycle.

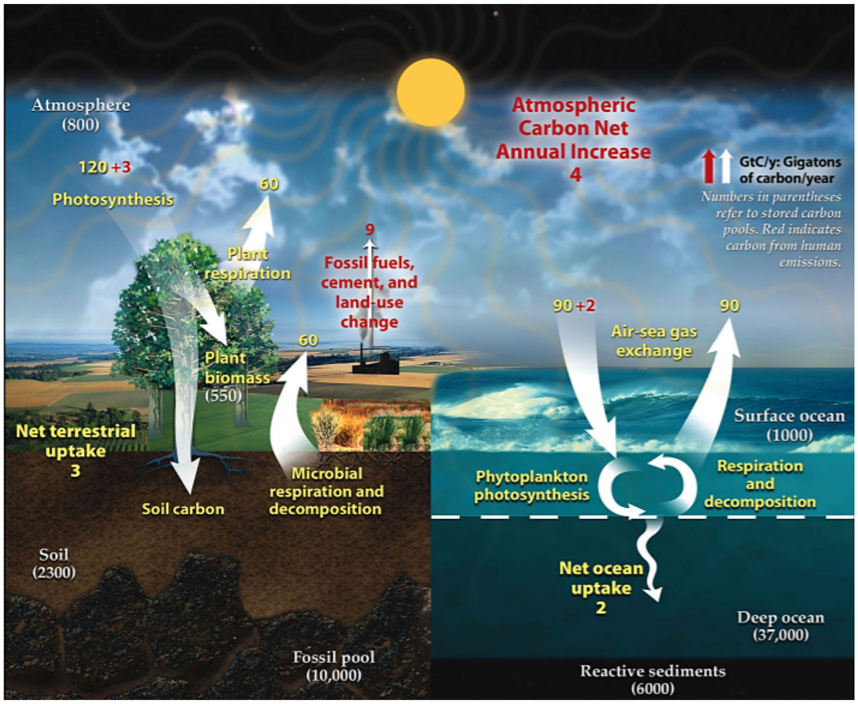


Figure 11.1.1: A simplified carbon cycle. Diagram adapted from U.S. DOE, Biological and Environmental Research Information System.

Some reservoirs hold on to carbon for only a short time. **Aerobic** (oxygen-using) organisms convert carbohydrates created by other organisms into **carbon dioxide** (CO_2) almost instantaneously, which they exhale into the atmosphere. When considering the flux of respiration, living organisms are the **source** of carbon, and the atmosphere is the **sink**. The carbon stays in the reservoir of living organisms for a relatively short time, depending on their life span, from hours and days to years and decades. In contrast, the **residence time** of carbon in the fossil pool is dramatically different. Fossil fuels form over a course of 300-400 million years, forming from ancient plants and animals that decomposed slowly under very specific, **anaerobic** (without oxygen) conditions in wetland environments. Their bodies were gradually transformed by the heat and pressure of the Earth's crust into the fossil fuels that we mine today to provide petroleum oil, natural gas, and coal (see more on this in chapter 4).

Reservoirs and fluxes of importance

The two largest reservoirs of carbon on Earth are the oceans, which cover the majority of Earth's surface, and the **lithosphere** (the mineral fraction of Earth: soils, rocks, and sediments). Each of these reservoirs holds more carbon than all of the other reservoirs combined. Much of the carbon stored in these reservoirs, especially deep in the lithosphere or in deep ocean environments, has an extremely long residence time, and does not actively participate in rapid fluxes. The notable exceptions here, of course, are fossil fuels, which are mined by humans and converted into gaseous forms of carbon through combustion.

Biomass, which is biological material derived from living, or recently living organisms, is a much smaller reservoir of carbon. The amount of carbon stored in all of the terrestrial vegetation (550 Gt C) (Gt = gigatonne = 10^9 metric tons = 10^{15} g) is just a fraction of that stored in the oceans (38,000 Gt C) and lithosphere (18,000 Gt C). All of the carbon that is currently stored in all of the vegetation on Earth got there through the process of **photosynthesis**. Plants and other photosynthetic organisms are called **primary producers**, because they "fix" atmospheric CO_2 into organic carbon, such as sugar, a form that is usable by animals and other organisms that need to consume their carbon molecules.

Photosynthetic organisms, such as plants, algae, and cyanobacteria, bring in CO_2 from the atmosphere and, using energy from the sun, convert CO_2 and water into glucose molecules (organic carbon). The products of photosynthesis are oxygen and glucose (Equation $11.1.1$). These glucose molecules are simple sugars that **autotrophs** ("self-feeders") can "burn" for energy, or transform into other usable carbon molecules through the process of cellular respiration (described in the next paragraph), or to build plant biomass. Photosynthesis takes place in organelles called **chloroplasts**, shown in Figure $11.1.2$. Photosynthesis accounts for 123 Gt of C per year that is removed from the atmosphere and stored in plant biomass. Such a massive amount of photosynthesis occurs on Earth that no other single flux moves as much carbon in the same timeframe.

$$6CO_2 + 6H_2O + \text{solar energy} \rightarrow C_6H_{12}O_6 + 6O_2 \tag{11.1.1}$$

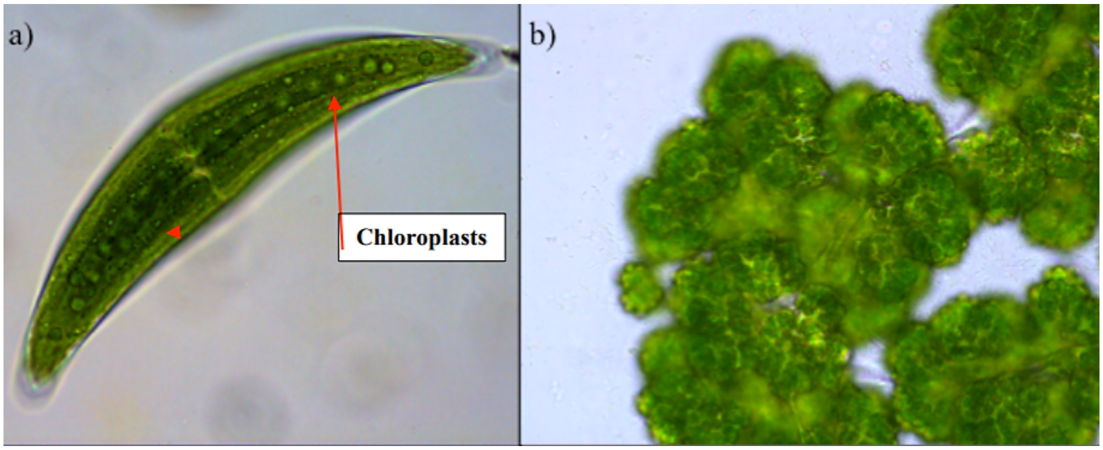


Figure 11.1.2: Chloroplasts visible in freshwater algae. Chloroplasts are green in color due to the chlorophyll a they contain, and are the site of photosynthesis. Chlorophyll a is the green pigment that allows plants, algae, and cyanobacteria to absorb the energy they need for photosynthesis from sunlight. a) Closterium moniliferum Ralfs, (Chlorophyta) green coccoid algae; b) Botryococcus braunii Kützing, (Chlorophyta) green coccoid algae with discoid chloroplasts. Image credit: K. Manoylov, Lake Sinclair, GA

Biomass in the carbon cycle, including plants and animals, is the reservoir of carbon that we are most likely most familiar with, and also the reservoir that is most readily available to us. We all participate in the flux of **consumption** of carbon when we eat food. All of our food is simply plant and/or animal biomass. Our body takes the carbon molecules contained in this biomass, and uses them, along with the oxygen we breathe in, for **cellular respiration** to create the adenosine triphosphate (ATP) we need for energy. The products of cellular respiration include the CO_2 we exhale, water, and energy that is stored in ATP (Equation $11.1.2$). Our bodies also builds additional biomass out of the carbon molecules in this food, allowing us to create new cells for growth or replenishment. This is the only way we, and all other **heterotrophs** ("other-eaters"), can bring in the carbon we need to build and maintain our bodies. Remember, you are what you eat!

$$C_6H_{12}O_6 + CO_2 \rightarrow 6CO_2 + 6H_2O + \text{energy} \tag{11.1.2}$$

Cellular respiration is an important flux in the carbon cycle, and one that contributes carbon to the atmosphere. Remember that animals and other heterotrophs complete cellular respiration using the carbon molecules that they bring in through their food. Plants and other photosynthetic autotrophs complete cellular respiration using the carbon molecules they formed from CO_2 through photosynthesis. Any carbon molecules that are left over after the organism has acquired sufficient energy through cellular respiration make up the biomass of the plant. As plants and animals die and decompose, their bodies are consumed by decomposer organisms such as fungi and bacteria. Through the flux of **decomposition**, some decaying biomass is converted into atmospheric carbon by the decomposers, while most of the biomass is buried into the soil, contributing to soil carbon. In oxygen-rich environments, decomposers rapidly consume dead and decaying biomass using the same process of aerobic cellular respiration described above. In oxygen-deficient environments, decomposers complete other metabolic pathways, and very slowly consume the organic matter. Some of the gases produced from anaerobic decomposition include **methane** (CH_4), **nitrous oxide** (N_2O), and the foul-smelling hydrogen sulfide (H_2S).

The biomass reservoir of the carbon cycle is also important to us as a source of energy. Through the flux of **combustion**, we convert the **potential energy** held in biomass into heat energy that we can use, and release carbon dioxide in the process. If you have ever burned logs on a campfire, or even burned food on the stove, you have completed this flux of biomass combustion. Of course, this happens naturally as well, the best example being natural forest fires caused by lightning strikes. The chemical reaction for combustion is identical to the chemical reaction for cellular respiration. The difference is that in cellular respiration, energy is released in a controlled fashion, and captured in ATP molecules. In combustion, all of this energy is released rapidly in the form of light and heat.

As all of the fluxes we've discussed so far involve the atmosphere, we have not yet discussed the flux that connects the atmosphere to the oceans. Carbon can enter the oceans through two primary fluxes: first through photosynthesis by algae or cyanobacteria (also called phytoplankton in Figure 11.1.1), and second through the chemical reaction of **ocean-atmosphere exchange**. The ocean, as with all surface water bodies, always contains some dissolved CO_2. This CO_2 is in **equilibrium** with the CO_2 in the air. Some atmospheric CO_2 is constantly dissolving into the ocean, while some dissolved CO_2 is constantly diffusing into the atmosphere. Under normal conditions, these two fluxes will be happening at equal rates. As you can see in Figure 11.1.1, however, this is no longer the case. In the section Human impacts on the carbon cycle, we will discuss why this is the case.

Activity: Better understanding the carbon cycle

To further review the carbon cycle, and better understand the human impacts on it, use this interactive graphic from Woods Hole laboratories: http://www.whoi.edu/feature/carboncycle/. As you will see, the information described in this text is only a small portion of the total carbon cycle on Earth. Finally, complete Table 11.1.1 as a way to review the sink/source relationship within this cycle. See if you can correctly identify the source and sink of carbon for each of these important fluxes in the carbon cycle.

Table 11.1.1: Practice understanding the sink/source relationship with cycles

Carbon flux	Carbon source	Carbon sink
Cellular respiration	Carbohydrates in living organisms	CO_2 in the atmosphere
Photosynthesis		
Consumption		
Combustion		
Decomposition		
Ocean/atmosphere exchange		
Fossil fuel formation		

11.2: Human Impacts on The Carbon Cycle

Learning Objectives

- Identify key human activities impacting the carbon cycle.
- Explain the greenhouse effect.

Humans, just like all other living organisms, have impacted the global carbon cycle since the dawn of our species. However, the magnitude of our impacts has changed dramatically throughout history. The **Industrial Revolution**, which occurred around the turn of the 19th century, began to make major changes in the use of resources around the world. Beginning in Britain, industrialization eventually affected the whole world. The development of coal-fueled steam power, and later transportation following the discovery of large oil deposits, had enormous influence on the economic and social structure of the world. As the world accelerated in the production and transportation of manufactured goods, the production and consumption of fossil fuels grew. As economic growth continued to increase, so did the production of carbon dioxide through fossil fuel combustion. See Figure $11.2.2$later in this text.

Some of the human impacts on the carbon cycle have been quantified for you in Figure 7.1.1. Changes to fluxes in the carbon cycle that humans are responsible for include: increased contribution of CO_2 and other **greenhouse gases** to the atmosphere through the combustion of fossil fuels and biomass; increased contribution of CO_2 to the atmosphere due to land-use changes; increased CO_2 dissolving into the ocean through ocean-atmosphere exchange; and increased terrestrial photosynthesis. The first two impacts, both contributing excess CO_2 to the atmosphere at a rate of 4 Gt of carbon per year have, by far, the largest impact on our planet. For this reason, this is the change that we will most often focus on throughout this section. The excess CO_2 in the atmosphere is responsible for the increased CO_2 dissolving into the ocean, which we will discuss later in this section. This is also, in part, responsible for the increased terrestrial photosynthesis that can be observed, as additional CO_2 is available to plants for photosynthesis. However, intensive agricultural and forestry practices also contribute to the change in this flux.

One characteristic example of a human impact on the carbon cycle is illustrated in Figure $11.2.1$. Throughout most of our recent human history, people have been physically altering the landscape around them in order to have more control over their surroundings and increase their odds of survival. One way that people have done this is through agriculture. In order for most forms of agriculture to be successful, native vegetation is eliminated or minimized. Resources from this native vegetation, such as wood, may be used for combustion to provide heat, sanitation, or fuel for cooking. Combustion may also be used as an efficient way to clear the land and make way for crops or grazing lands for livestock. Often, settlements are formed around these newly fashioned agricultural fields, and the land is used in a similar fashion for many years in the future.

Let's identify the ways in which humans are impacting the carbon cycle in this scenario of agricultural establishment. You should be able to identify from the above paragraph that the flux of combustion will release CO_2 previously held in vegetation into the atmosphere. In addition, remember that the land that used to house native vegetation is now home to agricultural lands. In most controlled agricultural environments, there is less total vegetative biomass than there would be under natural conditions. This decreased biomass leads to lower total photosynthesis rates, thereby decreasing the amount of CO_2 that is removed from the atmosphere and turned into plant biomass. Also, open soil on the fields between crops, during the winter months, or as a result of overgrazing allows for the air to penetrate deep into the soil structure. This provides the environment necessary for enhanced aerobic respiration by soil microorganisms. This decreases soil carbon, which can lead to erosion and soil degradation, and also releases additional CO_2 to the atmosphere.

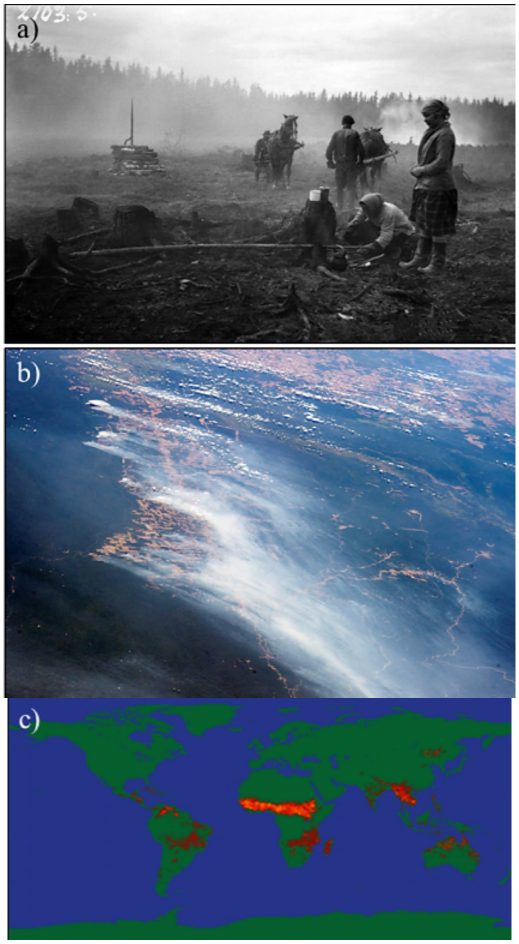


Figure 11.2.1: Impacts of slash and burn agriculture and biomass burning. a) Slash and burn agriculture in Maaninka, Finland. This kind of agriculture was still in use in Finland at the end of the 1920s. b) Slash and burn agriculture at the margins of the Amazon Rainforest in South America captured by astronauts on the International Space Station in August, 2014. Credit: Astronaut photograph ISS040-E-103496. c) Global distribution of biomass fires, represented by red, orange, and yellow dots (lighter colors indicates more fires), based on nighttime measurements obtained by the DMSP Operational Linescan System. Credit: Julia Cole, NASA Earth Observatory

As you learned in Chapter 5, biomass is an important form of energy to human civilization. Prior to the Industrial Revolution, this was essentially the only form of fuel to which most people on Earth had access. In many **less-industrialized countries**, combustion of biomass such as wood or animal dung is still the primary energy source that many citizens, particularly in rural areas, depend on for domestic use (heating, sanitation, and cooking) as it is inexpensive, relatively efficient, and readily available. Figure 11.2.1c shows the global distribution of biomass fires in the world. While the burning of biomass for domestic use contributes to some of these fires, it is the so-called **slash-and-burn agriculture** that makes up a larger contribution. Take a minute to compare the areas highlighted in Figure 11.2.1c to the countries of the world that are currently experiencing rapid population growth (Chapter 3). If you need a refresher, use the CIA World Factbook website to view current global population growth values by country: www.cia.gov/library/publicat.../2002rank.html.

While biomass burning still has a significant impact on the global carbon cycle, human impacts on fluxes such as fossil fuel extraction and combustion continue to grow. For a review of the impacts of non-renewable energy sources such as fossil fuels, see Chapter 4. Burning of any fossil fuel (coal, natural gas, crude oil) moves carbon from a previously-sequestered state deep within the Earth's crust into carbon dioxide in the atmosphere. As countries become more industrialized, their reliance on and combustion of fossil fuels tends to increase. Look at the graph in Figure 11.2.2, which compares CO_2 emissions from fossil fuels of regions across the globe.

How has the use and distribution of fossil fuels changed throughout the past 250 years?

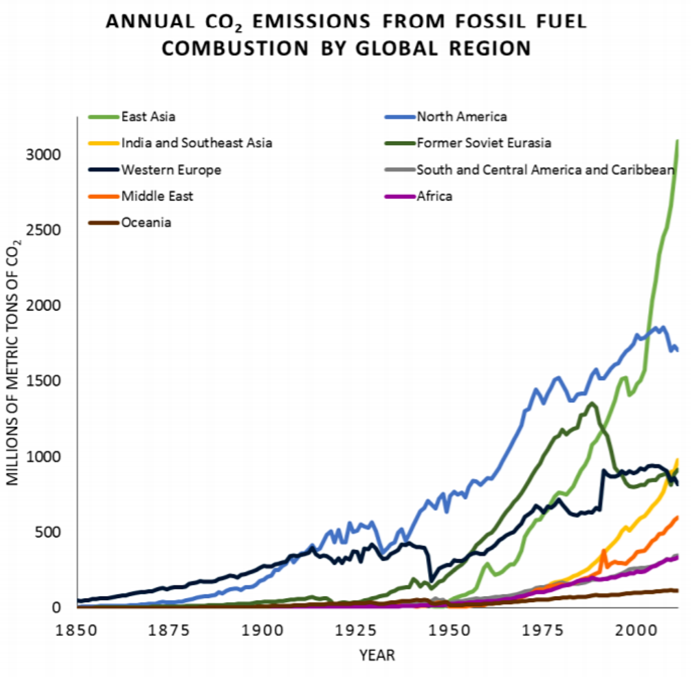


Figure 11.2.2: Total annual CO_2 emissions from fossil fuel combustion 1850-2011 by global region. Data from the Carbon Dioxide Information Analysis Center (CDIAC) cdiac.esd.ornl.gov/

The data shown in Figure 11.2.2 reveals much about the regions of the world it depicts. The effects of historic events such as the Great Depression of 1929-1939, World Wars, the fall of the Soviet Union in 1991, and the Kuwait oil fires of 1991 can be seen. Furthermore, between 1850 and 2011, different regions have gone in and out of the lead position as top producer of CO_2 from fossil fuel emissions. Population is one reason why fossil fuel use has changed throughout time. This is particularly apparent when comparing the data for Western Europe to that of India and Southeast Asia.

As countries industrialize, their relationship with agriculture also changes. **More-industrialized countries** rely very little on slash-and-burn agriculture. Their agricultural practices, however, are no less impactful on the environment. The growing population (Chapter 3) in many countries has required agriculture to become industrialized in order to meet demand. As a person living in the United States, **industrialized agriculture** probably produces the vast majority of the food you eat, including grains, fruits and vegetables, dairy and eggs, meats, and even fish. Industrialized agriculture can refer to a variety of practices, but has several main components: the use of motorized machinery; the use of chemicals such as fertilizers, pesticides, hormones, and/or antibiotics; and the intense and efficient production of one product across a large area of land.

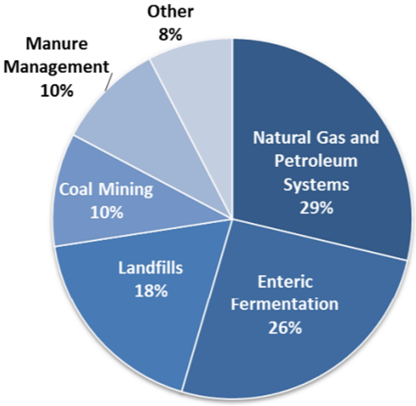


Figure 11.2.3: US methane emissions by source. 36% of the US methane production comes from agriculture: enteric fermentation (production of methane by anaerobic bacteria within the ruminant gut) and manure management. All emissions estimates from the Inventory of US Greenhouse Gas Emissions and Sinks: 1990-2013. US EPA

One example of the impacts of industrialized agriculture is the production of methane (CH_4), a potent greenhouse gas. You will learn more methaneabout methane later in this section. As you saw earlier, methane is a common product of anaerobic metabolisms. The gut of **ruminant animals** (such as sheep, cattle, and goats) has evolved to allow the animals to digest the very tough carbon molecules, such as cellulose, in grass. They do this through symbiosis, or cooperation, with anaerobic bacteria who live in the gut tract. These anaerobic bacteria produce methane and other gases as a result of their metabolism when they break down molecules like cellulose. This is sometimes called enteric fermentation. The methane gas is excreted from the animal, and this contributes significantly to total methane emissions (Figure 11.2.3). A similar type of bacteria live in the fecal matter, or manure, of livestock. As the manure is handled or stored for future use, methane is also released to the environment.

The methane excretions of one cow or a few sheep would be miniscule and insignificant. If you were a small farmer with only enough livestock to feed your family, your contribution to total methane emissions would be close to zero. However, the demand for animal protein from meat, dairy, and eggs is very large in the United States. As of January 2015, the United States had a total cattle inventory of 89.9 million animals, and in 2014, 25.5 billion pounds of beef was consumed in the United States (statistics: National Cattlemen's Beef Association). The impacts of enteric fermentation and manure management for almost 90 million animals are very significant, as seen in Figure 11.2.3. In both cases, carbon that was previously stored in biomass (cattle feed) is moved into the atmosphere, this time in the form of CH_4. This is another example of how humans have impacted the carbon cycle.

Previously in this chapter, you identified other ways the carbon cycle is impacted by human agriculture. Through industrialized agriculture, we must also account for the fossil fuels that are used. In order to deliver agricultural products to consumers, fossil fuels are used numerous times: deliveries of fertilizer, feed, and/or seed to farms; farm machinery; delivery of products to processers; food processing; delivery of foods to supermarkets; etc.

As animal products, especially meat, are expensive, the demand is typically greater in more industrialized countries than it is in less-industrialized countries. This makes industrialized agriculture, and especially industrialized animal agriculture, one of the major contributors to greenhouse gas emissions in more-industrialized countries.

Knowledge check – answer these questions on your own to further explore the impacts of biomass and fossil fuel burning on the global carbon cycle.

1. Why is there a correlation between population growth rate and global distribution of biomass fires?

2. Do you think this correlation is more likely due to personal biomass fires for activities such as cooking, or due to slash-and-burn agriculture? Why?

3. Given any other knowledge you might have about the areas highlighted in Figure 11.2.1c, what other environmental impacts may be occurring here besides carbon cycle alterations?

4. Compare the production of CO_2 emissions from fossil fuel combustion across world regions in 1900, 1950, and 2011 in Figure 11.2.2. What has accounted for these differences?

5. Has the total worldwide production of CO_2 from fossil fuels increased evenly relative to human population growth during the time period displayed in Figure 11.2.2? Why or why not?

6. What are the differences in contributions of greenhouse gas emissions from more industrialized countries and less-industrialized countries? What are the similarities?

Resources

Carbon Dioxide Information Analysis Center cdiac.esd.ornl.gov/

NOAA Earth System Research Laboratory: Carbon Cycle Science www.esrl.noaa.gov/research/themes/carbon/

Sass, Ronald. Q2: What are the Causes of Global Climate Change? OpenStax CNX. Sep 22, 2009 http://cnx.org/contents/5d263a29-7bd...c233619bca33@3

USDA Climate Change and Agriculture in the United States: Effects and Adaptation www.usda.gov/oce/climate_chan...griculture.htm

US EPA Overview of Greenhouse Gases: Methane http://epa.gov/climatechange/ghgemis...gases/ch4.html

Woods Hole Oceanographic Institution: Carbon Around the Earth http://www.whoi.edu/feature/carboncycle/

- **7.2: Human Impacts on the Carbon Cycle** by Caralyn Zehnder, Kalina Manoylov, Samuel Mutiti, Christine Mutiti, Allison VandeVoort, & Donna Bennett is licensed CC BY-NC-SA 4.0. Original source: https://oer.galileo.usg.edu/cgi/viewcontent.cgi?article=1003&context=biology-textbooks.

11.3: The Science of Climate Change

Learning Objectives

- Explain the natural greenhouse effect and its role in maintaining Earth's temperature.
- Discuss the methods used to study climate change, including ice core analysis.
- Analyze historical data on atmospheric CO2 concentrations and global temperature trends.

What is causing global climate change?

Scientists have identified the source of our current global climate change as being the increased human-caused emissions of greenhouse gases such as carbon dioxide (CO_2), methane (CH_4), and nitrous oxide (N_2O), since the industrial revolution. Greenhouse gases are defined as large (at least three atoms) gas molecules that participate in the **greenhouse effect**. While you already know about the “big three” greenhouse gases (CO_2, CH_4, and N_2O), 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, is it still an essential component of the natural greenhouse effect that occurs in our atmosphere.

The Earth receives energy from the sun and in turn radiates energy back into space. When these two energies are equal, a stable temperature of the Earth is achieved. This temperature can be calculated from basic physics and is equal to about -18°C (0°F). This thermal equilibrium temperature is obviously much colder than that of the surface of the Earth. The actual average value of the Earth’s surface temperature is about 15°C (59°F). The difference between these temperatures is due primarily to the natural greenhouse gas concentrations in the atmosphere, causing the greenhouse effect. If the Earth had no naturally occurring atmospheric greenhouse gases, the temperature at the surface of the Earth would equal the thermal equilibrium temperature. The influence of these greenhouse gases, mainly water and some CO_2, moderates the Earth’s climate and makes life possible (Figure $11.3.1$).

As **solar radiation** reaches the Earth’s atmosphere, there are a variety of possibilities for its fate. Some solar radiation is reflected by the Earth and its atmosphere, and does not contribute to warming. Some passes through the atmosphere and reaches the surface of the Earth. When this solar radiation is absorbed by objects on Earth’s surface, it is re-emitted as infrared radiation (heat) that escapes to space. However, some of this heat is intercepted in the atmosphere by greenhouse gases. These gases absorb and re-emit the radiation in all directions. This creates a warming impact on the Earth’s surface Radiation can be bounced around from one greenhouse gas molecule to another, becoming trapped, and increasing its warming potential. For this reason, an increased greenhouse gas concentration causes an increase in the overall warming potential of the Earth’s atmosphere.

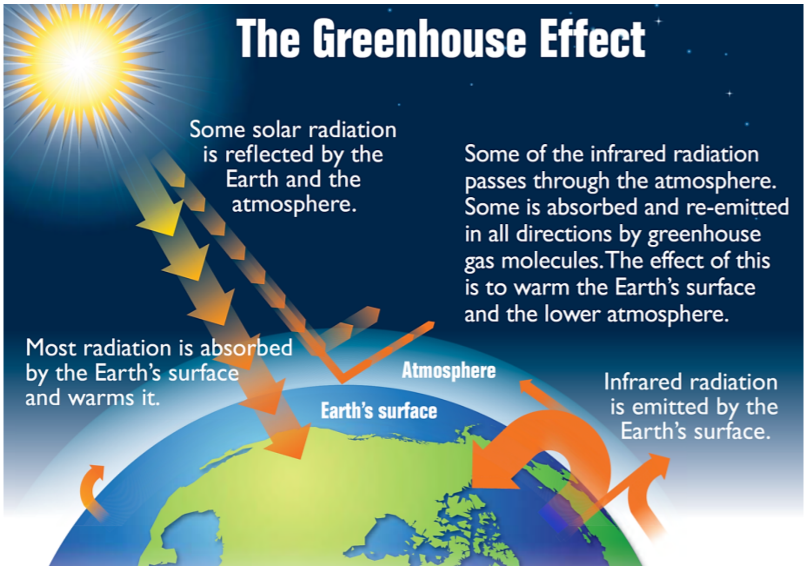


Figure 11.3.1: This diagram shows the Earth's greenhouse effect. The Earth absorbs some of the energy it receives from the sun and radiates the rest back toward space. However, certain gases in the atmosphere, called greenhouse gases, absorb some of the energy radiated from the Earth and trap it in the atmosphere. These gases essentially act as a blanket, making the Earth's surface warmer than it otherwise would be. While this greenhouse effect occurs naturally, making life as we know it possible, human activities in the past century have substantially increased the amount of greenhouse gases in the atmosphere, causing the atmosphere to trap more heat and leading to changes in the Earth's temperature. Credit: US EPA

What is Global Warming?

Global warming refers to the recent and ongoing rise in global average temperature near Earth's surface. It is caused mostly by increasing concentrations of greenhouse gases in the atmosphere. Global warming is causing climate patterns to change. However, global warming itself represents only one aspect of climate change. **Climate change** refers to any significant change in the measures of climate lasting for an extended period of time. In other words, climate change includes major changes in temperature, precipitation, or wind patterns, among other effects, that occur over several decades or longer.

On a geological time scale, the climate has changed many times in the past, even before the presence of humans. These changes occurred naturally because man had not yet evolved. A well-known example of past climate change is the occurrence of ice ages. Ice ages have occurred repeatedly throughout Earth's history, the most severe ice age of which scientists have reliable data occurred around 650,000 years ago. During this time, glacial ice covered much of Canada, the northern United States, and northern Europe; the level of the ocean decreased 120 m, and the temperature decreased by 5°C.

A geologic history of ice events is preserved in the ice sheets covering Antarctica and Greenland. This history has been uncovered over the past decades by scientists who have cored deeply into the ice and deciphered the temperature and atmospheric composition records stored in the ice. This process of obtaining **ice cores** is shown in Figure 11.3.2. The temperature at which the ice originally formed can be obtained from an interpretation of the measured ratio of the **stable isotopes** (see Chapter 1 supplement for a description of isotopes) of oxygen in the molecules of water forming the ice. The atmospheric gas composition is taken from air bubbles trapped in the ice at the time of formation. From these data, scientists have gathered a set of reliable data that track atmospheric temperature and gas concentrations that dates back 800,000 years. These data helped scientists come to the conclusion that the Earth's temperature and greenhouse gas concentrations are directly correlated to one another (Figure 11.3.3). During the ice age 650,000 years ago, the Earth was experiencing depressed temperature and atmospheric CO_2 concentrations below 200 **parts per million** (ppm). We can also see from these data, that CO_2 concentrations can be naturally elevated to as high as 300 ppm, correlating with increased temperatures.

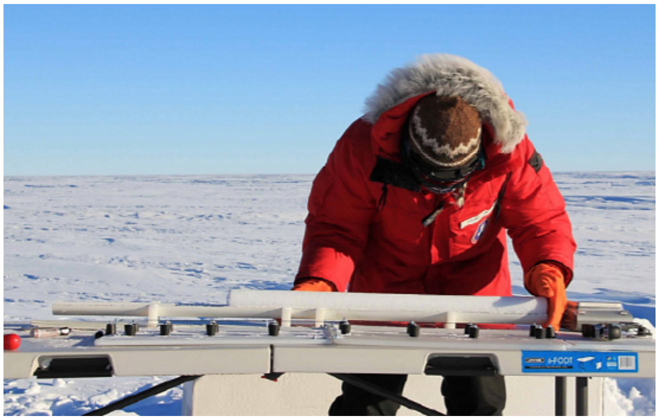


Figure 11.3.2: On Dec. 8, 2010, Michelle Koutnik, of the University of Copenhagen's Center for Ice and Climate, prepared ore of Antarctic ice to be wrapped and put into core tubes for transport back to labs at Brigham Young University in Utah. But first, Koutnik measured the core's length, diameter and weight. The traverse was the first of two field campaigns to study snow accumulation on the West Antarctic Ice Sheet and tie the information back to larger-scale data collected from satellites. 120 m, and the global average Credit: NASA/Lora Koenig.

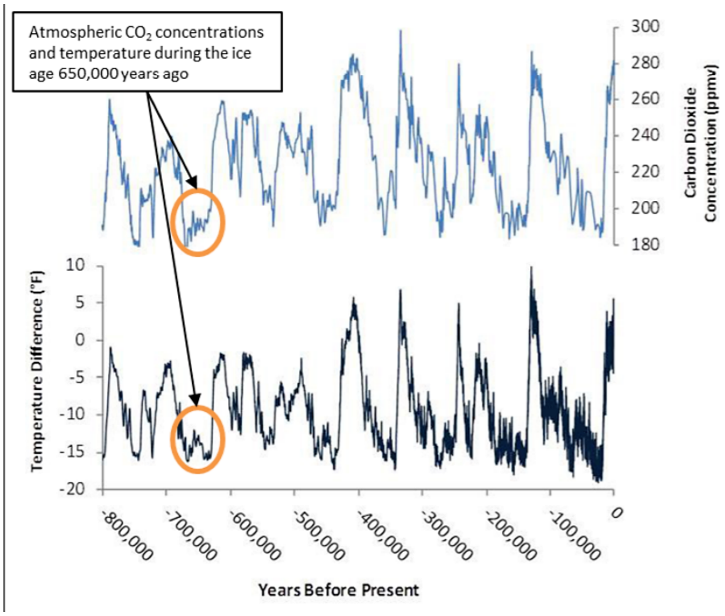
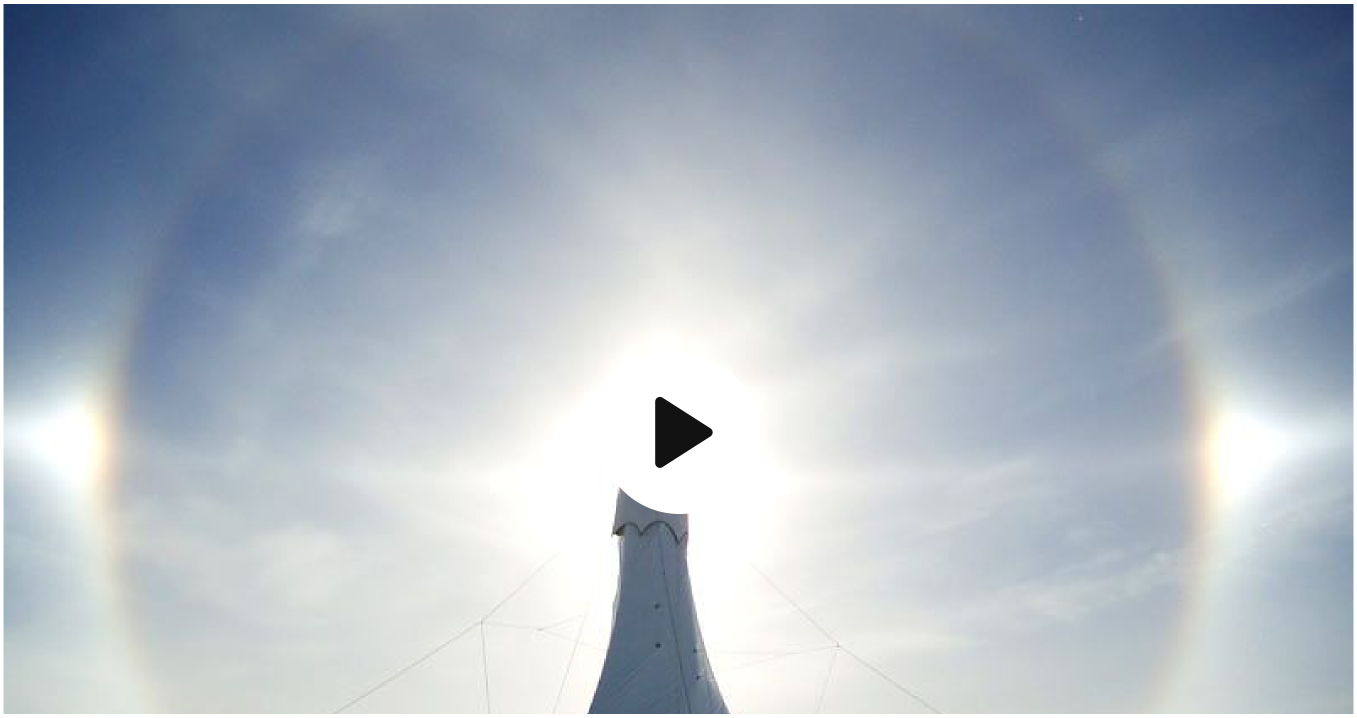


Figure 11.3.3: Estimates of the Earth’s changing CO_2 concentration (top) and Antarctic temperature (bottom), based on analysis of ice core data extending back 800,000 years. Until the past century, natural factors caused atmospheric CO_2 concentrations to vary within a range of about 180 to 300 ppm. Warmer periods coincide with periods of relatively high CO_2 concentrations. NOTE: The past century’s temperature changes and rapid CO_2 rise (to 400 ppm in 2015) are not shown here. Source: Based on data appearing in NRC (2010).

The 100,000 year major cycle of the ice ages and some variations within the cycles agree very well with predicted periodic relationships between the Earth’s orbit around the sun, generally referred to as the **Milankovitch cycles**. Milankovitch cycles describe the very slight “wobbles” that occur in the Earth’s tilt and path as it moves around the sun. The Earth is always slightly tilted on its axis with respect to the sun. The angle of this tilt, however, changes periodically, varying from about 22° to about 25°. A less severe tilt will cause milder summers and winters close to the poles, preventing full summer ice melt in the northern- and southernmost regions, and allowing for a buildup of ice from year to year.

The path through which the Earth travels on its journey around the sun also changes from a more circular to a more elongated shape. Again, a round orbit will cause milder summers and winters close to the poles. These are very long term changes, and the results of the Milankovitch cycles can be observed in the changes in temperature and atmospheric CO_2 concentration shown in Figure 11.3.3. The climate change event that scientists are currently documenting is occurring much more rapidly than could be explained by Milankovitch cycles. Therefore, scientists agree that the cause of our currently changing climate is due to human impacts and not natural forces.



11.4: Greenhouse Gases

Learning Objectives

- Identify the main types of greenhouse gases and their sources.
- Describe how greenhouse gases contribute to climate change and global warming.

We will be covering the four major categories of greenhouse gases that have been impacted by humans the most. See Table 11.4.1 for a numeric comparison of these greenhouse gases.

• Carbon dioxide, CO_2

• Methane, CH_4

• Nitrous oxide, N_2O

• Synthetic **fluorinated gases**, including **hydrofluorocarbons** (HFCs), **perfluorocarbons** (PFCs), and **sulfur hexafluoride** (SF_6)

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 (Chapter 4). To get an idea of how CO_2 concentration has changed over time, watch this video compiled by the National Oceanic and Atmospheric Administration (NOAA): http://www.esrl.noaa.gov/gmd/ccgg/trends/history.html. 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. 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. Similar to the action of the naturally existing greenhouse gases, any additional greenhouse gases leads to an increase in the surface temperature of the Earth.

How does carbon dioxide trap heat? from PBS Studios

The video explains how greenhouse gases like carbon dioxide trap heat by absorbing and transferring infrared energy, unlike oxygen or nitrogen. Other greenhouse gases, such as water vapor, methane, and nitrous oxide, work in the same way, trapping heat and causing global warming.

We're sorry, but there was an error playing this video. Please try again later.

While CO_2 is produced by aerobic cellular respiration, gases such as CH_4 and N_2O are often the products of anaerobic metabolisms. Agriculture is a major contributor to CH_4 emissions, as you saw in section 7.1. 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. For a review of how fossil fuels are mined, see Chapter 4. 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 and 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.

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 7.9 shows the increases in all three gases following the industrial revolution. Ice core data shows us that the atmospheric CO_2 concentration never exceeded 300 ppm before the industrial revolution. As of early 2015, the current atmospheric CO_2 concentration is 400 ppm. Comparing Figure 11.4.1 to Figure 7.3.3, above, what is likely to happen to global temperature following this unprecedented rise in greenhouse gas levels?

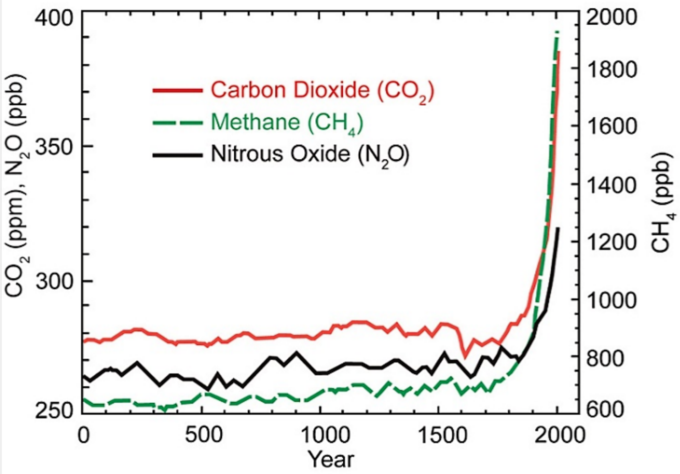


Figure 11.4.1: Increase in greenhouse gas concentrations in the atmosphere over the last 2,000 years. Increases in concentrations of these gases since 1750 are due to human activities in the industrial era. Concentration units are parts per million (ppm) or parts per billion (ppb), indicating the number of molecules of the greenhouse gas per million or billion molecules of air. Source: USGCRP (2009)

One class of greenhouse gas chemicals that has no natural sources is the fluorinated gases. These include HFCs, PFCs, and SF6, 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 11.4.1. Some of these chemicals, as well as the older **chlorofluorocarbons** (CFCs), have been phased out by international environmental legislation under the Montreal Protocol (see Chapter 6). 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.

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 11.4.1. 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. How does this impact the comparison of the environmental effects of agricultural practices in lessindustrialized and more-industrialized countries that we completed in section 7.1?

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	SF6	3,200 years	22,800

Other climate influencers

In addition to greenhouse gases, other manmade changes may be forcing climate change. Increases in near-surface ozone from internal combustion engines, aerosols such as carbon black, mineral dust and aviation-induced exhaust are acting to raise the surface temperature. This primarily occurs due to a decrease in the **albedo** of light-colored surfaces by the darker-colored carbon black, soot, dust, or particulate matter. As you know, it is more comfortable to wear a white shirt on a hot summer day than a black shirt. Why is this? Because the lighter-colored material bounces more solar radiation back toward space than the darker-colored material does, allowing it to stay cooler. The darker-colored material absorbs more solar radiation, increasing its temperature. Just as the white shirt has a higher albedo than the black shirt, light-colored objects in nature (such as snow) have a higher albedo than dark-colored objects (such as soot or dust). As humans increase the amount of carbon black, soot, dust, and particulates in the atmosphere, we decrease the albedo of light-colored surfaces, causing them to absorb more solar radiation and become warmer than they would without human influence. An example of this can be seen in the snow on Figure 11.4.2

Figure 11.4.2: A photograph of the extreme dust deposition from the deserts of the Colorado Plateau onto the Colorado Rockies snowpack in 2009. Taken from the high point of the Senator Beck Basin in the San Juan Mountains, it captures the extent of the impact of darkening in which the snow albedo dropped to about 30%, more than doubling the absorption of sunlight. Credit: S. McKenzie Skiles, Snow Optics Laboratory, NASA/JPL

11.5: Fossil Fuels and Greenhouse Gases

Fossil fuels are made up mainly of hydrogen and carbon. When burned, the carbon combines with oxygen to create carbon dioxide (CO_2). The amount of CO_2 produced depends on the carbon content of the fuel. For example, for the same amount of energy produced, natural gas produces about half and petroleum produces about three-fourths of the amount of CO_2 produced by coal. Energy-related CO_2 emissions, resulting from the combustion of coal, petroleum, and natural gas, account for about 80% of total U.S. human-caused (anthropogenic) greenhouse gas (GHG) emissions. There are many sources of non-energy CO_2 emissions, but those emissions account for a relatively small share of total GHG emissions. See chapter 7 for a discussion of the results of GHG emissions.

Energy use is largely driven by economic growth and by weather patterns that affect heating and cooling needs. The fuels used in electricity generation also have an impact on the amount of GHG emissions. In the United States, most of the electricity generated comes from coal power plants and consequently, the majority of the carbon dioxide emission resulting from electricity generation is from coal combustion (Figure 11.5.1). Although the industrial sector is the largest consumer of energy (including direct fuel use and purchased electricity), the transportation sector emits more carbon dioxide because of its near complete dependence on petroleum fuels. The residential and commercial sectors have lower emission levels (most of which comes from fossil energy combustion to produce electricity) than the transportation and industry sectors.

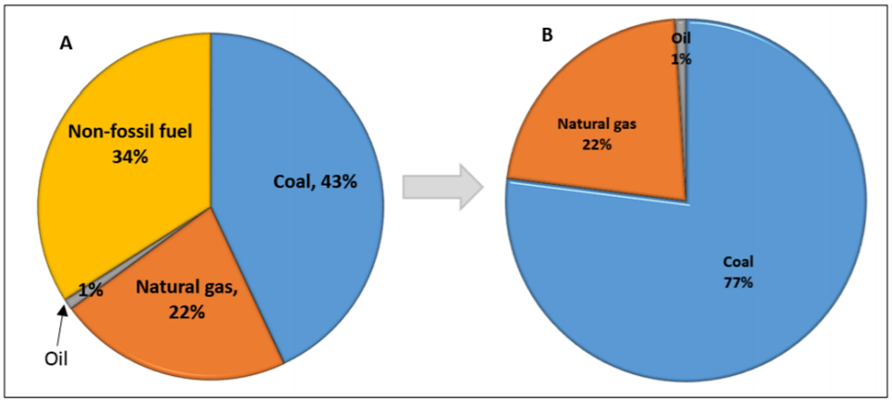


Figure 11.5.1: *A) Major fuel/energy sources for U.S. electricity generation, 2013. B) Resulting carbon dioxide emissions from electricity generation by fuel type, 2013. Based on data from U.S. Energy Information Administration*

- **4.6: Fossil Fuels and Greenhouse Gases** by Caralyn Zehnder, Kalina Manoylov, Samuel Mutiti, Christine Mutiti, Allison VandeVoort, & Donna Bennett is licensed CC BY-NC-SA 4.0. Original source: https://oer.galileo.usg.edu/cgi/viewcontent.cgi?article=1003&context=biology-textbooks.

11.6: Other climate influencers

Learning Objectives

- Define both "climate" and "weather" and explain how the two are related
- Discuss the role and mechanisms of the major controls on Earth's climate using the concepts of insolation and albedo.
- Identify and describe the mechanisms by which major external and internal changes to the climate (including solar output variation, volcanoes, biological processes, changes in glacial coverage, and meteorite impacts) operate
- Know that the Earth's climate has changed greatly over its history as a result of changes in insolation, albedo, and atmospheric composition

Introduction

The Earth's climate is continually changing. If we are to understand the current climate and predict the climate of the future, we need to be able to account for the processes that control the climate. One hundred million years ago, much of North America was arid and hot, with giant sand dunes common across the continent's interior. Six hundred and fifty million years ago it appears that the same land mass—along with the rest of the globe—was covered in a layer of snow and ice. What drives these enormous changes through Earth's history? If we understand these fundamental processes we can explain why the climate of today may also change.

Weather describes the short term state of the atmosphere. This includes such conditions as wind, air pressure, precipitation, humidity and temperature. Climate describes the typical, or average, atmospheric conditions. Weather and climate are different as the short term state is always changing but the long-term average is not. On The 1st of January, 2011, Chicago recorded a high temperature of 6 ^{o}C; this is a measure of the weather. Measurements of climate include the averages of the daily, monthly, and yearly weather patterns, the seasons, and even a description of how often extraordinary events, such as hurricanes, occur. So if we consider the average Chicago high temperature for the 1st of January (a colder 0.5 ^{o}C) or the average high temperature for the entire year (a warmer 14.5 ^{o}C) we are comparing the city's weather with its climate. The climate is the average of the weather.

Insolation, Albedo and Greenhouse Gases

What controls the climate? The average temperature of the Earth is about 15 ^{o}C (which is the yearly average temperature for the city of San Francisco), so most of the Earth's water is in a liquid state. The average temperature of Mars is about -55 ^{o}C (about the same as the average winter temperature of the South Pole), so all of the water on the Martian surface is frozen. This is a big difference! One reason Earth is so much hotter than Mars is that Earth is closer to the Sun. Mars receives less than half as much energy from the Sun per unit area as Earth does. This difference in insolation, which is the measure of the amount of solar radiation falling on a surface, is a very important factor in determining the climate of the Earth.

On Earth, we notice the effects of varying insolation on our climate. Sunlight falls most directly on the equator, and only obliquely (at an angle) on the poles. This means that the sunlight is more concentrated at the equator. As shown in Figure 11.6.2, the same amount of sunlight covers twice as much area when it strikes a surface at an angle of 30^{o} compared to when it strikes a surface directly: the same energy is spread more thinly, weakening its ability to warm the Earth.

Firgure 11.6.2 Insolation angle. Insolation is the effect of incidence angle on sunlight intensity. Note that the same amount of sunlight is spread out over twice the area when it strikes the surface at a 30-degree angle. *Source: Wikipedia*

As a consequence, the tropics receive about twice the insolation as the area inside the Arctic Circle – see Figure 11.6.3. This difference in energy explains why the equator has a hot climate and the poles have a cold climate. Differences in insolation also explain the existence of seasons. The Earth's axis is tilted at 23^{o} compared to its orbit, and so over the course of the year each hemisphere alternates between directly facing the Sun and obliquely facing the Sun. When the Northern hemisphere is most directly facing the Sun (the months of May, June and July) insolation is thus higher, and the climate is warmer. This variation in insolation explains why summer and winter occur (we get less energy from the Sun in winter then we do in summer), and why the timing of the seasons is opposite in the Southern and Northern hemispheres.

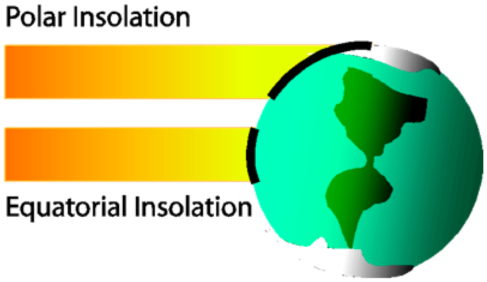


Figure 11.6.3 Insolation effect. A cartoon of how latitude is important in determining the amount of insolation. The same amount of sunlight (yellow bars) is spread out over twice the planet's surface area when the rays strike the Earth at an angle (compare the length of the dark lines at the equator and at the poles). *Source:* [Jonathan H. Tomkin.]

Figure 11.6.4 shows both the equatorial and seasonal impacts of insolation. High levels of insolation are shown in warm colors (red and pink) and low levels of insolation are shown in cold colors (blue). Notice that in January (top map) the maximum levels of insolation are in the Southern Hemisphere, as this is when the Southern Hemisphere is most directly facing the sun. The Arctic receives very little insolation at this time of year, as it experiences its long polar night. The reverse is true in April (bottom map).

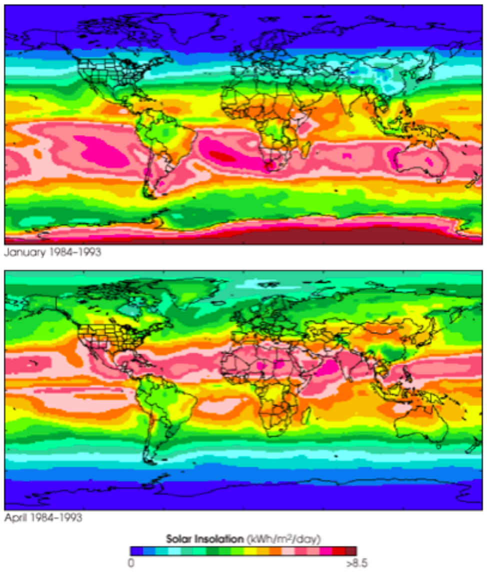


Figure 11.6.4 Insolation. Average insolation over ten years for the months of January (top) and April (bottom). *Source:* [Roberta DiPasquale, Surface Meteorology and Solar Energy Project, NASA Langley Research Center, and the ISCCP Project. Courtesy of NASA's Earth Observatory.]

The equator always receives plenty of sunlight, however, and has a much higher average temperature as a consequence; compare the average temperature of the equator with that of the poles in Figure 11.6.5.

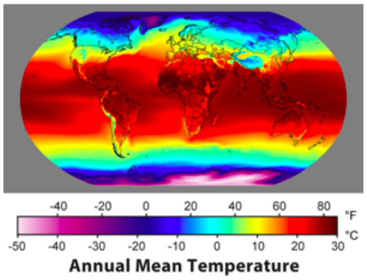


Figure 11.6.5 The Earth's average annual temperature. *Source:* Robert A. Rohde for Global Warming Art.

The level of insolation affecting Earth depends on the amount of light (or solar radiation) emitted by the Sun. Over the current geologic period, this is very slowly changing—solar radiation is increasing at a rate of around 10% every billion years. This change is much too slow to be noticeable to humans. The sun also goes through an 11-year solar cycle, in which the amount of solar radiation increases and decreases. At the solar cycle peak, the total solar radiation is about 0.1% higher than it is at the trough.

The Earth's orbit is not perfectly circular, so sometimes the Earth is closer to or further from the Sun than it is on average. This also changes the amount of insolation, as the closer the Earth is to the Sun the more concentrated the solar radiation. As we shall see in the next section, these orbital variations have made a big difference in conditions on the Earth during the period in which humans have inhabited it.

In addition to considering how much energy enters the Earth system via insolation, we also need to consider how much energy leaves. The climate of the Earth is controlled by the Earth's energy balance, which is the movement of energy into and out of the Earth system. Energy flows into the Earth from the Sun and flows out when it is radiated into space. The Earth's energy balance is determined by the amount of sunlight that shines on the Earth (the insolation) and the characteristics of the Earth's surface and atmosphere that act to reflect, circulate and re-radiate this energy. The more energy in the system the higher the temperature, so either increasing the amount of energy arriving or decreasing the rate at which it leaves would make the climate hotter.

One way to change how quickly energy exits the Earth system is to change the reflectivity of the surface. Compare the difference in dark surface of tilled soil (Figure 11.6.6a) with the blinding brightness of snow-covered ice (Figure 11.6.6b).

6a: Tilled soil. *Source:* *Tim Hallam.*

6b: The snow surface at Dome C Station, Antarctica *Source:* *Stephen Hudson*

Figure 11.6.6 The reflectivity of the surface of the earth.

The dark soil is absorbing the sun's rays and in so doing is heating the Earth surface, while the brilliant snow is reflecting the sunlight back into space. Albedo is a measure of how reflective a surface is. The higher the albedo the more reflective the material: a perfectly black surface has zero albedo, while a perfectly white surface has an albedo of 1 - it reflects 100% of the incident light. If a planet has a high albedo, much of the radiation from the Sun is reflected back into space, lowering the average temperature. Today, Earth has an average albedo of just over 30%, but this value depends on how much cloud cover there is and what covers the surface. Covering soil with grass increases the amount of light reflected from 17% to 25%, while adding a layer of fresh snow can increase the amount reflected to over 80%. Figure 11.6.7 is a composite photograph of the Earth with the cloud cover removed. As you can see, forests and oceans are dark (low albedo) while snow and deserts are bright (high albedo).

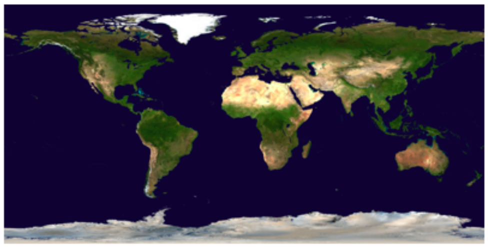


Figure 11.6.7 The surface of the Earth with cloud cover removed. The poles and deserts are much brighter than the oceans and forests. *Source: NASA Goddard Space Flight Center Image by Reto Stöckli. Courtesy of NASA's Earth Observatory.*

Changes in albedo can create a positive feedback that reinforces a change in the climate. A positive feedback is a process which amplifies the effect of an initial change. If the climate cools, (the initial change), snow covers more of the surface of the land, and sea-ice covers more of the oceans. Because snow has a higher albedo than bare ground, and ice has a higher albedo than water, this initial cooling increases the amount of sunlight that is reflected back into space, cooling the Earth further (the amplification, or positive feedback). Compare the brightness of Figure 11.6.7with a similar photo montage from February (Figure 11.6.7): the extra snow has increased the Earth's albedo. Imagine what would happen if the Earth produced even more snow and ice as a result of this further cooling. The Earth would then reflect more sunlight into space, cooling the planet further and producing yet more snow. If such a loop continued for long enough, this process could result in the entire Earth being covered in ice! Such a feedback loop is known as the Snowball Earth hypothesis, and scientists have found much supporting geological evidence. The most recent period in Earth's history when this could have occurred was around 650 Million years ago. Positive feedbacks are often described as "runaway" processes; once they are begun they continue without stopping.

Figure 11.6.8 This image shows the surface of the Earth in February (the Northern Hemisphere winter) with cloud cover removed. The seasonal snow cover is brighter (and so has a higher albedo) than the land surface it covers. *Source: NASA Goddard Space Flight Center Image by Reto Stöckli. Courtesy of NASA's Earth Observatory*

Albedo does not explain everything, however. The Earth and the Moon both receive the same amount of insolation. Although the Moon is only slightly more reflective than the Earth, it is much colder. The average temperature on Earth is 15 ^{o}C, while the Moon's average temperature is -23 ^{o}C. Why the difference? A planet's energy balance is also regulated by its atmosphere. A thick atmosphere can act to trap the energy from sunlight, preventing it from escaping directly into space. Earth has an atmosphere while the Moon does not. If the Earth did not have an atmosphere, it would have an average temperature of -18 ^{o}C; slightly warmer than the Moon since it has a lower albedo.

How does the atmosphere trap the energy from the Sun? Shouldn't the Earth's atmosphere reflect as much incoming radiation as it traps? It is true the atmosphere reflects incoming solar radiation—in fact, only around half the insolation that strikes the top of the atmosphere reaches the Earth's surface. The reason an atmosphere generally acts to warm a planet is that the nature of light radiation changes as it reaches the planet's surface. Atmospheres trap more light than they reflect.

Humans see the Earth's atmosphere as largely transparent; that is, we can see a long way in air. This is because we see light in the visible spectrum, which is the light radiation in the range of wavelengths the human eye is able to perceive, and visible light is able to travel a long way through the Earth's atmosphere before it is absorbed. Light is also transmitted in wavelengths we can't see, such as in the infrared spectrum, which is sometimes referred to as infrared light, heat, or thermal radiation. Compared to visible light, infrared light cannot travel very far in the Earth's atmosphere before it is absorbed. Solar radiation striking the Earth is largely in the visible part of the spectrum. The surface of the Earth absorbs this energy and re-radiates it largely in the infrared part of the spectrum. This means that solar radiation enters the Earth in the form of visible light, unhindered, but tries to leave in the form of infrared light, which is trapped. Thicker atmospheres keep this infrared radiation trapped for longer, and so warm the Earth—just like an extra blanket makes you warmer in bed.

This effect is shown in Figure 11.6.9. The visible light radiation enters the atmosphere, and quickly exits as infrared radiation if there is no atmosphere (top Earth). With our atmosphere (the middle Earth), visible light enters unhindered but the infrared light is partially reflected back to the surface, increasing the amount of energy and thus the temperature at the Earth's surface. If the atmosphere is made thicker (bottom Earth) the infrared radiation is trapped for longer, further warming the planet's surface.

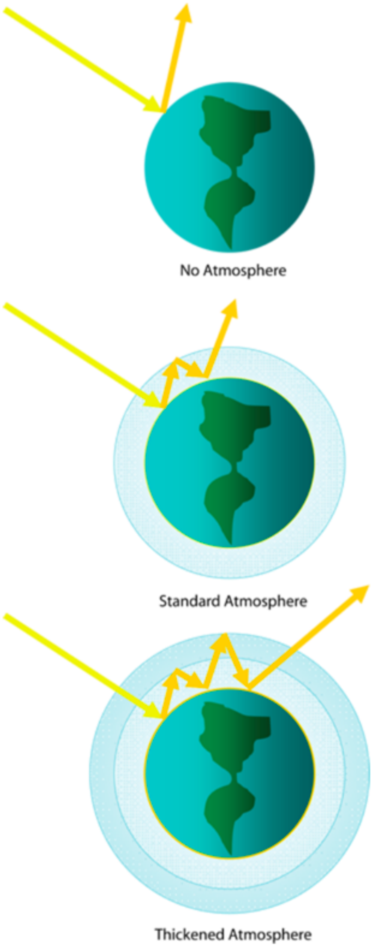


Figure 11.6.9 A cartoon of the greenhouse effect. (Top) Visible light radiation emitted by the sun (yellow arrows) strikes the Earth and reflects as infrared radiation (orange arrow); (middle) an atmosphere reflects some of the infrared radiation back toward the planet; (bottom) a thickened atmosphere reflects greater amounts of infrared radiation. *Source: Jonathan H. Tomkin.*

Earth's Changing Atmosphere

The composition of Earth's atmosphere has changed over geologic time. The atmosphere has largely come from volcanic venting of gas from Earth's interior (see Figure 11.6.10), but biology has also made important changes by producing oxygen and removing carbon dioxide. Greenhouse gases currently make up only a small fraction of the Earth's atmosphere—99% of air consists of nitrogen and oxygen molecules.

Figure 11.6.10 The Mt. Bromo volcano in Indonesia emitting gas into the atmosphere. *Source: Jan-Pieter Nap, taken on July 11, 2004.*

While volcanoes can warm the Earth by adding carbon dioxide to the atmosphere, which produces a greenhouse effect, they can also cool the Earth by injecting ash and sulfur into the atmosphere. These additions raise the albedo of the atmosphere, allowing less sunlight to reach the surface of the Earth. The effect lasts until the particles settle out of the atmosphere, typically within a few years. Volcanic eruptions have impacted human societies throughout history; the Mt. Tambora eruption in 1815 cooled the Earth so much that snow fell during June in New England, and the more recent Mt. Pinatubo eruption in 1991 (see Figure 11.6.11) ejected so much sulfuric acid into the atmosphere that global temperatures were lowered by about 0.5 ^{o}C in the following year.

Figure 11.6.11 The 1991 eruption of Mt. Pinatubo. *Source: U.S. Geological Survey photograph, by Richard P. Hoblitt.*

Evidence from the geologic past indicates that similar events have caused mass extinctions wherein a significant fraction of all species on Earth were wiped out in a relatively short amount of time. Sustained outgassing from continuous volcanic eruptions is thought to have produced so much ash and aerosols that light sufficient to support photosynthesis in plants was unable to penetrate the atmosphere, causing the food chain to collapse. The ash particles produced by extended eruptions would also have increased the Earth's albedo, making conditions inhospitably cool for plants and animals adapted to a warmer environment.

Asteroid impacts can also cause the climate to suddenly cool. When large asteroids strike the Earth, ash is ejected into the atmosphere, which increases albedo in the same way as volcanic eruptions. Everyday clouds (made up of water droplets) both cool and warm the Earth. They can cool the Earth by increasing the planet's albedo, reflecting sunlight into space before it reaches the surface. Clouds can also warm the Earth, by reflecting infrared radiation emitted by the surface back towards the planet. Different types of clouds, and different conditions, determine which effect predominates. On a hot summer's day, for example, clouds cool us by shielding us from the sun's rays, but on a winter's night a layer of cloud can act as a warming blanket.

The composition of the Earth's atmosphere is not fixed; greenhouse gases can be added to and removed from the atmosphere over time. For example, carbon dioxide is added by volcanoes and the decay or burning of organic matter. It is removed by photosynthesis in plants, when it is dissolved in the oceans and when carbonate sediments (a type of rock) are produced. Over geologic time, these processes have significantly reduced the proportion of carbon dioxide in the atmosphere. Atmospheric carbon dioxide levels just prior to the industrial revolution are thought to have been only one twentieth of those of 500 million years ago. Natural processes also remove carbon dioxide added by human activity, but only very slowly. It is estimated that it would take the Earth around a thousand years to naturally remove most of the carbon dioxide released by the industrial consumption of fossil fuels up to the present.

Greenhouse gases other than carbon dioxide are shorter-lived: methane is removed from the atmosphere in around a decade, and chlorofluorocarbons break down within a century. Individual water molecules spend only a few days at a time in the atmosphere, but unlike the other greenhouse gases, the total amount of water vapor in the atmosphere remains constant. Water evaporated from the oceans replaces water lost by condensation and precipitation.

Changing the composition of the Earth's atmosphere also changes the climate. Do you remember the snowball earth? — how increasing ice cover also increased the Earth's albedo, eventually covering the entire planet in ice and snow? Today's climate is temperate—so we must have escaped this frozen trap. But how? The leading hypothesis is that the composition of the Earth's atmosphere changed, with volcanoes slowly adding more and more carbon dioxide to it. Without access to the oceans, plants, or surface rocks, this carbon dioxide was not removed from the atmosphere and so continued to build up over millions of years. Eventually, the additional warming caused by the increase in greenhouse gases overcame the cooling caused by the snow's high albedo, and temperatures rose enough to melt the ice, freeing the Earth.

For most of Earth's history, carbon dioxide concentrations have been higher than they are today. As a consequence, past climates have often been very warm. During the late stage of the dinosaur era (the cretaceous era, a period that lasted between 65 and 145 million years ago), carbon dioxide levels were about 5 times higher than they are today, and the average global temperatures were more than 10 ^{o}C higher than today's. There were no large ice sheets, and dinosaur fossils from this period have been found as far north as Alaska. These animals would not survive the cold conditions found in the arctic today. Further south, fossil crocodiles from 60 million years ago have been found in North Dakota. The modern average winter temperature in North Dakota is around -10 ^{o}C –but being cold-blooded, crocodiles are most at home when the air temperature is around 30 ^{o}C! The climate was warmer in the past when the amount of carbon dioxide was higher.

Resources

The National Aeronautical and Space Administration (NASA) Earth Observatory website has an array of climate resources. For a more in-depth discussion of Earth's energy budget, go to http://earthobservatory.nasa.gov/Features/EnergyBalance/

Are you interested in finding more about the controversial Snowball Earth hypothesis? The National Science foundation and Harvard University have set up a website with more about the hypothesis and the evidence. Go to http://www.snowballearth.org/

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11.7: Quaternary Climate — Information From Ice Cores

Learning Objectives

After reading this module, students should be able to

- describe the changing climate of the Quaternary
- explain why Milankovitch cycles explain the variations of climate over the Quaternary, in terms of the similar periods of orbital variations and glacial cycles
- explain how the glacier/climate system is linked via albedo feedbacks
- describe how sediment and ice cores provide information about past climates
- use the mechanisms that cause stable isotope fractionation to predict the impact of changing climate on stable isotope records

Introduction

We saw the major drivers of the climate—the energy that comes from the Sun (insolation) and the properties of the planet that determine how long that energy stays in the Earth system (albedo, greenhouse gases). In this section, we will look at the recent natural changes in Earth's climate, and we will use these drivers to understand why the climate has changed.

The most recent period of Earth's geologic history—spanning the last 2.6 million years—is known as the quaternary period. This is an important period for us because it encompasses the entire period over which humans have existed—our species evolved about 200,000 years ago. We will examine how the climate has changed over this period in detail. By understanding recent natural processes of climate change, we will be able to better understand why scientists attribute the currently observed changes in global climate as being the result of human activity.

Quaternary Climate — Information From Ice Cores

How do we know about the Quaternary climate? After all, most of the period predates human existence, and we have only been recording the conditions of climate for a few centuries. Scientists are able to make informed judgments about the climates of the deep past by using proxy data. Proxy data is information about the climate that accumulates through natural phenomena. In the previous module, for example, we discussed how 60-million-year-old crocodile fossils have been found in North Dakota. This gives us indirect information about the climate of the period—that the climate of the region was warmer than it is today. Although not as precise as climate data recorded by instruments (such as thermometers), proxy data has been recovered from a diverse array of natural sources, and provides a surprisingly precise picture of climate change through deep time.

One highly detailed record of past climate conditions has been recovered from the great ice sheets of Greenland and Antarctica. These ice sheets are built by snow falling on the ice surface and being covered by subsequent snowfalls. The compressed snow is transformed into ice. It is so cold in these polar locations that the ice doesn't melt even in the summers, so the ice is able to build up over hundreds of thousands of years. Because the ice at lower depths was produced by progressively earlier snowfalls, the age of the ice increases with depth, and the youngest ice is at the surface. The Antarctic ice sheet is up to three miles thick. It takes a long time to build up this much ice, and the oldest ice found at the bottom of the Antarctica ice sheet is around 800,000 years old.

Scientists drill into these ice sheets to extract *ice cores*, which record information about past climates. Figure 11.7.1 shows what these cores look like when they are cut open. Like tree rings, ice cores indicate years of growth. Note how the middle core (which required over a mile of drilling to extract!) has distinct layers—this is because the seasons leave an imprint in the layers of snow. Scientists can use this imprint to help calculate the age of the ice at different depths, although the task becomes more difficult the deeper the core sample, since the ice layers become more compressed. The ice records several different types of climate information: the temperature of the core, the properties of the water that make up the ice, trapped dust, and tiny entombed bubbles of ancient atmosphere.

Firgure 11.7.1 Ice cores. Three different sections of an ice core. The seasonal layers are most clear in the middle section (note the dark and light bands). The deepest section (bottom core) is taken from almost two miles down and is colored brown by rocky debris from the ground under the ice. *Source: National Ice Core Laboratory*

The water molecules that make up the ice record information about the temperature of the atmosphere. Each water molecule is made up of two hydrogen atoms and one oxygen atom (and so has the chemical name H_2O). Not all oxygen atoms are the same however; some are "light" and some are "heavy". These different types of oxygen are called isotopes, which are atoms that have same number of protons but different numbers of neutrons. The heavy isotope of oxygen (oxygen-18, or ^{18}O) is more than 10% heavier than the light isotope (oxygen-16 or ^{16}O). This means that some water molecules weigh more than others. This is important because lighter water molecules are more easily evaporated from the ocean, and once in the atmosphere, heavier water molecules are more likely to condense and fall as precipitation. As we can see from Figure 11.7.2, the water in the ice sheets is lighter (has a higher proportion of ^{16}O relative to ^{18}O) than the water in the oceans.

The process of differentiation between heavy and light water molecules is temperature dependent. If the atmosphere is warm, there is more energy available to evaporate and hold the heavier ^{18}O water in the atmosphere, so the snow that falls on the polar ice sheets is relatively higher in ^{18}O. When the atmosphere is cold, the amount of energy is less, and so less ^{18}O makes it to the poles to be turned into glacial ice. We can compare the amount of ^{18}O in different parts of the ice core to see how the atmosphere's temperature—the climate—has changed.

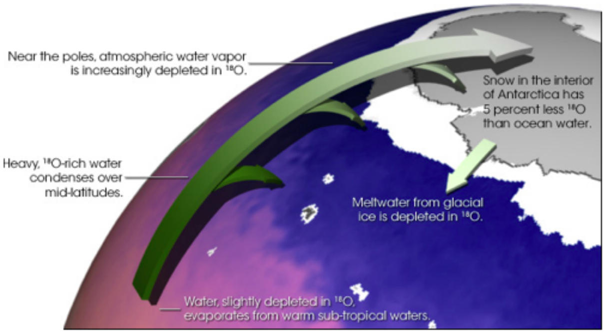


Figure 11.7.2 Water becomes lighter as it travels toward the poles. The heavy (18O) water drops out of the atmosphere (as rain or snow) before reaching the ice sheet. This means that the snow that forms the glacial ice is lighter than the ocean water (has more 16O than 18O, compared to ocean water). *Source:* [Robert Simmon, NASA GSFC, NASA Earth Observatory]

Figure 11.7.3 shows what this record looks like over the last 400,000 years. The blue and green lines depict two different Antarctic ice cores (taken from ice about 350 miles apart) and the variations in oxygen isotopes are converted into temperature changes. The y-axis shows temperature change; today's climate is at zero—the dashed line. Notice that the Earth's climate has not been stable! Sometimes the temperature is higher than it is today—the blue and green lines are higher than the dashed about 120,000 years ago, for example. Most of the time the climate is much colder than today's, however: the most common value is around -6 °C (-13 °F). On average, the earth's temperature between 25,000 and 100,000 years ago was about 6 °C lower than it is today. These changes can be double-checked by measuring the temperature of the ice in the cores directly. Ice that is 30,000 years old is indeed colder than the ice made today, just as the isotope data predicts.

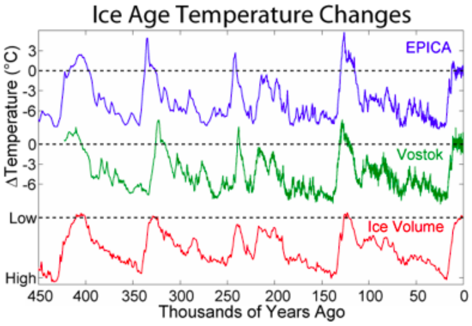


Figure 11.7.3 Ice age temperature. The blue and green lines depict two different Antarctic ice cores (taken from ice about 350 miles apart) and the variations in oxygen isotopes are converted into temperature changes. The red line depicts global ice volume. The y-axis shows temperature change; today's climate is at zero – the dashed line. *Source:* Robert A. Rohde

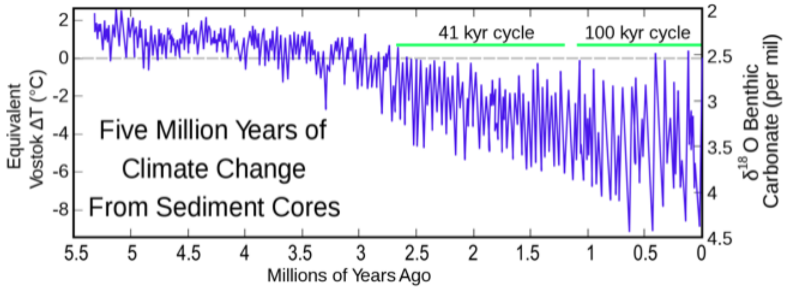


Figure 11.7.4 A comparison of the age of sediment (x-axis) and the change in temperature over time (left y-axis) as derived from oxygen isotope ratios (right y-axis). The dashed line shows today's climate. Note that the climate is cooling over the last few million years, but it is highly variable. In the last one million years the climate alternates between warm and cool conditions on a 100,000-year time scale ("100 kyr cycle"), before this it alternated on a 41,000 year cycle. Both these period lengths are the same as Milankovitch cycles. These cores suggest that today's temperature is higher than almost all of that of the Quaternary (the last 2.6 Million years). *Source: Jo Weber*

The changes in climate recorded in ice sheets are thought to be worldwide. The same climate changes observed in Antarctica are also found in cores taken from Greenland, which is on the other side of the Earth. Isotope data can also be taken from sediment cored from the ocean floor—all over the planet—and these cores also show the same changes in climate, alternating between cold and warm. Because ocean sediment is deposited over millions of years, the sediment can give an indication of the climate across the whole of the Quaternary and beyond. Figure 11.7.4 shows how temperature has changed over time (blue line), compared with today (dashed line). The temperature has, on average, gotten colder over the Quaternary, but it also appears to oscillate between warm and cold periods. We'll investigate these periodic changes in the next section of this chapter.

As falling snow accumulates on the ground, tiny bubbles of air become trapped in it. These bubbles are retained as the snow transforms to ice, and constitute tiny samples of the ancient atmosphere that can be analyzed to find out if the changes in temperature (as recorded in the oxygen isotopes) are related to changes in the atmosphere. The temperature recorded by the isotopes in the ice is directly related to the amount of carbon dioxide in the trapped air (Figure 11.7.5): the times with higher carbon dioxide are also times of high temperature.

Falling snow also captures and entombs atmospheric dust, which is topsoil born aloft by the wind, and which is especially prevalent during droughts. The fact that more dust occurs in the ice accumulated during cold periods suggests that the glacial climate was dry, as well as cold.

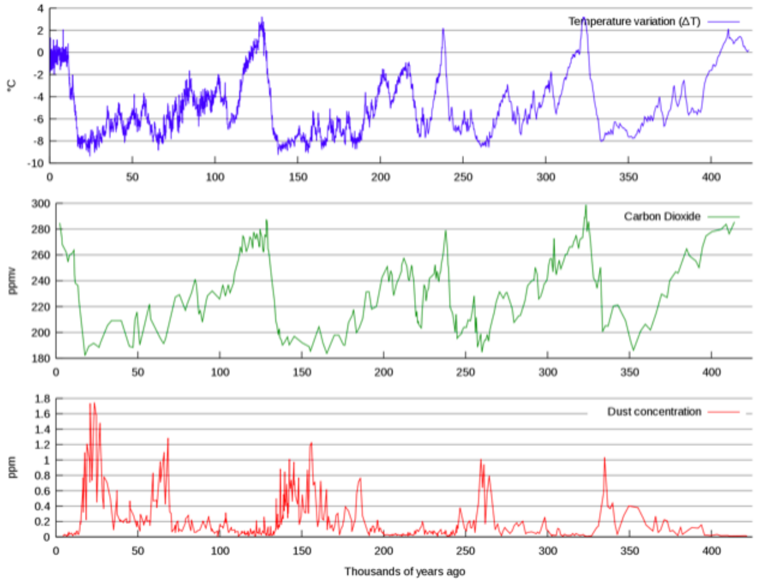


Figure 11.7.5 Vostok Petit Data. These graphs depict how changes in temperature—inferred from changes in isotope ratios (blue line)--correspond to changes in atmospheric carbon dioxide (green line) and dust (red line) over the last 400,000 years as recorded in an ice core extracted from Antarctica. Carbon dioxide varies directly with temperature – the warmer the climate the higher the carbon dioxide level. Atmospheric dust is highest during the coolest periods (such as 25,000 and 150,000 years ago). *Source: William M. Connolley produced figure using data from the National Oceanic and Atmospheric Administration, U.S. Department of Commerce, Paleoclimatology branch, Vostok Ice Core Data.*

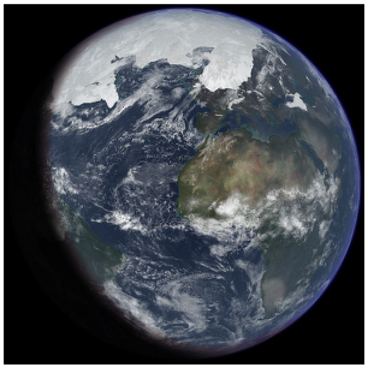


Figure 11.7.6 Ice Age Earth- an artist's impression of the Earth during an ice age. Note that the Northern parts of North America and Europe (including Canada and Scandinavia) are entirely covered by ice-sheets. *Source: Itti*

During the Quaternary, the Earth has cycled between glacial periods (sometimes referred to as "ice ages") and interglacial periods. The ice was at its most recent extreme around 20,000 years ago in a period known as the last glacial maximum, or LGM. As we can see from the ice core record, the Quaternary climate is usually cold (see Figure 11.7.3), with long periods of cold punctuated with shorter (10,000 year long, or so) periods of warmer conditions, like those we experience today. In many ways, our current climate is exceptional—for most of human existence, the Earth has been a much colder place.

What was the Earth like during these glacial periods? Almost all the world was cold; average temperatures were around 6 ^oC (-13 ^oF) colder than today. Such conditions allow ice sheets to grow—much of North America, Asia and Europe were covered under mile-thick ice (see Figure 11.7.3). Because this ice was made of water that was once in the oceans, sea levels were much lower. At the LGM, sea level was about 120 meters (or about 400 feet) lower than it is today. As the seas retreated, the continents grew larger, creating land bridges that joined Asia with North America, Britain with Europe, and Australia with Papua New Guinea.

During glacial periods the climate was also much drier, as evidenced by the increase in atmospheric dust (Figure 11.7.5). The lands at and near the poles were covered with ice, and dry grasslands occupied areas where temperate forests occur today. Deserts were much larger than they are now, and tropical rainforests, having less water and less warmth, were small. The animals and plants of glacial periods were different in their distribution than they are today, as they were adapted to these different conditions. Fossils of Mastodons (Figure 11.7.7) have been found from all across what is now the United States, including from Florida, which currently enjoys a subtropical climate.

Figure 11.7.7 An artist's impression of a Knight Mastodon, an elephant-like mammal with a thick wooly coat. Mastodon fossils dating from past glacial periods have been found across North America—from Florida to Alaska. *Source: Charles R. Knight*

During glacial periods humans would have been unable to occupy the globe as they do today because all landmasses experienced different climactic conditions. Some countries of the present could not exist, as they would be almost completely covered by ice. As examples, look for Canada, Iceland and The United Kingdom in Figure 11.7.9.

Milankovitch Cycles

Why has the Earth cycled through hot and cold climates throughout the Quaternary? As we learned in the previous module, the Earth's climate is controlled by several different factors—insolation, greenhouse gases, and albedo are all important. Scientists believe that changes in insolation are responsible for these climate swings, and the insolation varies as a result of wobbles in the Earth's orbit.

The Earth's orbit is not fixed – it changes regularly over time. These periodic changes in Earth's orbit named are referred to as Figure 11.7.8, and are illustrated in Figure 11.7.8. Changes in the Earth's orbit alter the pattern of insolation that the Earth receives. There are three principle ways in which the Earth's orbit varies:

1. Eccentricity (or *Orbital shape*): The Earth's orbit is not perfectly circular, but instead follows an ellipse. This means that the Earth is, through the course of the year, sometimes closer and sometimes further away from the Sun. Currently, the Earth is closest to the Sun in early January, and furthest from the Sun in Early July. This changes the amount of insolation by a few percent, so Northern Hemisphere seasons are slightly milder than they would be if the orbital was circular (cooler summers and warmer winters). The orbital shape changes over time: the Earth moves between being nearly circular and being mildly elliptical. There are two main periods over which this change occurs, one takes around 100,000 years (this is the time over which the orbit goes from being circular, to elliptic, and back to circular), another takes around 400,000 years.
2. Axial Tilt (or Obliquity): The Earth axis spins at an angle to its orbit around the Sun – currently this angle is 23.5 degrees (this angle is known as the *axial tilt*). This difference in orbit creates the seasons (as each hemisphere takes turns being tilted towards and away from the Sun over the course of the year). If the axis of spin lined up with the direction of the Earth's orbit (so that the tilt angle was zero) there would be no seasons! This axial tilt also changes over time, varying between 22.1 and 24.5 degrees. The larger the angle, the larger the temperature difference between summer and winter. It takes about 41,000 year for the axial tilt to change from one extreme to the other, and back again. Currently, the axial tilt is midway between the two extremes and is decreasing—which will make the seasons weaker (cooler summers and warmer winters) over the next 20,000 years.
3. Axial Precession: The direction of Earth's axis of rotation also changes over time relative to the stars. Currently, the North Pole points towards the star Polaris, but the axis of rotation cycles between pointing to that star and the star Vega. This impacts the Earth's climate as it determines when the seasons occur in Earth's orbit. When the axis is pointing at Vega, the Northern Hemisphere's peak summer is in January, not July. If this were true today, it would mean that the Northern Hemisphere would experience more extreme seasons, because January is when the Earth is closest to the Sun (as discussed above in eccentricity). This cycle takes around 20,000 years to complete.

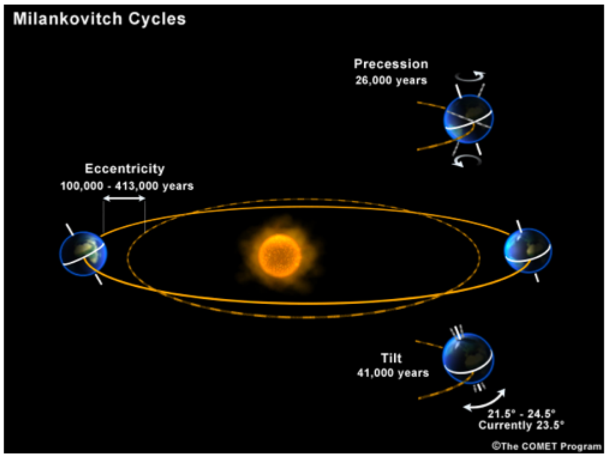


Figure 8. Milankovitch Cycles Illustration of the three variables in Earth's orbit, with periods of variation marked. *Source:* *COMET® at the University Corporation for Atmospheric Research (UCAR) pursuant to a Cooperative Agreements with the National Oceanic and Atmospheric Administration, U.S. Department of Commerce. ©1997-2009 University Corporation for Atmospheric Research. All Rights Reserved.*

The three cycles described above have different periods, all of which are long by human standards: 20,000, 40,000, 100,000 and 400,000 years. If we look at the temperature data from ice and sediment cores, we see that these periods are reflected in Earth's climate. In the last million or so years, the 100,000-year eccentricity in the orbit has determined the timing of glaciations, and before that the 40,000-year axial tilt was dominant (Figure 11.7.8). These cycles have been important for a long time; geologists have even found evidence of these periods in rocks that are hundreds of millions of years old.

But how do the Milankovitch Cycles change our climate? These orbital cycles do not have much impact on the *total* insolation the Earth receives: they change only the *timing* of that insolation. Since the total insolation does not change, these orbital variations have the power to make the Earth's seasons stronger or weaker, but the average annual temperature should stay the same. The best explanation for long term changes in average annual temperature is that the Milankovitch cycles initiate a positive feedback that amplifies the small change in insolation.

Insolation and the Albedo Feedback

Today, the Earth's orbit is not very eccentric (it is almost circular), but at the beginning of each of the recent ice age periods, the orbit was much more elliptical. This meant that the Earth was further away from the sun during the northern hemisphere summers, reducing the total insolation. Lower insolation meant that the summer months were milder than they would otherwise be, with cooler temperatures. Summer temperatures were also lower when the Earth's axial tilt was smaller, so the two different orbital parameters could reinforce one another's effects, in this case producing especially mild summers.

It is thought that these mild northern summers produced an albedo feedback that made the whole planet slip into an ice age. The northern hemisphere has continents near the poles—Europe, Asia, and North America. Today, these continents have largely temperate climates. During the winter, snow falls across much of the land (see Figure $11.7.9$ in the previous module) only to melt during the summer months. If the summers are not hot enough to melt all the snow and ice, glaciers can advance, covering more of the land. Because ice has a high albedo, more sunlight is reflected than before, and the Earth is made cooler. This creates a positive feedback, as the cooler conditions allow the ice to advance further—which, in turn, increases the albedo and cools the Earth! Eventually, a large proportion of the northern continents became covered in ice (Figure $11.7.9$).

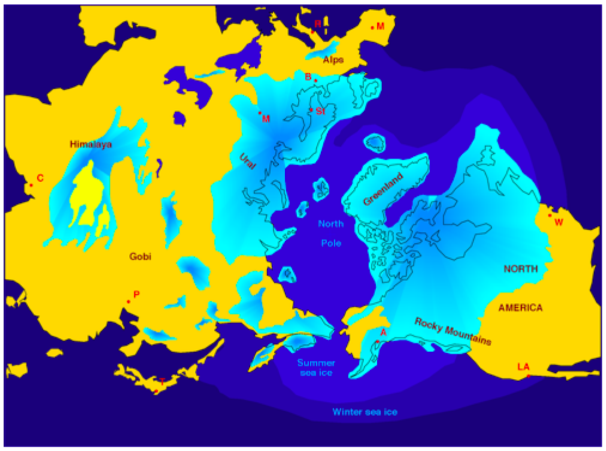


Figure 11.7.9 Glacial coverage (light blue) of the northern hemisphere during the ice ages. *Source:* Hannes Grobe

This positive feedback process works in the other direction, as well. The interglacial periods are ushered in when the orbital parameters create summers that are unusually warm, which melts some of the ice. When the ice sheets shrink, the Earth's albedo decreases, which further warms the system. The giant northern ice sheets shriveled up in a few thousand years as warm summers and decreasing albedo worked together.

These cycles of alternating cooling and warming are also related to changes in the amount of greenhouse gases in the atmosphere. As we observed in Figure $11.7.5$, the climate contains higher levels of carbon dioxide during interglacial periods. Although this appears to make sense—carbon dioxide is a greenhouse gas, and so should produce warmer climates—it is also a puzzle, because it is not clear how changes in Milankovitch cycles lead to higher levels of carbon dioxide in the atmosphere. It is clear that these changes in carbon dioxide are important in making the change in temperature between interglacial and glacial periods so extreme. Several different hypotheses have been proposed to explain why glacial periods produce lower levels of carbon dioxide (it may be related to how the physical changes influence the Earth's ecosystems ability to absorb carbon dioxide: perhaps lower sea levels increase the nutrient supply in the ocean, or the drop in sea level destroys coral reefs, or iron-rich dust from new deserts fertilizes the oceans) but further work on this question remains to be done.

It is a concern for all of us that there are gaps in our understanding of how the feedbacks between insolation, albedo and greenhouse gases operate, as it makes it hard to predict what the consequences of any changes in the climate system might lead to. The current level of atmospheric carbon dioxide is unprecedented in human experience; it is at the highest level ever recorded in the Quaternary. Will the current increase in greenhouse gases lead to a positive feedback, warming the Earth even more?

Review Questions

1. In the text, we discuss how polar ice has a smaller ^{18}O to ^{16}O ratio (that is, it has proportionally less heavy isotope water) than ocean water does. Hydrogen also has isotopes, the two most common being hydrogen-1 (^{1}H) and hydrogen-2 (^{2}H, also known as deuterium). Water is made up of both hydrogen and oxygen, and scientists analyze both elements when examining ice cores.

Do you predict that polar ice sheets would have a higher ratio or a lower ratio of ^{1}H to ^{2}H than ocean water? Will colder global temperatures increase or decrease the amount of ^{2}H in polar ice?

2. In the text, we discuss how polar ice has a smaller ^{18}O to ^{16}O ratio (that is, it has proportionally less heavy-isotope water) when the climate is cooler. We also discuss how changes in the ratio of ^{18}O to ^{16}O ratio in sediment cores can also be used to determine the climate's average temperature. In ocean sediments, the ratio of ^{18}O to ^{16}O increases when the climate is cooler (that is, it has proportionally more heavy isotope water). Explain why isotope ratios in ocean sediment have the opposite reaction to those in polar ice.
3. There are three different ways in which the Earth's orbit changes through time. What combination of orbital parameters would be most likely to start an ice age? (Hint: Ice ages require cool northern summers.)

Resources

Do you want to know more about how ice cores are extracted and analyzed? NASA's Earth Observatory has details about the practical issues of drilling ice cores (deep ice needs to "relax" for as long as a year at the surface before being cut open – or it can shatter!) and how chemical data is interpreted. Go to http://earthobservatory.nasa.gov/Features/Paleoclimatology_IceCores/ for an in-depth article with great links.

11.8: 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.

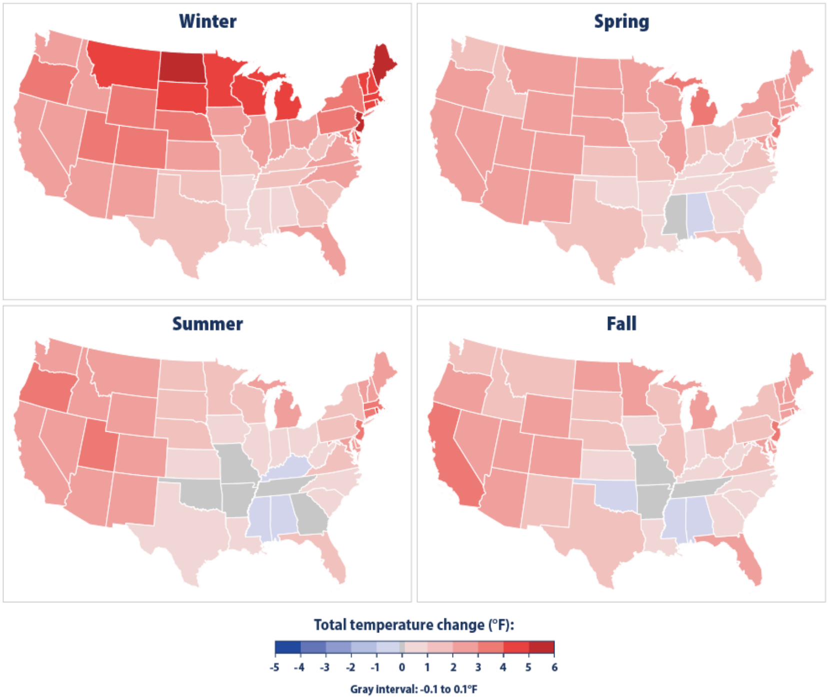
As an example, heat energy from sunlight and from hot masses of air is converted to latent heat by the evaporation of ocean water off the west coast of Africa. Prevailing winds drive masses of air laden with water vapor westward across the ocean. Rainfall forms, releasing the energy from the latent heat 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.

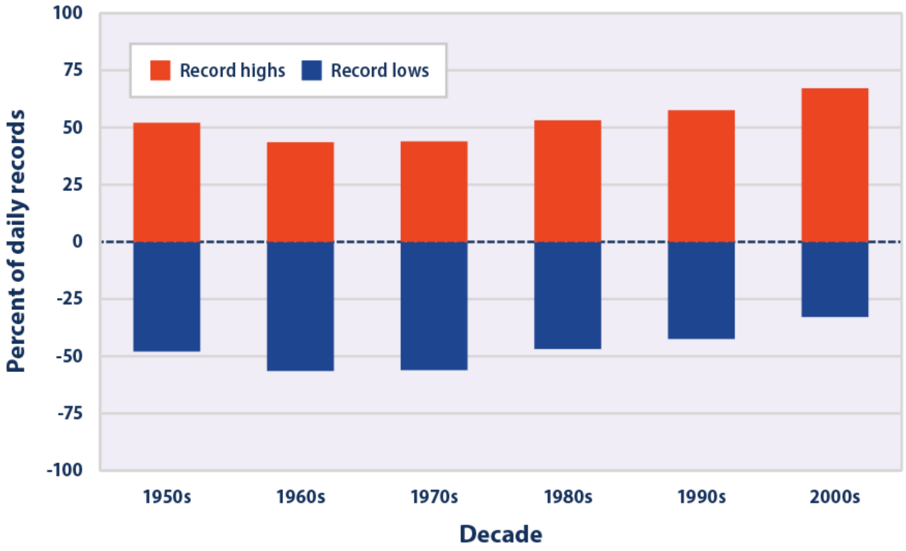
When the long-term trends in weather begin to significantly shift, the climate is said to change. Of all the things that humans may be doing that could change climate, emissions of large quantities of gases that cause atmospheric warming are the most serious. Although there could be some benefits of 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 $1.5–5^{\circ}$ C by 2100 (Video11.8.1). 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 11.8.1: 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 11.8.1 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 11.8.1). 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


100
75
50
25
0
-25
-50
-75
-100
Percent of daily records
🟥 Record highs
🟦 Record lows
1950s
1960s
1970s
1980s
1990s
2000s
Decade

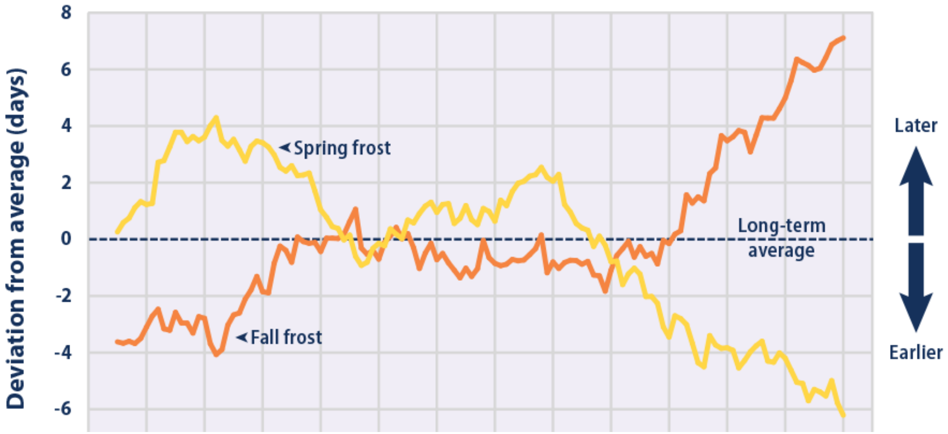

8
6
4
2
0
-2
-4
-6
Deviation from average (days)
Spring frost
Fall frost
Long-term average
Later
Earlier

Figure 11.8.1: 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 11.8.2, 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?

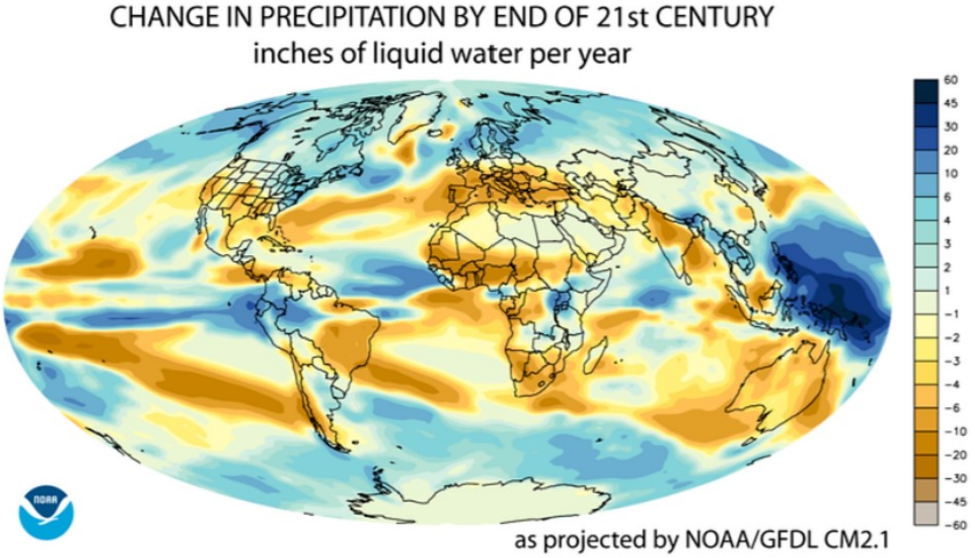


Figure 11.8.2. The plotted precipitation differences were computed as the difference between the 2081 to 2100 20-year averages minus the 1951 to 2000 50-year average. Blue areas project increases in precipitation; brown areas project decreases.

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 $11.8.3$).

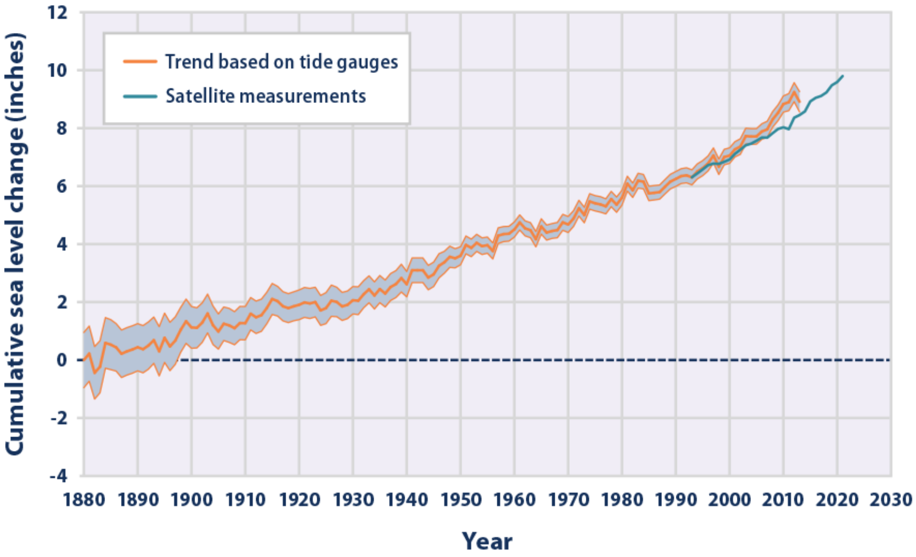


Figure $11.8.3$: Cumulative changes in sea level for the world's oceans since 1880. The sea level change refers to the height of the ocean surface, regardless of whether nearby land is rising or falling. The shaded band shows the likely range of values based on the number of measurements collected and the precision of the methods used. (NOAA via epa.gov)

From the data in Figure $11.8.3$, 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 $11.8.4$) 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 11.8.4: 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 11.8.5 shows the relative rise in sea level along the coasts of the US. The change is particularly drastic along the eastern coast. Figure 11.8.6 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.

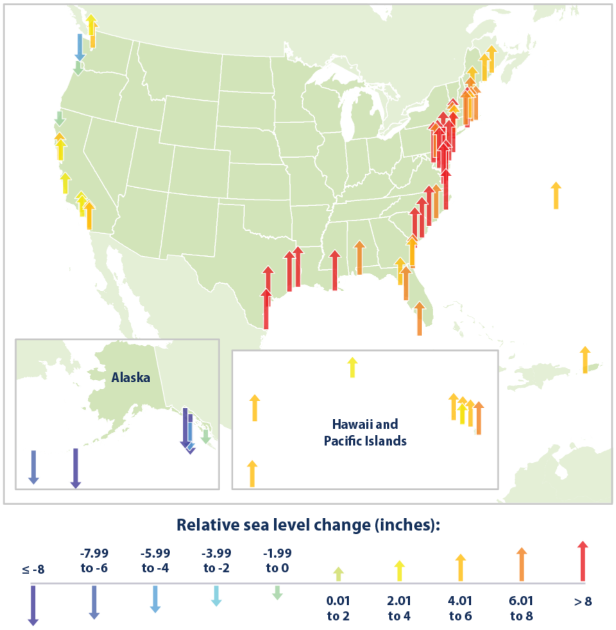


Figure 11.8.5: Cumulative changes in relative sea level from 1960 to 2021 at tide gauge stations along the US coast. Relative sea level reflects changes in sea level as well as land elevation. (NOAA via epa.gov)

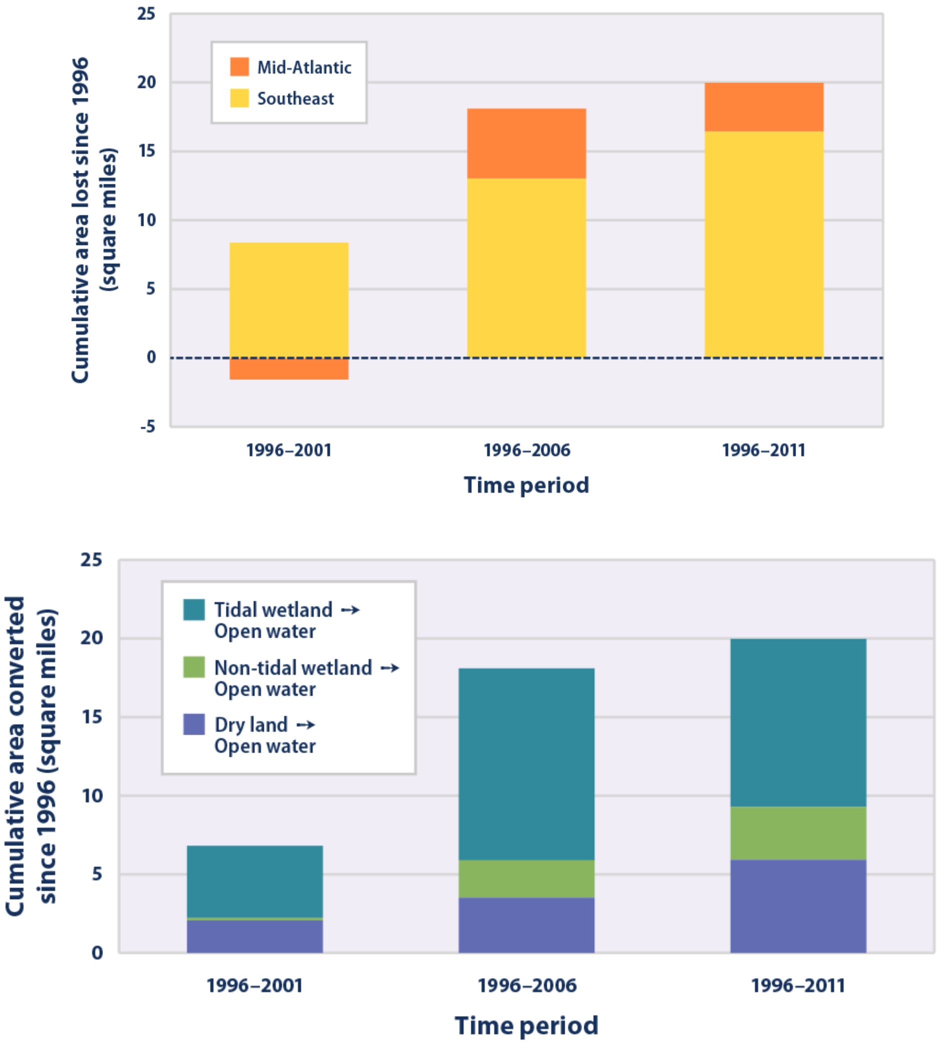


Figure 11.8.6: The first graph shows the net amount of land converted to open water along the Atlantic coast during three time periods. The second graph breaks down the types of land that have been converted. (NOAA via epa.gov)

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 11.8.1 and 11.8.2).

$$\mathrm{CO_2} + \mathrm{H_2O} \leftrightarrow \mathrm{H^+} + \mathrm{HCO_3^-} \tag{11.8.1}$$

$$\mathrm{HCO_3^-} \leftrightarrow \mathrm{H^+} + \mathrm{CO_3^{2-}} \tag{11.8.2}$$

As you can see, both Equations 11.8.1 and 11.8.2 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 11.8.7.

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/.

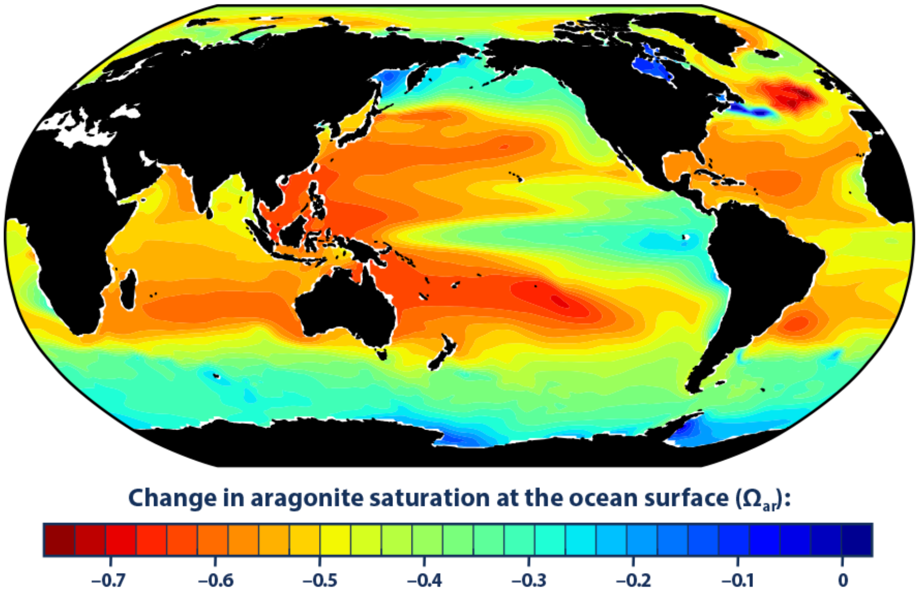


Figure 11.8.7: This map shows changes in the aragonite saturation level of ocean surface waters between the 1880s and the most recent decade (2006–2015). Aragonite is a form of calcium carbonate that many marine animals use to build their skeletons and shells. The lower the saturation level, the more 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.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)

- Libretext: Introduction to Environmental Science (Zendher et al.)
- US EPA
- National Oceanic and Atmospheric Administration

- Stanley E. Manahan (University of Missouri)
- Vicki MacMurdo (Anoka-Ramsey Community College)

11.9: Looking Forward- Climate 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 11.9.1. 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.

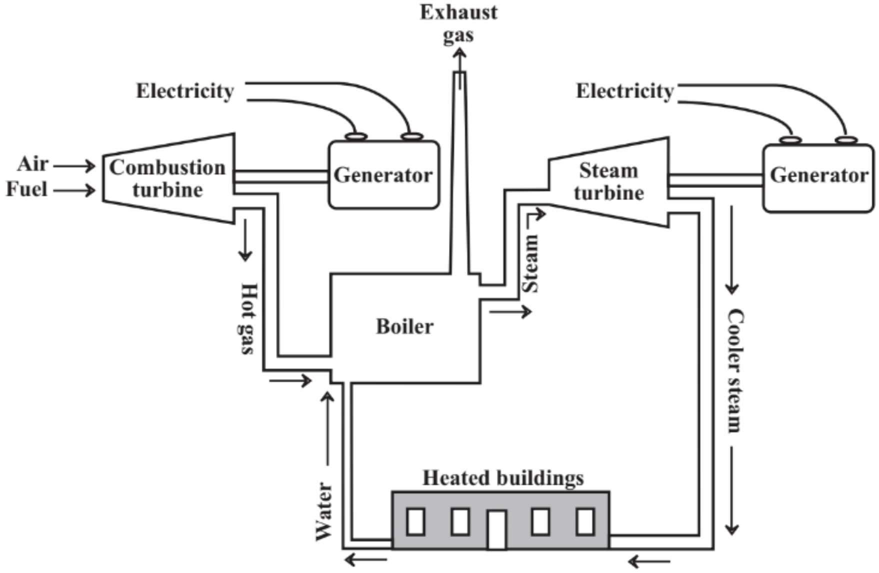


Figure 11.9.1: A combined power cycle in which natural gas or oil is first used to fire a gas turbine connected to an electrical generator. The hot exhaust gases from this turbine are fed to a boiler to produce steam, which drives a steam turbine, also connected to a generator. The still hot exhaust 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 11.9.2, 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.

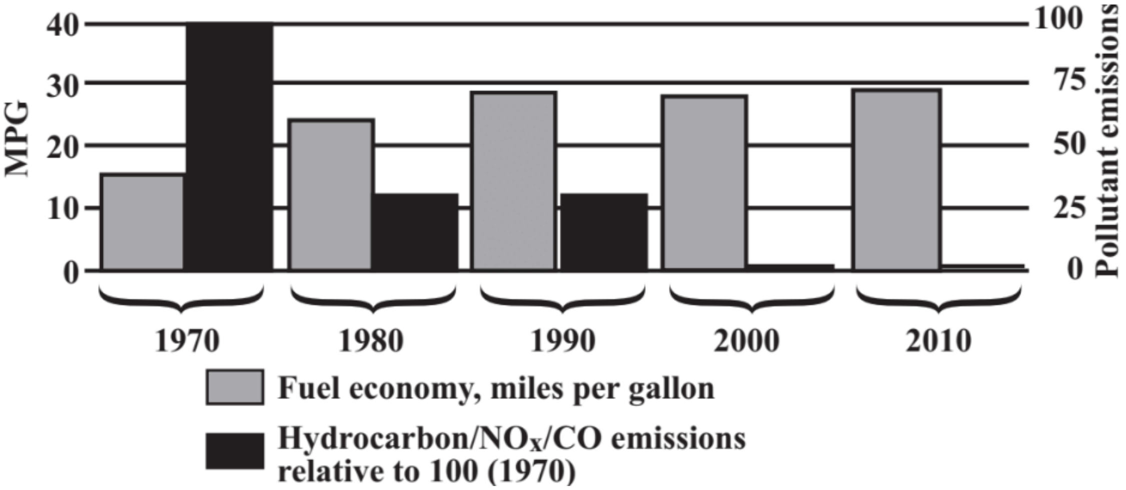


Figure 11.9.2: U.S. auto fleet fuel economy and emissions over 4 decades, showing a trend of generally higher fuel economy with greatly reduced emissions.

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 11.9.3). 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.

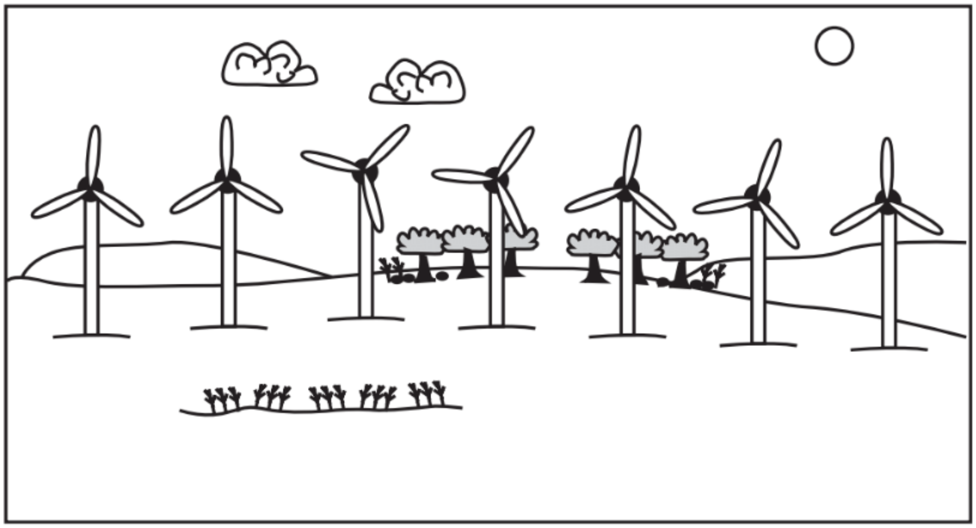


Figure 11.9.3: Wind-powered electrical generators mounted on towers are becoming increasingly common sights in the world in areas where consistent wind makes this nonpolluting source of renewable energy practical.

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 $11.9.4$).

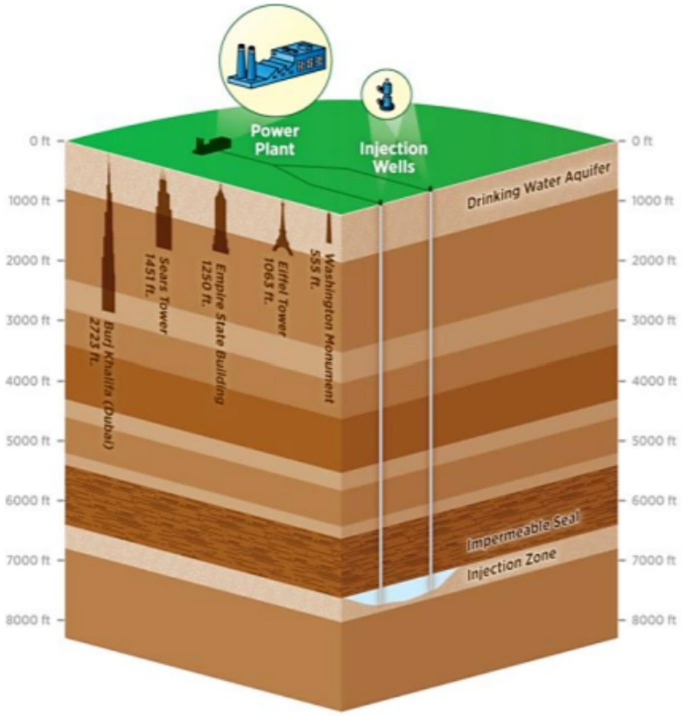


Figure 11.9.4: Carbon capture and sequestration schematic with landmarks shown to scale for depth reference. (US EPA)

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:

$$2\ C_{(s)} + O_{2(g)} + 2\ H_2O_{(g)} \rightarrow 2\ CO_{2(g)} + 2H_{2(g)} \tag{11.9.1}$$

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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Index

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Sample Word 1 | Sample Definition 1

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