

# Exploring the Physical World: An Updated and Expanded Introduction To the Physical Sciences



# EXPLORING THE PHYSICAL WORLD: AN UPDATED AND EXPANDED INTRODUCTION TO THE PHYSICAL SCIENCES

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# INTRODUCTION

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## About This Book

This textbook was created through Connecting the Pipeline: Libraries, OER, and Dual Enrollment from Secondary to Postsecondary, a \$1.3 million project funded by [LOUIS: The Louisiana Library Network](#) and the [Institute of Library and Museum Services](#). This project supports the extension of access to high-quality post-secondary opportunities to high school students across Louisiana and beyond by creating materials that can be adopted for dual enrollment environments. Dual enrollment is the opportunity for a student to be enrolled in high school and college at the same time.

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## Cover Image

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## Introduction

This is a physical science text intended for non-science majors that covers introductory chemistry and physics topics. Below are descriptions of the content found under different headings throughout the text, and how it is recommended that the reader approaches these sections.

**Link To Learning** – Sections that begin with this subtitle include links to simulations or websites that are included to apply and reinforce your knowledge of the topic. It is recommended that you follow directions (if given) on what to try within the simulations, but also that you investigate and explore the simulations on your own. Try to find concepts in the simulation that make more sense after reading about the physical or chemical science behind it.

**Examples** – Most chapters have at least one example problem in which the solution is provided step-by-step along with the basic concepts that will assist in solving similar problems.

**Misconception Alert** – These sections provide clarification of common misunderstandings related to the topic.



**Take-Home Experiments Or Everyday Life Activities –**

These at-home learning opportunities relate the scientific concepts presented to observations and experiences in daily life.

**Check Your Understanding** – Sections that begin with this subtitle are meant to gauge your understanding of the material.



1.

# THE NATURE OF SCIENCE

## Scientific Method and Measurement

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### Chapter Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

- List and Use the steps of the Scientific Method(CLO2)(CLO4)
- Estimate Physical Parameters (CLO2)(CLO3)(CLO5)
- Solve simple problems using derived quantities (CLO2)(CLO5)
- Recognize and use SI Base Units and Prefixes

(CLO2)(CLO5)

- Convert from one unit system to another.

(CLO2)(CLO5)

- Use dimensional analysis to check the consistency of your work. (CLO5)

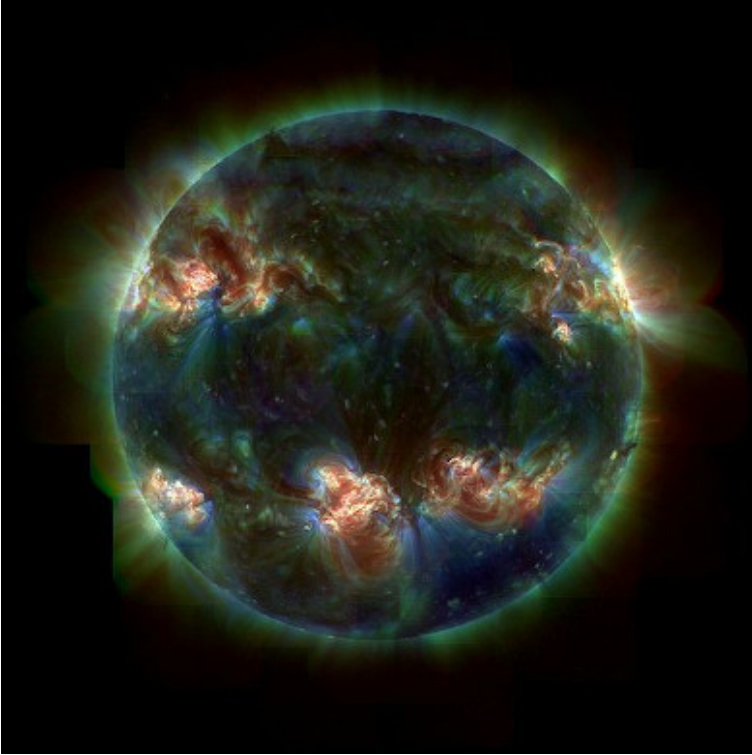


Fig 1.1 An image of the Sun taken in UV (Ultraviolet light) taken by the Transition Region and Coronal Explorer (TRACE) and combined with the data from the Stanford Lockheed Institute for Space Research.

## The Nature of Science

Science is an overarching category of many different disciplines that involve hypothesizing, observation, testing and evidence. It is an ever-changing and fluid field because regardless of the results of research or experimentation, there will always be

more questions to ask. Some examples of scientific fields are Biology, Botany, Psychology, Chemistry, Astronomy and Physics.

This book deals with the physical sciences – physics and chemistry. All the sciences are based in the use of **experiment** and testing to understand the world around us better. The **scientific method** requires us to constantly re-examine our understanding, by testing new evidence with our current **theories** and making changes to those theories if the evidence does not meet the test. The scientific method therefore is the powerful tool you will use throughout the physical sciences.

In this chapter you will learn how to gather evidence using the **scientific method**. These skills will then be used throughout this textbook to test scientific theories and practices.

## Development of a Scientific Theory

The most important, and most exciting, thing about science and **scientific theories** is that they are not fixed. Hypotheses are formed and carefully tested, leading to scientific theories that explain those **observations** and predict results. The results are not made to fit the hypotheses. If new information comes to light with the use of better equipment, or the results of other experiments, this new information is used to improve and expand current theories. If a **theory** is found to have been

inaccurate it is changed to fit this new information. The data should never be made to fit the theory, if the data does not fit the theory then the theory is reworked or discarded. Although this changing of opinion is often taken for inconsistency, it is this very willingness to adapt that makes science useful, and allows new discoveries to be made.

Remember that the term **theory** has a different meaning in science. A **scientific theory** is not like your theory of about why you can only ever find one sock. A scientific theory is one that has been tested and proven through repeated experiment and data. Scientists are constantly testing the data available, as well as commonly held beliefs, and it is this constant testing that allows progress and improved theories.

## Gravity

The theory of gravity has been slowly developing since the beginning of the 16th century. Galileo Galilei is credited with some of the earliest work. At the time it was widely believed that heavier objects accelerated faster toward the earth than light objects did. Galileo had a **hypothesis** that this was not true and performed **experiments** to prove this.

Galileo's work allowed Sir Isaac Newton to hypothesize not only a theory of gravity on earth, but that gravity is what held the planets in their orbits. Newton's theory was used by John Couch Adams and Urbain Le Verrier to predict the planet Neptune in the solar system and this prediction was proved

experimentally when Neptune was discovered by Johann Gottfried Galle.

Although a large majority of gravitational motion could be explained by Newton's theory of gravity, there were things that did not fit. But although a newer theory that better fit the facts was eventually proved by Albert Einstein, Newton's gravitational theory is still successfully used in many applications where the masses, speeds and energies are not too large.

## Thermodynamics

The principles of the three rules of thermodynamics describe how energy works, on all size levels (from the workings of the Earth's core, to a car engine). The basis for these three rules started as far back as 1650 with Otto von Guericke. He had a **hypothesis** that a vacuum pump could be made, and proved this by making one. In 1656 Robert Boyle and Robert Hooke used this information and built an air pump.

Boyle's **law** came about from his air pump **experiments**, where he discovered that pressure is inversely proportional to volume at a constant temperature ( $p \propto 1/V \propto 1/V$  at constant  $T$ ).

Over the next hundreds of years the **theory** was expanded on and improved. Denis Papin built a steam pressurizer and release valve, and designed a piston cylinder and engine, which Thomas Savery and Thomas Newcomen built. These engines inspired the study of heat capacity and latent heat. Joseph



Black and James Watt increased the steam engine efficiency and it was their work that Sadi Carnot (considered the father of thermodynamics) studied before publishing a discourse on heat, power, energy and engine efficiency in 1824.

This work by Carnot was the beginning of modern thermodynamics as a science, with the first thermodynamics textbook written in 1859, and the first and second **laws** of thermodynamics being determined in the 1850s. Scientists such as Lord Kelvin, Max Planck, J. Willard Gibbs among many others studied thermodynamics. Over the course of 350 years thermodynamics has developed from the building of a vacuum pump to some of the most important fundamental **laws** of energy.

## Scientific method

The **scientific method** is the basic skill process in the world of science. Since the beginning of time humans have been curious as to why and how things happen in the world around us. The scientific method provides scientists with a well-structured scientific platform to help find the answers to their questions. Using the scientific method there is no limit as to what we can investigate. The scientific method can be summarized as follows:

1. Ask a question about the world around you.
2. Do background research on your questions.

3. Make a **hypothesis** about the event that gives a sensible result. You must be able to test your hypothesis through **experiment**.
4. Design an experiment to test the hypothesis. These methods must be repeatable and follow a logical approach.
5. Collect data accurately and interpret the data. You must be able to take measurements, collect information, and present your data in a useful format (drawings, explanations, tables and graphs).
6. Draw conclusions from the results of the experiment. Your **observations** must be made objectively, never force the data to fit your hypothesis.
7. Decide whether your hypothesis explains the data collected accurately.
8. If the data fits your hypothesis, verify your results by repeating the experiment or getting someone else to repeat the experiment.
9. If your data does not fit your hypothesis perform more background research and make a new hypothesis.

Remember that in the development of both the gravitational theory and thermodynamics, scientists expanded on information from their predecessors or peers when developing their own theories. It is therefore very important to communicate findings to the public in the form of scientific publications, at conferences, in articles or TV or radio

programs. It is important to present your experimental data in a specific format, so that others can read your work, understand it, and repeat the experiment.

1. Aim: A brief sentence describing the purpose of the experiment.
2. Apparatus: A list of the apparatus.
3. Method: A list of the steps followed to carry out the experiment.
4. Results: Tables, graphs, and observations about the experiment.
5. Discussion: What your results mean.
6. Conclusion: A brief sentence concluding whether the aim was met.

## Hypothesis

A **hypothesis** should be specific and should relate directly to the question you are asking. For example, if your question about the world was, why do rainbows form, your hypothesis could be: *Rainbows form (**dependent variable**) because of light shining through water droplets (**independent variable**).* After formulating a hypothesis, it needs to be tested through experiment under controlled conditions. **Controls** are kept constant throughout an experiment and provide a constant standard of comparison.

A clear pattern of correlation between the dependent and

independent variables must be established.

An incorrect prediction does not mean that you have failed. It means that the experiment has brought some new facts to light that you might not have thought of before.

In science we never ‘prove’ a hypothesis through a single experiment because there is a chance that you made an error somewhere along the way. What you can say is that your results **SUPPORT** the original hypothesis.

# Analysis of the Scientific Method

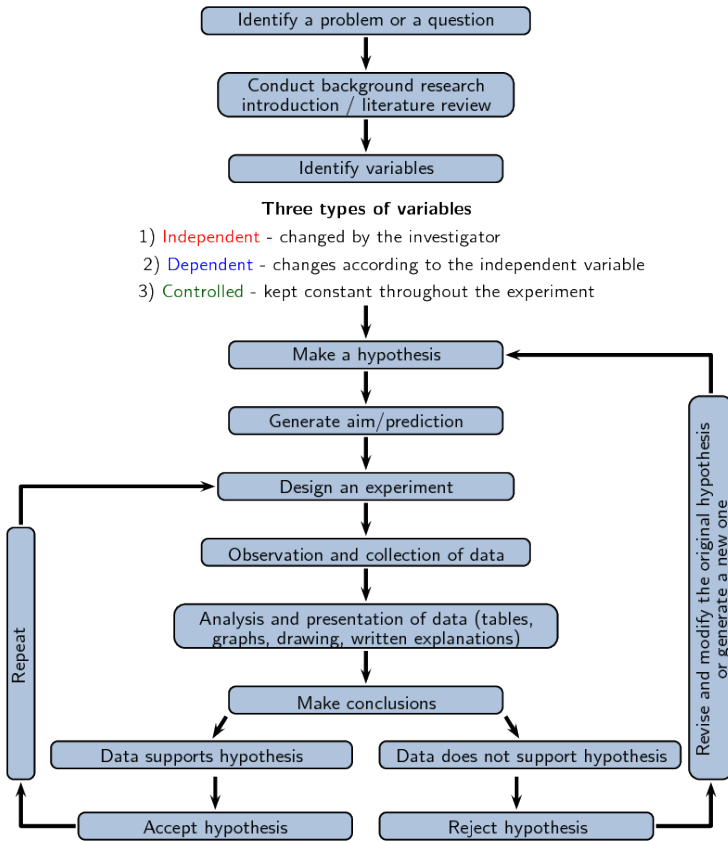


Fig 1.2 Flow chart diagram of the scientific method

Study the flow diagram provided, then discuss the questions that follow.

1. Once you have a problem you would like to study, why is

- it important to conduct background research before doing anything else?
2. What is the difference between a dependent, independent, and controlled variable and why is it important to identify them?
  3. What is the purpose of a control in an experiment?
  4. What is the difference between identifying a problem, a hypothesis, and a scientific theory?
  5. Why is it important to repeat your experiment if the data fits the hypothesis?

## Introduction to Measurements

Measurements provide much of the information that informs the hypotheses, theories, and laws describing 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**. (Scientific notation is also known as exponential notation; a review of this topic can be found through this link here: [Exponential notation review and examples](#).) For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as  $2.98 \times 10^5$  kg. The **mass** of the average mosquito is about 0.0000025 kilograms, which can be written as  $2.5 \times 10^{-6}$  kg.

**Units**, such as liters, pounds, and centimeters, are standards of comparison for measurements. A 2-liter bottle of a soft drink contains a volume of beverage that is twice that of the accepted volume of 1 liter. The meat used to prepare a 0.25-pound hamburger weighs one-fourth as much as the accepted **weight** of 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.

The measurement units for seven fundamental **properties** ("**base units**") are listed in Table 1.1. 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. Units for other properties may be derived from these seven **base units**.

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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

---

**Table 1.1**

Everyday measurement units are often defined as fractions or multiples of other units. Milk is commonly packaged in containers of 1 gallon (4 quarts), 1 quart (0.25 gallon), and one pint (0.5 quart). This same approach is used with **SI units**, 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 1.2.

---

Common Unit Prefixes		
Prefix	Symbol	Factor
femto	f	$10^{-15}$
pico	p	$10^{-12}$
nano	n	$10^{-9}$
micro	$\mu$	$10^{-6}$
milli	m	$10^{-3}$
centi	c	$10^{-2}$
deci	d	$10^{-1}$
kilo	k	$10^3$
mega	M	$10^6$
giga	G	$10^9$
tera	T	$10^{12}$

---

**Table 1.2**

### Link to Learning

Need more practice with scientific notation? This site covers the basics of scientific notation (opens in a new window): [Math Skills Review Scientific Notation](#).

## 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. SI stands for *Système International*, which means international system in French. 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 39.37 inches longer than a yard (Figure 1.1); one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers ( $1 \text{ km} = 1000 \text{ m} = 10^3 \text{ m}$ ), whereas shorter distances can be reported in centimeters ( $1 \text{ cm} = 0.01 \text{ m} = 10^{-2} \text{ m}$ ) or millimeters ( $1 \text{ mm} = 0.001 \text{ m} = 10^{-3} \text{ m}$ ).

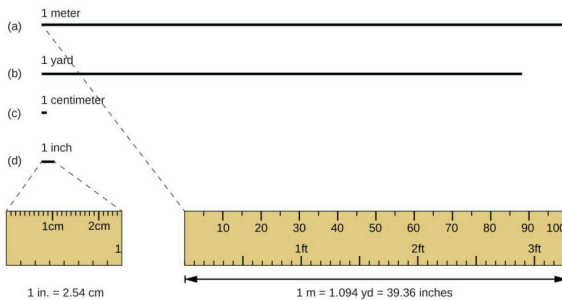


Fig. 1.3  
The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

## Mass

The **standard unit** of **mass** in the SI system is the **kilogram**

(kg). The kilogram was previously defined by the International Union of Pure and Applied Chemistry (IUPAC) as the mass of a specific reference object. This object was originally one liter of pure water, and more recently it was a metal cylinder made from a platinum-iridium alloy with a height and diameter of 39 mm (Figure 1.2). In May 2019, this definition was changed to one that is based instead on precisely measured values of several fundamental physical constants. 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).



Fig. 1.4 This replica prototype kilogram as previously defined is housed at the National Institute of Standards and Technology (NIST) in Maryland.

## Temperature

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 ( $^{\circ}$ ). The degree **Celsius** ( $^{\circ}\text{C}$ ) 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 (Figure 1.5). Water freezes at 273.15 K ( $0^{\circ}\text{C}$ ) and boils at 373.15 K ( $100^{\circ}\text{C}$ ) by definition, and normal human body temperature is approximately 310 K ( $37^{\circ}\text{C}$ ). Kelvin is used by scientists to better reflect the amount of thermal energy in any **measurement**, with the absence of energy being 0 K ( $-273.15^{\circ}\text{C}$ ), also known as **absolute zero**. The **conversion** between these two units and the **Fahrenheit** scale will be discussed later in this chapter.

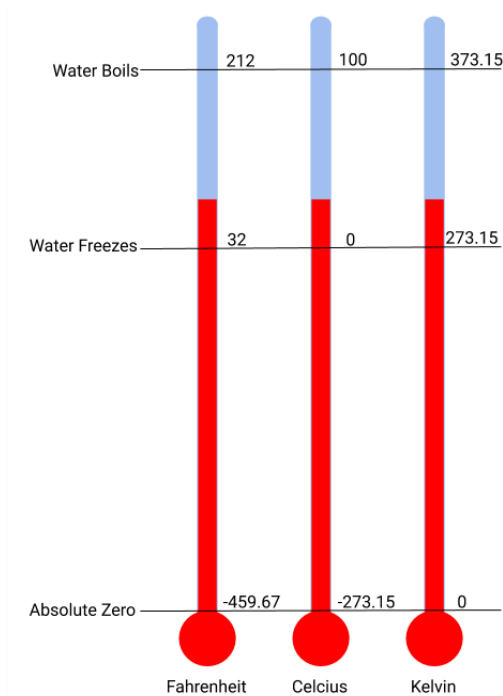


Fig 1.5 An illustration of the different temperature units with their respective benchmarks of where water boils and freezes as well as absolute zero.

## 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 =  $0.000003 \text{ s} = 3 \times 10^{-6}$  and 5 megaseconds =  $5,000,000 \text{ s} = 5 \times 10^6 \text{ s}$ . 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 multiply the **base unit** of **length** (m) by itself three times to define a unit of **volume**,  $\text{m}^3$ . We can also use the base units of **mass** (kg) and **length** (m) to define a unit of **density**, which is mass over volume or  $\text{kg}/\text{m}^3$ .

### 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 1.4). The standard volume is a cubic meter ( $\text{m}^3$ ), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

A more commonly used unit of volume is derived from the decimeter (0.1 m, or 10 cm). A cube with edge lengths of exactly one decimeter contains a volume of one cubic decimeter ( $\text{dm}^3$ ). A liter (L) is the more common name for the cubic decimeter. One liter is about 1.06 quarts.

A cubic centimeter ( $\text{cm}^3$ ) is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for cubic centimeter) is often used by health professionals. A cubic centimeter is equivalent to a milliliter (mL) and is 1/1000 of a liter.



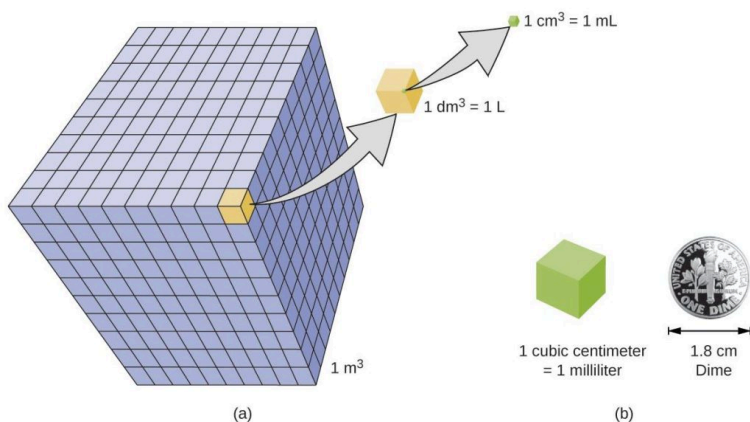


Fig 1.6 (a) The relative volumes are shown for cubes of 1 m<sup>3</sup>, 1 dm<sup>3</sup> (1 L), and 1 cm<sup>3</sup> (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm<sup>3</sup> (1-mL) cube.

## 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<sup>3</sup>). For many situations, however, this is an inconvenient **unit**, and we often use grams per cubic centimeter (g/cm<sup>3</sup>) 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<sup>3</sup> (the density of gasoline) to 19 g/

$\text{cm}^3$  (the density of gold). The density of air is about 1.2 g/L. Table 1.3 shows the densities of some common substances.

Densities of Common Substances		
Solids	Liquids	Gases (at 25 °C)
ice (at 0 °C) 0.92 g/cm <sup>3</sup>	water 1.0 g/cm <sup>3</sup>	dry air 1.20 g/L
oak (wood) 0.60–0.90 g/cm <sup>3</sup>	ethanol 0.79 g/cm <sup>3</sup>	oxygen 1.31 g/L
iron 7.9 g/cm <sup>3</sup>	acetone 0.79 g/cm <sup>3</sup>	nitrogen 1.14 g/L
copper 9.0 g/cm <sup>3</sup>	glycerin 1.26 g/cm <sup>3</sup>	carbon dioxide 1.8 g/L
lead 11.3 g/cm <sup>3</sup>	olive oil 0.92 g/cm <sup>3</sup>	helium 0.16 g/L
silver 10.5 g/cm <sup>3</sup>	gasoline 0.70–0.77 g/cm <sup>3</sup>	neon 0.83 g/L
gold 19.3 g/cm <sup>3</sup>	mercury 13.6 g/cm <sup>3</sup>	radon 9.1 g/L

**Table 1.3**

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.

**density = mass/volume**

## Accuracy and Precision

Scientists typically make repeated **measurements** of a quantity to ensure the quality of their findings and to evaluate 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 1.7).

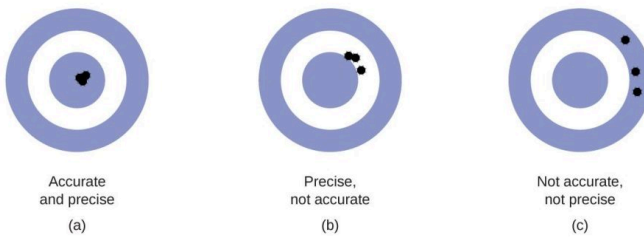


Fig 1.7 (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

## Unit Conversion

It is often the case that a quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships. For example, consider measuring the average speed of an athlete running sprints. This is typically accomplished by measuring the *time* required for the athlete to run from the starting line to the finish line, and the *distance* between these two lines, and then computing *speed* from the equation that relates these three properties:  $\text{speed} = \text{distance} / \text{time}$

An Olympic-quality sprinter can run 100 m in approximately 10 s, corresponding to an average speed of  $100 \text{ m} / 10 \text{ s} = 10 \text{ m/s}$

Note that this simple arithmetic involves dividing the numbers of each measured quantity to yield the number of the computed quantity ( $100/10 = 10$ ) and also dividing the **units** of each measured quantity to yield the unit of the computed quantity (meter divided by second = meter per second). Now, consider using this same relation to predict the time required for a person running at this speed to travel a distance of 25 m. The same relation among the three **properties** is used, but in this case, the two quantities provided are a speed (10 m/s) and a distance (25 m). To yield the sought property, time, the equation must be rearranged appropriately:  $\text{time} = \text{distance} / \text{speed}$

The time can then be computed as:  $25 \text{ m}/(10 \text{ m/s}) = 2.5 \text{ s}$

Again, arithmetic on the numbers ( $25/10 = 2.5$ ) was accompanied by the same arithmetic on the units ( $\text{m}/(\text{m/s}) = \text{s}$ ) to yield the number and unit of the result, 2.5 s. Note that, just as for numbers, when a unit is divided by an identical unit (in this case,  $\text{m}/\text{m}$ ), the result is “1”—or, as commonly phrased, the units “cancel.”

These calculations are examples of a versatile mathematical approach known as **dimensional analysis** (or the factor-label method). Dimensional analysis is based on this premise: *the units of quantities must be subjected to the same mathematical operations as their associated numbers*. This method can be applied to computations ranging from simple unit **conversions** to more complex, multi-step calculations involving several different quantities.

## Conversion Factors and Dimensional Analysis

A ratio of two equivalent quantities expressed with different measurement units can be used as a unit conversion factor. For example, the lengths of 2.54 cm and 1 in. are equivalent (by definition), and so a unit conversion factor may be derived from the ratio  $2.54 \text{ cm}/1 \text{ in.}$ , ( $2.54 \text{ cm} = 1 \text{ in.}$ ), or  $2.54 \text{ cm}/\text{in.}$

Several other commonly used conversion factors are given in Table 1.4.

Common Conversion Factors		
Length	Volume	Mass
1 m = 1.0936 yd	1 L = 1.0567 qt	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft <sup>3</sup> = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1 tbsp = 14.787 mL	1 (troy) oz = 31.103 g

**Table 1.4**

When a quantity (such as distance in inches) is multiplied by an appropriate unit conversion factor, the quantity is converted to an equivalent value with different units (such as distance in centimeters). For example, a basketball player’s vertical jump of 34 inches can be converted to centimeters by:

$$34\text{in.} \times (2.54\text{ cm}/1\text{in.}) = 86\text{ cm}$$

Since this simple arithmetic involves *quantities*, the premise of **dimensional analysis** requires that we multiply both *numbers and units*. The numbers of these two quantities are multiplied to yield the number of the product quantity, 86, whereas the **units** are multiplied to yield in.×(cm/in.) Just as for numbers, a ratio of identical units is also numerically equal to one, in./in.=1, and the unit product thus simplifies to *cm*. (When identical units divide to yield a factor of 1, they are said to “cancel.”) Dimensional analysis may be used to confirm the

proper application of unit conversion factors as demonstrated in the following example.

Beyond simple unit conversions, the factor-label method can be used to solve more complex problems involving computations. Regardless of the details, the basic approach is the same—all the *factors* involved in the calculation must be appropriately oriented to ensure that their *labels* (units) will appropriately cancel and/or combine to yield the desired unit in the result. As your study of chemistry continues, you will encounter many opportunities to apply this approach.

### Link to Learning

Need a refresher or more practice with dimensional analysis? Visit this site for the basics of dimensional analysis (opens in a new window): [Math Skills Review Dimensional Analysis](#).

### Example 1.1

## CONVERTING GRAMS TO OUNCES

Using a Unit Conversion Factor the mass of a competition frisbee is 125 g. Convert its mass to ounces using the unit conversion factor derived from the relationship 1 oz = 28.349 g.

Solution:

Given the conversion factor, the mass in ounces may be derived using an equation similar to the one used for converting length from inches to centimeters.

$$x \text{ oz.} = 125 \text{ g} \times \text{unit conversion factor}$$

The unit conversion factor may be represented as: 1 oz./28.349 g and 28.349 g/1 oz.

The correct unit conversion factor is the ratio that cancels the units of grams and leaves ounces.

$$x \text{ oz.} = 125\text{g} \times (1 \text{ oz.}/28.349\text{g}) = 125/28.349\text{oz.} = 4.41 \text{ oz.}$$



## Example 1.2

## CACULATING FUEL ECONOMY AND FUEL COSTS

While being driven from Philadelphia to Atlanta, a distance of about 1250 km, a 2014 Lamborghini Aventador Roadster uses 213 L of gasoline.

- (a) What (average) fuel economy, in miles per gallon, did the Roadster get during this trip?
- (b) If gasoline costs \$3.80 per gallon, what was the fuel cost for this trip?

Solution:

- (a) First convert distance from kilometers to miles:

$$1250\text{km} \times (0.62137 \text{ mi}/1\text{km}) = 777 \text{ mi}$$

and then convert volume from liters to gallons:

$$213\text{L} \times (1.0567\text{qt}/1\text{L}) \times (1 \text{ gal}/4\text{qt}) = 56.3 \text{ gal}$$

Finally,

$$\begin{aligned} \text{(average) mileage} &= 777 \text{ mi}/56.3 \text{ gal} = 13.8 \text{ miles/} \\ &\text{gallon} = 13.8 \text{ mpg} \end{aligned}$$

Alternatively, the calculation could be set up in a way that uses all the conversion factors sequentially, as follows:

$$(1250\text{km}/213\text{L}) \times (0.62137 \text{ mi}/1\text{km}) \times (1\text{L}/1.0567\text{qt}) \times (4\text{qt}/1 \text{ gal}) = 13.8 \text{ mpg}$$

(b) Using the previously calculated volume in gallons, we find:

$$56.3\text{gal} \times (\$3.80/1\text{gal}) = \$214$$

## Check Your Understanding



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## 2.

# LINEAR MOTION

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### Chapter Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

- Discuss the requirements for determining motion (CLO2) (CLO3)
- Define position, speed, velocity, and acceleration (CLO2)
- Define and compute distance (CLO2)(CLO5)
- Discuss the relationship between position, velocity, and acceleration of an object in motion (CLO2)(CLO3)
- Recognize and use proper units for position, speed, velocity, and acceleration (CLO2)(CLO5)

- Analyze motion graphs. (CLO2)(CLO5)
- Perform calculations using the kinematics equations for constant acceleration situations.(CLO2)(CLO5)
- Define free fall and solve free fall problems. (CLO2)(CLO5)



Fig. 2.1 These cyclists in Vietnam can be described by their position relative to buildings and a canal. Their motion can be described by their change in position, or displacement, in the frame of reference.

# Introduction to Kinematics

Objects are in motion everywhere we look. Everything from a tennis game to a space-probe flyby of the planet Neptune involves motion. When you are resting, your heart moves blood through your veins. And even in inanimate objects, there is continuous motion in the vibrations of atoms and molecules. Questions about motion are interesting in and of themselves: How long will it take for a space probe to get to Mars? Where will a football land if it is thrown at a certain angle? But an understanding of motion is also key to understanding other concepts in physics. An understanding of **acceleration**, for example, is crucial to the study of **force**.

A formal study of physics begins with **kinematics** which is defined as the *study of motion without considering its causes*. The word “kinematics” comes from a Greek term meaning motion and is related to other English words such as “cinema” (movies) and “kinesiology” (the study of human motion). In this chapter, we examine the simplest type of motion—namely, motion along a straight line, or one-dimensional motion.



# Kinematics – Describing Motion

## Position

In order to describe the motion of an object, you must first be able to describe its **position**—where it is at any particular time. More precisely, you need to specify its position relative to a convenient reference frame. Earth is often used as a reference frame, and we often describe the position of an object as it relates to stationary objects in that reference frame. For example, a rocket launch would be described in terms of the position of the rocket with respect to the Earth as a whole, while a professor's position could be described in terms of where she is in relation to the nearby white board. (See Figure 2.2) In other cases, we use reference frames that are not stationary but are in motion relative to the Earth. To describe the position of a person in an airplane, for example, we use the airplane, not the Earth, as the reference frame. (See Figure 2.3)

## Displacement

If an object moves relative to a reference frame (for example, if a professor moves to the right relative to a white board or a passenger moves toward the rear of an airplane), then the object's position changes. This change in position is known

as **displacement**. The word “displacement” implies that an object has moved, or has been displaced.

Displacement is the *change in position* of an object:  $\Delta \mathbf{x} = \mathbf{x}_f - \mathbf{x}_0$ , where  $x_f$  is the final position, and  $x_0$  is the initial position.

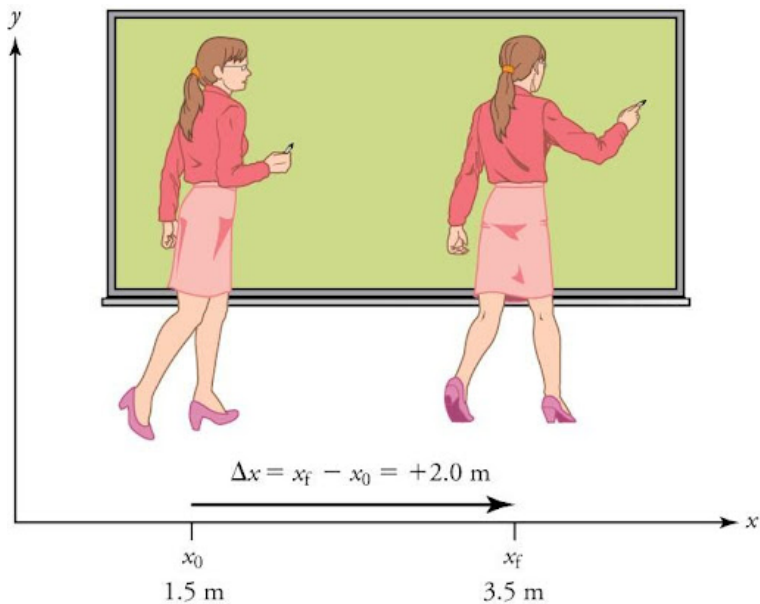


Fig 2.2 A professor paces left and right while lecturing. Her position relative to Earth is given by  $x$ . The  $+2.0\text{m}$  displacement of the professor relative to Earth is represented by an arrow pointing to the right.

Note that the **SI unit** for displacement is the **meter** (m), but sometimes kilometers, miles, feet, and other **units** of length are used. Keep in mind that when units other than the meter are

used in a problem, you may need to convert them into meters to complete the calculation.

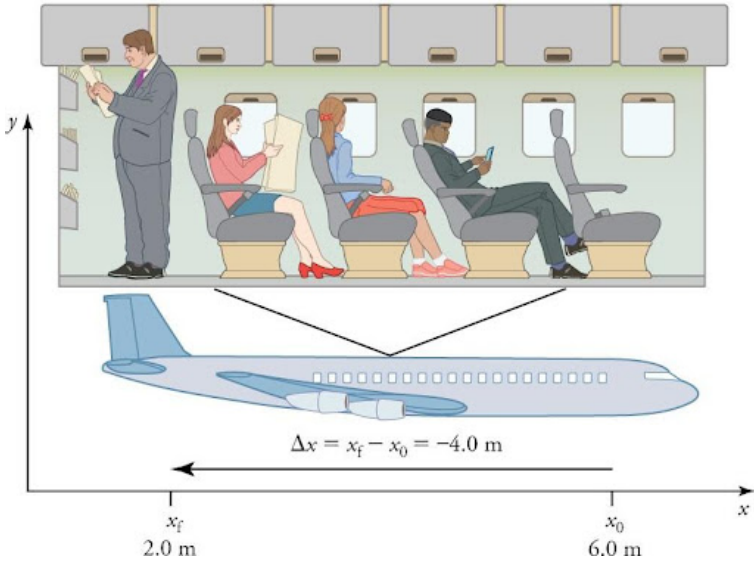


Fig. 2.3 A passenger moves from his seat to the back of the plane. His location relative to the airplane is given by  $x$ . The  $-4.0\text{m}$  displacement of the passenger relative to the plane is represented by an arrow toward the rear of the plane. Notice that the arrow representing his displacement is twice as long as the arrow representing the displacement of the professor (he moves twice as far) in Figure 2.2

Note that displacement has a direction as well as a **magnitude**. The professor's displacement is  $2.0 \text{ m}$  to the right, and the airline passenger's displacement is  $4.0 \text{ m}$  toward the rear. In one-dimensional motion, direction can be specified with a plus or minus sign. When you begin a problem, you should select

which direction is positive (usually that will be to the right or up, but you are free to select positive as being any direction). The professor's initial position is  $x_0 = 1.5\text{m}$  and her final position is  $x_f = 3.5\text{m}$ . Thus her displacement is  $\Delta x = x_f - x_0 = 3.5\text{ m} - 1.5\text{ m} = +2.0\text{ m}$ .

In this coordinate system, motion to the right is positive, whereas motion to the left is negative. Similarly, the airplane passenger's initial position is  $x_0 = 6.0\text{ m}$  and his final position is  $x_f = 2.0\text{ m}$ .

$$\Delta x = x_f - x_0 = 2.0\text{ m} - 6.0\text{ m} = -4.0\text{ m}.$$

His displacement is negative because his motion is toward the rear of the plane, or in the negative direction in our coordinate system.

## Distance

Although displacement is described in terms of direction, distance is not. **Distance** is defined to be *the magnitude or size of displacement between two positions*. Note that the distance between two positions is not the same as the distance traveled between them. Distance traveled is *the total length of the path traveled between two positions*. Distance has no direction and, thus, no sign. For example, the distance the professor walks is 2.0 m. The distance the airplane passenger walks is 4.0 m.



Fig 2.4 The motion of these snails can be described by their speeds and their velocities.

There is more to motion than distance and displacement. Questions such as, “How long does a foot race take?” and “What was the runner’s speed?” cannot be answered without an understanding of other concepts. In this section we add definitions of time, velocity, and speed to expand our description of motion.

### Misconception Alert

It is important to note that the distance traveled, however, can be greater than the magnitude of the displacement (by magnitude, we mean just the size of the displacement without regard to its direction; that is, just a number with a unit). For example, the professor could pace back and forth many times, perhaps walking a distance of 150 m during a lecture, yet still end up only 2.0 m to the right of her starting point. In this case her displacement would be +2.0 m, the magnitude of her displacement would be 2.0 m, but the distance she traveled would be 150 m. In kinematics we nearly always deal with displacement and magnitude of displacement, and almost never with distance traveled. One way to think about this is to assume you marked the start of the motion and the end of the motion. The displacement is simply the difference in the position of the two marks and is independent of the path taken in traveling between the two marks. The distance traveled,

however, is the total length of the path taken between the two marks.

## Time

As discussed in Chapter 1, Measurements, the most fundamental physical quantities are defined by how they are measured. This is the case with **time**. Every measurement of time involves measuring a change in some physical quantity. It may be a number on a digital clock, a heartbeat, or the position of the Sun in the sky. In physics, the definition of time is simple—time is *change*, or the interval over which change occurs. It is impossible to know that time has passed unless something changes.

The amount of time or change is calibrated by comparison with a standard. The SI unit for time is the **second** (s). We might, for example, observe that a certain pendulum makes one full swing every 0.75 s. We could then use the pendulum to measure time by counting its swings or by connecting the pendulum to a clock mechanism that registers time on a dial. This allows us to not only measure the amount of time, but also to determine a sequence of events.

How does time relate to motion? We are usually interested in elapsed time for a particular motion, such as how long it takes an airplane passenger to get from his seat to the back of the plane. To find elapsed time, we note the time at the beginning and end of the motion and subtract the two. For example, a lecture may start at 11:00 A.M. and end at 11:50 A.M., so that the elapsed time would be 50 min. Elapsed time  $\Delta t$  is the difference between the ending time and beginning time,

$\Delta t = t_f - t_0$ , where  $\Delta t$  is the change in time or elapsed time,  $t_f$  is the time at the end of the motion, and  $t_0$  is the time at the beginning of the motion. (As usual, the delta symbol,  $\Delta$ , means the change in the quantity that follows it.)

Life is simpler if the beginning time  $t_0$  is taken to be zero, as when we use a stopwatch. If we were using a stopwatch, it would simply read zero at the start of the lecture and 50 min at the end. If  $t_0 = 0$ , then  $\Delta t = t_f \equiv t$ .

## Velocity

Your notion of **velocity** is probably the same as its scientific definition. You know that if you have a large displacement in a small amount of time you have a large velocity, and that velocity has units of distance divided by time, such as miles per hour or kilometers per hour.



## Average Velocity

**Average velocity** is *displacement (change in position) divided by the time of travel*,

$$\text{average } \mathbf{v} = \Delta \mathbf{x} / \Delta t$$

Notice that this definition indicates that *velocity is a vector because displacement is a vector*. It has both magnitude and direction. The SI unit for velocity is meters per second or m/s, but many other units, such as km/h, mi/h (also written as mph), and cm/s, are in common use. Suppose, for example, an airplane passenger took 5 seconds to move  $-4$  m (the negative sign indicates that displacement is toward the back of the plane). His average velocity would be:

$$\text{average } v = \Delta x / \Delta t = -4\text{m} / 5\text{s} = -0.8 \text{ m/s.}$$

The minus sign indicates the average velocity is also toward the rear of the plane.

The average velocity of an object does not tell us anything about what happens to it between the starting point and ending point, however. For example, we cannot tell from average velocity whether the airplane passenger stops momentarily or backs up before he goes to the back of the plane. To get more details, we must consider smaller segments of the trip over smaller time intervals.

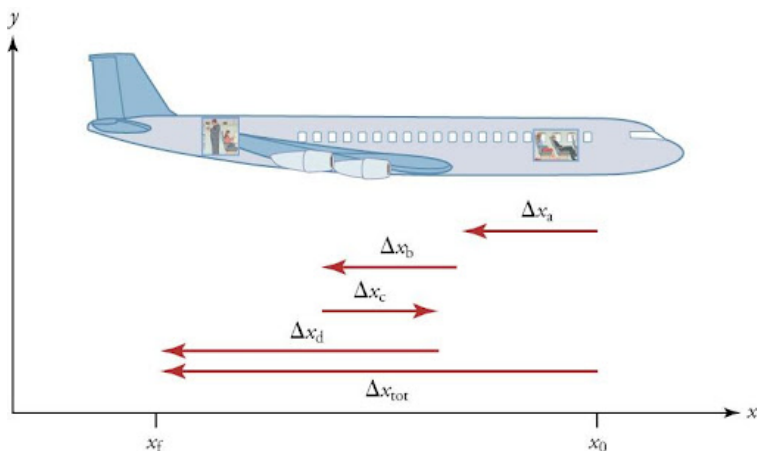


Fig. 2.5 A more detailed record of an airplane passenger heading toward the back of the plane, showing smaller segments of his trip.

The smaller the time intervals considered in a motion, the more detailed the information. When we carry this process to its logical conclusion, we are left with an infinitesimally small interval. Over such an interval, the average velocity becomes the *instantaneous velocity* or the *velocity at a specific instant*. A car's speedometer, for example, shows the magnitude (but not the direction) of the instantaneous velocity of the car. (Police give tickets based on instantaneous velocity, but when calculating how long it will take to get from one place to another on a road trip, you need to use average velocity.)

## Speed

In everyday language, most people use the terms “speed” and

“velocity” interchangeably. In physics, however, they do not have the same meaning and they are distinct concepts. One major difference is that **speed** has no direction. Thus *speed is a scalar*. Just as we need to distinguish between instantaneous velocity and average velocity, we also need to distinguish between instantaneous speed and average speed.

Instantaneous speed is the magnitude of instantaneous velocity. For example, suppose the airplane passenger at one instant had an instantaneous velocity of  $-3.0$  m/s (the minus meaning toward the rear of the plane). At that same time his instantaneous speed was  $3.0$  m/s. Or suppose that at one time during a shopping trip your instantaneous velocity is  $40$  km/h due north. Your instantaneous speed at that instant would be  $40$  km/h—the same magnitude but without a direction. Average speed, however, is very different from average velocity. Average speed is the distance traveled divided by elapsed time.

We have noted that distance traveled can be greater than the magnitude of displacement. So average speed can be greater than average velocity, which is displacement divided by time. For example, if you drive to a store and return home in half an hour, and your car’s odometer shows the total distance traveled was  $6$  km, then your average speed was  $12$  km/h. Your average velocity, however, was zero, because your displacement for the round trip is zero. (Displacement is change in position, and thus, is zero for a round trip.) Thus average speed is *not* simply the magnitude of average velocity.

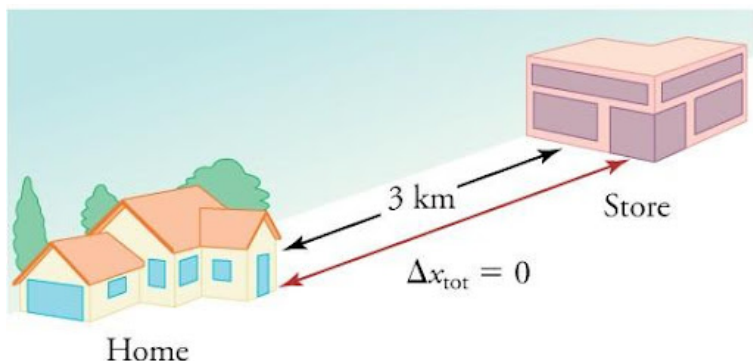


Fig 2.6 During a 30-minute round trip to the store, the total distance traveled is 6 km. The average speed is 12 km/h. The displacement for the round trip is zero, since there was no net change in position. Thus the average velocity is zero.

Another way of visualizing the motion of an object is to use a graph. A plot of position or of velocity as a function of time can be very useful. For example, for this trip to the store, the position, velocity, and speed-vs.-time graphs are displayed in Figure 2.7. (Note that these graphs depict a very simplified model of the trip. We are assuming that speed is constant during the trip, which is unrealistic given that we'll probably stop at the store. But for simplicity's sake, we will model it with no stops or changes in speed. We are also assuming that the route between the store and the house is a perfectly straight line.)

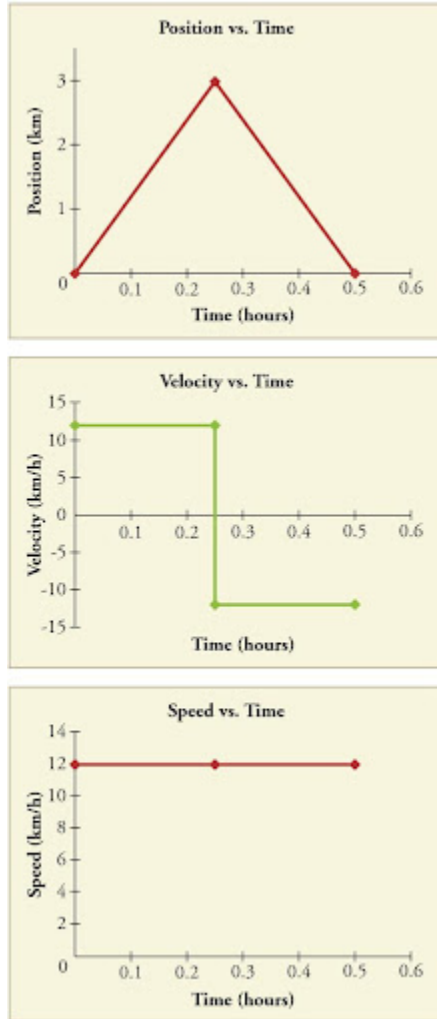


Fig 2.7 Position vs. time, velocity vs. time, and speed vs. time on a trip. Note that the velocity for the return trip is negative



Fig 2.8 A plane decelerates, or slows down, as it comes in for landing in St. Maarten. Its acceleration is opposite in direction to its velocity.

In everyday conversation, to accelerate means to speed up. The accelerator in a car can in fact cause it to speed up. The greater the **acceleration**, the greater the change in velocity over a given time. The formal definition of acceleration is consistent with these notions, but more inclusive.

## Average Acceleration

**Average Acceleration** is the rate at which velocity changes,

$$\text{average } \mathbf{a} = \Delta \mathbf{v} / \Delta t = (\mathbf{v}_f - \mathbf{v}_0) / (t_f - t_0),$$

Because acceleration is **velocity** in m/s divided by time in s, the SI units for acceleration are meters per second squared

or meters per second per second, which literally means by how many meters per second the velocity changes every second.

Recall that velocity is a **vector**—it has both magnitude and direction. This means that a change in velocity can be a change in magnitude (or speed), but it can also be a change in *direction*. For example, if a car turns a corner at constant speed, it is accelerating because its direction is changing. The quicker you turn, the greater the acceleration. So there is an acceleration when velocity changes either in magnitude (an increase or decrease in speed) or in direction, or both.

## Acceleration as a Vector

Acceleration is a vector in the same direction as the *change* in velocity,  $\Delta v$ . Since velocity is a vector, it can change either in magnitude or in direction. **Acceleration** is therefore a change in either speed, direction, or both.

Keep in mind that although acceleration is in the direction of the *change* in velocity, it is not always in the direction of *motion*. When an object slows down, its acceleration is opposite to the direction of its motion. This is known as **deceleration**.



Fig 2.9 A subway train in Sao Paulo, Brazil, decelerates as it comes into a station. It is accelerating in a direction opposite to its direction of motion.

### Misconception Alert

Deceleration always refers to acceleration in the direction opposite to the direction of the velocity. Deceleration always reduces speed. Negative acceleration, however, is acceleration in the negative direction in the chosen coordinate



system. Negative acceleration may or may not be deceleration, and deceleration may or may not be considered negative acceleration. For example, consider Figure 2.10.

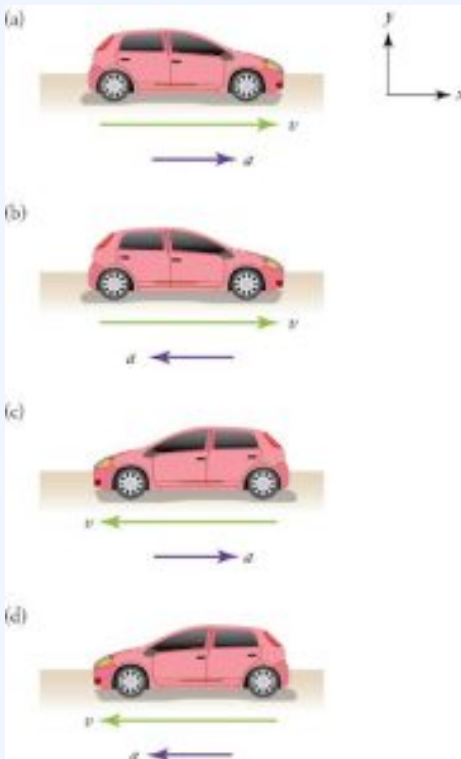


Fig. 2.10 (a) This car is speeding up as it moves toward the right. It has positive acceleration in our coordinate system. (b)

This car is slowing down as it moves toward the right. It has negative acceleration, because its acceleration is toward the left. The car is also decelerating: its acceleration is opposite to its motion. (c) This car is moving toward the left, but slowing down over time. Its acceleration is positive in our coordinate system because it is toward the right. However, the car is decelerating because its acceleration is opposite to its motion. (d) This car is speeding up as it moves toward the left. It has negative acceleration because it is accelerating toward the left. Because its acceleration is in the same direction as its motion, it is not decelerating.

## Instantaneous Acceleration

Instantaneous acceleration  $a$ , or the acceleration at a specific instant in time, is similar to instantaneous velocity —that is, considering an infinitely small interval of time. How do we find instantaneous acceleration using only algebra? The answer is that we choose an average acceleration that is

representative of the motion. Figure 2.11 shows graphs of instantaneous acceleration versus time for two very different motions. In Figure 2.11(a), the acceleration varies slightly and the average over the entire interval is nearly the same as the instantaneous acceleration at any time. In this case, we should treat this motion as if it had a constant acceleration equal to the average (in this case about  $1.8 \text{ m/s}^2$ ). In Figure 2.11(b), the acceleration varies drastically over time. In such situations it is best to consider smaller time intervals and choose an average acceleration for each. For example, we could consider motion over the time intervals from 0 s to 1.0 s and from 1.0 s to 3.0 s as separate motions with accelerations of  $+3.0 \text{ m/s}^2$  and  $-2.0 \text{ m/s}^2$ , respectively.

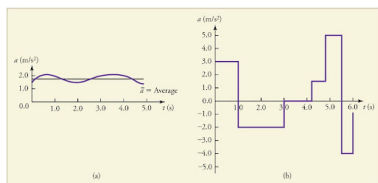


Fig. 2.11 Graphs of instantaneous acceleration versus time for two different one-dimensional motions. (a) Here acceleration varies only slightly and is always in the same direction, since it is positive. The average over the interval is nearly the same as the acceleration at any given time. (b) Here the acceleration varies greatly, perhaps representing a package on a post office conveyor belt that is accelerated forward and backward as it bumps along. It is necessary to consider small time intervals (such as from 0 s to 1.0 s) with constant or nearly constant acceleration in such a situation.

The next several examples consider the motion of the subway train shown in Figure 2.12. In (a) the shuttle moves to the right, and in (b) it moves to the left. The examples are designed to further illustrate aspects of motion and to illustrate some of the reasoning that goes into solving problems.

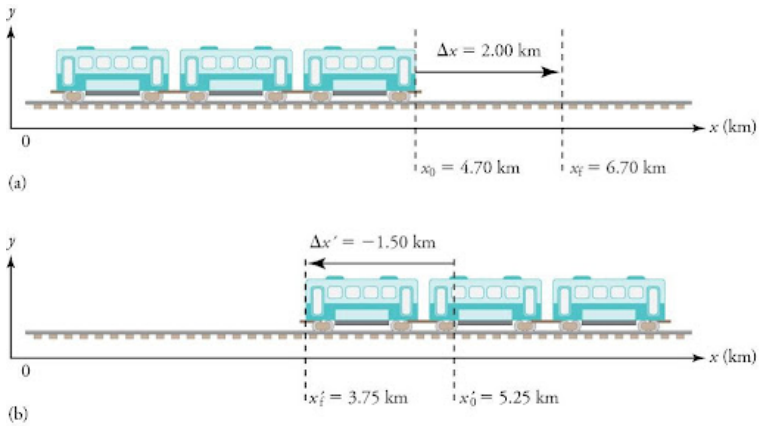


Fig. 2.12 One-dimensional motion of a subway train considered in Example 7.2, Example 7.3, Example 7.4, Example 7.5. Here we have chosen the  $x$ -axis so that  $+$  means to the right and  $-$  means to the left for displacements, velocities, and accelerations. (a) The subway train moves to the right from  $x_0$  to  $x_f$ . Its displacement  $\Delta x$  is  $+2.0 \text{ km}$ . (b) The train moves to the left from  $x'_i$  to  $x'_f$ . Its displacement  $\Delta x'$  is  $-1.5 \text{ km}$ . (Note that the prime symbol ( $'$ ) is used simply to distinguish between displacement in the two different situations. The distances of travel and the size of the cars are on different scales to fit everything into the diagram.)

## Example 2.1

## CALCULATING ACCELERATION: A RACEHORSE LEAVES THE GATE

A racehorse coming out of the gate accelerates from rest to a velocity of  $15.0 \text{ m/s}$  due west in  $1.80 \text{ s}$ . What is its average acceleration?



Fig. 2.13

### Strategy

First we draw a sketch and assign a coordinate system to the problem. This is a simple problem, but it always helps to visualize it. Notice that we assign

east as positive and west as negative. Thus, in this case, we have negative velocity.

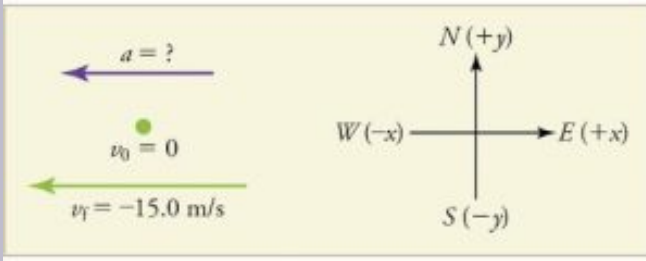


Fig. 2.14

We can solve this problem by identifying  $\Delta v$  and  $\Delta t$  from the given information and then calculating the average acceleration directly from the equation  $a = \Delta v / \Delta t = (v_f - v_0) / (t_f - t_0)$ .

### Solution

1. Identify the knowns.  $v_0 = 0 \text{ m/s}$ ,  $v_f = -15.0 \text{ m/s}$  (the negative sign indicates direction toward the west),  $\Delta t = 1.80 \text{ s}$ .
2. Find the change in velocity. Since the horse is going from zero to  $-15.0 \text{ m/s}$ , its change in velocity equals its final velocity:  $\Delta v = v_f = -15.0$

m/s.

3. Plug in the known values ( $\Delta v$  and  $\Delta t$ ) and solve for the unknown  $a$ .

average  $a = \Delta v / \Delta t = (-15.0 \text{ m/s}) / (1.80 \text{ s}) = -8.33 \text{ m/s}^2$ .

### Discussion

The negative sign for acceleration indicates that acceleration is toward the west. An acceleration of  $8.33 \text{ m/s}^2$  due west means that the horse increases its velocity by  $8.33 \text{ m/s}$  due west each second, that is,  $8.33 \text{ meters per second per second}$ , which we write as  $8.33 \text{ m/s}^2$ . This is truly an average acceleration, because the ride is not smooth. We shall see later that an acceleration of this magnitude would require the rider to hang on with a force nearly equal to his weight.

### Example 2.2



## CALCULATING DISPLACEMENT: A SUBWAY TRAIN

What are the magnitude and sign of displacements for the motions of the subway train shown in parts (a) and (b) of Figure 2.12?

### Strategy

A drawing with a coordinate system is already provided, so we don't need to make a sketch, but we should analyze it to make sure we understand what it is showing. Pay particular attention to the coordinate system. To find displacement, we use the equation  $\Delta x = x_f - x_0$ . This is straightforward since the initial and final positions are given.

### Solution

1. Identify the knowns. In the figure we see that  $x_f = 6.70 \text{ km}$  and  $x_0 = 4.70 \text{ km}$  for part (a), and  $x_f' = 3.75 \text{ km}$  and  $x_0' = 5.25 \text{ km}$  for part (b).
2. Solve for displacement in part (a).

$$\Delta x = x_f - x_0 = 6.70 \text{ km} - 4.70 \text{ km} = +2.00 \text{ km}$$

3. Solve for displacement in part (b).

$$\Delta x' = x_f' - x_0' = 3.75 \text{ km} - 5.25 \text{ km} = -1.50 \text{ km}$$

**Discussion**

The direction of the motion in (a) is to the right and therefore its displacement has a positive sign, whereas motion in (b) is to the left and thus has a negative sign.

**Example 2.3****COMPARING DISTANCE TRAVELED  
WITH DISPLACEMENT: A SUBWAY  
TRAIN**

What are the distances traveled for the motions shown in parts (a) and (b) of the subway train in Figure 2.12?

**Strategy**

To answer this question, think about the definitions of distance and distance traveled, and how they are

related to displacement. Distance between two positions is defined to be the magnitude of displacement, which was found in Example 7.2. Distance traveled is the total length of the path traveled between the two positions. In the case of the subway train shown in Figure 2.12, the distance traveled is the same as the distance between the initial and final positions of the train.

### **Solution**

1. The displacement for part (a) was  $+2.00$  km. Therefore, the distance between the initial and final positions was  $2.00$  km, and the distance traveled was  $2.00$  km.
2. The displacement for part (b) was  $-1.5$  km. Therefore, the distance between the initial and final positions was  $1.50$  km, and the distance traveled was  $1.50$  km.

### Example 2.4

## CALCULATING AVERAGE VELOCITY: THE SUBWAY TRAIN

What is the average velocity of the train in part b of Example 2.2 (and shown again below) if it takes 5.00 min to make its trip?

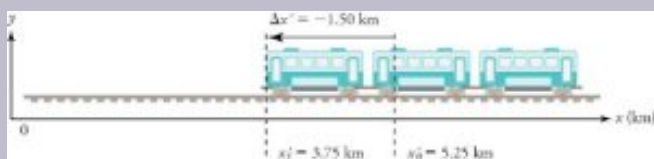


Fig. 2.15

### Strategy

Average velocity is displacement divided by time. It will be negative here, since the train moves to the left and has a negative displacement.

### Solution

1. Identify the knowns.  $x_f' = 3.75 \text{ km}$ ,  $x_0' = 5.25 \text{ km}$ ,  $\Delta t = 5.00 \text{ min}$ .
2. Determine displacement,  $\Delta x'$ . We found  $\Delta x'$  to be  $-1.5 \text{ km}$  in Example 7.2.
3. Solve for average velocity.

$$\text{average } v = \Delta x' / \Delta t = (-1.50 \text{ km}) / (5.00 \text{ min})$$

4. Convert units.

$$\text{average } v = \Delta x' / \Delta t = ((-1.50 \text{ km}) / (5.00 \text{ min})) * (60 \text{ min} / 1 \text{ h}) = -18.0 \text{ km/h}$$

### Discussion

The negative velocity indicates motion to the left.

## Example 2.5

### EXAMPLE 2.5 CALCULATING DECELERATION: THE SUBWAY TRAIN

Finally, suppose the train in Figure 2.15 slows to a stop from a velocity of 20.0 km/h in 10.0 s. What is its average acceleration?

### Strategy

Once again, let's draw a sketch:

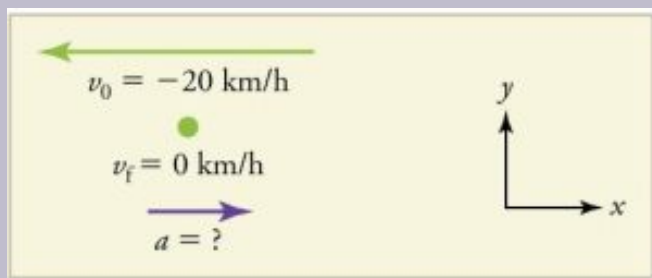


Fig 2.16

As before, we must find the change in velocity and the change in time to calculate average acceleration.

### Solution

1. Identify the knowns.  $v_0 = -20 \text{ km/h}$ ,  $v_f = 0 \text{ km/h}$ ,  $\Delta t = 10.0 \text{ s}$ .
2. Calculate  $\Delta v$ . The change in velocity here is actually positive, since  $\Delta v = v_f - v_0 = 0 - (-20 \text{ km/h}) = +20 \text{ km/h}$ .
3. Solve for acceleration.

$$a = \Delta v / \Delta t = (+20.0 \text{ km/h}) / (10.0 \text{ s})$$

4. Convert units.

$$a = ((+20.0 \text{ km/h}) / (10.0 \text{ s})) * (103 \text{ m} / 1 \text{ km}) * (1 \text{ h} / 3600 \text{ s}) = +0.556 \text{ m/s}^2$$

**Discussion**

The plus sign means that acceleration is to the right. This is reasonable because the train initially has a negative velocity (to the left) in this problem and a positive acceleration opposes the motion (and so it is to the right). Again, acceleration is in the same direction as the change in velocity, which is positive here. This acceleration can be called a deceleration since it is in the direction opposite to the velocity.

## Kinematic Relationships



Fig. 2.17 Kinematic equations can help us describe and predict the motion of moving objects such as these kayaks racing in Newbury, England.

We might know that the greater the acceleration of, say, a car moving away from a stop sign, the greater the displacement in a given time. But we have not developed a specific equation that relates **acceleration** and **displacement**. In this section, we develop some convenient equations for kinematic relationships, starting from the definitions of displacement, velocity, and acceleration already covered.



## Notation: $t$ , $x$ , $v$ , $a$

First, let us make some simplifications in notation. Taking the initial time to be zero, as if **time** is measured with a stopwatch, is a great simplification. Since elapsed time is  $\Delta t = t_f - t_0$ , taking  $t_0 = 0$  means that  $\Delta t = t_f$ , the final time on the stopwatch. When initial time is taken to be zero, we use the subscript 0 to denote initial values of **position** and **velocity**. That is,  $x_0$  *is the initial position* and  $v_0$  *is the initial velocity*. We put no subscripts on the final values. That is,  $t$  *is the final time*,  $x$  *is the final position*, and  $v$  *is the final velocity*. This gives a simpler expression for elapsed time—now,  $\Delta t = t$ . The expression for displacement is now  $\Delta x = x - x_0$ , and the expression for change in velocity is now  $\Delta v = v - v_0$ . We now make the important assumption that **acceleration is constant**. This assumption allows us to avoid using calculus to find instantaneous acceleration. Since acceleration is constant, the average and instantaneous accelerations are equal. That is, **average  $a = a = \text{constant}$** , so we use the symbol  $a$  for acceleration at all times. Assuming acceleration to be constant does not seriously limit the situations we can study nor degrade the accuracy of our treatment. For one thing, acceleration is constant in a great number of situations. Furthermore, in many other situations we can accurately describe motion by assuming a constant acceleration equal to the average acceleration for that motion. Finally, in motions where acceleration changes drastically, such

as a car accelerating to top speed and then braking to a stop, the motion can be considered in separate parts, each of which has its own constant acceleration.

## Solving For Displacement ( $\Delta x$ ) & Final Position ( $x$ ) From Average Velocity When Acceleration ( $a$ ) Is Constant

To get our first two new equations, we start with the definition of average velocity:

$$\text{average } v = \Delta x / \Delta t.$$

Substituting the simplified notation for  $\Delta x$  and  $\Delta t$  yields **average  $v = (x - x_0) / t$ .**

Solving for  $x$  yields

$$x = x_0 + (\text{average } v)t,$$

where the average velocity is

$$\text{average } v = (v_0 + v) / 2 \text{ (constant } a)$$

The equation **average  $v = (v_0 + v) / 2$**  reflects the fact that, when acceleration is constant, average  $v$  is just the simple average of the initial and final velocities. For example, if you steadily increase your velocity (that is, with constant acceleration) from 30 to 60 km/h, then your **average velocity** during this steady increase is 45 km/h. Using the equation **average  $v = (v_0 + v) / 2$**  to check this, we see that

$$\text{average } v = (v_0 + v) / 2 = (30 \text{ km/h} + 60 \text{ km/h}) / 2 = 45 \text{ km/h},$$

which seems logical.

## Example 2.6

### CALCULATING DISPLACEMENT: HOW FAR DOES THE JOGGER RUN?

A jogger runs down a straight stretch of road with an average velocity of  $4.00 \text{ m/s}$  for  $2.00 \text{ min}$ . What is his final position, taking his initial position to be zero?

#### Strategy

Draw a sketch.

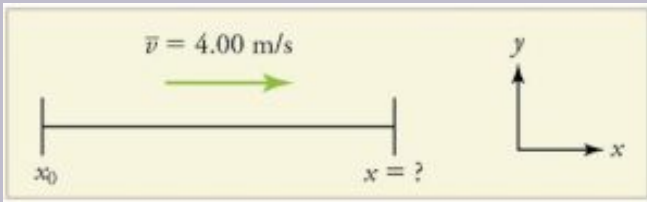


Fig. 2.18

The final position  $x$  is given by the equation

$$x = x_0 + (\text{average } v)t.$$

To find  $x$ , we identify the values of  $x_0$ , and  $t$  from

the statement of the problem and substitute them into the equation.

### Solution

1. Identify the knowns.  $v = 4.00 \text{ m/s}$ ,  $\Delta t = 2.00$ , and  $x_0 = 0 \text{ m}$ .
2. Enter the known values into the equation.

$$x = x_0 + (\text{average } v)t = 0 + (4.00 \text{ m/s})(120 \text{ s}) = 480 \text{ m}$$

### Discussion

Velocity and final displacement are both positive, which means they are in the same direction.

The equation  $x = x_0 + (\text{average } v)t$  gives insight into the relationship between **displacement**, **average velocity**, and **time**. It shows, for example, that displacement is a linear function of average velocity. On a car trip, for example, we will get twice as far in a given time if we average 90 km/h than if we average 45 km/h.

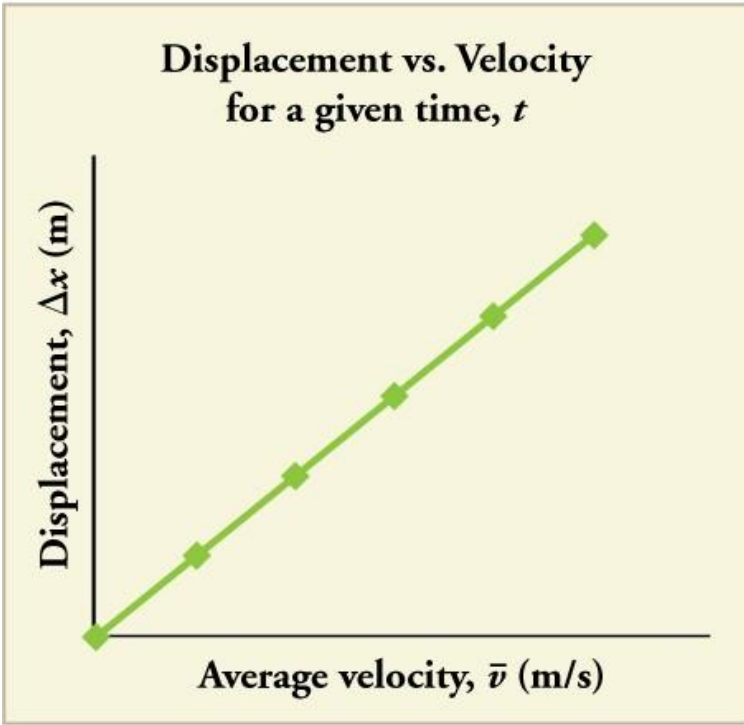


Fig. 2.19 There is a linear relationship between displacement and average velocity. For a given time  $t$ , an object moving twice as fast as another object will move twice as far as the other object.

## Solving For Final Velocity

We can derive another useful equation by manipulating the definition of acceleration.

$$\mathbf{a} = \Delta \mathbf{v} / \Delta t$$

Substituting the simplified notation for  $\Delta v$  and  $\Delta t$  gives us constant  $a = (v - v_0) / t$ .

Solving for  $v$  yields

$$\mathbf{v = v_0 + at \quad (constant\ a)}.$$

### Example 2.7

#### CALCULATING FINAL VELOCITY: AN AIRPLANE SLOWING DOWN AFTER LANDING

An airplane lands with an initial velocity of 70.0 m/s and then decelerates at  $1.50 \text{ m/s}^2$  for 40.0 s. What is its final velocity?

#### **Strategy**

Draw a sketch. We draw the acceleration vector in the direction opposite the velocity vector because the plane is decelerating.

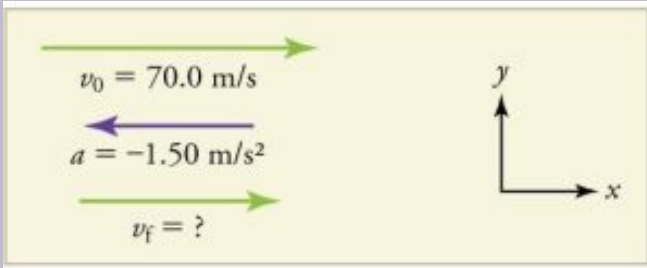


Fig 2.20

### Solution

1. Identify the knowns.  $v_0 = 70.0 \text{ m/s}$ ,  $a = -1.50 \text{ m/s}^2$ ,  $t = 40.0 \text{ s}$
2. Identify the unknown. In this case, it is final velocity,  $v_f$ .
3. Determine which equation to use. We can calculate the final velocity using the equation  $v = v_0 + at$ .
4. Plug in the known values and solve.

$$v = v_0 + at = (70.0 \text{ m/s}) + (-1.50 \text{ m/s}^2)(40.0 \text{ s}) = 10.0 \text{ m/s}$$

### Discussion

The final velocity is much less than the initial velocity, as desired when slowing down, but still positive. With jet engines, reverse thrust could be maintained long enough to stop the plane and start

moving it backward. That would be indicated by a negative final velocity, which is not the case here.



Fig. 2.21 The airplane lands with an initial velocity of 70.0 m/s and slows to a final velocity of 10.0 m/s before heading for the terminal. Note that the acceleration is negative because its direction is opposite to its velocity, which is positive.

In addition to being useful in problem solving, the equation  $v = v_0 + at$  gives us insight into the relationships among velocity, acceleration, and time. From it we can see, for example, that

- final velocity depends on how large the acceleration is and how long it lasts
- if the acceleration is zero, then the final velocity equals the initial velocity ( $v = v_0$ ), as expected (i.e., velocity is constant)
- if  $a$  is negative, then the final velocity is less than the initial velocity



(All of these observations fit our intuition, and it is always useful to examine basic equations in light of our intuition and experiences to check that they do indeed describe nature accurately.)

## Making Connections: Real-World Connection



Fig. 2.22 The Space Shuttle Endeavor blasts off from the Kennedy Space Center in February 2010.

An intercontinental ballistic missile (ICBM) has a larger average acceleration than the Space Shuttle and achieves a

greater velocity in the first minute or two of flight (actual ICBM burn times are classified—short-burn-time missiles are more difficult for an enemy to destroy). But the Space Shuttle obtains a greater **final velocity**, so that it can orbit the earth rather than come directly back down as an ICBM does. The Space Shuttle does this by accelerating for a longer time.

## Solving For Final Position When Velocity Is Not Constant ( $a \neq 0$ )

We can combine the equations above to find a third equation that allows us to calculate the final position of an object experiencing constant acceleration. We start with

$$\mathbf{v} = \mathbf{v}_0 + \mathbf{a}t.$$

Adding  $v_0$  to each side of this equation and dividing by 2 gives

$$(\mathbf{v}_0 + \mathbf{v})/2 = \mathbf{v}_0 + (1/2)\mathbf{a}t.$$

Since  $(\mathbf{v}_0 + \mathbf{v})/2 = \mathbf{v}$  for constant acceleration, then  $\mathbf{v} = \mathbf{v}_0 + (1/2)\mathbf{a}t^2$

Now we substitute this expression for  $v$  into the equation for displacement,  $\mathbf{x} = \mathbf{x}_0 + \mathbf{v}t$ , yielding

$$\mathbf{x} = \mathbf{x}_0 + \mathbf{v}_0t + (1/2)\mathbf{a}t^2 \text{ (constant } \mathbf{a}).$$

### Example 2.8

## CALCULATING DISPLACEMENT OF AN ACCELERATING OBJECT: DRAGSTERS

Dragsters can achieve average accelerations of  $26.0 \text{ m/s}^2$ . Suppose such a dragster accelerates from rest at this rate for  $5.56 \text{ s}$ . How far does it travel in this time?



Fig 2.23 U.S. Army Top Fuel pilot Tony “The Sarge” Schumacher begins a race with a controlled burnout.

### Strategy

Draw a sketch

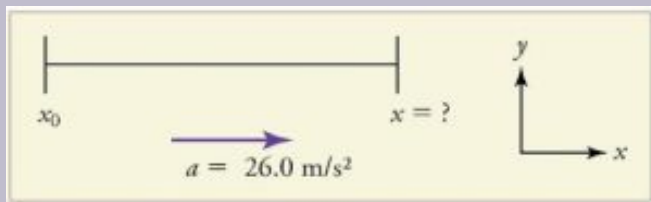


Fig 2.24

We are asked to find displacement, which is  $x$  if we take  $x_0$  to be zero. (Think about it like the starting line of a race. It can be anywhere, but we call it 0 and measure all other positions relative to it.) We can use the equation  $x = x_0 + v_0t + (1/2)at^2$  once we identify  $v_0$ , and  $t$  from the statement of the problem.

### Solution

1. Identify the knowns. Starting from rest means that  $v_0 = 0$ ,  $a$  is given as  $26.0 \text{ m/s}^2$  is given as  $5.56 \text{ s}$ .
2. Plug the known values into the equation to solve for the unknown  $x$ .  $x = x_0 + v_0t + (1/2)at^2$

Since the initial position and velocity are both zero, this simplifies to

$$x = (1/2)at^2.$$

Substituting the identified values of  $a$  and  $t$

gives

$$x = (1/2)(26.0 \text{ m/s}^2)(5.56 \text{ s})^2,$$

yielding

$$x = 402 \text{ m.}$$

### Discussion

If we convert 402 m to miles, we find that the distance covered is very close to one quarter of a mile, the standard distance for drag racing. So the answer is reasonable. This is an impressive displacement in only 5.56 s, but top-notch dragsters can do a quarter mile in even less time than this.

What else can we learn by examining the equation  $x = x_0 + v_0 t + (1/2)at^2$ ? We see that:

- displacement depends on the square of the elapsed time when acceleration is not zero. In Example 7.8, the dragster covers only one fourth of the total distance in the first half of the elapsed time
- if acceleration is zero, then the initial velocity equals average velocity ( $v_0 = v$ ) and  $x = x_0 + v_0 t + (1/2)at^2$  becomes  $x = x_0 + v_0 t$

## Solving For Final Velocity When Velocity Is Not Constant ( $a \neq 0$ )

A fourth useful equation can be obtained from another algebraic manipulation of previous equations.

If we solve  $\mathbf{v} = \mathbf{v}_0 + \mathbf{a}t$ , we get

$$\mathbf{t} = (\mathbf{v} - \mathbf{v}_0)/\mathbf{a}.$$

Substituting this and **average  $\mathbf{v} = (\mathbf{v}_0 + \mathbf{v})/2$**  into  $\mathbf{x} = \mathbf{x}_0 + (\mathbf{average\ v})t$ , we get

$$\mathbf{v}^2 = \mathbf{v}_0^2 + 2\mathbf{a}(\mathbf{x} - \mathbf{x}_0) \text{ (constant } \mathbf{a}).$$

An examination of the equation  $\mathbf{v}^2 = \mathbf{v}_0^2 + 2\mathbf{a}(\mathbf{x} - \mathbf{x}_0)$  can produce further insights into the general relationships among physical quantities:

1. The final velocity depends on how large the acceleration is and the distance over which it acts.
2. For a fixed deceleration, a car that is going twice as fast doesn't simply stop in twice the distance—it takes much further to stop. (This is why we have reduced speed zones near schools.)

## Summary Of Kinematic Equations (Constant $\mathbf{a}$ )

$$\mathbf{x} = \mathbf{x}_0 + (\mathbf{average\ v})t$$

$$\mathbf{average\ v} = (\mathbf{v}_0 + \mathbf{v})/2$$

$$\mathbf{v} = \mathbf{v}_0 + \mathbf{a}t$$

$$\mathbf{x} = \mathbf{x}_0 + \mathbf{v}_0t + (1/2)\mathbf{a}t^2$$

$$\mathbf{v}^2 = \mathbf{v}_0^2 + 2\mathbf{a}(\mathbf{x} - \mathbf{x}_0)$$

## Check Your Understanding



*An interactive H5P element has been excluded from this version of the text. You can view it online here:*

<https://louis.pressbooks.pub/physicalscience1/?p=64#h5p-13>

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3.

# FORCE, WORK, AND POWER

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## Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

- Explain and apply Newton's first, second, and third laws. (CLO1)(CLO2)(CLO3)
- Identify various forces in a diagram or problem (CLO2) (CLO3)
- Translate a free-body diagram into a mathematical representation. (CLO2)(CLO5)
- Calculate mass, weight, and force (CLO2)(CLO5)

- Explain what is meant by normal force. (CLO2)(CLO3)
- Explain the role that inertial plays in effecting motion (CLO2)(CLO3)
- Discuss the requirements for and consequences of balanced and unbalanced forces(CLO2) (CLO3)
- Calculate the net force acting on objects and their resulting accelerations. (CLO2)(CLO5)
- Identify action-reaction pairs (CLO2)(CLO3)
- Define and calculate work using appropriate SI units (CLO2)(CLO3)(CLO5)
- Discuss the components needed for work to occur (CLO2)(CLO3)
- Define and calculate power based on work and time. (CLO2)(CLO3)(CLO5)



Fig 3.1 Newton's laws of motion describe the motion of the dolphin's path.

## Dynamics and the Laws of Motion

Motion draws our attention. Motion itself can be beautiful, causing us to marvel at the forces needed to achieve spectacular motion, such as that of a dolphin jumping out of the water, or a pole vaulter, or the flight of a bird, or the orbit of a satellite. The study of motion is kinematics, but kinematics only *describes* the way objects move—their velocity and their acceleration. Dynamics considers the forces that affect the motion of moving objects and systems. Newton's laws of motion are the foundation of dynamics. These laws provide

an example of the breadth and simplicity of principles under which nature functions. They are also universal laws in that they apply to similar situations on Earth as well as in space.

**Dynamics** is the study of the forces that cause objects and systems to move. To understand this, we need a working definition of **force**. Our intuitive definition of force—that is, a push or a pull—is a good place to start. We know that a push or pull has both magnitude and direction (therefore, it is a **vector** quantity) and can vary considerably in each regard. For example, a cannon exerts a strong force on a cannonball that is launched into the air. In contrast, Earth exerts only a tiny downward pull on a flea. Our everyday experiences also give us a good idea of how multiple forces add together.

Experience suggests that an object at rest will remain at rest if left alone, and that an object in motion tends to slow down and stop unless some effort is made to keep it moving. What Newton’s first law of motion states, however, is the following:

## Newton's First Law of Motion

**Newton’s first law of motion** states that a body at rest remains at rest, or, if in motion, remains in motion at a constant velocity unless acted on by a net external force. This is also called the **law of inertia**.

Note the repeated use of the verb “remains.” We can think of this law as preserving the status quo of motion.

Rather than contradicting our experience, Newton’s first

law of motion states that there must be a cause (which is a net external force) *for there to be any change in **velocity** (either a change in magnitude or direction)*. We will define *net external force* in the next section. An object sliding across a table or floor slows down due to the net force of **friction** acting on the object. If friction disappeared, would the object still slow down?

The idea of cause and effect is crucial in accurately describing what happens in various situations. For example, consider what happens to an object sliding along a rough horizontal surface. The object quickly grinds to a halt. If we spray the surface with talcum powder to make the surface smoother, the object slides farther. If we make the surface even smoother by rubbing lubricating oil on it, the object slides even further. Extrapolating to a frictionless surface, we can imagine the object sliding in a straight line indefinitely. Friction is thus the *cause* of the slowing (consistent with Newton's first law). The object would not slow down at all if friction were completely eliminated. Consider an air hockey table. When the air is turned off, the puck slides only a short distance before friction slows it to a stop. However, when the air is turned on, it creates a nearly frictionless surface, and the puck glides long distances without slowing down. Additionally, if we know enough about the friction, we can accurately predict how quickly the object will slow down. Friction is an external force.

**Newton's first law** is completely general and can be

applied to anything from an object sliding on a table to a satellite in orbit to blood pumped from the heart. Experiments have thoroughly verified that any change in velocity (speed or direction) must be caused by an external force. The idea of *generally applicable or universal laws* is important not only here—it is a basic feature of all laws of physics. Identifying these laws is like recognizing patterns in nature from which further patterns can be discovered. The genius of Galileo, who first developed the idea for the first law, and Newton, who clarified it, was to ask the fundamental question, “What is the cause?” Thinking in terms of cause and effect is a worldview fundamentally different from the typical ancient Greek approach when questions such as “Why does a tiger have stripes?” would have been answered in Aristotelian fashion, “That is the nature of the beast.” True perhaps, but not a useful insight.

**Newton’s second law of motion** is closely related to Newton’s first law of motion. It mathematically states the cause and effect relationship between force and changes in motion. Newton’s second law of motion is more quantitative and is used extensively to calculate what happens in situations involving a force.

## Newton’s Second Law of Motion

Before we can write down Newton’s second law as a simple equation giving the exact relationship of **force**, **mass**, and



**acceleration**, we need to sharpen some ideas that have already been mentioned. First, what do we mean by a change in motion? The answer is that a change in motion is equivalent to a change in velocity. A change in velocity means, by definition, that there is an acceleration. Newton's first law says that a **net external force** causes a change in motion; thus, we see that a net external force causes acceleration. Another question immediately arises. What do we mean by an external force? An intuitive notion of external is correct—an external force acts from outside the system (object or collection of objects) of interest.

For example, in Figure 3.2(a) the system of interest is the wagon plus the child in it. The two forces exerted by the other children are external forces. An internal force acts between elements of the system. Again looking at Figure 3.2(a), the force the child in the wagon exerts to hang onto the wagon is an internal force between elements of the system of interest. Only external forces affect the motion of a system, according to Newton's first law. You must define the boundaries of the system before you can determine which forces are external.

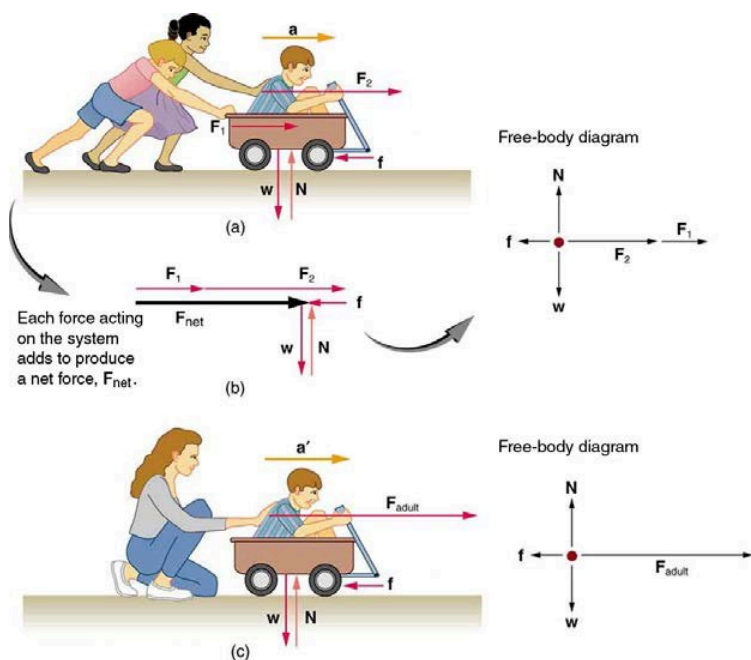


Fig. 3.2 Different forces exerted on the same mass produce different accelerations. (a) Two children push a wagon with a child in it. Arrows representing all external forces are shown. The system of interest is the wagon and its rider. The weight  $w$  of the system and the support of the ground  $N$  are also shown for completeness and are assumed to cancel. The vector  $f$  represents the friction acting on the wagon, and it acts to the left, opposing the motion of the wagon. (b) All of the external forces acting on the system add together to produce a net force,  $F_{net}$ . The free-body diagram shows all of the forces acting on the system of interest. The dot represents the center of mass of the system. Each force vector extends from this dot. Because there are two forces acting to the right, we draw the vectors collinearly. (c) A larger net external force produces a larger acceleration ( $a' > a$ ) when an adult pushes the child.

Now, it seems reasonable that acceleration should be directly proportional to and in the same direction as the net (total) external force acting on a system. This assumption has been verified experimentally and is illustrated in Figure 3.2. In part (a), a smaller force causes a smaller acceleration than the larger force illustrated in part (c). For completeness, the vertical forces are also shown; they are assumed to cancel since there is no acceleration in the vertical direction. The vertical forces are the weight  $\mathbf{w}$  and the support of the ground  $\mathbf{N}$ , and the horizontal force  $\mathbf{f}$  represents the force of **friction**. These will be discussed in more detail in later sections. For now, we will define friction as a force that opposes the motion past each other of objects that are touching. Figure 3.2(b) shows how vectors representing the external forces add together to produce a net force,  $\mathbf{F}_{\text{net}}$  or  $\Sigma\mathbf{F}$ .

**Newton's second law of motion** is that the acceleration of a system is directly proportional to and in the same direction as the net external force acting on the system, and inversely proportional to its mass.

In equation form, Newton's second law of motion is

$$\mathbf{a} = \Sigma\mathbf{F}/m.$$

This is often written in the more familiar form

$$\Sigma\mathbf{F} = m\mathbf{a}.$$

Although these two equations are really the same, the first gives more insight into what Newton's second law means. The law is a *cause and effect relationship* among three quantities

that is not simply based on their definitions. The validity of the second law is completely based on experimental verification.

## Units of Force

The equation,  $\Sigma \mathbf{F} = m\mathbf{a}$ , is used to define the units of force in terms of the three basic units for mass, length, and time. The **SI unit** of **force** is called the **newton** (abbreviated N) and is the force needed to accelerate a 1 kg system at the rate of  $1 \text{ m/s}^2$ . That is, since  $\Sigma \mathbf{F} = m\mathbf{a}$ ,

$$1 \text{ N} = 1 \text{ kg} \cdot \text{m/s}^2$$

While almost the entire world uses the newton for the unit of force, in the United States the most familiar unit of force is the **pound** (lb), where  $1 \text{ N} = 0.225 \text{ lb}$ .

***Net external force*** is the sum of all of the forces acting on an object. For example, if there is a force with a magnitude of 8 Newtons in the positive direction and a force with a magnitude of 7 Newtons in the negative direction on an object, the net external force would be  $8 + (-7) = 1$  Newton. If the forces acting on an object are of equal magnitude in opposite directions, the net external force would be equal to 0.

### Example 3.1

## WHAT ACCELERATION CAN A PERSON PRODUCE WHEN PUSHING A LAWN MOWER?

Suppose that the net external force (push minus friction) exerted on a lawn mower is 51 N (about 11 lb) parallel to the ground. The mass of the mower is 24 kg. What is its acceleration?



Fig. 3.3 The net force on a lawn mower is 51 N to the right. At what rate does the lawn mower accelerate to the right?

### Strategy

Since  $\Sigma F$  and  $m$  are given, the acceleration can be

calculated directly from Newton's second law as stated in  $\Sigma F = ma$ .

### **Solution**

The magnitude of the acceleration  $a$  is  $a = \Sigma F/m$ .

Entering known values gives

$$a = 51 \text{ N}/24 \text{ kg}$$

Substituting the units  $\text{kg}\cdot\text{m}/\text{s}^2$  for N yields

$$a = (51 \text{ kg}\cdot\text{m}/\text{s}^2)/224 \text{ kg} = 2.1 \text{ m}/\text{s}^2$$

### **Discussion**

The direction of the acceleration is the same direction as that of the net force, which is parallel to the ground. There is no information given in this example about the individual external forces acting on the system, but we can say something about their relative magnitudes. For example, the force exerted by the person pushing the mower must be greater than the friction opposing the motion (since we know the mower moves forward), and the vertical forces must cancel if there is to be no acceleration in the vertical direction (the mower is moving only horizontally). The acceleration found is small enough to be reasonable for a person pushing a mower. Such an effort would not last too long

because the person's top speed would soon be reached.

There is a passage in the musical *Man of la Mancha* that relates to Newton's third law of motion. Sancho, in describing a fight with his wife to Don Quixote, says, "Of course I hit her back, Your Grace, but she's a lot harder than me and you know what they say, 'Whether the stone hits the pitcher or the pitcher hits the stone, it's going to be bad for the pitcher.'" This is exactly what happens whenever one body exerts a **force** on another—the first also experiences a force (equal in magnitude and opposite in direction). Numerous common experiences, such as stubbing a toe or throwing a ball, confirm this. It is precisely stated in Newton's third law of motion.

## Newton's Third Law of Motion

Whenever one body exerts a force on a second body, the first body experiences a force that is equal in magnitude and opposite in direction to the force that it exerts. This is **Newton's third law of motion**.

This law represents a certain *symmetry in nature*: Forces always occur in pairs, and one body cannot exert a force on

another without experiencing a force itself. We sometimes refer to this law loosely as “action-reaction,” where the force exerted is the action and the force experienced as a consequence is the reaction. Newton’s third law has practical uses in analyzing the origin of forces and understanding which forces are external to a system.

We can readily see Newton’s third law at work by taking a look at how people move about. Consider a swimmer pushing off from the side of a pool, as illustrated in Figure 3.4. She pushes against the pool wall with her feet and accelerates in the direction *opposite* to that of her push. The wall has exerted an equal and opposite force back on the swimmer. You might think that two equal and opposite forces would cancel, but they do not *because they act on different systems*. In this case, there are two systems that we could investigate: the swimmer or the wall. If we select the swimmer to be the system of interest, as in the figure, then  $F_{\text{wall on feet}}$  is an external force on this system and affects its motion. The swimmer moves in the direction of  $F_{\text{wall on feet}}$ . In contrast, the force  $F_{\text{feet on wall}}$  acts on the wall and not on our system of interest. Thus  $F_{\text{feet on wall}}$  does not directly affect the motion of the system and does not cancel  $F_{\text{wall on feet}}$ . Note that the swimmer pushes in the direction opposite to that in which she wishes to move. The reaction to her push is thus in the desired direction.



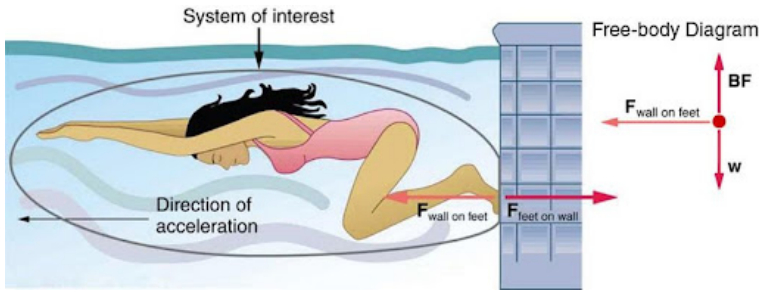


Fig. 3.4 When the swimmer exerts a force on the wall, she accelerates in the direction opposite to that of her push. This means the net external force on her is in the direction opposite to force of her feet on the wall. This opposition occurs because, in accordance with Newton's third law of motion, the wall exerts a force on her, equal in magnitude but in the direction opposite to the one she exerts on it. The line around the swimmer indicates the system of interest. Note that the force of the swimmer's feet on the wall does not act on this system (the swimmer) and, thus, does not cancel the force of the wall on her feet. Thus the free-body diagram shows only  $F_{\text{sub}}(\text{wall on feet})$ ,  $w$ , the gravitational force, and  $BF$ , the buoyant force of the water supporting the swimmer's weight. The vertical forces  $w$  and  $BF$  cancel since there is no vertical motion.

Other examples of Newton's third law are easy to find. As a professor paces in front of a whiteboard, she exerts a force backward on the floor. The floor exerts a reaction force forward on the professor that causes her to accelerate forward. Similarly, a car

[Link to Learning](#)

Use this

interactive simulation by PhET to see Newton's Laws in action (opens in new window): [Forces and Motion](#).

accelerates because the ground pushes forward on the drive wheels in reaction to the drive wheels pushing backward on the ground. You can see evidence of the wheels pushing backward when tires spin on a gravel road and throw rocks backward. In another example, rockets move forward by expelling gas backward at high velocity. This means the rocket exerts a large backward force on the gas in the rocket

combustion chamber, and the gas therefore exerts a large reaction force forward on the rocket. This reaction force is called thrust. It is a common misconception that rockets propel themselves by pushing on the ground or on the air behind them. They actually work better in a vacuum, where they can more readily expel the exhaust gases. Helicopters similarly create lift by pushing air down, thereby experiencing an upward reaction force. Birds and airplanes also fly by exerting force on air in a direction opposite to that of whatever force they need. For example, the wings of a bird force air downward and backward in order to get lift and move forward. An octopus propels itself in the water by ejecting water through a funnel from its body, similar to a jet ski. In a situation similar to Sancho's, professional cage fighters

experience reaction forces when they punch, sometimes breaking their hand by hitting an opponent's body.

## What It Means to Do Work

The scientific definition of **work** differs in some ways from its everyday meaning. Certain things we think of as hard work, such as writing an exam or carrying a heavy load on level ground, are not work as defined by a scientist. The scientific definition of work reveals its relationship to energy—whenever work is done, energy is transferred.

For work, in the scientific sense, to be done, a force must be exerted and there must be displacement in the direction of the force.

### What is Work?

The work done on a system by a constant force is *the product of the component of the force in the direction of motion times the distance through which the force acts*. For one-way motion in one

[Link to Learning](#)

Use this interactive simulation to view the

relationships between distance, work, and force (opens in new window): [PhET Balancing Act Simulation](#).

Here, the only acceleration that is acted on these objects is their gravitational acceleration. Knowing that they all have the same acceleration, we can view the variation of mass across objects to be

dimension, this is expressed in equation form as:

$$W = |F|(\cos\theta)|d|$$

where  $W$  is work,  $d$  is the displacement, and  $\theta$  is the angle between the force vector  $F$  and the displacement vector  $d$ . We can also write this as:

$$W = Fd \cos \theta$$

## What is Power?

*Power*—the word conjures up many images: a professional football player muscling aside his opponent, a dragster roaring away from the starting line, a volcano blowing its lava into the atmosphere, or a rocket blasting off, as in Figure 3.5.

**Power** is the rate at which work is done.

$$P = W/t$$

The **SI unit** for power is the **watt (W)**, where 1 watt equals 1 joule/second ( $1 \text{ W} = 1 \text{ J/s}$ )

Because **work** is energy

transfer, power is also the rate at which energy is expended. A 60-W light bulb, for example, expends 60 J of energy per second. Great power means a large amount of work or energy developed in a short time. For example, when a powerful car accelerates rapidly, it does a large amount of work and consumes a large amount of fuel in a short time.

These images of power have in common the rapid performance of work, consistent with the scientific definition of **power** ( $P$ ) as the rate at which work is done.

proportional to the force acted on those objects. Why does it make sense that when distance increases and work stays the same, the force decreases? If the work does not change, the force must decrease when the distance increases in order to maintain a balanced equation.



Fig. 3.5 This powerful rocket on the Space Shuttle Endeavor did work and consumed energy at a very high rate.

## Simple Machines

A **simple machine** is a mechanical device that allows you to do the same amount of work with reduced effort by increasing the distance traveled to decrease the force. The six basic simple machines are the wheel and axle, inclined plane (or ramp), pulley, lever, wedge and the screw.

## Check Your Understanding



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## 4.

# PROPERTIES OF MATERIALS, CHEMICAL REACTIONS, AND HEAT

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### Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

- Identify the states of matter (CLO 2)
- Write and balance chemical equations (CLO7)
- Define three common types of chemical reactions (precipitation, acid-base, and oxidation-reduction) (CLO6)
- Classify chemical reactions as one of these three types given appropriate descriptions or

- chemical equations (CLO6)(CLO7)
- Identify common acids and bases (CLO6)(CLO8)
  - Define and distinguish between heat and temperature. (CLO2)(CLO3)
  - Express temperature using Fahrenheit, Celsius, and Kelvin scales. (CLO2)(CLO5)
  - Discuss heat transfer applications and relate them to phase changes of matter. (CLO2)(CLO3)
  - Define methods of heat transfer: conduction, convection, and radiation. (CLO2)(CLO3)



(a)



(b)

Fig 4.1 Wax undergoes a physical change when solid wax is heated and forms liquid wax. (b) Steam condensing inside a cooking pot is a physical change, as water vapor is changed into liquid water.

## Properties of Materials

The characteristics that distinguish one substance from another are called **properties**. A **physical property** is a characteristic of matter that is not associated with a change in its chemical composition. Familiar examples of physical properties include density, color, hardness, melting and boiling points, and electrical conductivity. Some physical properties, such as density and color, may be observed without changing the physical state of the matter. Other physical properties, such as the melting temperature of iron or the freezing temperature of water, can only be observed as matter undergoes a physical change.

## States of Matter

Solids, liquids, and gases are the three **states of matter** commonly found on earth (Figure 4.2). A solid is rigid and possesses a definite shape. A liquid flows and takes the shape of its container, except that it forms a flat or slightly curved upper surface when acted upon by **gravity**. (In zero gravity, liquids assume a spherical shape.) Both liquid and solid samples have volumes that are very nearly independent of pressure. A gas takes both the shape and volume of its container.

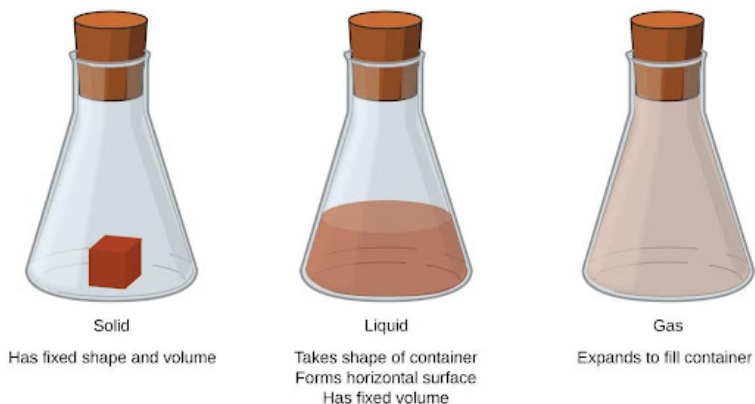


Fig. 4.2 The three most common states or phases of matter are solid, liquid, and gas.

A fourth state of matter, plasma, occurs naturally in the interiors of stars. A plasma is a gaseous state of matter that contains appreciable numbers of electrically charged particles (Figure 4.3). The presence of these charged particles imparts unique properties to plasmas that justify their classification as a state of matter distinct from gases. In addition to stars, plasmas are found in some other high-temperature environments (both natural and man-made), such as lightning strikes, certain television screens, and specialized analytical instruments used to detect trace amounts of metals.



Fig. 4.3 A plasma torch can be used to cut metal.

### Link to Learning

In a tiny cell in a plasma television, the plasma emits ultraviolet light,

Some samples of matter appear to have properties of solids, liquids, and/or gases at the same time. This can occur when the sample is composed of many small pieces. For example, we can pour sand as if it were a liquid because it is composed of many small grains of solid sand. Matter can also have **properties** of more than one state when it is a mixture,

such as with clouds. Clouds appear to behave somewhat like gases, but they are actually mixtures of air (gas) and tiny particles of water (liquid or solid).

## Classification of Matter

Matter can be classified into several categories. Two broad categories are mixtures and pure substances. A **pure substance** has a constant composition. All specimens of a pure substance have exactly the same makeup and properties. Any sample of sucrose (table sugar) consists of 42.1% carbon, 6.5% hydrogen, and 51.4% oxygen by mass. Any sample of sucrose also has the same physical properties, such as melting point, color, and sweetness, regardless of the source from which it is isolated.

Pure substances may be divided into two classes: elements and compounds. Pure substances that cannot be broken down

which in turn causes the display at that location to appear a specific color. The composite of these tiny dots of color makes up the image that you see. Watch this video to learn more about plasma and the places you encounter it (opens in new window):

[What is Plasma?](#)

into simpler substances by chemical changes are called **elements**. Iron, silver, gold, aluminum, sulfur, oxygen, and copper are familiar examples of the more than 100 known elements, of which about 90 occur naturally on the earth, and two dozen or so have been created in laboratories.

Pure substances that are comprised of two or more elements are called **compounds**. Compounds may be broken down by chemical changes to yield either elements or other compounds, or both.

The properties of combined elements are different from those in the free, or uncombined, state. For example, white crystalline sugar (sucrose) is a compound resulting from the chemical combination of the element carbon, which is a black solid in one of its uncombined forms, and the two elements hydrogen and oxygen, which are colorless gases when uncombined. Free sodium, an element that is a soft, shiny, metallic solid, and free chlorine, an element that is a yellow-green gas, combine to form sodium chloride (table salt), a compound that is a white, crystalline solid.

A **mixture** is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes, such as evaporation (you will learn more about this later). A mixture with a composition that varies from point to point is called a **heterogeneous mixture**. Italian dressing is an example of a heterogeneous mixture. Its composition can vary because it may be prepared from varying amounts of oil, vinegar, and herbs. It is not the same from



point to point throughout the mixture—one drop may be mostly vinegar, whereas a different drop may be mostly oil or herbs because the oil and vinegar separate and the herbs settle. Other examples of heterogeneous mixtures are chocolate chip cookies (we can see the separate bits of chocolate, nuts, and cookie dough) and granite (we can see the quartz, mica, feldspar, and more).

A **homogeneous mixture**, also called a **solution**, exhibits a uniform composition and appears visually the same throughout. An example of a solution is a sports drink, consisting of water, sugar, coloring, flavoring, and electrolytes mixed together uniformly. Each drop of a sports drink tastes the same because each drop contains the same amounts of water, sugar, and other components. Note that the composition of a sports drink can vary—it could be made with somewhat more or less sugar, flavoring, or other components, and still be a sports drink. Other examples of homogeneous mixtures include air, maple syrup, gasoline, and a solution of salt in water.

Although there are just over 100 elements, tens of millions of chemical compounds result from different combinations of these elements. Each compound has a specific composition and possesses definite chemical and physical properties that distinguish it from all other compounds. And, of course, there are innumerable ways to combine elements and compounds to form different mixtures.

## Atoms and Molecules

An **atom** is the smallest particle of an element that has the properties of that element and can enter into a chemical combination. Consider the element gold, for example. Imagine cutting a gold nugget in half, then cutting one of the halves in half, and repeating this process until a piece of gold remained that was so small that it could not be cut in half (regardless of how tiny your knife may be). This minimally sized piece of gold is an atom (from the Greek *atomos*, meaning “indivisible”). Such an atom would no longer be gold if it were divided further.

An atom is so small that its size is difficult to imagine. One of the smallest things we can see with our unaided eye is a single thread of a spider web: These strands are about 1/10,000 of a centimeter (0.0001 cm) in diameter. Although the cross-section of one strand is almost impossible to see without a microscope, it is huge on an atomic scale. A single carbon atom in the web has a diameter of about 0.000000015 centimeter, and it would take about 7000 carbon atoms to span the diameter of the strand. An atom is so light that its mass is also difficult to imagine. A billion lead atoms (1,000,000,000 atoms) weigh about  $3 \times 10^{-13}$  grams, a mass that is far too light to be weighed on even the world’s most sensitive balances. It would require over 300,000,000,000,000 lead atoms (300 trillion, or  $3 \times 10^{14}$ ) to be weighed, and they would weigh only 0.0000001 gram.

It is rare to find collections of individual atoms. Only a few elements, such as the gases helium, neon, and argon, consist of a collection of individual atoms that move about independently of one another. Other elements, such as the gases hydrogen, nitrogen, oxygen, and chlorine, are composed of units that consist of pairs of atoms. One form of the element phosphorus consists of units composed of four phosphorus atoms. The element sulfur exists in various forms, one of which consists of units composed of eight sulfur atoms. These units are called molecules. A **molecule** consists of two or more atoms joined by strong forces called chemical bonds. The atoms in a molecule move around as a unit, much like the cans of soda in a six-pack or a bunch of keys joined together on a single key ring. A molecule may consist of two or more identical atoms, as in the molecules found in the elements hydrogen, oxygen, and sulfur, or it may consist of two or more different atoms, as in the molecules found in water. Each water molecule is a unit that contains two hydrogen atoms and one oxygen atom. Each glucose molecule is a unit that contains 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. Like atoms, molecules are incredibly small and light. If an ordinary glass of water were enlarged to the size of the earth, the water molecules inside it would be about the size of golf balls.

## Physical and Chemical Changes

A **physical change** is a change in the state or properties of matter without any accompanying change in the chemical identities of the substances contained in the matter. Physical changes are observed when wax melts, when sugar dissolves in coffee, and when steam condenses into liquid water (Figure 4.1). Other examples of physical changes include magnetizing and demagnetizing metals (as is done with common anti-theft security tags) and grinding solids into powders (which can sometimes yield noticeable changes in color). In each of these examples, there is a change in the physical state, form, or properties of the substance, but no change in its chemical composition.

The change of one type of matter into another type (or the inability to change) is a chemical property. Examples of chemical properties include flammability, toxicity, acidity, and many other types of reactivity. Iron, for example, combines with oxygen in the presence of water to form rust; chromium does not oxidize (Figure 4.4). Nitroglycerin is very dangerous because it explodes easily; neon poses almost no hazard because it is very unreactive.

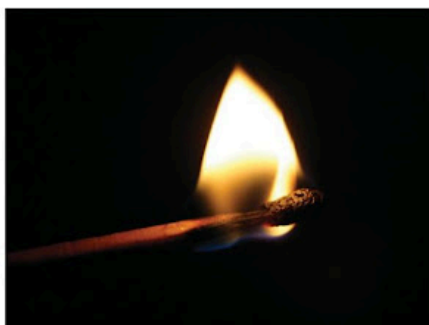


Fig. 4.4 (a) One of the chemical properties of iron is that it rusts; (b) one of the chemical properties of chromium is that it does not.

A **chemical change** always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure 4.5).



(a)



(b)



(c)



(d)

Fig 4.5 (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form.

## Balancing Equations

The **chemical equation** described in section 4.1 is balanced, meaning that equal numbers of atoms for each element involved in the reaction are represented on the reactant (left)

and product (right) sides. This is a requirement the equation must satisfy to be consistent with the law of conservation of matter. It may be confirmed by simply summing the numbers of atoms on either side of the arrow and comparing these sums to ensure they are equal. Note that the number of atoms for a given element is calculated by multiplying the coefficient of any formula containing that element by the element's subscript in the formula. If an element appears in more than one formula on a given side of the equation, the number of atoms represented in each must be computed and then added together. For example, both product species in the example reaction, CO<sub>2</sub> and H<sub>2</sub>O, contain the element oxygen, and so the number of oxygen atoms on the product side of the equation is

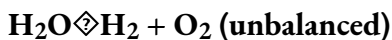
$$(1 \text{ CO}_2 \text{ molecule} \times 2 \text{ O atoms per CO}_2 \text{ molecule}) + (2 \text{ H}_2\text{O molecules} \times 1 \text{ O atom per H}_2\text{O molecule}) = 4 \text{ O atoms}$$

The equation for the reaction between methane and oxygen to yield carbon dioxide and water is confirmed to be balanced per this approach, as shown here:



Element	Reactants	Products	Balanced?
C	1×1 = 1	1×1 = 1	1 = 1, yes
H	4×1 = 4	2×2 = 4	4 = 4, yes
O	2×2 = 4	(1×2) + (2×1) = 4	4 = 4, yes

A **balanced chemical equation** often may be derived from a qualitative description of some chemical reaction by a fairly simple approach known as balancing by inspection. Consider as an example the decomposition of water to yield molecular hydrogen and oxygen. This process is represented qualitatively by an unbalanced chemical equation:

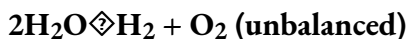


Comparing the number of H and O atoms on either side of this equation confirms its imbalance:

Element	Reactants	Products	Balanced?
H	$1 \times 2 = 2$	$1 \times 2 = 2$	$2 = 2$ , yes
O	$1 \times 1 = 1$	$1 \times 2 = 2$	$1 \neq 2$ , no

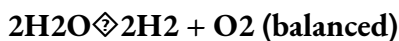
The numbers of H atoms on the reactant and product sides of the equation are equal, but the numbers of O atoms are not. To achieve balance, the coefficients of the equation may be changed as needed. Note that the coefficients must be whole numbers and cannot be fractions or decimals. Keep in mind, of course, that the formula subscripts define, in part, the identity of the substance, and so these cannot be changed without altering the qualitative meaning of the equation. For example, changing the reactant formula from  $\text{H}_2\text{O}$  to  $\text{H}_2\text{O}_2$  would yield balance in the number of atoms, but doing so also changes the reactant's identity (it's now hydrogen peroxide and not water). The O atom balance may be achieved by changing the coefficient for  $\text{H}_2\text{O}$  to 2.





Element	Reactants	Products	Balanced?
H	$2 \times 2 = 4$	$1 \times 2 = 2$	$4 \neq 2$ , no
O	$2 \times 1 = 2$	$1 \times 2 = 2$	$2 = 2$ , yes

The H atom balance was upset by this change, but it is easily reestablished by changing the coefficient for the  $\text{H}_2$  product to 2.



Element	Reactants	Products	Balanced?
H	$2 \times 2 = 4$	$2 \times 2 = 4$	$4 = 4$ , yes
O	$2 \times 1 = 2$	$1 \times 2 = 2$	$2 = 2$ , yes

These coefficients yield equal numbers of both H and O atoms on the reactant and product sides, and the balanced equation is, therefore:



Example 4.1

Write a balanced equation for the decomposition of ammonium nitrate to form molecular nitrogen, molecular oxygen, and water. (Hint: Balance oxygen last, since it is present in more than one molecule on the right side of the equation.)



[Link to Learning](#)

Use this interactive tutorial by PhET for additional practice balancing

## Types of Chemical Reactions

Humans interact with one another in various and complex ways, and we classify these interactions according to common patterns of behavior. When two humans exchange information, we say they are communicating. When they hit

each other with their fists, we say they are fighting. Faced with a wide range of varied interactions between chemical substances, scientists have likewise found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. This module will provide an introduction to three

of the most prevalent types of **chemical reactions**: precipitation, acid-base, and oxidation-reduction.

equations  
(opens in new  
window):

[Balancing  
Chemical  
Equations.](#)

## Precipitation Reactions

A **precipitation reaction** is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as double displacement, double replacement, or metathesis reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals.

The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its solubility,

defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be soluble. A substance will precipitate when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be insoluble, and these are the substances that readily precipitate from solution.

Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow (Figure 4.6). The properties of pure  $\text{PbI}_2$  crystals make them useful for fabrication of X-ray and gamma ray detectors.



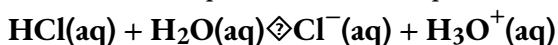
Fig 4.6 A precipitate of  $PbI_2$  forms when solutions containing  $Pb^{2+}$  and  $I^-$  are mixed.

## Acid-Base Reactions

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.

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,  $\text{H}_3\text{O}^+$ . As an example, consider the equation shown here:



The process represented by this equation confirms that hydrogen chloride is an acid. When dissolved in water,  $\text{H}_3\text{O}^+$  ions are produced by a chemical reaction in which  $\text{H}^+$  ions are transferred from HCl molecules to  $\text{H}_2\text{O}$  molecules (Figure 4.7).

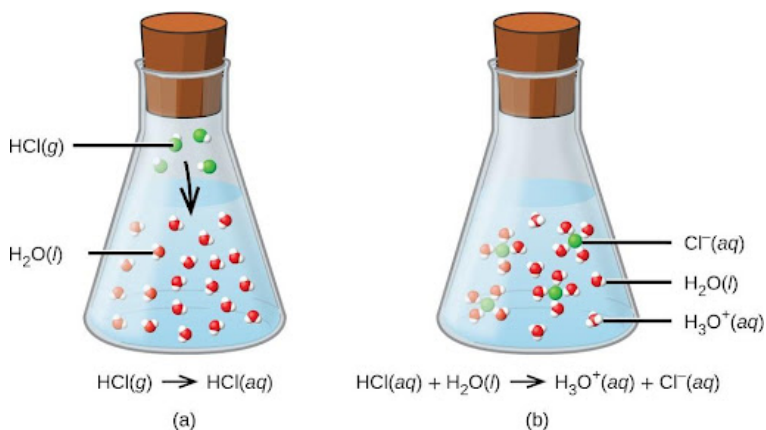


Fig. 4.7 When hydrogen chloride gas dissolves in water, (a) it reacts as an acid, transferring protons to water molecules to yield (b) hydronium ions (and solvated chloride ions).

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. 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 (Figure 4.8).

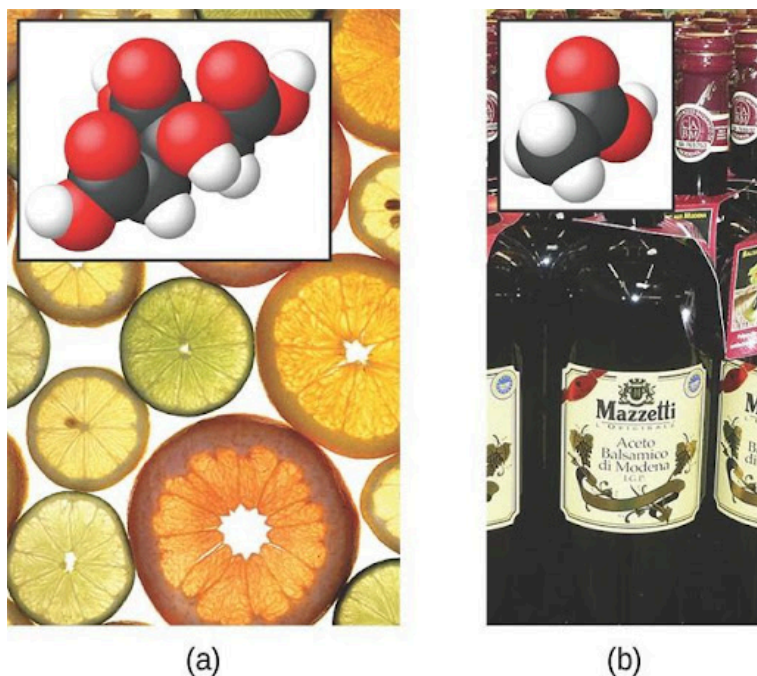


Fig 4.8 Fruits such as oranges, lemons, and grapefruit contain the weak acid citric acid. (b) Vinegars contain the weak acid acetic acid.

A **base** is a substance that will dissolve in water to yield hydroxide ions,  $\text{OH}^-$ . 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,  $\text{NaOH}$  and  $\text{Ca}(\text{OH})_2$ . Unlike the acid compounds discussed previously, these compounds do not react chemically with water; instead they dissolve and dissociate, releasing hydroxide ions directly into the solution. For example,  $\text{KOH}$  and  $\text{Ba}(\text{OH})_2$  dissolve in water and dissociate completely to



produce cations ( $\text{K}^+$  and  $\text{Ba}^{2+}$ , respectively) and hydroxide ions,  $\text{OH}^-$ . 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:



This equation confirms that sodium hydroxide is a base. When dissolved in water,  $\text{NaOH}$  dissociates to yield  $\text{Na}^+$  and  $\text{OH}^-$  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,  $\text{NaOH}$  and other ionic hydroxides are all classified as strong bases.

### Chemistry in Everyday Life: Stomach Antacids

Our stomachs contain a solution of roughly 0.03 M  $\text{HCl}$  (molarity), 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,  $\text{CaCO}_3$ . The reaction:  $\text{CaCO}_3 + 2\text{HCl} \rightleftharpoons \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2$  not only neutralizes stomach acid, it also produces  $\text{CO}_2$ , which may result in a satisfying belch.

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure 4.9). 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.



Fig 4.9 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.

### Chemistry in Everyday Life: Culinary Chemistry

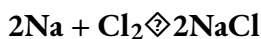
Examples of acid-base chemistry are abundant in the culinary world. One example is the use of baking soda, or sodium bicarbonate, in baking.  $\text{NaHCO}_3$  is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from 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.

## Oxidation-Reduction Reactions

Earth's atmosphere contains about 20% molecular oxygen, O<sub>2</sub>, a chemically reactive gas that plays an essential role in the metabolism of aerobic organisms and in many environmental processes that shape the world. The term oxidation was originally used to describe chemical reactions involving O<sub>2</sub>, but its meaning has evolved to refer to a broad and important reaction class known as oxidation-reduction (redox)

reactions. A few examples of such reactions will be used to develop a clear picture of this classification.

Some **redox reactions** involve the transfer of electrons between reactant species to yield ionic products, such as the reaction between sodium and chlorine to yield sodium chloride:



For redox reactions of this sort, the loss and gain of electrons define the complementary processes that occur:

**Oxidation** = loss of electrons

**Reduction** = gain of electrons

### Link to Learning

Explore the microscopic view of strong and weak acids and bases (opens in new window): [PhET Acid-Base Solutions](#).

In this reaction, then, sodium is oxidized and chlorine undergoes reduction. Viewed from a more active perspective, sodium functions as a reducing agent (reductant), since it provides electrons to (or reduces) chlorine. The chlorine functions as an oxidizing agent (oxidant), as it effectively removes electrons from (oxidizes) sodium.

Reducing agent = species that is oxidized

Oxidizing agent = species that is reduced

## Thermal Energy, Temperature, and Heat

**Thermal energy** is **kinetic energy** associated with the random motion of atoms and molecules. **Temperature** is a quantitative measure of “hot” or “cold.” When the atoms and molecules in an object are moving or vibrating quickly, they have a higher average kinetic energy (KE), and we say that the object is “hot.” When the atoms and molecules are moving slowly, they have lower average KE, and we say that the object is “cold” (Figure 4.10). Assuming that no **chemical reaction** or phase change (such as melting or vaporizing) occurs, increasing the amount of thermal energy in a sample of matter will cause its temperature to increase. And, assuming that no chemical reaction or phase change (such as condensation or freezing) occurs, decreasing the amount of thermal energy in a sample of matter will cause its temperature to decrease.

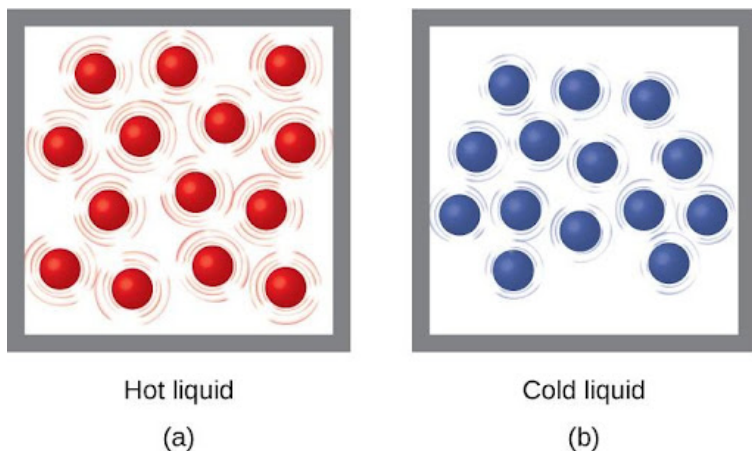


Fig. 4.10 (a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.

Most substances expand as their temperature increases and contract as their temperature decreases. This property can be used to measure temperature changes, as shown in Figure 4.11. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

### Link to Learning

Click on this interactive simulation to view the effects of temperature on molecular

motion (opens  
in new  
window): [PhET  
States of  
Matter](#).

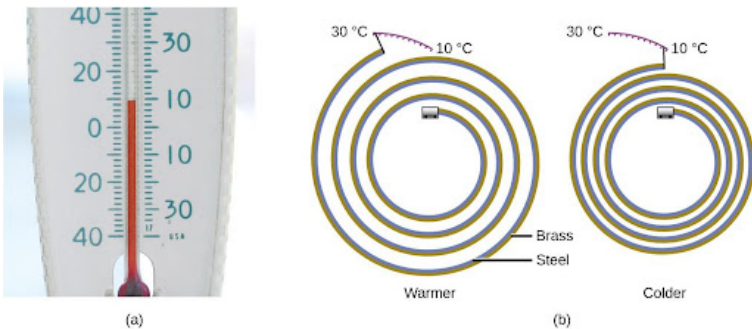


Fig. 4.11 (a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. When heated or cooled, one of the metals (brass) expands or contracts more than the other metal (steel), causing the strip to coil or uncoil. Both types of thermometers have a calibrated scale that indicates the temperature.



## Heat Transfer – Conduction, Convection and Radiation

Heat transfer can happen three different ways; through **conduction**, **convection** and **radiation**. Conduction is heat transfer through contact between atoms. The conductivity of materials is influenced by their atomic structure and bonds. For example, metals are very conductive because they have free flowing electrons that can easily transfer thermal energy. Convection is heat transfer through the movement of gases or liquids. When these fluids are heated and move away from the heat source, they carry the thermal energy with them. A great example of convection is a space heater; the air is heated and then blown out of the heater, sending thermal energy with it out into the room and heating it up. Radiation is the heat transfer through electromagnetic waves. We experience heat from radiation every day from the sun, which transfers heat to the earth through solar radiation. Radiation is the only type of heat transfer that can travel through empty space, because both conduction and convection require a medium to transfer the thermal energy through.

**Heat** ( $q$ ) is the transfer of thermal energy between two bodies at different temperatures. **Heat flow** (a redundant term, but one commonly used) increases the thermal energy of

[Link to Learning](#)

The following demonstration allows one to view the effects of heating and cooling a coiled bimetallic strip (opens in new window):

[Bimetallic strip gif.](#)

one body and decreases the thermal energy of the other. Suppose we initially have a high temperature (and high thermal energy) substance (H) and a low temperature (and low thermal energy) substance (L). The atoms and molecules in H have a higher average KE than those in L. If we place substance H in contact with substance L, the thermal energy will flow spontaneously from substance H to substance L. The temperature of substance H will decrease, as will the average KE of

its molecules; the temperature of substance L will increase, along with the average KE of its molecules. Heat flow will continue until the two substances are at the same temperature and therefore reach thermal equilibrium(Figure 4.12).

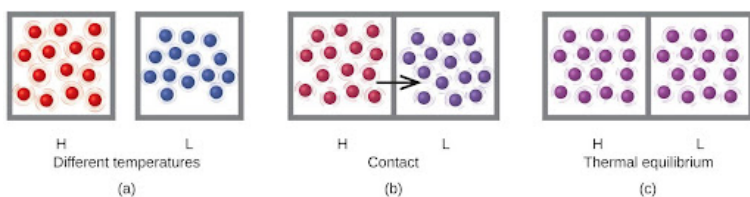


Fig. 4.12 (a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they contact each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two objects reach “thermal equilibrium” when both substances are at the same temperature and their molecules have the same average kinetic energy.

Matter undergoing **chemical reactions** and physical changes can release or absorb heat. A change that releases heat is called an **exothermic process**. For example, the combustion reaction that occurs when using an oxyacetylene torch is an exothermic process—this process also releases energy in the form of light as evidenced by the torch’s flame. A reaction or change that absorbs heat is an **endothermic process**. A cold pack used to treat

### Link to Learning

Click on this PhET simulation to explore energy forms and changes (opens in new window):

[PhET Energy](#)

### Forms and Changes.

Visit the Energy Systems tab to create combinations of energy sources, transformation methods, and outputs. Click on Energy Symbols to visualize the transfer of energy.

muscle strains provides an example of an endothermic process. When the substances in the cold pack (water and a salt like ammonium nitrate) are brought together, the resulting process absorbs heat, leading to the sensation of cold.

## Check Your Understanding



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<https://louis.pressbooks.pub/physicalscience1/?p=58#h5p-18>



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can view it online here:*

[https://louis.pressbooks.pub/  
physicalscience1/?p=58#h5p-19](https://louis.pressbooks.pub/physicalscience1/?p=58#h5p-19)

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5.

# WAVES, ELECTROMAGNETIC RADIATION, AND SOUND

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## Chapter Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

1. Recognize that Waves Carry Energy, but not Matter (CLO 1)
2. Define Mechanical Waves (CLO 1)
3. Distinguish between Transverse Waves and Compressional Waves (CLO 1)



4. Compare and Contrast Transverse and Compressional Waves (CLO 3)
5. Describe the Relationship between Frequency and Wavelength (CLO 3)
6. Explain How a Wave's Amplitude is Related to the Wave's Energy (CLO 3)
7. Calculate the Speed of a Wave (CLO 5)
8. Identify What Influences The Speed of Sound (CLO 3)
9. Discuss and Apply Various Properties of Sound (CLO 3)
10. Identify the Properties of Electromagnetic Radiation (CLO 3)



Fig 5.1 Waves in the ocean behave similarly to all other types of waves.

## Introduction To Waves

What do we mean when we say something is a **wave**? The most intuitive and easiest wave to imagine is the familiar water wave. More precisely, a wave is a disturbance that propagates, or moves from the place it was created. For water waves, the disturbance is on the surface of the water, perhaps created by a rock thrown into a pond or by a swimmer splashing the surface repeatedly. For **sound waves**, the disturbance is a change in air pressure, perhaps created by the oscillating cone inside a speaker. For earthquakes, there are several types of disturbances, including disturbance of Earth's surface and

pressure disturbances under the surface. Even radio waves are most easily understood using an analogy with water waves. Visualizing water waves is useful because there is more to it than just a mental image. Water waves exhibit characteristics common to all waves, such as **amplitude**, **period**, **frequency** and **energy**. All wave characteristics can be described by a small set of underlying principles.

### Misconception Alert

Many people think that water waves push water from one direction to another. In fact, the particles of water tend to stay in one location, save for moving up and down due to the energy in the wave. The energy moves forward through the water, but the water stays in one place. If you feel yourself pushed in an ocean, what you feel is the energy of the wave, not a rush of water.

## Wave Characteristics

The simplest waves repeat themselves for several cycles and

are associated with **simple harmonic motion**. Let us start by considering the simplified water wave in Figure 5.2. The wave is an up and down disturbance of the water surface. It causes a seagull to move up and down in simple harmonic motion as the wave **crests** and **troughs** (peaks and valleys) pass under the bird. The time for one complete up and down motion is the wave's period  $T$ .

The wave's frequency is  $f = 1/T$ . The wave itself moves to the right in the figure. This movement of the wave is actually the disturbance moving to the right, not the water itself (or the bird would move to the right). We define **wave velocity**  $v_w$  to be the speed at which the disturbance moves. Wave velocity is sometimes also called the *propagation velocity or propagation speed*, because the disturbance propagates from one location to another.

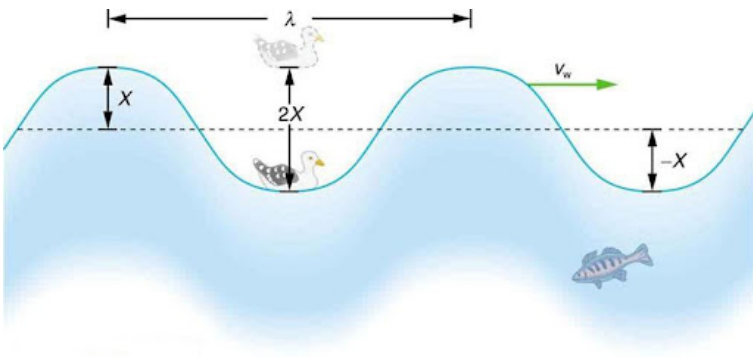


Fig. 5.2 An idealized ocean wave passes under a sea gull that bobs up and down in simple harmonic motion. The wave has a wavelength  $\lambda$ , which is the distance between adjacent identical parts of the wave.

### Take-Home Experiment: Waves in a Bowl

Fill a large bowl or basin with water and wait for the water to settle so there are no ripples. Gently drop a cork into the middle of the bowl. Estimate the wavelength and period of oscillation of the water wave that propagates away from the cork. Remove the cork from the bowl and wait for the water to settle again. Gently drop the cork at a height that is different from the first drop. Does the wavelength

depend upon how high above the water the cork is dropped?

### Example 5.1

#### **CALCULATE THE VELOCITY OF WAVE PROPAGATION: GULL IN THE OCEAN**

Calculate the wave velocity of the ocean wave in Figure 5.2 if the distance between wave crests is 10.0 m and the time for a seagull to bob up and down is 5.00 s.

#### **Strategy**

We are asked to find  $v_w$ . The given information tells us that  $\lambda = 10.0\text{m}$  and  $T =$

5.00s. Therefore, we can use  $v_w = \lambda/T$  to find the wave velocity.

### **Solution**

1. Enter the known values into  $v_w = \lambda/T$   
 $v_w = 10.0 \text{ m}/5.00 \text{ s} = 2.00 \text{ m/s}$

### **Discussion**

This slow speed seems reasonable for an ocean wave. Note that the wave moves to the right in the figure at this speed, not the varying speed at which the seagull moves up and down.

# Electromagnetic Spectrum

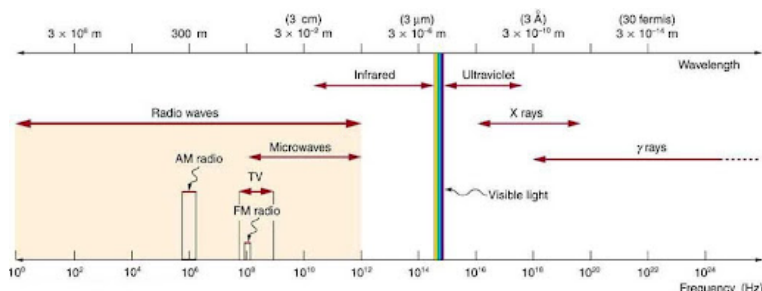


Fig 5.3 Various types of electromagnetic waves are categorized according to their wavelengths and frequencies.

The **electromagnetic spectrum**, showing the major categories of **electromagnetic waves**. The range of **frequencies** and **wavelengths** is remarkable. The dividing line between some categories is distinct, whereas other categories overlap.

Figure 5.3 shows how the various types of electromagnetic waves are categorized according to their wavelengths and frequencies—that is, it shows the electromagnetic spectrum. Many of the characteristics of the various types of electromagnetic waves are related to their frequencies and wavelengths, as we shall see. We can get a good understanding of electromagnetic waves by considering how they are produced. Whenever a current varies, associated **electric fields**



and **magnetic fields** vary, moving out from the source like waves.

## Electromagnetic Spectrum: Rules of Thumb

Three rules that apply to electromagnetic waves in general are as follows:

- High-frequency electromagnetic waves are more energetic and are more able to penetrate than low-frequency waves.
- High-frequency electromagnetic waves can carry more information per unit time than low-frequency waves.
- The shorter the wavelength of any electromagnetic wave probing a material, the smaller the detail it is possible to resolve.

Note that there are exceptions to these rules of thumb.

## Transmission, Reflection, and Absorption

What happens when an electromagnetic wave impinges on a material? If the material is **transparent** to the particular frequency, then the wave can largely be transmitted. If the

material is **opaque** to the frequency, then the wave can be totally reflected. The wave can also be absorbed by the material, indicating that there is some interaction between the wave and the material, such as the thermal agitation of molecules.

Of course it is possible to have partial **transmission**, **reflection**, and **absorption**. We normally associate these properties with visible light, but they do apply to all electromagnetic waves. What is not obvious is that something that is transparent to light may be opaque at other frequencies. For example, ordinary glass is transparent to visible light but largely opaque to ultraviolet radiation. Human skin is opaque to visible light—we cannot see through people—but transparent to X-rays.

## Radio and TV Waves

The broad category of radio waves is defined to contain any electromagnetic wave produced by currents in wires and circuits. Its name derives from their most common use as a carrier of audio information (i.e., radio). The name is applied to electromagnetic waves of similar frequencies regardless of source. Radio waves from outer space, for example, do not come from alien radio stations. They are created by many astronomical phenomena, and their study has revealed much about nature on the largest scales.

There are many uses for radio waves, and so the category

is divided into many subcategories, including microwaves and those electromagnetic waves used for AM and FM radio, cellular telephones, and TV.

The lowest commonly encountered radio frequencies are produced by high-voltage AC power transmission lines at frequencies of 50 or 60 Hz. (See Figure 5.4.) These extremely long wavelength electromagnetic waves (about 6000 km!) are one means of energy loss in long-distance power transmission.



Fig. 5.4 This high-voltage traction power line running to Eutingen Railway Substation in Germany radiates electromagnetic waves with very long wavelengths.

There is an ongoing controversy regarding potential health hazards associated with exposure to these electromagnetic

fields. Some people suspect that living near such transmission lines may cause a variety of illnesses, including cancer. But demographic data are either inconclusive or simply do not support the hazard theory. Recent reports that have looked at many European and American epidemiological studies have found no increase in risk for cancer due to exposure to electromagnetic fields.

There are three ways in which light can travel from a source to another location. (See Figure 5.5.) It can come directly from the source through empty space, such as from the Sun to Earth. Or light can travel through various media, such as air and glass, to the person. Light can also arrive after being reflected, such as by a mirror. In all of these cases, light is modeled as traveling in straight lines called rays. Light may change direction when it encounters objects (such as a mirror) or in passing from one material to another (such as in passing from air to glass), but it then continues in a straight line or as a **ray**. It is acceptable to visualize light rays as laser rays (or even science fiction depictions of ray guns).

## Ray

The word “ray” comes from mathematics and here means a straight line that originates at some point.

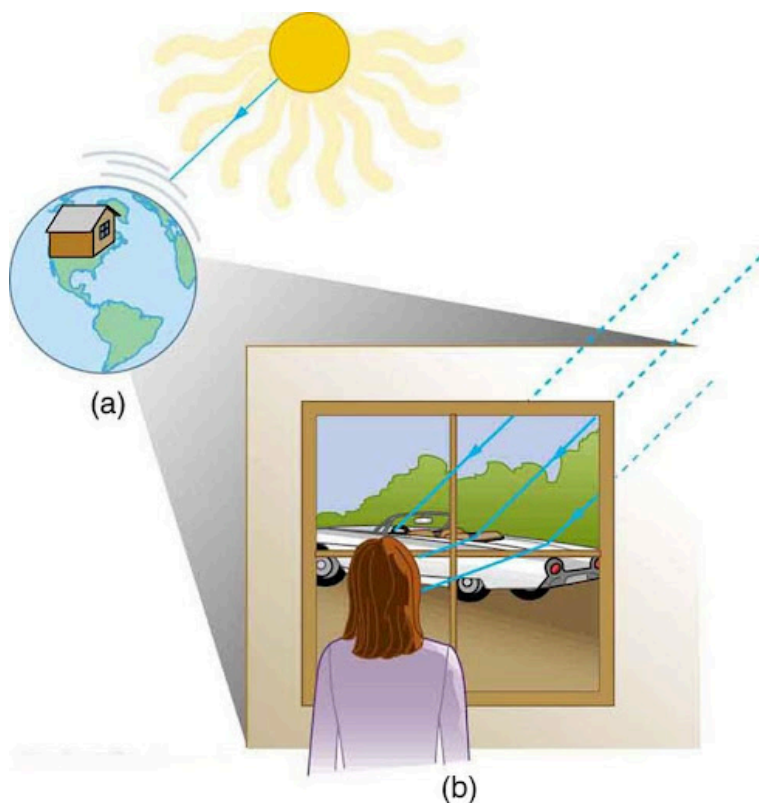


Fig. 5.5 Three methods for light to travel from a source to another location. (a) Light reaches the upper atmosphere of Earth traveling through empty space directly from the source. (b) Light can reach a person in one of two ways. It can travel through media like air and glass. It can also reflect from an object like a mirror. In the situations shown here, light interacts with objects large enough that it travels in straight lines, like a ray.

Experiments, as well as our own experiences, show that when light interacts with objects several times as large as its wavelength, it travels in straight lines and acts like a ray. Its

wave characteristics are not pronounced in such situations. Since the wavelength of light is less than a micron (a thousandth of a millimeter), it acts like a ray in the many common situations in which it encounters objects larger than a micron. For example, when light encounters anything we can observe with unaided eyes, such as a mirror, it acts like a ray, with only subtle wave characteristics. We will concentrate on the ray characteristics.

Since light moves in straight lines, changing directions when it interacts with materials, it is described by geometry and simple trigonometry. This part of optics, where the ray aspect of light dominates, is therefore called geometric optics. There are two laws that govern how light changes direction when it interacts with matter. These are the **law of reflection**, for situations in which light bounces off matter, and the law of **refraction**, for situations in which light passes through matter.

## Geometric Optics

The part of optics dealing with the ray aspect of light is called geometric optics.

Whenever we look into a mirror, or squint at sunlight glinting from a lake, we are seeing a reflection. When you look at this page, too, you are seeing light reflected from it. Large telescopes use reflection to form an image of stars and other astronomical objects.

The law of reflection is illustrated in Figure 5.6, which also shows how the angles are measured relative to the perpendicular to the surface at the point where the light ray strikes. We expect to see reflections from smooth surfaces, but Figure 5.7 illustrates how a rough surface reflects light. Since the light strikes different parts of the surface at different angles, it is reflected in many different directions, or diffused. Diffused light is what allows us to see a sheet of paper from any angle, as illustrated in Figure 5.8. Many objects, such as people, clothing, leaves, and walls, have rough surfaces and can be seen from all sides. A mirror, on the other hand, has a smooth surface (compared with the wavelength of light) and reflects light at specific angles, as illustrated in Figure 5.9. When the moon reflects from a lake, as shown in Figure 5.10, a combination of these effects takes place.

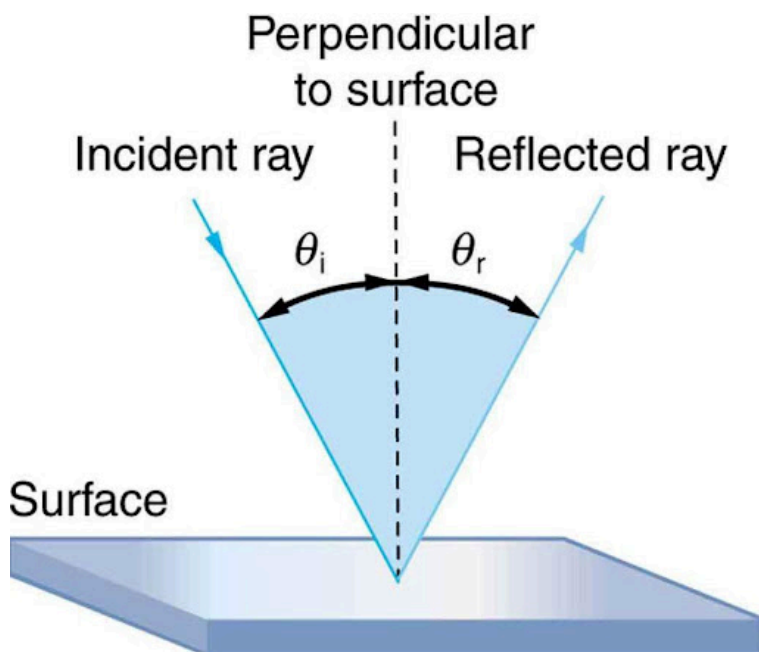


Fig. 5.6 The law of reflection states that the angle of reflection equals the angle of incidence— $\theta_r = \theta_i$ . The angles are measured relative to the perpendicular to the surface at the point where the ray strikes the surface.



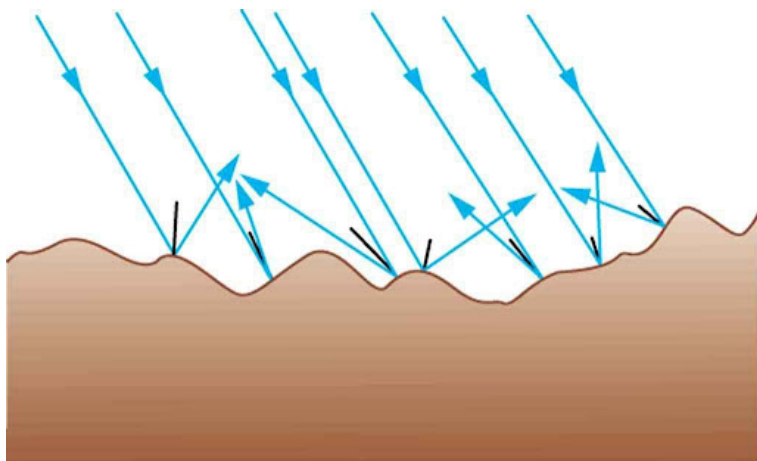


fig 5.7 Light is diffused when it reflects from a rough surface. Here many parallel rays are incident, but they are reflected at many different angles since the surface is rough.

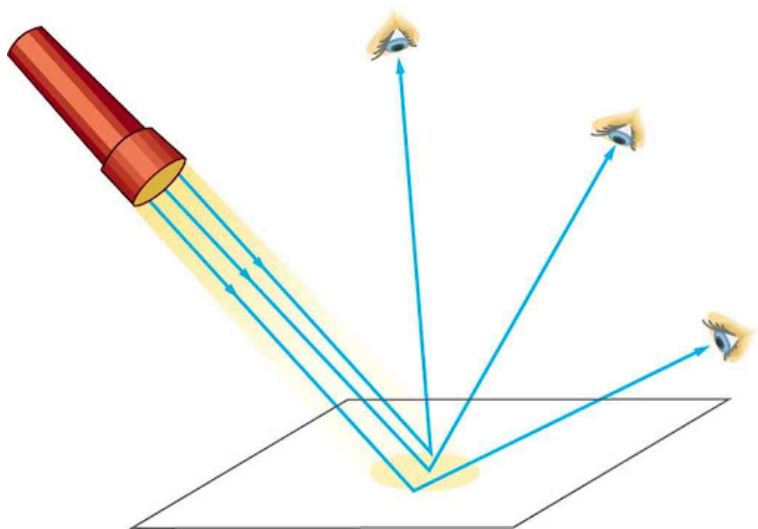


Fig. 5.8 When a sheet of paper is illuminated with many parallel incident rays, it can be seen at many different angles, because its surface is rough and diffuses the light.

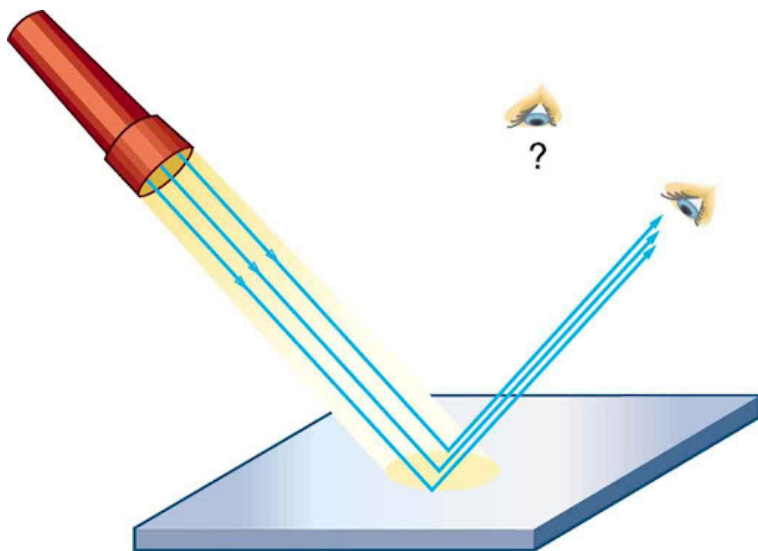


Fig 5.9 A mirror illuminated by many parallel rays reflects them in only one direction, since its surface is very smooth. Only the observer at a particular angle will see the reflected light.



Fig 5.10 Moonlight is spread out when it is reflected by the lake, since the surface is shiny but uneven.

The law of reflection is very simple: The **angle of reflection** ( $\theta_r$ ) equals the **angle of incidence** ( $\theta_i$ ).

## Refraction

The changing of a light ray's direction (loosely called bending) when it passes through variations in matter is called refraction. Refraction is responsible for a tremendous range of optical phenomena, from the action of lenses to voice transmission through optical fibers.

It is easy to notice some odd things when looking into a fish

tank. For example, you may see the same fish appearing to be in two different places. (See Figure 5.11.) This is because light coming from the fish to us changes direction when it leaves the tank, and in this case, it can travel two different paths to get to our eyes.

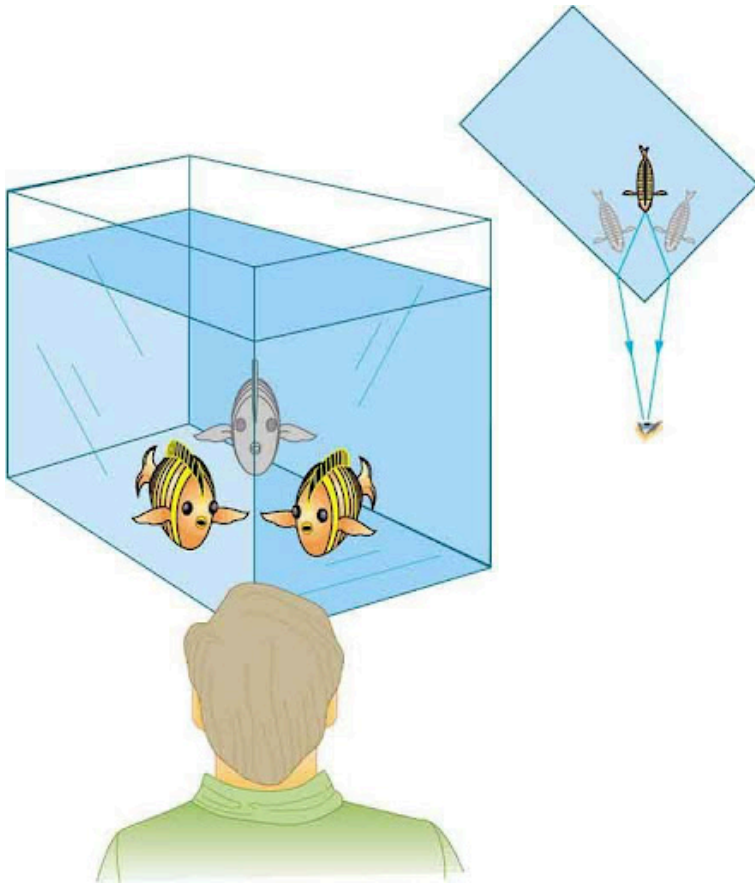


Fig 5.11 Looking at the fish tank as shown, we can see the same fish in two different locations, because light changes directions when it passes from water to air. In this case, the light can reach the observer by two different paths, and so the fish seems to be in two different places. This bending of light is called refraction and is responsible for many optical phenomena.

## Speed of Light

The speed of light ( $c$ ) not only affects refraction, it is one of the central concepts of Einstein's theory of relativity. As the accuracy of the measurements of the speed of light were improved,  $c$  was found not to depend on the velocity of the source or the observer. However, the speed of light does vary in a precise manner with the material it traverses. Early attempts to measure the speed of light, such as those made by Galileo, determined that light moved extremely fast, perhaps instantaneously. The first real evidence that light traveled at a finite speed came from the Danish astronomer Ole Roemer in the late 17th century.

The speed of light is now known to great precision. In fact, the speed of light in a vacuum  $c$  is so important that it is accepted as one of the basic physical quantities and has the fixed value

**Formula does not parse**

$$c = 2.99792458 \times 10^8 \text{ m/s} \approx 3.00 \times 10^8 \text{ m/s},$$

where the approximate value of

$3.00 \times 10^8 \text{ m/s}$  is used whenever three-digit accuracy is sufficient.

## Take-Home Experiment: Bending Light

Explore bending of light between two media with different indices of refraction. See how changing from air to water to glass changes the bending angle. Play with prisms of different shapes and make rainbows.

### Link to Learning

Use this interactive simulation by PhET to explore these two laws of light (Website opens



## Lenses

Lenses are found in a huge array of optical instruments, ranging from a simple magnifying glass to the eye to a camera's zoom lens.

In this section, we will use the

**law of refraction** to explore the

properties of lenses and how they form images. The word *lens* derives from the Latin word for a lentil bean, the shape of which is similar to the convex lens in Figure 5.12. The **convex lens** shown has been shaped so that all light rays that enter it parallel to its axis cross one another at a single point on the opposite side of the lens. (The axis is defined to be a line normal, or perpendicular, to the lens at its center, as shown in Figure 5.12.) Such a lens is called a converging (or convex) lens for the converging effect it has on light rays. An expanded view of the path of one ray through the lens is shown, to illustrate how the ray changes direction both as it enters and as it leaves the lens. Since the **index of refraction** of the lens is greater than that of air, the ray moves towards the perpendicular as it enters and away from the perpendicular as it leaves. (This is in accordance with the law of refraction.)

Due to the lens's shape, light is thus bent toward the axis at both surfaces. The point at which the rays cross is defined to be the **focal point**  $F$  of the lens. The distance from the center of the lens to its focal point is defined to be the **focal length**

in a new  
window):  
[Bending Light.](#)

(F) of the lens. Figure 5.13 shows how a **converging lens**, such as that in a magnifying glass, can converge the nearly parallel light rays from the sun to a small spot.

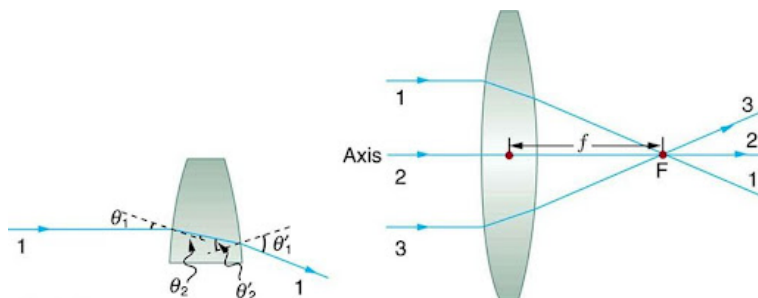


Fig. 5.12 Rays of light entering a converging lens parallel to its axis converge at its focal point F. (Ray 2 lies on the axis of the lens.) The distance from the center of the lens to the focal point is the lens's focal length  $f$ . An expanded view of the path taken by ray 1 shows the perpendiculars and the angles of incidence and refraction at both surfaces.



Fig 5.13 A small light source, like a light bulb filament, placed at the focal point of a convex lens, results in parallel rays of light emerging from the other side. The paths are exactly the reverse of those shown in the preceding image. This technique is used in lighthouses and sometimes in traffic lights to produce a directional beam of light from a source that emits light in all directions.

The greater effect a lens has on light rays, the more powerful

it is said to be. For example, a powerful converging lens will focus parallel light rays closer to itself and will have a smaller focal length than a weak lens. The light will also focus into a smaller and more intense spot for a more powerful lens.

A lens that causes the light rays to bend away from its axis is called a **diverging lens**. The concave lens is a **diverging lens**, because it causes the light rays to bend away (diverge) from its axis.

As noted in the initial discussion of the **law of refraction**, the paths of light rays are exactly reversible. This means that the direction of the arrows could be reversed for all of the rays in Figure 5.12. For example, if a point light source is placed at the focal point of a **convex lens**, as shown in Figure 5.14, parallel light rays emerge from the other side.

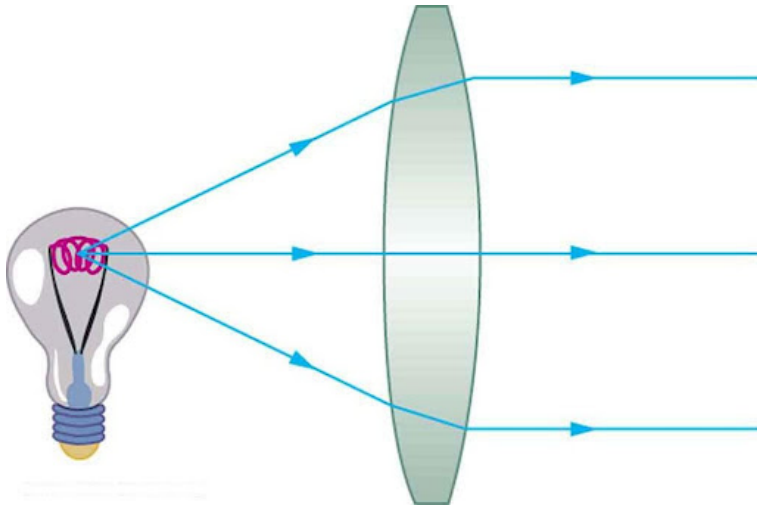


Fig. 5.14 A small light source, like a light bulb filament, placed at the focal point of a convex lens, results in parallel rays of light emerging from the other side. The paths are exactly the reverse of those shown in Figure 5.12. This technique is used in lighthouses and sometimes in traffic lights to produce a directional beam of light from a source that emits light in all directions.

## Ray Tracing and Thin Lenses

Ray tracing is the technique of determining or following (tracing) the paths that light rays take. For rays passing through matter, the law of refraction is used to trace the paths. Here we use ray tracing to help us understand the action of lenses in situations ranging from forming images on film to magnifying small print to correcting nearsightedness. While ray tracing for complicated lenses, such as those found in sophisticated

cameras, may require computer techniques, there is a set of simple rules for tracing rays through thin lenses. A thin lens is defined to be one whose thickness allows rays to refract, as illustrated in Figure 5.12, but does not allow properties such as dispersion and aberrations. An ideal thin lens has two refracting surfaces but the lens is thin enough to assume that light rays bend only once. A thin symmetrical lens has two focal points, one on either side and both at the same distance from the lens. (See Figure 5.15.) Another important characteristic of a **thin lens** is that light rays through its center are deflected by a negligible amount, as seen in Figure 5.16.

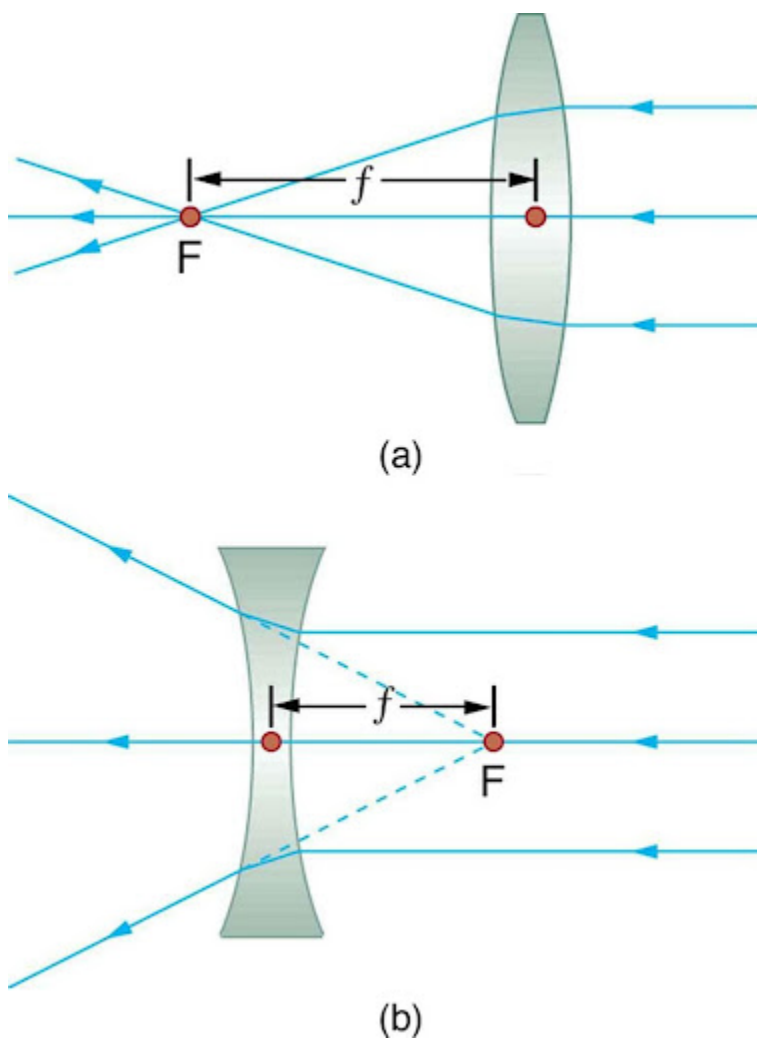


Fig. 5.15 Thin lenses have the same focal length on either side. (a) Parallel light rays entering a converging lens from the right cross at its focal point on the left. (b) Parallel light rays entering a diverging lens from the right seem to come from the focal point on the right.

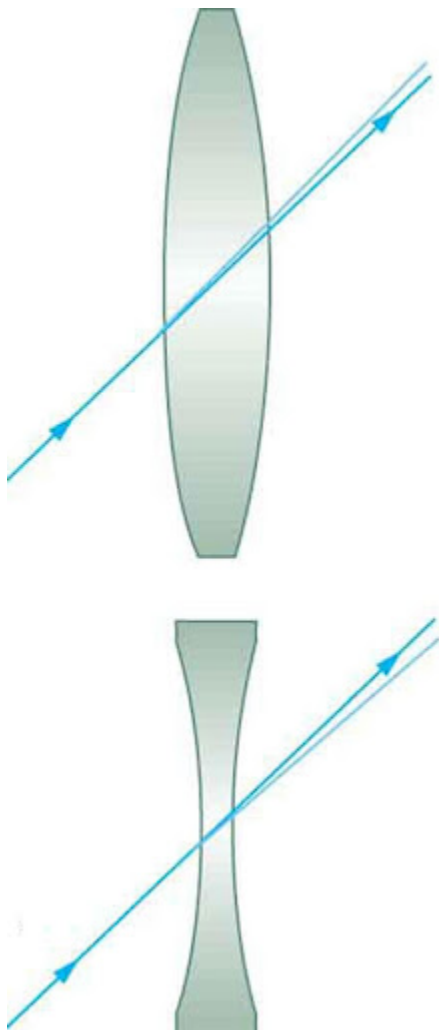


Fig. 5.16 The light ray through the center of a thin lens is deflected by a negligible amount and is assumed to emerge parallel to its original path (shown as a shaded line).



## Rules for Ray Tracing

A ray entering a converging lens parallel to its axis passes through the focal point  $F$  of the lens on the other side.

A ray entering a diverging lens parallel to its axis seems to come from the focal point  $F$ .

A ray passing through the center of either a converging or a diverging lens does not change direction.

A ray entering a converging lens through its focal point exits parallel to its axis.

A ray that enters a diverging lens by heading toward the focal point on the opposite side exits parallel to the axis.

Using paper, pencil, and a straight edge, ray tracing can accurately describe the operation of a lens. The rules for ray tracing for thin lenses are based on the preceding illustrations.

## Image Formation by Thin Lenses

In some circumstances, a lens forms an obvious image, such as when a movie projector casts an image onto a screen. In other cases, the image is less obvious. Where, for example, is the image formed by eyeglasses? We will use ray tracing for thin lenses to illustrate how they form images, and we will develop equations to describe the image formation quantitatively.

Consider an object some distance away from a converging lens, as shown in Figure 5.17. To find the location and size of the image formed, we trace the paths of selected light rays

originating from one point on the object, in this case the top of the person's head. The figure shows three rays from the top of the object that can be traced using the ray tracing rules given above. (Rays leave this point going in many directions, but we concentrate on only a few with paths that are easy to trace.) The first ray is one that enters the lens parallel to its axis and passes through the focal point on the other side (rule 1). The second ray passes through the center of the lens without changing direction (rule 3). The third ray passes through the nearer focal point on its way into the lens and leaves the lens parallel to its axis (rule 4). The three rays cross at the same point on the other side of the lens. The image of the top of the person's head is located at this point. All rays that come from the same point on the top of the person's head are refracted in such a way as to cross at the point shown. Rays from another point on the object, such as her belt buckle, will also cross at another common point, forming a complete image, as shown. Although three rays are traced in Figure 5.17, only two are necessary to locate the image. It is best to trace rays for which there are simple ray tracing rules. Before applying ray tracing to other situations, let us consider the example shown in Figure 5.17 in more detail.

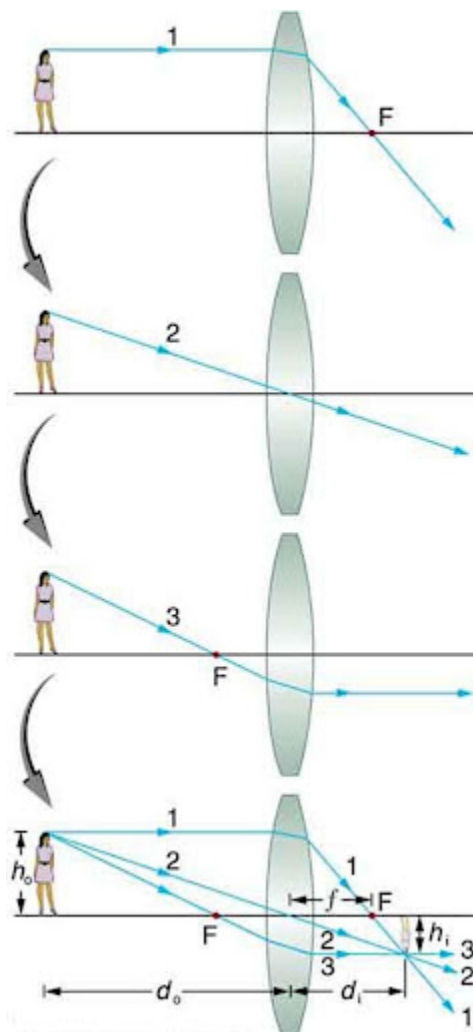


Fig. 5.17 Ray tracing is used to locate the image formed by a lens. Rays originating from the same point on the object are traced—the three chosen rays each follow one of the rules for ray tracing, so that their paths are easy to determine. The

image is located at the point where the rays cross. In this case, a real image—one that can be projected on a screen—is formed.

The image formed in Figure 5.17 is a real image, meaning that it can be projected. That is, light rays from one point on the object actually cross at the location of the image and can be projected onto a screen, a piece of film, or the retina of an eye, for example. Figure 5.18 shows how such an image would be projected onto film by a camera lens. This figure also shows how a real image is projected onto the retina by the lens of an eye. Note that the image is there whether it is projected onto a screen or not.

### Take Home Experiment: A Visit to the Optician

Look through your eyeglasses (or those of a friend) backward and forward and comment on whether they act like thin lenses.

## Real Image

The image in which light rays from one point on the object actually cross at the location of the image and can be projected onto a screen, a piece of film, or the retina of an eye is called a **real image**.

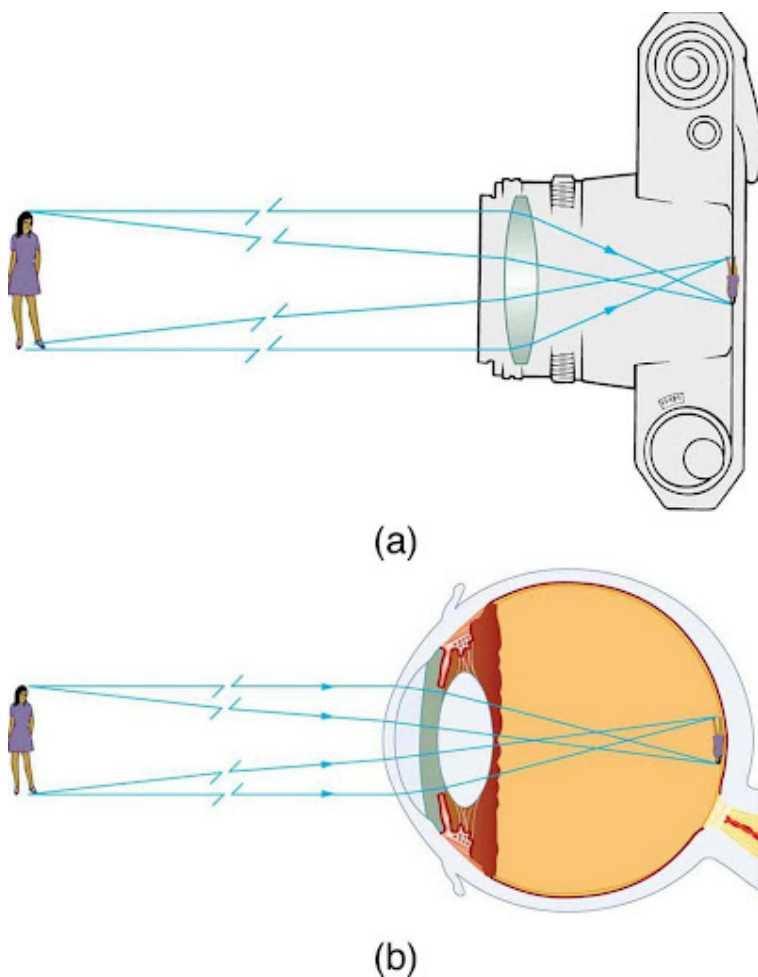


Fig. 5.18 Real images can be projected. (a) A real image of the person is projected onto film. (b) The converging nature of the multiple surfaces that make up the eye result in the projection of a real image on the retina.

Real images are formed by converging lenses whenever an

object is farther from the lens than its focal length. This is true for movie projectors, cameras, and the eye.

A different type of image is formed when an object, such as a person's face, is held close to a convex lens. The image is upright and larger than the object, as seen in Figure 5.19 (b), and so the lens is called a magnifier. If you slowly pull the magnifier away from the face, you will see that the magnification steadily increases until the image begins to blur. Pulling the magnifier even farther away produces an inverted image as seen in Figure 5.19 (a). The distance at which the image blurs, and beyond which it inverts, is the focal length of the lens. To use a convex lens as a magnifier, the object must be closer to the converging lens than its focal length.



(a)



(b)

Fig. 5.19 (a) When a converging lens is held farther away from the face than the lens's focal length, an inverted image is formed. Note that the image is in focus but the face is not, because the image is much closer to the camera taking this photograph than the face. (b) A magnified image of a face is produced by placing it closer to the converging lens than its focal



length.

Figure 5.20 uses ray tracing to show how an image is formed when an object is held closer to a converging lens than its focal length. Rays coming from a common point on the object continue to diverge after passing through the lens, but all appear to originate from a point at the location of the image. The image is on the same side of the lens as the object and is farther away from the lens than the object. This image, like all case 2 images, cannot be projected and, hence, is called a virtual image. Light rays only appear to originate at a virtual image; they do not actually pass through that location in space. A screen placed at the location of a virtual image will receive only diffuse light from the object, not focused rays from the lens. Additionally, a screen placed on the opposite side of the lens will receive rays that are still diverging, and so no image will be projected on it. We can see the magnified image with our eyes, because the lens of the eye converges the rays into a real image projected on our retina. Finally, we note that a virtual image is upright and larger than the object, meaning that the magnification is positive and greater than 1.

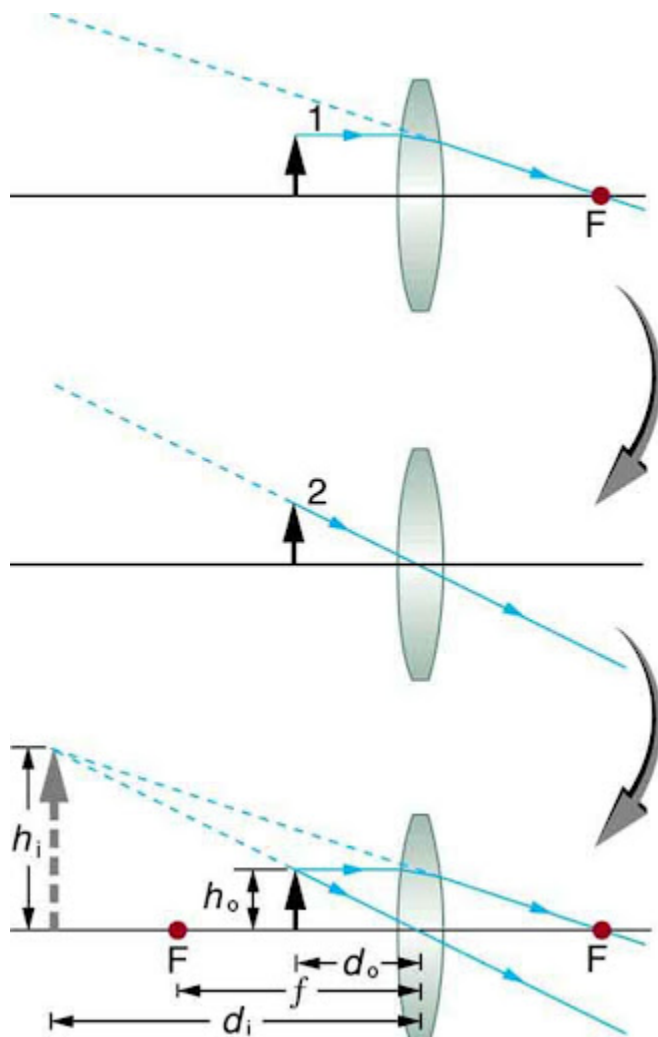


Fig 5.20 Ray tracing predicts the image location and size for an object held closer to a converging lens than its focal length. Ray 1 enters parallel to the axis and exits through the focal point on the opposite side, while ray 2 passes through the center of the lens without changing path. The two rays continue to diverge on the other side of the lens, but both

appear to come from a common point, locating the upright, magnified, virtual image. This is a case 2 image.

## Virtual Image

An image that is on the same side of the lens as the object and cannot be projected on a screen is called a **virtual image**.

A third type of image is formed by a diverging or concave lens. Try looking through eyeglasses meant to correct nearsightedness. (See Figure 5.21.) You will see an image that is upright but smaller than the object. This means that the magnification is positive but less than 1.



Fig. 5.21 A car viewed through a concave or diverging lens looks upright. This is a case 3 image.

The ray diagram in Figure 5.22 shows that the image is on the same side of the lens as the object and, hence, cannot be projected—it is a virtual image. Note that the image is closer to the lens than the object. This is a case 3 image, formed for any object by a negative focal length or diverging lens.

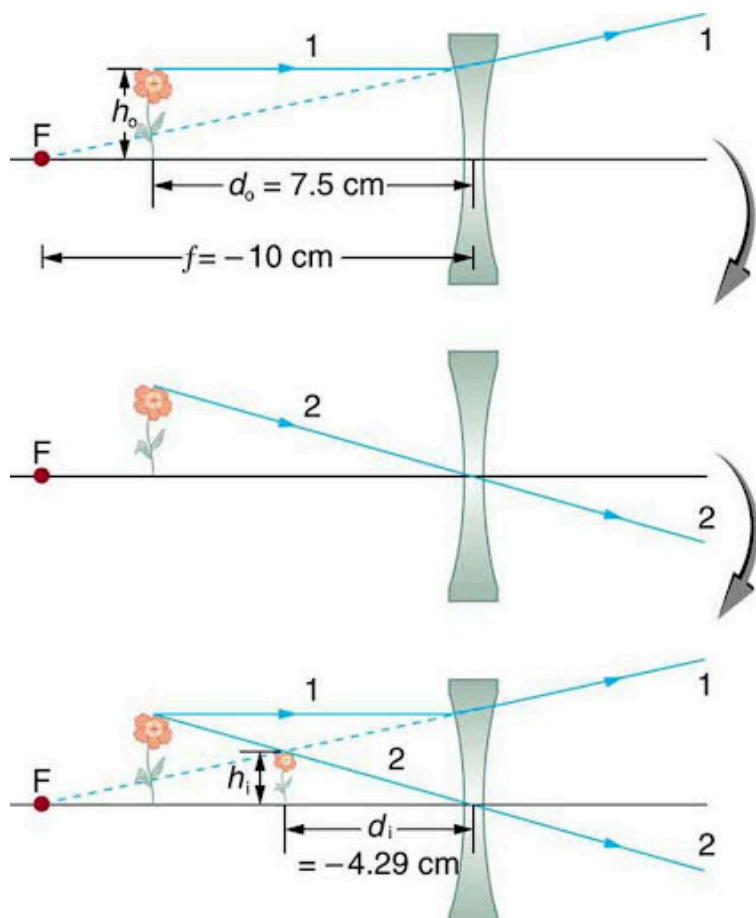


Fig. 5.22 Ray tracing predicts the image location and size for a concave or diverging lens. Ray 1 enters parallel to the axis and is bent so that it appears to originate from the focal point. Ray 2 passes through the center of the lens without changing path. The two rays appear to come from a common point, locating the upright image.

## Sound



Fig. 5.23 This glass has been shattered by a high-intensity sound wave of the same frequency as the resonant frequency of the glass. While the sound is not visible, the effects of the sound prove its existence.

Sound can be used as a familiar illustration of waves. Because hearing is one of our most important senses, it is interesting to see how the physical properties of sound correspond to our perceptions of it. Hearing is the perception of sound, just as vision is the perception of visible light. But sound has important applications beyond hearing. Ultrasound, for

example, is not heard but can be employed to form medical images and is also used in treatment.

The physical phenomenon of sound is defined to be a disturbance of matter that is transmitted from its source outward. Sound is a wave. On the atomic scale, it is a disturbance of atoms that is far more ordered than their thermal motions. In many instances, sound is a periodic wave, and the atoms undergo simple harmonic motion.

A vibrating string produces a sound wave as illustrated in Figure 5.24, Figure 5.25, and Figure 5.26. As the string oscillates back and forth, it transfers energy to the air, mostly as thermal energy created by turbulence. But a small part of the string's energy goes into compressing and expanding the surrounding air, creating slightly higher and lower local pressures. These compressions (high pressure regions) and rarefactions (low pressure regions) move out as longitudinal pressure waves having the same frequency as the string—they are the disturbance that is a **sound wave**. (Sound waves in air and most fluids are longitudinal, because fluids have almost no shear strength. In solids, sound waves can be both transverse and longitudinal.) Figure 5.26 shows a graph of gauge pressure versus distance from the vibrating string.

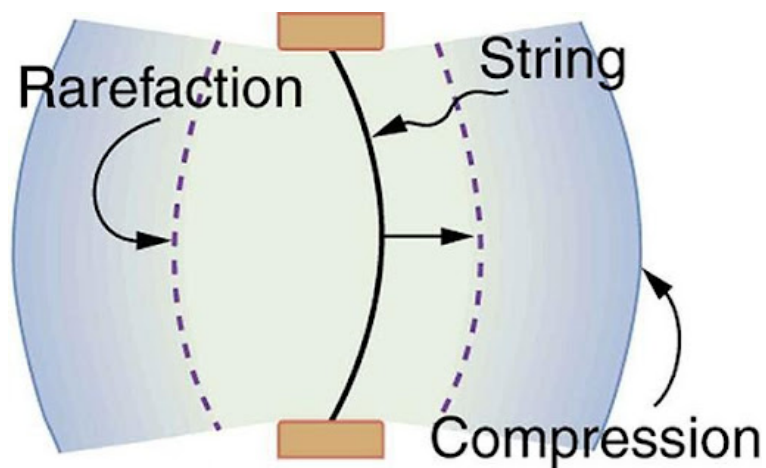


Fig. 5.24 A vibrating string moving to the right compresses the air in front of it and expands the air behind it.



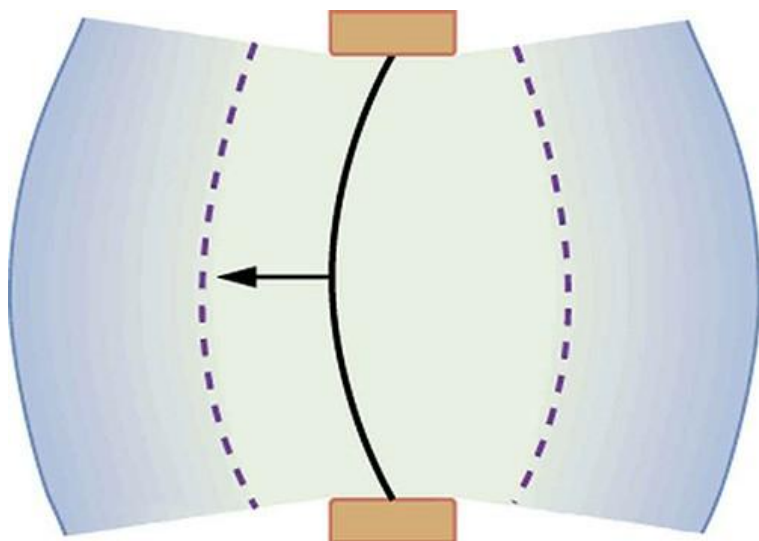


Fig 5.25 As the string moves to the left, it creates another compression and rarefaction as the ones on the right move away from the string.

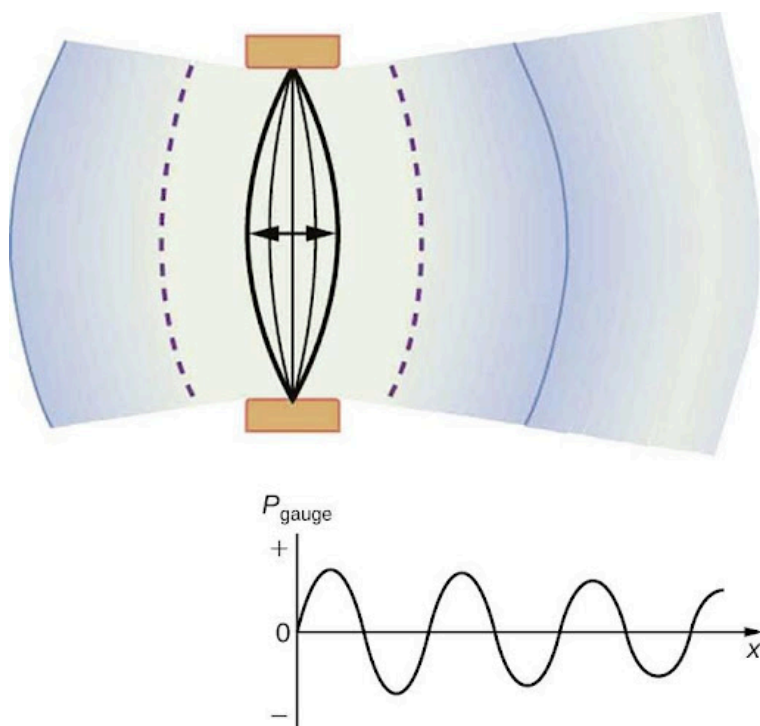


Fig 5.26 After many vibrations, there are a series of compressions and rarefactions moving out from the string as a sound wave. The graph shows gauge pressure versus distance from the source. Pressures vary only slightly from atmospheric for ordinary sounds. Note that sound waves consist of these changes in density or pressure of a medium and thus are *\*not\** electromagnetic radiation.

## Check Your Understanding



*An interactive H5P element has been excluded from this version of the text. You can view it online here:*

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*An interactive H5P element has been excluded from this version of the text. You can view it online here:*

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6.

# THE ATOM AND ELEMENTS

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## Chapter Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

1. Describe the Electron Configuration of an Atom by using the Bohr Model (CLO 2)
2. Compute the Atomic Mass of an Atom (CLO 4)
3. Identify Isotopes of Common Elements (CLO 4)
4. Use the Periodic Table to Obtain Information (CLO 4)



Fig. 6.1 Atoms are so tiny that we cannot see them with our eyes, but we know they interconnect forming chains, rings and bonds. When looking at the seed head of this false dandelion through a macro lens, one might be able to connect the dots and imagine the atoms



with their chains, rings and bonds. Atoms truly are everywhere.

## Atoms – Building Blocks of Matter

The earliest recorded discussion of the basic structure of matter comes from ancient Greek philosophers, the scientists of their day. In the fifth century BC, Leucippus and Democritus argued that all matter was composed of small, finite particles that they called *atomos*, a term derived from the Greek word for “indivisible.” They thought of **atoms** as moving particles that differed in shape and size, and which could join together. Later, Aristotle and others came to the conclusion that matter consisted of various combinations of the four “elements”—fire, earth, air, and water—and could be infinitely divided. Interestingly, these philosophers thought about atoms and “elements” as philosophical concepts, but apparently never considered performing experiments to test their ideas.

## Dalton's Atomic Theory

The Aristotelian view of the composition of matter held sway for over two thousand years, until English schoolteacher John Dalton helped to revolutionize chemistry with his hypothesis

that the behavior of **matter** could be explained using an atomic theory. First published in 1807, many of Dalton's hypotheses about the microscopic features of matter are still valid in modern atomic theory. Here are the postulates of Dalton's atomic theory.

1. Matter is composed of exceedingly small particles called atoms. An atom is the smallest unit of an element that can participate in a chemical change.
2. An element consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element. A macroscopic sample of an element contains an incredibly large number of atoms, all of which have identical chemical properties.
3. Atoms of one element differ in properties from atoms of all other elements.
4. A compound consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the numbers of atoms of each of its elements are always present in the same ratio.
5. Atoms are neither created nor destroyed during a chemical change, but are instead rearranged to yield substances that are different from those present before the change.

Dalton's atomic theory provides a microscopic explanation of

the many macroscopic properties of matter that you've learned about. For example, if an element such as copper consists of only one kind of **atom**, then it cannot be broken down into simpler substances, that is, into substances composed of fewer types of atoms. And if atoms are neither created nor destroyed during a chemical change, then the total mass of matter present when matter changes from one type to another will remain constant (the **law of conservation of matter**).

## Atomic Structure

The development of modern atomic theory revealed much about the inner structure of atoms. An atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The **nucleus** contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of  $10^{-10}$  m, whereas the diameter of the nucleus is roughly  $10^{-15}$  m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (Figure 6.2).

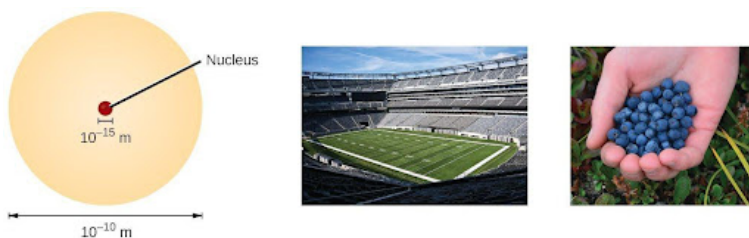


Fig. 6.2 If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry.

Atoms—and the **protons**, **neutrons**, and **electrons** that compose them—are extremely small. For example, a carbon atom weighs less than  $2 \times 10^{-23}$  g, and an electron has a charge of less than  $2 \times 10^{-19}$  C (coulomb). When describing the properties of tiny objects such as atoms we use appropriately small units of measure, such as the **atomic mass unit** (amu) and the fundamental unit of charge (e). The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant **isotope** of carbon, atoms of which are assigned masses of exactly 12 amu. (This isotope is known as “carbon-12” as will be discussed later in this module.) Thus, one amu is exactly a 12th of the mass of one carbon-12 atom:  $1 \text{ amu} = 1.6605 \times 10^{-24} \text{ g}$ . The **fundamental unit of charge** (also called the elementary charge) equals the magnitude of the charge of an electron (e) with  $e = 1.602 \times 10^{-19} \text{ C}$ .

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087

amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of 1<sup>−</sup> and is a much lighter particle with a mass of about 0.00055 amu (it would take about 1800 electrons to equal the mass of one proton). The properties of these fundamental particles are summarized in Table 2.1. (An observant student might notice that the sum of an atom’s subatomic particles does not equal the atom’s actual mass: The total mass of six protons, six neutrons, and six electrons is 12.0993 amu, slightly larger than 12.00 amu. This “missing” mass is known as the **mass defect**.)

Properties of Subatomic Particles					
Name	Location	Charge (C)	Unit Charge	Mass (amu)	
electron	outside nucleus	$-1.602 \times 10^{-19}$	1 <sup>−</sup>	0.00055	0.00
proton	nucleus	$1.602 \times 10^{-19}$	1 <sup>+</sup>	1.00727	1.67
neutron	nucleus	0	0	1.00866	1.67

**Table 6.1**

The number of protons in the nucleus of an atom is its **atomic number** (Z). This is the defining trait of an element: Its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges,

so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its **mass number** ( $A$ ). The number of neutrons is therefore the difference between the mass number and the atomic number:  $A - Z = \text{number of neutrons}$ .

Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an **ion**. The charge of an atom is defined as follows:

**Atomic charge = number of protons – number of electrons**

As will be discussed in more detail, atoms (and molecules) typically acquire charge by gaining or losing electrons. A cation (a positive ion) forms when a neutral atom loses one or more electrons from its valence shell, and an anion (a negative ion) forms when a neutral atom gains one or more electrons in its valence shell. For example, a neutral sodium atom ( $Z = 11$ ) has 11 electrons. If this atom loses one electron, it will become a cation with a  $1+$  charge ( $11 - 10 = 1+$ ). A neutral oxygen atom ( $Z = 8$ ) has eight electrons, and if it gains two electrons it will become an anion with a  $2-$  charge ( $8 - 10 = 2-$ ).

## Examples 6.1

## Composition of an Atom

Iodine is an essential trace element in our diet; it is needed to produce thyroid hormone. Insufficient iodine in the diet can lead to the development of a goiter, an enlargement of the thyroid gland (Figure 6.3).

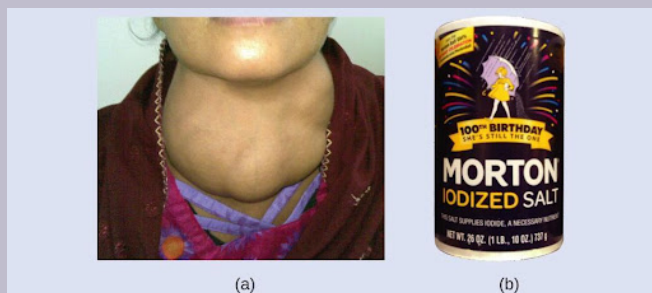


Fig. 6.3 (a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high.

The addition of small amounts of iodine to table salt (iodized salt) has essentially eliminated this health concern in the United States, but as much as 40% of the world's population is still at risk of iodine deficiency. The iodine atoms are added as anions, and each has a  $1^-$  charge, a mass number of 127, and the atomic number 53. Determine the numbers of protons, neutrons, and electrons in one of these iodine anions.

### **Solution**

The atomic number of iodine (53) tells us that a neutral iodine atom contains 53 protons in its nucleus and 53 electrons outside its nucleus. Because the sum of the numbers of protons and neutrons equals the mass number, 127, the number of neutrons is 74 ( $127 - 53 = 74$ ). Since the iodine is added as a  $1^-$  anion, the number of electrons is 54 [ $53 - (1^-) = 54$ ].

Check Your Learning: An ion of platinum has a mass number of 195 and contains 74 electrons. How many protons and neutrons does it contain, and what is its charge?

**ANSWER: 78 protons; 117 neutrons; charge is  $4+$**



## Chemical Symbols and Formulas

A **chemical symbol** is an abbreviation that we use to indicate an element or an atom of an element. For example, the symbol for mercury is Hg (Figure 6.4). We use the same symbol to indicate one atom of mercury (microscopic domain) or to label a container of many atoms of the element mercury (macroscopic domain).

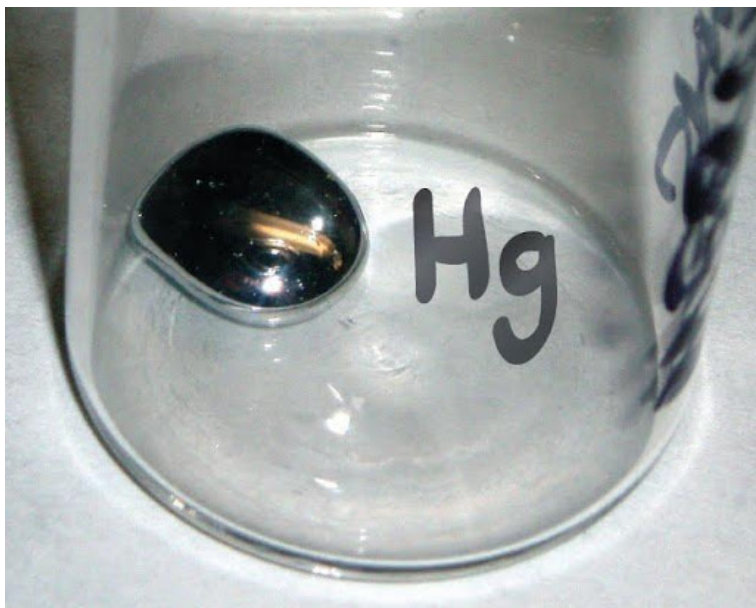


Fig 6.4 The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

The symbols for several common elements and their atoms are listed in Table 2.2. Some symbols are derived from the common name of the element; others are abbreviations of the name in another language. Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the notation for the **compound** carbon monoxide, which contains atoms of the

elements carbon (C) and oxygen (O). All known elements and their symbols are in the **periodic table**.

---

Some Common Elements and Their Symbols			
Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from <i>ferrum</i> )
bromine	Br	lead	Pb (from <i>plumbum</i> )
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from <i>hydrargyrum</i> )
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from <i>kalium</i> )
copper	Cu (from <i>cuprum</i> )	silicon	Si
fluorine	F	silver	Ag (from <i>argentum</i> )
gold	Au (from <i>aurum</i> )	sodium	Na (from <i>natrium</i> )
helium	He	sulfur	S
hydrogen	H	tin	Sn (from <i>stannum</i> )
iodine	I	zinc	Zn

---

**Table 6.2**

Traditionally, the discoverer (or discoverers) of a new element names the element. However, until the name is recognized by the International Union of Pure and Applied Chemistry (IUPAC), the recommended name of the new element is based on the Latin word(s) for its atomic number. For example, element 106 was called unnilhexium (Unh), element 107 was called unnilseptium (Uns), and element 108 was called unniloctium (Uno) for several years. These elements are now named after scientists (or occasionally locations); for example, element 106 is now known as *seaborgium* (Sg) in honor of Glenn Seaborg, a Nobel Prize winner who was active in the discovery of several heavy elements.

### Link to Learning

Visit the IUPAC, the International Union of Pure and Applied Chemistry, website to learn more and explore its

## Isotopes

The symbol for a specific **isotope** of any element is written by placing the mass number as a superscript to the left of the element symbol (Figure 2.5). The atomic number is sometimes written as a subscript preceding the symbol, but since this number defines the element's identity, as does its symbol, it is often omitted. For example, magnesium exists as a mixture of three

isotopes, each with an **atomic number** of 12 and with mass numbers of 24, 25, and 26, respectively. These isotopes can be identified as  $^{24}\text{Mg}$ ,  $^{25}\text{Mg}$ , and  $^{26}\text{Mg}$ . These isotope symbols are read as “element, mass number” and can be symbolized consistent with this reading. For instance,  $^{24}\text{Mg}$  is read as “magnesium 24,” and can be written as “magnesium-24” or “Mg-24.”  $^{25}\text{Mg}$  is read as “magnesium 25,” and can be written as “magnesium-25” or “Mg-25.” All magnesium atoms have 12 protons in their nucleus. They differ only because a  $^{24}\text{Mg}$  atom has 12 neutrons in its nucleus, a  $^{25}\text{Mg}$  atom has 13 neutrons, and a  $^{26}\text{Mg}$  has 14 neutrons.

periodic table  
listed here:  
[IUPAC](#).

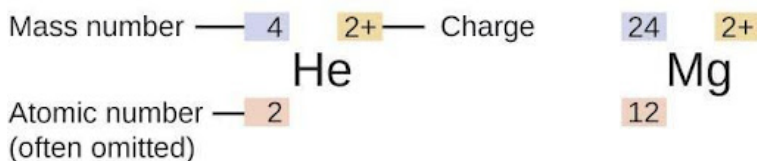


Fig. 6.5 The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge as a right superscript.

### Link to Learning

Use this building an atom simulator to build atoms of the first 10 elements, see which isotopes exist, check nuclear stability, and gain experience with isotope symbols: [PhET Build an Atom simulator](#).

## Atomic Mass

Because each proton and each neutron contribute approximately one amu to the mass of an atom, and each electron contributes far less, the **atomic mass** of a single atom is approximately equal to its mass number (a whole number). However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more **isotopes**.

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element.

For example, the element boron is composed of two isotopes: About 19.9% of all boron atoms are  $^{10}\text{B}$  with a mass of 10.0129 amu, and the remaining 80.1% are  $^{11}\text{B}$  with a mass of 11.0093 amu. The average atomic mass is then calculated by multiplying the individual atomic

masses of the isotopes to their respective percent abundances and adding their products together.

For boron, as 19.9% and 80.1% are equal to 0.199 and 0.801, we have

$$(10.0129 \text{ amu} \times 0.199) + (11.0093 \text{ amu} \times 0.801) = 10.8 \text{ amu}$$

It is important to understand that no single boron atom weighs exactly 10.8 amu; 10.8 amu is the average mass of all boron atoms, and individual boron atoms weigh either approximately 10 amu or 11 amu.

The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a **mass spectrometer**.

Mass spectrometry (MS) is widely used in chemistry, forensics, medicine, environmental science, and many other fields to analyze and help identify the substances in a sample of material. In a typical mass spectrometer (Figure 6.6), the sample is vaporized and exposed to a high-energy electron beam that causes the sample's **atoms** (or **molecules**) to become electrically charged, typically by

### Link to Learning

Use this simulation to make mixtures of the main isotopes of the first 18 elements, gain experience with average atomic mass, and check naturally

occurring isotope ratios using the Isotopes and Atomic Mass simulation: [PhET Isotopes and Atomic Mass simulation](#).

losing one or more electrons. These **cations** then pass through a (variable) electric or magnetic field that deflects each cation's path to an extent that depends on both its mass and charge. The **ions** are detected, and a plot of the relative number of ions generated versus their mass-to-charge ratios (a *mass spectrum*) is made. The height of each vertical feature or peak in a mass spectrum is proportional to the fraction of cations with the specified mass-to-

charge ratio. Since its initial use during the development of modern atomic theory, MS has evolved to become a powerful tool for chemical analysis in a wide range of applications.



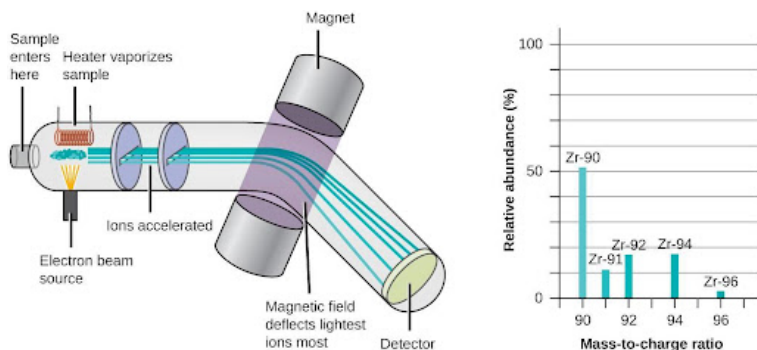


Fig. 6.6 Analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr.

## Molecular, Structural, and Empirical Formulas

A **molecular formula** is a representation of a molecule that uses chemical symbols to indicate the types of atoms followed by subscripts to show the number of atoms of each type in the molecule. (A subscript is used only when more than one atom of a given type is present.) Molecular formulas are also used as

### Link to Learning

See this video from Bozeman Science that explains mass spectrometry: [Mass Spectrometry](#). Watch this video from

the Royal  
Society for  
Chemistry for a  
brief  
description of  
the rudiments  
of mass  
spectrometry:  
[Mass](#)  
[Spectrometry](#)  
[MS](#).

abbreviations for the names of compounds.

The **structural formula** for a compound gives the same information as its molecular formula (the types and numbers of atoms in the molecule) but also shows how the atoms are connected in the molecule. The structural formula for methane contains symbols for one C atom and four H atoms, indicating the number of atoms in the molecule (Fig. 6.7). The lines represent bonds that hold the atoms

together. (A **chemical bond** is an attraction between atoms or ions that holds them together in a molecule or a crystal.) We will discuss chemical bonds and see how to predict the arrangement of atoms in a molecule later. For now, simply know that the lines are an indication of how the atoms are connected in a molecule. A ball-and-stick model shows the geometric arrangement of the atoms with atomic sizes not to scale, and a space-filling model shows the relative sizes of the atoms.

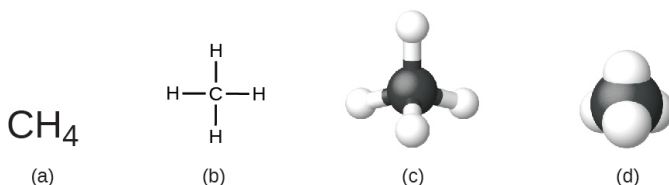


Fig 6.7 A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.

Although many **elements** consist of discrete, individual atoms, some exist as molecules made up of two or more atoms of the element chemically bonded together. For example, most samples of the elements hydrogen, oxygen, and nitrogen are composed of molecules that contain two atoms each (called diatomic molecules) and thus have the molecular formulas  $\text{H}_2$ ,  $\text{O}_2$ , and  $\text{N}_2$ , respectively. Other elements commonly found as diatomic molecules are fluorine ( $\text{F}_2$ ), chlorine ( $\text{Cl}_2$ ), bromine ( $\text{Br}_2$ ), and iodine ( $\text{I}_2$ ). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is  $\text{S}_8$ .

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example,  $\text{H}_2$  and  $2\text{H}$  represent distinctly different species.  $\text{H}_2$  is a molecular formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together. The expression  $2\text{H}$ , on

the other hand, indicates two separate hydrogen atoms that are not combined as a unit. The expression  $2\text{H}_2$  represents two molecules of diatomic hydrogen (Fig. 6.8).

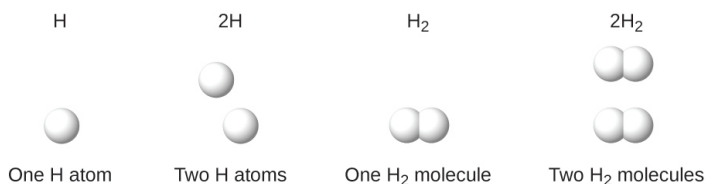


Fig 6.8 The symbols H, 2H, H<sub>2</sub>, and 2H<sub>2</sub> represent very different entities.

**Compounds** are formed when two or more elements chemically combine, resulting in the formation of bonds. For example, hydrogen and oxygen can react to form water, and sodium and chlorine can react to form table salt. We sometimes describe the composition of these compounds with an **empirical formula**, which indicates the types of atoms present and *the simplest whole-number ratio of the number of atoms (or ions) in the compound*. For example, titanium dioxide (used as pigment in white paint and in the thick, white, blocking type of sunscreen) has an empirical formula of  $\text{TiO}_2$ . This identifies the elements titanium (Ti) and oxygen (O) as the constituents of titanium dioxide, and indicates the presence of twice as many atoms of the element oxygen as atoms of the element titanium.

As discussed previously, we can describe a compound with

a **molecular formula**, in which the subscripts indicate the *actual numbers of atoms* of each element in a molecule of the compound. In many cases, the molecular formula of a substance is derived from experimental determination of both its empirical formula and its **molecular mass** (the sum of atomic masses for all atoms composing the molecule).

It is important to be aware that it may be possible for the same atoms to be arranged in different ways: Compounds with the same molecular formula may have different atom-to-atom bonding and therefore different structures. Compounds with the same chemical formula but different molecular structures are called **isomers**.

## The Periodic Table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K): These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive

than are Ca, Sr, and Ba; Li, Na, and K form compounds with oxygen in a ratio of two of their atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one of their atoms to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of any of the elements above.

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

By the twentieth century, it became apparent that the periodic relationship involved **atomic numbers** rather than **atomic masses**. The modern statement of this relationship, the periodic law, is as follows: the properties of the elements are periodic functions of their atomic numbers. A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in

the same vertical column (Figure 6.9). Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. The elements are arranged in seven horizontal rows, called periods or series, and 18 vertical columns, called groups. Groups are labeled at the top of each column. In the United States, the labels traditionally were numerals with capital letters. However, IUPAC recommends that the numbers 1 through 18 be used, and these labels are more common. For the table to fit on a single page, parts of two of the rows, a total of 14 columns, are usually written below the main body of the table.

Periodic Table of the Elements

Period

Group

1	1	2																	13	14	15	16	17	18																																																																	
1	H 1.008 hydrogen																		2	He 4.003 helium																																																																					
2	3	4							5	6	7	8	9	10	11	12																																																																									
	3	Li 6.94 lithium	4	Be 9.012 beryllium							5	B 10.81 boron	6	C 12.01 carbon	7	N 14.01 nitrogen	8	O 16.00 oxygen	9	F 19.00 fluorine	10	Ne 20.18 neon																																																																			
3	11	Na 22.99 sodium	12	Mg 24.31 magnesium							13	Al 26.98 aluminum	14	Si 28.09 silicon	15	P 30.97 phosphorus	16	S 32.06 sulfur	17	Cl 35.45 chlorine	18	Ar 39.95 argon																																																																			
4	19	K 39.10 potassium	20	Ca 40.08 calcium	21	Sc 44.96 scandium	22	Ti 47.87 titanium	23	V 50.94 vanadium	24	Cr 52.00 chromium	25	Mn 54.94 manganese	26	Fe 55.85 iron	27	Co 58.93 cobalt	28	Ni 58.69 nickel	29	Cu 63.55 copper	30	Zn 65.38 zinc	31	Ga 69.72 gallium	32	Ge 72.63 germanium	33	As 74.92 arsenic	34	Se 78.97 selenium	35	Br 79.90 bromine	36	Kr 83.80 krypton																																																					
5	37	Rb 85.47 rubidium	38	Sr 87.62 strontium	39	Y 88.91 yttrium	40	Zr 91.22 zirconium	41	Nb 92.91 niobium	42	Mo 95.95 molybdenum	43	Tc [97] technetium	44	Ru 101.1 ruthenium	45	Rh 101.9 rhodium	46	Pd 106.4 palladium	47	Ag 107.9 silver	48	Cd 112.4 cadmium	49	In 114.8 indium	50	Sn 118.7 tin	51	Sb 121.8 antimony	52	Te 127.6 tellurium	53	I 126.9 iodine	54	Xe 131.3 xenon																																																					
6	55	Cs 132.9 cesium	56	Ba 137.3 barium	57-71 La-Lu lanthanum series		72	Hf 178.5 hafnium	73	Ta 180.9 tantalum	74	W 183.8 tungsten	75	Re 186.2 rhenium	76	Os 190.2 osmium	77	Ir 192.2 iridium	78	Pt 195.1 platinum	79	Au 197.0 gold	80	Hg 200.6 mercury	81	Tl 204.4 thallium	82	Pb 207.2 lead	83	Bi 209.0 bismuth	84	Po [209] polonium	85	At [210] astatine	86	Rn [222] radon																																																					
7	87	Fr [223] francium	88	Ra [226] radium	89-103 Ac-Lr actinide series		104	Rf [261] rutherfordium	105	Db [271] dubnium	106	Sg [271] seaborgium	107	Bh [271] bohrium	108	Hs [277] hassium	109	Mt [276] meitnerium	110	Ds [281] darmstadtium	111	Rg [281] roentgenium	112	Cn [285] copernicium	113	Uut [285] ununtrium	114	Fl [289] flerovium	115	Uup [289] ununpentium	116	Lv [293] livermorium	117	Uus [294] ununseptium	118	Uuo [294] ununoctium																																																					
																				<table> <tr> <td>57</td><td>La 138.9 lanthanum</td><td>58</td><td>Ce 140.1 cerium</td><td>59</td><td>Pr 140.9 praseodymium</td><td>60</td><td>Nd 144.2 neodymium</td><td>61</td><td>Pm [145] promethium</td><td>62</td><td>Sm 150.4 samarium</td><td>63</td><td>Eu 152.0 europium</td><td>64</td><td>Gd 157.3 gadolinium</td><td>65</td><td>Tb 158.9 terbium</td><td>66</td><td>Dy 162.5 dysprosium</td><td>67</td><td>Ho 164.9 holmium</td><td>68</td><td>Er 167.3 erbium</td><td>69</td><td>Tm 168.9 thulium</td><td>70</td><td>Yb 173.1 ytterbium</td><td>71</td><td>Lu 175.0 lutetium</td></tr> <tr> <td>89</td><td>Ac [227] actinium</td><td>90</td><td>Th 232.0 thorium</td><td>91</td><td>Pa 231.0 protactinium</td><td>92</td><td>U 238.0 uranium</td><td>93</td><td>Np [237] neptunium</td><td>94</td><td>Pu [244] plutonium</td><td>95</td><td>Am [243] americium</td><td>96</td><td>Cm [247] curium</td><td>97</td><td>Bk [247] berkelium</td><td>98</td><td>Cf [251] californium</td><td>99</td><td>Es [252] einsteinium</td><td>100</td><td>Fm [257] fermium</td><td>101</td><td>Md [258] mendelevium</td><td>102</td><td>No [259] nobelium</td><td>103</td><td>Lr [262] lawrencium</td></tr> </table>										57	La 138.9 lanthanum	58	Ce 140.1 cerium	59	Pr 140.9 praseodymium	60	Nd 144.2 neodymium	61	Pm [145] promethium	62	Sm 150.4 samarium	63	Eu 152.0 europium	64	Gd 157.3 gadolinium	65	Tb 158.9 terbium	66	Dy 162.5 dysprosium	67	Ho 164.9 holmium	68	Er 167.3 erbium	69	Tm 168.9 thulium	70	Yb 173.1 ytterbium	71	Lu 175.0 lutetium	89	Ac [227] actinium	90	Th 232.0 thorium	91	Pa 231.0 protactinium	92	U 238.0 uranium	93	Np [237] neptunium	94	Pu [244] plutonium	95	Am [243] americium	96	Cm [247] curium	97	Bk [247] berkelium	98	Cf [251] californium	99	Es [252] einsteinium	100	Fm [257] fermium	101	Md [258] mendelevium	102	No [259] nobelium	103	Lr [262] lawrencium
57	La 138.9 lanthanum	58	Ce 140.1 cerium	59	Pr 140.9 praseodymium	60	Nd 144.2 neodymium	61	Pm [145] promethium	62	Sm 150.4 samarium	63	Eu 152.0 europium	64	Gd 157.3 gadolinium	65	Tb 158.9 terbium	66	Dy 162.5 dysprosium	67	Ho 164.9 holmium	68	Er 167.3 erbium	69	Tm 168.9 thulium	70	Yb 173.1 ytterbium	71	Lu 175.0 lutetium																																																												
89	Ac [227] actinium	90	Th 232.0 thorium	91	Pa 231.0 protactinium	92	U 238.0 uranium	93	Np [237] neptunium	94	Pu [244] plutonium	95	Am [243] americium	96	Cm [247] curium	97	Bk [247] berkelium	98	Cf [251] californium	99	Es [252] einsteinium	100	Fm [257] fermium	101	Md [258] mendelevium	102	No [259] nobelium	103	Lr [262] lawrencium																																																												

Atomic number → 1

Symbol → H

Atomic mass → 1.008

Name → hydrogen

Color Code

Metal	Solid
Metalloid	Liquid
Nonmetal	Gas

Fig. 6.9 Elements in the periodic table are organized according to their properties.

Even after the periodic nature of elements and the table itself were widely accepted, gaps remained. Mendeleev had predicted, and others including Henry Moseley had later confirmed, that there should be elements below Manganese in Group 7. German chemists Ida Tacke and Walter Noddack set out to find the elements, a quest being pursued by scientists around the world. Their method was unique in that they did not only consider the properties of manganese, but also the elements horizontally adjacent to the missing elements 43 and 75 on the table. Thus, by investigating ores containing minerals of ruthenium (Ru), tungsten (W), osmium (Os), and so on, they were able to identify naturally occurring elements that helped complete the table. Rhenium, one of their discoveries, was one of the last natural elements to be discovered and is the last stable element to be discovered. (Francium, the last natural element to be discovered, was identified by Marguerite Perey in 1939.)

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



conductors of heat and electricity—shaded yellow); **nonmetals** (elements that appear dull, poor conductors of heat and electricity—shaded green); and **metalloids** (elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals—shaded purple).

The elements can also be classified into the main-group elements (or representative elements) in the columns labeled 1, 2, and 13–18; the transition metals in the columns labeled 3–12; and inner transition metals in the two rows at the bottom of the table (the top-row elements are called lanthanides and the bottom-row elements are actinides

## Electron Structure of Atoms and the Aufbau Principle

**Electrons** in successive atoms on the periodic table tend to fill low-energy orbitals first. The arrangement of electrons in the orbitals of an atom is called the **electron configuration** of the atom.

To determine the electron configuration for any particular atom, we can “build” the structures in the order of atomic numbers. Beginning with hydrogen, and continuing across the periods of the periodic table, we add one proton at a time to the nucleus and one electron to the proper subshell until we have described the electron configurations of all the elements. This procedure is called the **Aufbau principle**, from the

German word *Aufbau* (“to build up”). Each added electron occupies the subshell of lowest energy available. Electrons enter higher-energy subshells only after lower-energy subshells have been filled to capacity.

## Electron Configurations and the Periodic Table

The periodic table arranges atoms based on increasing atomic number so that elements with the same chemical properties recur periodically. Because they are in the outer shells of an atom, **valence electrons** play the most important role in chemical reactions (Figure 6.10). The outer electrons have the highest energy of the electrons in an atom and are more easily lost or shared than the core electrons. Valence electrons are also the determining factor in some physical properties of the elements.

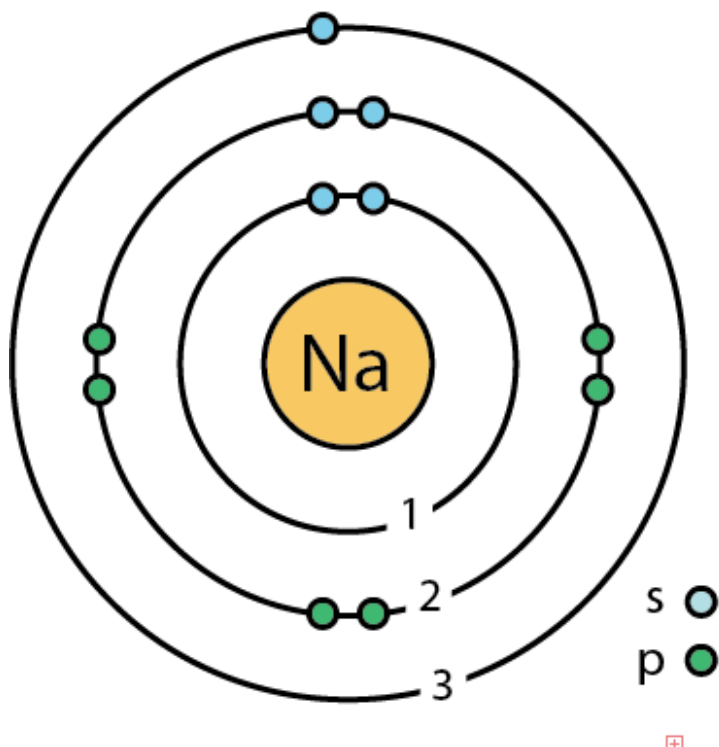


Fig 6.10 A Bohr model of an atom of sodium (Na) illustrating core electrons at lower energy levels (which are shells closer to the nucleus) and its single valence electron at its outermost (3rd) shell.

Elements in any one group (or column) have the same number of valence electrons; the alkali metals lithium and sodium each have only one valence electron, the alkaline earth metals beryllium and magnesium each have two, and the halogens fluorine and chlorine each have seven valence electrons. The similarity in chemical properties among elements of the same group occurs because they have the same number of valence

electrons. It is the loss, gain, or sharing of valence electrons that defines how elements react.

It is important to remember that the periodic table was developed on the basis of the chemical behavior of the elements, well before any idea of their atomic structure was available. Now we can understand why the periodic table has the arrangement it has—the arrangement puts elements whose atoms have the same number of valence electrons in the same group.

## Check Your Understanding



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7.

# IONIC AND MOLECULAR COMPOUNDS AND BONDING

---

## Chapter Learning Objectives

After successful completion of the assignments in this section, you will be able to demonstrate competency in the following areas:

1. Define the Components and Types of Ionic Compounds (CLO6)
2. Model Electron Configuration by using Lewis Structures (CLO6)(CLO7)
3. Apply the Octet Rule for Lewis Structures (CLO6)(CLO7)
4. Discuss everyday situations related to

molecular structures (CLO8)

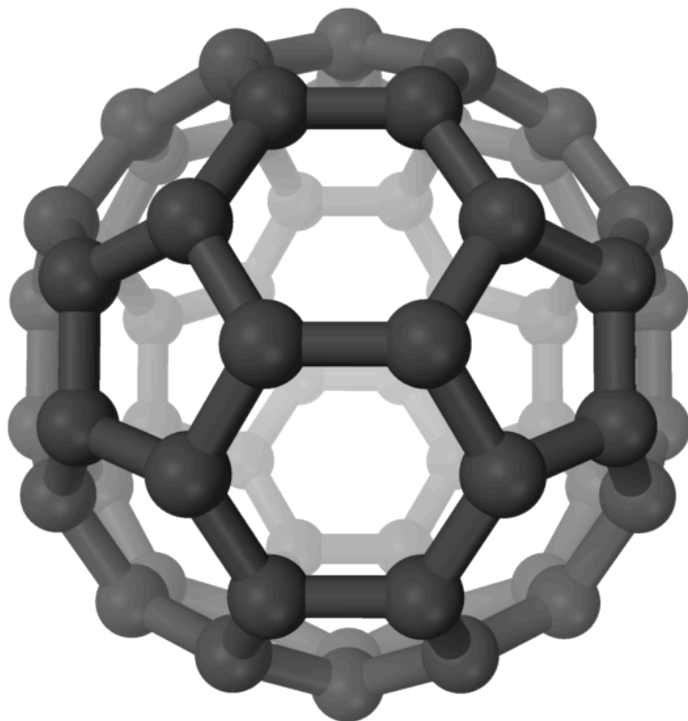


Fig 7.1 Buckminsterfullerene is a type of fullerene. Each of its 60 carbon atoms is bonded to its three neighbors.



# Ionic and Molecular Compounds

In ordinary chemical reactions, the nucleus of each atom (and thus the identity of the element) remains unchanged. Electrons, however, can be added to **atoms** by transfer from other atoms, lost by transfer to other atoms, or shared with other atoms. The transfer and sharing of electrons among atoms govern the chemistry of the elements. During the formation of some **compounds**, atoms gain or lose electrons, and form electrically charged particles called ions.

As you have learned, **ions** are atoms or molecules bearing an electrical charge. An atom that gains one or more electrons will exhibit a negative charge and is called an **anion**. Positively charged atoms called **cations** are formed when an atom loses one or more electrons. You can use the **periodic table** to predict whether an atom will form an anion or a cation, and you can often predict the charge of the resulting ion. Atoms of many main-group metals lose enough electrons to leave them with the same number of electrons as an atom of the preceding noble gas.

## Ionic Compounds

When an element composed of atoms that readily lose electrons (a metal) reacts with an element composed of atoms

that readily gain electrons (a nonmetal), a transfer of electrons usually occurs, producing ions. The compound formed by this transfer is stabilized by the electrostatic attractions (ionic bonds) between the ions of opposite charge present in the compound.

A compound that contains ions and is held together by **ionic bonds** is called an **ionic compound** (or salts). The periodic table can help us recognize many of the compounds that are ionic: When a metal is combined with one or more nonmetals, the compound is usually ionic. This guideline works well for predicting ionic compound formation for most of the compounds typically encountered in an introductory chemistry course. However, it is not always true (for example, aluminum chloride,  $\text{AlCl}_3$ , is not ionic).

In every ionic compound, the total number of positive charges of the **cations** equals the total number of negative charges of the **anions**. Thus, ionic compounds are electrically neutral overall, even though they contain positive and negative ions. We can use this observation to help us write the formula of an ionic compound. The formula of an ionic compound must have a ratio of ions such that the numbers of positive and negative charges are equal.

The properties of ionic compounds shed some light on the nature of ionic bonds. Ionic solids exhibit a crystalline structure and tend to be rigid and brittle; they also tend to have high melting and boiling points, which suggests that ionic bonds are very strong. Ionic solids are also poor conductors

of electricity for the same reason—because electricity is the movement of electrons, the strength of ionic bonds prevents ions from moving freely in the solid state. Most ionic solids, however, dissolve readily in water. Once dissolved or melted, ionic compounds are excellent conductors of electricity and heat because the ions can move about freely.

## Molecular Compounds

Many compounds do not contain ions but instead consist solely of discrete, neutral **molecules**. These **molecular compounds** (covalent compounds) result when atoms share, rather than transfer (gain or lose), electrons. We can often identify molecular compounds on the basis of their physical properties. Under normal conditions, molecular compounds often exist as gases, low-boiling liquids, and low-melting solids, although many important exceptions exist.

Whereas ionic compounds are usually formed when a metal and a nonmetal combine, covalent compounds are usually formed by a combination of nonmetals. Thus, the periodic table can help us recognize many of the compounds that are covalent. While we can use the positions of a compound's elements in the periodic table to predict whether it is ionic or covalent at this point in our study of chemistry, you should be aware that this is a very simplistic approach that does not account for a number of interesting exceptions. Shades of gray

exist between ionic and molecular compounds, and you'll learn more about those later.

Neutral atoms and their associated ions have very different physical and chemical **properties**. Sodium atoms form sodium metal, a soft, silvery-white metal that burns vigorously in air and reacts explosively with water. Chlorine *atoms* form chlorine gas,  $\text{Cl}_2$ , a yellow-green gas that is extremely corrosive to most metals and very poisonous to animals and plants. The vigorous reaction between the elements sodium and chlorine forms the white, crystalline compound sodium chloride, common table salt, which contains sodium *cations* and chloride *anions* (Figure 7.2). The compound composed of these ions exhibits properties entirely different from the properties of the elements sodium and chlorine. Chlorine is poisonous, but sodium chloride is essential to life; sodium atoms react vigorously with water, but sodium chloride simply dissolves in water.



Fig. 7.2 (a) Sodium is a soft metal that must be stored in mineral oil to prevent reaction with air or water. (b) Chlorine is a pale yellow-green gas. (c) When combined, they form white crystals of sodium chloride (table salt)

## The Formation of Ionic Compounds

**Binary ionic compounds** are composed of just two elements: a metal (which forms the **cations**) and a nonmetal (which forms the **anions**). For example, NaCl is a binary ionic compound. We can think about the formation of such compounds in terms of the periodic properties of the elements. Many metallic elements have relatively low ionization potentials (the energy necessary to remove an electron from an atom) and lose electrons easily. These elements lie to the left in a period or near the bottom of a group on the periodic table. Nonmetal atoms have relatively high electron affinities and thus readily gain electrons lost by metal atoms, thereby filling their valence shells. Nonmetallic elements are found in the upper-right corner of the periodic table.

As all substances must be electrically neutral, the total number of positive charges on the cations of an ionic compound must equal the total number of negative charges on its anions.

It is important to note that the formula for an ionic compound does *not* represent the physical arrangement of its ions. It is incorrect to refer to a sodium chloride (NaCl) “molecule” because there is not a single **ionic bond**, per se, between any specific pair of sodium and chloride ions. The attractive forces between ions are isotropic—the same in all

directions—meaning that any particular ion is equally attracted to all of the nearby ions of opposite charge. This results in the ions arranging themselves into a tightly bound, three-dimensional lattice structure. Sodium chloride, for example, consists of a regular arrangement of equal numbers of  $\text{Na}^+$  cations and  $\text{Cl}^-$  anions (Figure 7.3).

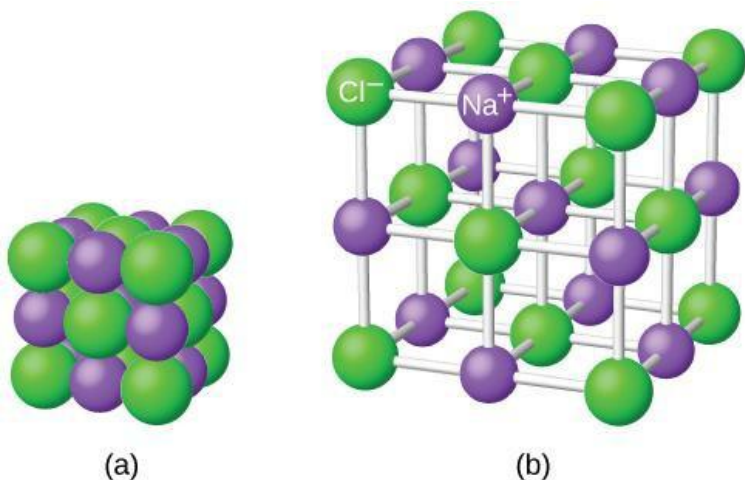


Fig. 7.3 The atoms in sodium chloride (common table salt) are arranged to (a) maximize opposite charges interacting. The smaller spheres (purple) represent sodium ions, the larger ones (green) represent chloride ions. In the expanded view (b), the geometry can be seen more clearly. Note that each ion is “bonded” to all of the surrounding ions—six in this case.

The strong electrostatic attraction between  $\text{Na}^+$  and  $\text{Cl}^-$  ions holds them tightly together in solid  $\text{NaCl}$ . It requires 769 kJ

of energy to dissociate one mole of solid NaCl into separate gaseous  $\text{Na}^+$  and  $\text{Cl}^-$  ions.

## Covalent Bonds

Ionic bonding results from the electrostatic attraction of oppositely charged ions that are typically produced by the transfer of electrons between metallic and nonmetallic atoms. Covalent bonding results from the mutual attraction of atoms for a “shared” pair of electrons and are typically formed between two nonmetals. **Covalent bonds** are formed between two atoms when both have similar tendencies to attract electrons to themselves (i.e., when both atoms have identical or fairly similar ionization energies and electron affinities). For example, two hydrogen atoms bond covalently to form an  $\text{H}_2$  molecule; each hydrogen atom in the  $\text{H}_2$  molecule has two electrons stabilizing it, giving each atom the same number of valence electrons (2) as the noble gas He.

Compounds that contain covalent bonds exhibit different physical properties than ionic compounds. Because the attraction between molecules, which are electrically neutral, is weaker than that between electrically charged ions, covalent compounds generally have much lower melting and boiling points than ionic compounds. In fact, many covalent compounds are liquids or gases at room temperature, and in their solid states they are typically much softer than ionic solids. Furthermore, whereas ionic compounds are good

conductors of electricity when dissolved in water, most covalent compounds are insoluble in water; since they are electrically neutral, they are poor conductors of electricity in any state.

## Pure vs. Polar Covalent Bonds

If the atoms that form a covalent bond are identical, as in  $\text{H}_2$ ,  $\text{Cl}_2$ , and other diatomic molecules, then the electrons in the bond must be shared equally. We refer to this as a pure covalent bond. Electrons shared in **pure covalent bonds** have an equal probability of being near each nucleus.

In the case of  $\text{Cl}_2$ , each atom starts off with seven valence electrons, and each Cl shares one electron with the other, forming one covalent bond:  $\text{Cl} + \text{Cl} \rightarrow \text{Cl}_2$

(See Lewis structure diagram for  $\text{Cl}_2$  on pg 28)

The total number of electrons around each individual atom consists of six nonbonding electrons and two shared (i.e., bonding) electrons for eight total electrons, matching the number of valence electrons in the noble gas argon. Since the bonding atoms are identical,  $\text{Cl}_2$  also features a pure covalent bond.

When the atoms linked by a covalent bond are different, the bonding electrons are shared, but no longer equally. Instead, the bonding electrons are more attracted to one atom than the other, giving rise to a shift of electron density toward that atom. This unequal distribution of electrons is known as a



**polar covalent bond**, characterized by a partial positive charge on one atom and a partial negative charge on the other. The atom that attracts the electrons more strongly acquires the partial negative charge and vice versa. For example, the electrons in the H–Cl bond of a hydrogen chloride molecule spend more time near the chlorine atom than near the hydrogen atom. Thus, in an HCl molecule, the chlorine atom carries a partial negative charge and the hydrogen atom has a partial positive charge. Figure 3.3 shows the distribution of electrons in the H–Cl bond. Note that the shaded area around Cl is much larger than it is around H, compared to the even distribution of electrons in the H<sub>2</sub> nonpolar bond.

We sometimes designate the positive and negative atoms in a polar covalent bond using a lowercase Greek letter “delta,”  $\delta$ , with a plus sign or minus sign to indicate whether the atom has a partial positive charge ( $\delta^+$ ) or a partial negative charge ( $\delta^-$ ). This symbolism is shown for the H–Cl molecule in Figure 7.4.

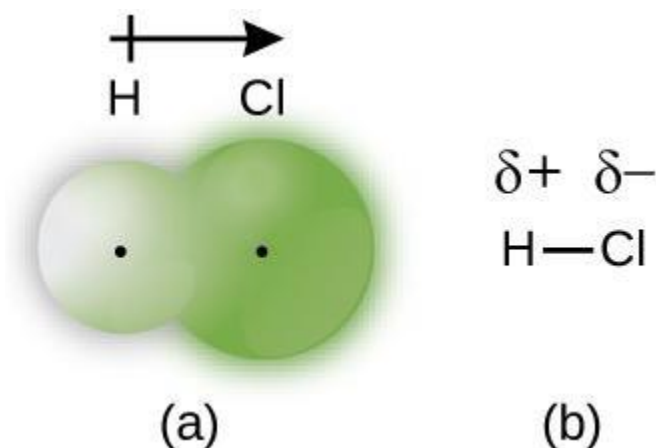


Fig. 7.4 The distribution of electron density in the HCl molecule is uneven. The electron density is greater around the chlorine nucleus. The small, black dots indicate the location of the hydrogen and chlorine nuclei in the molecule. (b) Symbols  $\delta+$  and  $\delta-$  indicate the polarity of the H-Cl bond.

## Metallic Bonds

When metal atoms bond in **metallic bonds** their **valence electrons** are shared among many atoms. These valence electrons can move around and are not fixed to any particular atom or pair of atoms, unlike both ionic and covalent bonds. Metals have unique properties like their conductivity and malleability because these electrons shared between the metal atoms can move easily. Atoms in metals are arranged in a crystalline structure that is held together by the strong bonds between the cations and the free “sea” of electrons.

## Electronegativity

Whether a bond is nonpolar or polar covalent is determined by a property of the bonding atoms called **electronegativity**. Electronegativity is a measure of the tendency of an atom to attract electrons (or electron density) towards itself. It determines how the shared electrons are distributed between the two atoms in a bond. The more strongly an atom attracts the electrons in its bonds, the larger its electronegativity. Electrons in a **polar covalent bond** are shifted toward the more electronegative atom; thus, the more electronegative atom is the one with the partial negative charge. The greater the difference in electronegativity, the more polarized the electron distribution and the larger the partial charges of the atoms.

Figure 7.5 shows the electronegativity values of the elements as proposed by one of the most famous chemists of the twentieth century: Linus Pauling. In general, electronegativity increases from left to right across a period in the periodic table and decreases down a group. The nonmetals, which lie in the upper right, tend to have the highest electronegativities, with fluorine the most electronegative element of all ( $EN = 4.0$ ). Metals tend to be less electronegative elements, and the group 1 metals have the lowest electronegativities. Note that noble gases are excluded from this figure because these atoms usually do not share electrons with other atoms since they have a full valence shell. (While noble gas compounds such as  $XeO_2$  do

exist, they can only be formed under extreme conditions, and thus they do not fit neatly into the general model of electronegativity.)

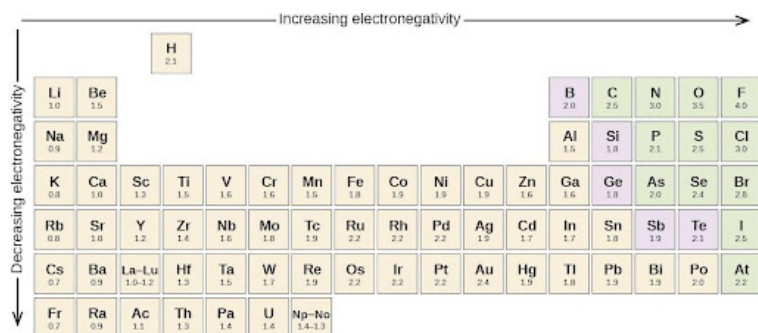


Fig. 7.5 The electronegativity values derived by Pauling follow predictable periodic trends, with the higher electronegativities toward the upper right of the periodic table.

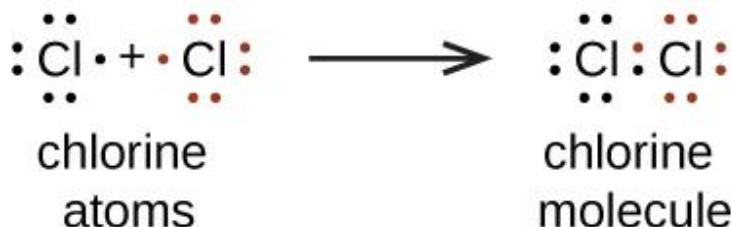
## Lewis Symbols

We use Lewis symbols to describe valence electron configurations of atoms and monatomic ions. A **Lewis symbol** consists of an elemental symbol surrounded by one dot for each of its valence electrons:



## Lewis Structures

We also use Lewis symbols to indicate the formation of covalent bonds, which are shown in **Lewis structures**, drawings that describe the bonding in molecules and polyatomic ions. For example, when two chlorine atoms form a chlorine molecule, they share one pair of electrons:



The Lewis structure indicates that each Cl atom has three pairs of electrons that are not used in bonding (called lone pairs) and one shared pair of electrons (written between the atoms). A dash (or line) is sometimes used to indicate a shared pair of electrons:

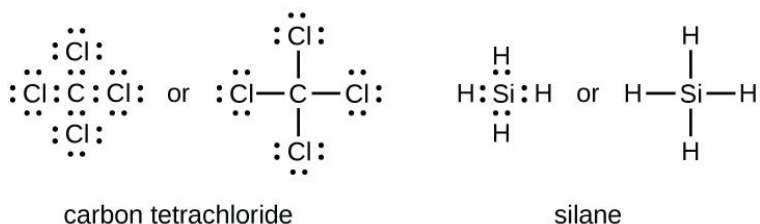


A single shared pair of electrons is called a single bond. Each Cl atom interacts with eight valence electrons: the six in the lone pairs and the two in the single bond.

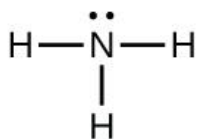
## The Octet Rule

The other halogen molecules ( $F_2$ ,  $Br_2$ ,  $I_2$ , and  $At_2$ ) form bonds like those in the chlorine molecule: one single bond between atoms and three lone pairs of electrons per atom. This allows each halogen atom to have a noble gas electron configuration. The tendency of main group atoms to form enough bonds to obtain eight valence electrons is known as the **octet rule**.

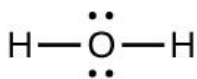
The number of bonds that an atom can form can often be predicted from the number of electrons needed to reach an octet (eight valence electrons); this is especially true of the nonmetals of the second period of the periodic table (C, N, O, and F). For example, each atom of a group 14 element has four electrons in its outermost shell and therefore requires four more electrons to reach an octet. These four electrons can be gained by forming four covalent bonds, as illustrated here for carbon in  $CCl_4$  (carbon tetrachloride) and silicon in  $SiH_4$  (silane). Because hydrogen only needs two electrons to fill its valence shell, it is an exception to the octet rule. The transition elements and inner transition elements also do not follow the octet rule:



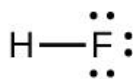
Group 15 elements such as nitrogen have five valence electrons in the atomic Lewis symbol: one lone pair and three unpaired electrons. To obtain an octet, these atoms form three covalent bonds, as in  $\text{NH}_3$  (ammonia). Oxygen and other atoms in group 16 obtain an octet by forming two covalent bonds:



ammonia

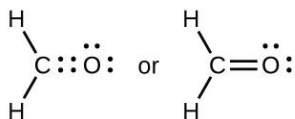


Water

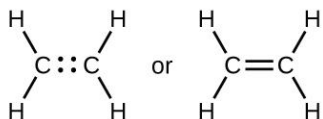
hydrogen  
fluoride

## Double and Triple Bonds

As previously mentioned, when a pair of atoms shares one pair of electrons, we call this a **single bond**. However, a pair of atoms may need to share more than one pair of electrons in order to achieve the necessary octet. A **double bond** forms when two pairs of electrons are shared between a pair of atoms, as between the carbon and oxygen atoms in  $\text{CH}_2\text{O}$  (formaldehyde) and between the two carbon atoms in  $\text{C}_2\text{H}_4$  (ethylene):

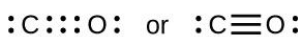


formaldehyde

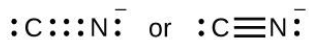


ethylene

A **triple bond** forms when three electron pairs are shared by a pair of atoms, as in carbon monoxide (CO) and the cyanide ion (CN<sup>-</sup>):



carbon monoxide



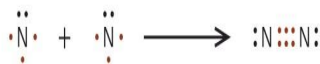
cyanide ion

### Link to Learning

Use this simulation to explore 3D models of various molecules, comparing them to their Lewis Structures (opens in a new window): [PhET Molecule Shapes](#).

## Writing Lewis Structures with the Octet Rule

For very simple molecules and molecular ions, we can write the **Lewis structures** by merely pairing up the unpaired electrons on the constituent atoms. See these examples:



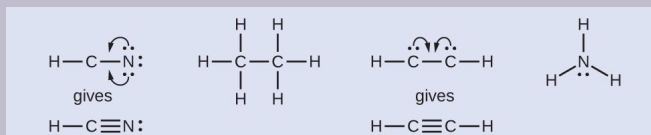


## Example 7.1

### WRITING LEWIS STRUCTURES

NASA's Cassini-Huygens mission detected a large cloud of toxic hydrogen cyanide (HCN) on Titan, one of Saturn's moons. Titan also contains ethane ( $\text{H}_3\text{CCH}_3$ ), acetylene ( $\text{HCCH}$ ), and ammonia ( $\text{NH}_3$ ). What are the Lewis structures of these molecules?

#### Solution



**Check Your Learning:** Both carbon monoxide,  $\text{CO}$ , and carbon dioxide,  $\text{CO}_2$ , are products of the combustion of fossil fuels. Both of these gases also cause problems:  $\text{CO}$  is toxic and  $\text{CO}_2$  has been implicated in global climate change. What are the Lewis structures of these two molecules?

#### ANSWER:



## Check Your Understanding



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8.

# MATTER AND ENERGY

## Transformation and Conservation

---

### Learning Objectives

After successful completion of this section, you will be able to demonstrate competency in the following areas:

1. Explain and apply the Law of Conservation of Energy (CLO1)(CLO2)(CLO3)(CLO5)(CLO6)
2. Define energy, kinetic energy, and gravitational potential energy (CLO2)(CLO3)
3. Analyze situations that involve kinetic and potential energy. (CLO2)(CLO3)
4. Calculate kinetic and potential energies. (CLO2)(CLO5)

5. Discuss the conservation of mass  
(CLO1)(CLO6)



Fig 8.1 How many forms of energy can you identify in this photograph of a wind farm in Iowa?

## Introduction to Energy

**Energy** plays an essential role both in everyday events and in scientific phenomena. You can no doubt name many forms of energy, from that provided by our foods, to the energy we use to run our cars, to the sunlight that warms us on the beach. You can also cite examples of what people call energy that may not be scientific, such as someone having an energetic

personality. Energy not only has many interesting forms; it is involved in almost all phenomena and is one of the most important concepts of physics. What makes it even more important is that the total amount of energy in the universe is constant. Energy can change forms, but it cannot appear from nothing or disappear without a trace. Energy is thus one of a handful of physical quantities that we say is conserved.

Conservation of energy (as physicists like to call the principle that energy can neither be created nor destroyed) is based on experiment. Even as scientists discovered new forms of energy, conservation of energy has always been found to apply. Perhaps the most dramatic example of this was supplied by Einstein when he suggested that mass is equivalent to energy (his famous equation  $E=mc^2$ ).

From a societal viewpoint, energy is one of the major building blocks of modern civilization. Energy resources are key limiting factors to economic growth. The world use of energy resources, especially oil, continues to grow, with ominous consequences economically, socially, politically, and environmentally.

There is no simple, yet accurate, scientific definition for energy. Energy is characterized by its many forms and the fact that it is conserved. We can loosely define energy as the ability to do **work**, admitting that in some circumstances not all energy is available to do work. Work is intimately related to energy and how energy moves from one system to another or changes form.

## Chemical Changes to Matter

Matter undergoes two types of changes: physical changes which do not change the composition of the original substance only its appearance and other physical attributes, and chemical changes which result in the formation of a new substance with different chemical and physical properties from the original substance.

**Chemical changes** and their accompanying changes in **energy** are important parts of our everyday world. The macronutrients in food (proteins, fats, and carbohydrates) undergo metabolic reactions that provide the energy to keep our bodies functioning. We burn a variety of fuels (gasoline, natural gas, coal) to produce energy for transportation, heating, and the generation of electricity. Industrial chemical reactions use enormous amounts of energy to produce raw materials (such as iron and aluminum). Energy is then used to manufacture those raw materials into useful products, such as cars, skyscrapers, and bridges.





Fig 8.2 The energy involved in chemical changes is important to our daily lives: (a) A cheeseburger for lunch provides the energy you need to get through the rest of the day; (b) the combustion of gasoline provides the energy that moves your car (and you) between home, work, and school; and (c) coke, a processed form of coal, provides the energy needed to convert iron ore into iron, which is essential for making many of the products we use daily.

Over 90% of the energy we use comes originally from the sun in the form of **radiant energy**. Every day, the sun provides the earth with almost 10,000 times the amount of energy necessary to meet all of the world's energy needs for that day. Our challenge is to find ways to convert and store incoming solar energy so that it can be used in reactions or chemical processes that are both convenient and non-polluting. Plants and many bacteria capture solar energy through photosynthesis. We release the energy stored in plants when we burn wood or plant products such as ethanol. We also use this energy to fuel our bodies by eating food that comes directly from plants or from animals that got their energy by eating plants. Burning coal and petroleum also releases stored solar energy: These fuels are fossilized plant and animal matter.

## Forms of Energy

Energy can be defined as the capacity to supply heat or do **work**. One type of **work** ( $w$ ) is the process of causing matter to move against an opposing force. For example, we do work when we inflate a bicycle tire—we move matter (the air in the pump) against the opposing force of the air already in the tire.

Like matter, energy comes in different types. One scheme classifies energy into two types; potential and kinetic energy. **Potential energy** is the energy an object has because of its relative position, composition, or condition. **Kinetic energy** is the energy that an object possesses because of its motion. Water at the top of a waterfall or dam has **potential energy** because of its position; when it flows downward through generators, it has **kinetic energy** that can be used to do work and produce electricity in a hydroelectric plant. A battery has **potential energy** because the chemicals within it can produce electricity that can do work.

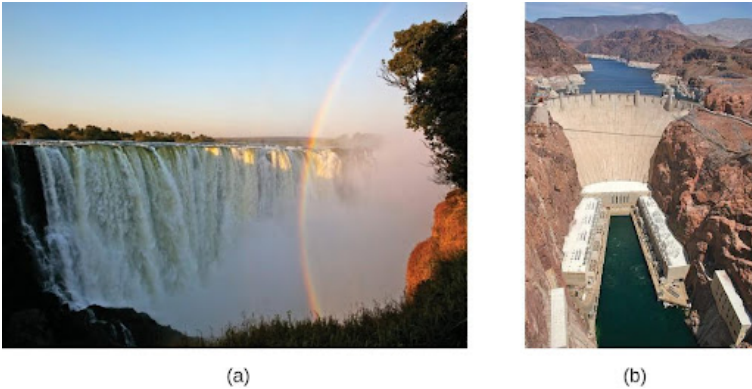


Fig 8.3 (a) Water at a higher elevation, for example, at the top of Victoria Falls, has a higher potential energy than water at a lower elevation. As water falls, some of its potential energy is converted into kinetic energy. (b) If the water flows through generators at the bottom of a dam, such as the Hoover Dam shown here, its kinetic energy is converted into electrical energy.

The energy associated with translational motion (**kinetic energy**) is expressed  $\mathbf{KE} = (1/2) \mathbf{mv}^2$ . Kinetic energy is a form of energy associated with the motion of a particle, single body, or system of objects moving together. The unit of measure for kinetic energy is the **joule**. We are aware that it takes energy to get an object, like a car, up to speed, but it may be a bit surprising that kinetic energy is proportional to velocity squared. This proportionality means, for example, that a car traveling at 100 km/h has four times the kinetic energy it has at 50 km/h, helping to explain why high-speed collisions are so

devastating. We will now consider an example to illustrate the relationship between work and energy.

### Example 8.1

Suppose a 30.0-kg package on a roller belt conveyor system is moving at 0.500 m/s. What is its kinetic energy?

- Strategy: Because the mass ( $m$ ) and speed or velocity ( $v$ ) are given, the kinetic energy can be calculated from its definition as given in the equation  $KE = (1/2) mv^2$ .
- Solution: The kinetic energy is given by  $KE = (1/2) mv^2$ . Entering known values gives  $KE = 0.5 * (30.0\text{kg})*(0.500\text{m/s})^2$  which yields  $KE=3.75\text{kg}\cdot\text{m}^2/\text{s}^2=3.75\text{ J}$ .
- Discussion: Note that the unit of kinetic energy is the joule ( $\text{kg}\cdot\text{m}^2/\text{s}^2$ ). It is interesting that, although this is a fairly massive package, its kinetic energy is not large at this relatively low speed. This fact is consistent with the observation that people can move packages

like this without exhausting themselves.

In Figure 8.4, if the object is lifted straight up we refer to the energy the object has due to gravity as **potential energy** (PE) gained by the object, recognizing that this is energy stored in the gravitational field of Earth.

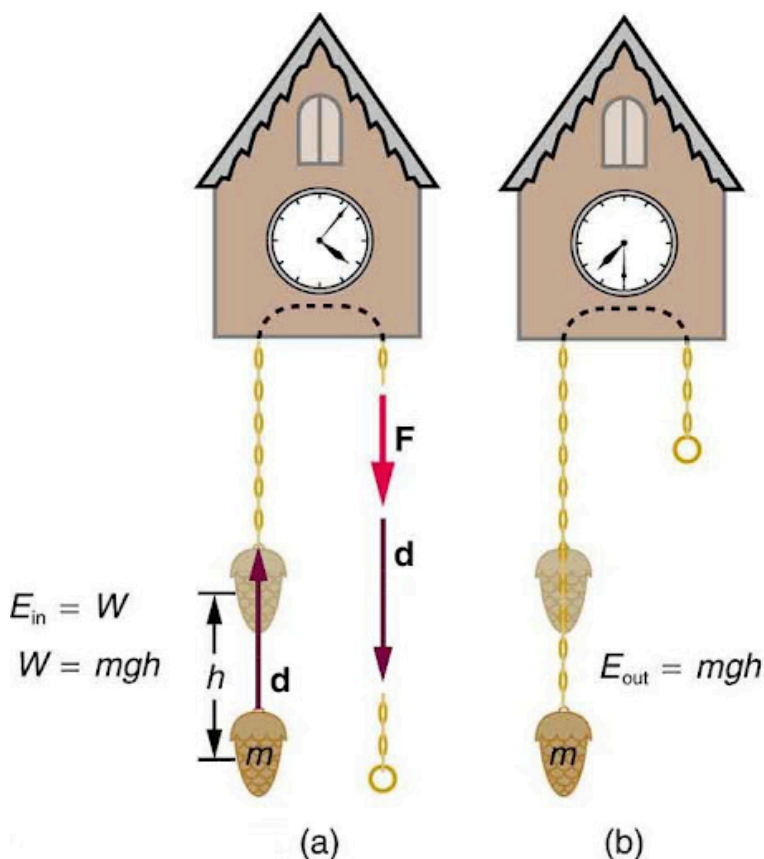


Fig 8.4 (a) The work done to lift the weight is stored in the mass-Earth system as gravitational potential energy. (b) As the weight moves downward, this gravitational potential energy is transferred to the cuckoo clock.

Why do we use the word “system” when discussing potential energy? Potential energy is a property of a system rather than of a single object—due to its physical position. An object’s gravitational potential energy is due to its position relative to

the surroundings within the Earth-object system. The force applied to the object is an external force, from outside the system. When it does positive work it increases the gravitational potential energy of the system. Because gravitational potential energy depends on relative position, we need a reference level at which to set the potential energy equal to 0. We usually choose this point to be Earth's surface, but this point is arbitrary; what is important is the *difference* in gravitational potential energy, because this difference is what relates to the work done. The difference in gravitational potential energy of an object (in the Earth-object system) between two rungs of a ladder will be the same for the first two rungs as for the last two rungs.

## Converting Between Potential Energy and Kinetic Energy

Gravitational potential energy may be converted to other forms of energy, such as kinetic energy. If we release the mass, gravitational force will do an amount of work equal to the  $mgh$  on it, thereby increasing its kinetic energy by that same amount (ie, converting E to KE).

More precisely, we define the change in gravitational potential energy  $\Delta PE$  to be  $\Delta PE = mgh$ . You may also see the change in gravitational potential energy written as  $\Delta U_g = mgy$ . These two equations both represent the potential energy due to gravity and can be used interchangeably. In these equations

m represents the mass of the object and g represents the gravitational acceleration ( $9.81\text{m/s}^2$ ). Both y and h represent the change in height, but y more specifically refers to the change in distance in the y direction.

For example, if a 0.500-kg mass hung from a cuckoo clock is raised 1.00 m, then its change in gravitational potential energy is  $\Delta\text{PE} = mgh = (0.500\text{ kg})(9.80\text{m/s}^2)(1.00\text{ m}) = 4.90\text{ kg}\cdot\text{m}^2/\text{s}^2 = 4.90\text{ J}$ .

## Mechanical Energy

Mechanical energy is kinetic energy and any potential energy related to the potential for an object to move. We deal with all other forms of energy by lumping them into a single group called other energy (OE). Then we can state the conservation of energy in equation form as

$$\mathbf{KE_i + PE_i + OE_i = KE_f + PE_f + OE_f}$$

where subscript i stands for initial and subscript f stands for final.

All types of energy and work can be included in this very general statement of conservation of energy. Kinetic energy is KE, mechanical potential energy is represented by PE, and all other energies are included as OE. This equation applies to all previous examples; in those situations OE was constant, and so it subtracted out and was not directly considered.



## Calculating Potential and Kinetic Energy

To calculate the **potential energy** of an object, we use the following formula:

Potential Energy (PE) = mass times the acceleration due to gravity times height

$$\text{PE} = mgh = \text{N} \cdot \text{h} \quad (g = 9.8 \text{ m/s}^2)$$

$$1 \text{ Newton (N)} = 1\text{kg} \cdot 1\text{m/s}^2 \text{ or } 1\text{kg} \cdot \text{m/s}^2$$

Where:

PE is potential energy in N

m is the mass in Kg

g is the acceleration due to gravity in  $\text{m/s}^2$

h is the height of the object in m

To calculate the **kinetic energy** in a system we use this formula:

**Kinetic Energy (KE) =  $\frac{1}{2}$  mass times velocity squared**

$$\text{KE} = \frac{1}{2} mv^2$$

Where:

KE is kinetic energy in N

m is the mass in Kg

$v^2$  is the velocity of the object in  $\text{m/s squared}$

### Link to Learning:

Use this simulation to see how a skate park applies the conservation of energy and its various properties (website opens in a new window): [PhET Energy Skate Park](#).

## Conservation of Energy & Matter

When one substance is converted into another, there is always an associated conversion of one form of energy into another. **Heat** is usually released or absorbed, but sometimes the conversion involves light, **electrical energy**, or some other form of energy. For example, **chemical energy** (a type of **potential energy**) is stored in the molecules that compose gasoline. When gasoline is combusted within the cylinders of a car's engine, the rapidly expanding gaseous products of this chemical reaction generate **mechanical energy** (a type of **kinetic energy**) when they move the cylinders' pistons. Other forms of energy include **sound energy**, **nuclear energy**, **magnetic energy**, **gravitational energy**, **radiant energy**, and **thermal energy**.

Energy, as we have noted, is conserved, making it one of the most important physical quantities in nature. The **law of conservation of energy** can be stated as follows:

*Total energy is constant in any process. It may change in form or be transferred from one system to another, but the total remains the same.*

## Making Connections: Usefulness of the Energy Conservation Principle

The fact that **energy** is conserved and has many forms makes it very important. You will find that energy is discussed in many contexts, because it is involved in all processes. It will also become apparent that many situations are best understood in terms of energy and that problems are often most easily conceptualized and solved by considering energy.

When does OE play a role? One example occurs when a person eats. Food is oxidized with the release of carbon dioxide, water, and energy. Some of this chemical energy is converted to kinetic energy when the person moves, to potential energy when the person changes altitude, and to thermal energy OE.

## Problem-Solving Strategies for Energy

You will find the following problem-solving strategies useful whenever you deal with energy. The strategies help in organizing and reinforcing energy concepts. In fact, they are used in the examples presented in this chapter. The familiar general problem-solving strategies presented earlier—involving identifying physical principles, knowns,

unknowns, checking units, and so on—continue to be relevant here.

**Step 1.** Determine the system of interest, identify what information is given and what quantity is to be calculated. A sketch will help.

**Step 2.** Examine all the forces involved and determine whether you know or are given the potential energy from the work done by the forces. Then use step 3 or step 4.

**Step 3.** If you know that you can apply conservation of mechanical energy simply in terms of potential and kinetic energy, the equation expressing conservation of energy is  $KE_i + PE_i = KE_f + PE_f$ .

**Step 4.** You have already identified the types of energy involved (in step 2). Before solving for the unknown, *eliminate terms wherever possible* to simplify the algebra. For example, choose  $h=0$  at either the initial or the final point, so the gravitational potential energy ( $PE_g$ ) is zero there. Then solve for the unknown in the customary manner.

**Step 5.** Check the answer to see if it is reasonable. Once you have solved a problem, reexamine the forms of work and energy to see if you have set up the conservation of energy equation correctly. For example, work done against friction should be negative, potential energy at the bottom of a hill should be less than that at the top, and so on. Also check to see that the numerical value obtained is reasonable. For example, the final speed of a skateboarder who coasts down a 3-m-high ramp could reasonably be 20 km/h, but *not* 80 km/h.

## Transformation of Energy

The transformation of energy from one form into others is happening all the time. The chemical energy in food is converted into thermal energy through metabolism; light energy is converted into chemical energy through photosynthesis. In a larger example, the chemical energy contained in coal is converted into thermal energy as it burns to turn water into steam in a boiler. This thermal energy in the steam in turn is converted to mechanical energy as it spins a turbine, which is connected to a generator to produce electrical energy. (In all of these examples, not all of the initial energy is converted into the forms mentioned.)

Another example of energy conversion occurs in a solar cell. Sunlight impinging on a solar cell (see Figure 8.5) produces electricity, which in turn can be used to run an electric motor. Energy is converted from the primary source of solar energy into electrical energy and then into mechanical energy.



Fig 8.5 Solar energy is converted into electrical energy by solar cells, which is used to run a motor in this solar- power aircraft.

# Conservation of Mass

The **mass** of an object is a measure of the amount of matter in it. One way to measure an object's mass is to measure the force it takes to accelerate the object. It takes much more force to accelerate a car than a bicycle because the car has much more mass. A more common way to determine the mass of an object is to use a balance to compare its mass with a standard mass.

Although **weight** is related to mass, it is not the same thing. Weight refers to the force that gravity exerts on an object. This

force is directly proportional to the mass of the object. The weight of an object changes as the force of gravity changes, but its mass does not. An astronaut's mass does not change just because she goes to the moon. But her weight on the moon is only one-sixth her earth-bound weight because the moon's gravity is only one-sixth that of the earth's. She may feel "weightless" during her trip when she experiences negligible external forces (gravitational or any other), although she is, of course, never "massless."

The **law of conservation of matter** summarizes many scientific observations about matter. It states:

There is no detectable change in the total quantity of matter present when matter converts from one type to another (a chemical change) or changes among solid, liquid, or gaseous states (a physical change).

Brewing beer and the operation of batteries provide examples of the conservation of matter. During the brewing of beer, the ingredients (water, yeast, grains, malt, hops, and sugar) are converted into beer (water, alcohol, carbonation, and flavoring substances) with no actual loss of substance. This is most clearly seen during the bottling process, when glucose turns into ethanol and carbon dioxide, and the total mass of the substances does not change. This can also be seen in a lead-acid car battery: The original substances (lead, lead oxide, and sulfuric acid), which are capable of producing electricity, are changed into other substances (lead sulfate and

water) that do not produce electricity, with no change in the actual amount of matter.

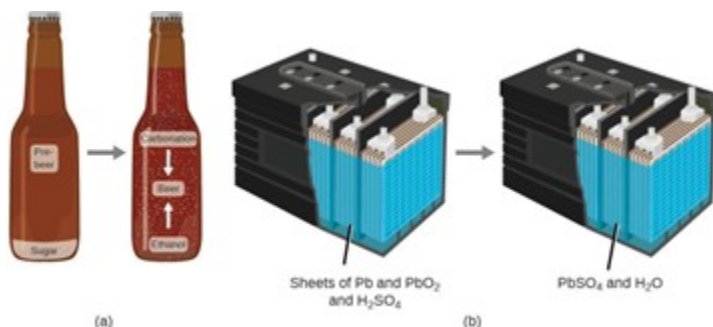


Fig. 8.6 (a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbon dioxide. (b) The mass of the lead, lead oxide, and sulfuric acid consumed by the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

Although this conservation law holds true for all conversions of matter, convincing examples are few and far between because, outside of the controlled conditions in a laboratory, we seldom collect all of the material that is produced during a particular conversion. For example, when you eat, digest, and assimilate food, all of the matter in the original food is preserved. But because some of the matter is incorporated into your body, and much is excreted as various types of waste, it is challenging to verify by measurement.



## Check Your Understanding



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9.

# ELECTRICITY AND ELECTRICAL CIRCUITS

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## Learning Objectives

After successful completion of this section, you will be able to demonstrate competency in the following areas.

- Define: static electricity (CLO2)(CLO3)
- Determine the type of electric charge present (CLO2)(CLO3)(CLO4)
- State units of charge and electric force in SI units. (CLO2)(CLO5)
- Explain various electrostatic interactions (CLO2)(CLO3)
- Define: electrical force, and electric field

(CLO2)

- State Coulomb's Law (CLO1)(CLO2)
- Discuss the relationships between the variables in Coulomb's Law equation. (CLO2)(CLO3)(CLO5)
- Describe interactions between charges and electric fields. (CLO2)(CLO3)
- Use Coulomb's Law to calculate electric force (CLO5)
- Compare the Coulomb Force to the Gravitational Force (CLO2)(CLO3)(CLO5)
- Define electric current, electric resistance, and voltage. (CLO2)(CLO3)
- State Ohm's Law. (CLO1)
- Use the correct units of current, voltage, and resistance in the metric systems. (CLO5)
- Explain the relationships between the variables in the Ohm's Law formula (CLO2)(CLO3)
- Use Ohm's Law to solve simple circuit problems (CLO5)
- Determine the electric resistance, voltage, electric current, and power in a circuit (CLO2)(CLO3)(CLO5)
- Discuss the relationship between energy,

power, voltage, current, and resistance in a circuit. (CLO2)(CLO3)

- Describe and differentiate between parallel and series electrical circuits. (CLO2)(CLO3)
- Determine the equivalent resistance for series and parallel combinations of resistors. (CLO2)(CLO5)



Fig. 9.1 Static electricity from this plastic slide causes the child's hair to stand on end. The sliding motion stripped electrons away from the child's body, leaving an excess of positive charges, which repel each other along each strand of hair.

# Introduction to Electricity

The image of American politician and scientist Benjamin Franklin (1706–1790) flying a kite in a thunderstorm is familiar to many schoolchildren. (See Figure 9.2) In this experiment, Franklin demonstrated a connection between lightning and **static electricity**. Sparks were drawn from a key hung on a kite string during an electrical storm. These sparks were like those produced by static electricity, such as the spark that jumps from your finger to a metal doorknob after you walk across a wool carpet. What Franklin demonstrated in his dangerous experiment was a connection between phenomena on two different scales: one the grand power of an electrical storm, the other an effect of more human proportions. Connections like this one reveal the underlying unity of the laws of nature, an aspect we humans find particularly appealing.

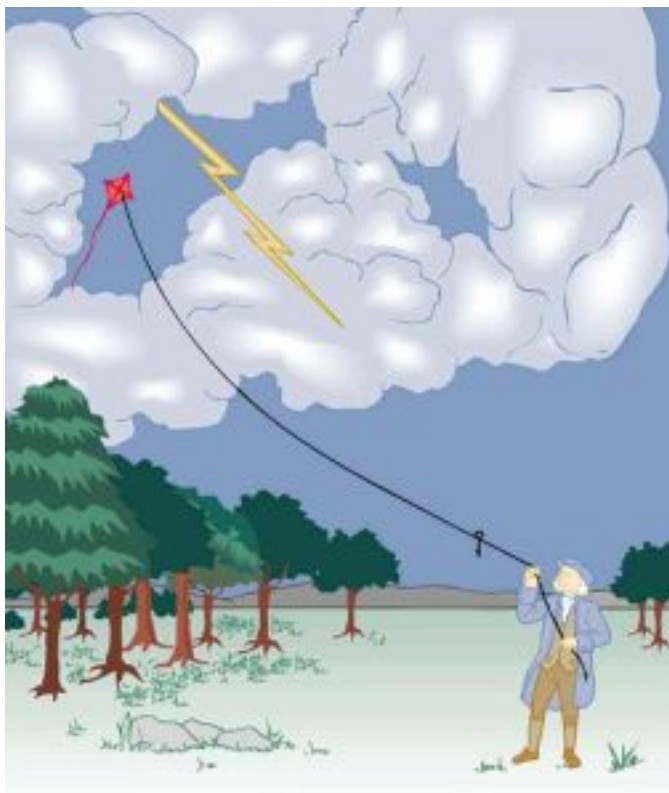


Fig. 9.2 When Benjamin Franklin demonstrated that lightning was related to static electricity, he made a connection that is now part of the evidence that all directly experienced forces except the gravitational force are manifestations of the electromagnetic force.

Much has been written about Franklin. His experiments were only part of the life of a man who was a scientist, inventor, revolutionary, statesman, and writer. Franklin's experiments were not performed in isolation, nor were they the only ones to reveal connections. For example, the Italian scientist Luigi



Galvani (1737–1798) performed a series of experiments in which static electricity was used to stimulate contractions of leg muscles of dead frogs, an effect already known in humans subjected to static discharges. But Galvani also found that if he joined two metal wires (say copper and zinc) end to end and touched the other ends to muscles, he produced the same effect in frogs as static discharge. Alessandro Volta (1745–1827), partly inspired by Galvani's work, experimented with various combinations of metals and developed the battery. During the same era, other scientists made progress in discovering fundamental connections. The **periodic table** was developed as the systematic **properties** of the elements were discovered. This influenced the development and refinement of the concept of atoms as the basis of **matter**. Such submicroscopic descriptions of matter also help explain a great deal more. Atomic and molecular interactions, such as the forces of friction, cohesion, and adhesion, are now known to be manifestations of the **electromagnetic force**. Static electricity is just one aspect of the electromagnetic force, which also includes moving electricity and **magnetism**. All the macroscopic forces that we experience directly, such as the sensations of touch and the tension in a rope, are due to the **electromagnetic force**, one of the four fundamental forces in nature. The gravitational force, another fundamental force, is actually sensed through the electromagnetic interaction of molecules, such as between those in our feet and those on the top of a bathroom scale.

(The other two fundamental forces, the strong nuclear force and the weak nuclear force, cannot be sensed on the human scale.)

## Static Electricity and Conservation of Charge



Fig. 9.3 Borneo amber was mined in Sabah, Malaysia, from shale-sandstone-mudstone veins. When a piece of amber is rubbed with a piece of silk, the amber gains more electrons, giving it a net negative charge. At the same time, the silk, having lost electrons, becomes positively charged.

What makes plastic wrap cling? **Static electricity**. Not only are applications of static electricity common these days, its

existence has been known since ancient times. The first record of its effects dates to ancient Greeks who noted more than 500 years B.C. that polishing amber temporarily enabled it to attract bits of straw (see Figure 9.3). The very word *electric* derives from the Greek word for amber (*electron*).

Many of the characteristics of static electricity can be explored by rubbing things together. Rubbing creates the spark you get from walking across a wool carpet, for example. Static cling generated in a clothes dryer and the attraction of straw to recently polished amber also result from rubbing. Similarly, lightning results from air movements under certain weather conditions. You can also rub a balloon on your hair, and the static electricity created can then make the balloon cling to a wall. We also have to be cautious of static electricity, especially in dry climates. When we pump gasoline, we are warned to discharge ourselves (after sliding across the seat) on a metal surface before grabbing the gas nozzle. Attendants in hospital operating rooms must wear booties with a conductive strip of aluminum foil on the bottoms to avoid creating sparks which may ignite flammable anesthesia gases combined with the oxygen being used.

Some of the most basic characteristics of static electricity include:

- The effects of static electricity are explained by a physical quantity not previously introduced, called electric charge.

- There are only two types of charge, one called positive and the other called negative.
- Like charges repel, whereas unlike charges attract.
- The force between charges decreases with distance.

How do we know there are two types of **electric charge**? When various materials are rubbed together in controlled ways, certain combinations of materials always produce one type of charge on one material and the opposite type on the other. By convention, we call one type of charge “positive”, and the other type “negative.” For example, when glass is rubbed with silk, the glass becomes positively charged and the silk negatively charged. Since the glass and silk have opposite charges, they attract one another like clothes that have rubbed together in a dryer. Two glass rods rubbed with silk in this manner will repel one another, since each rod has positive charge on it. Similarly, two silk cloths so rubbed will repel, since both cloths have negative charge. Figure 9.4 shows how these simple materials can be used to explore the nature of the force between charges.

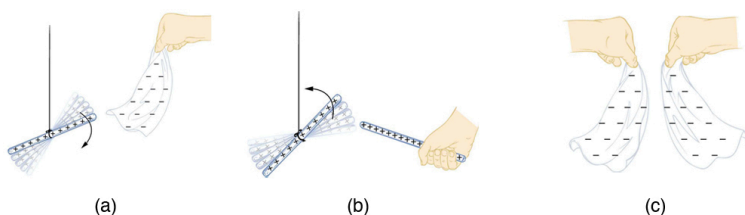


Fig 9.4 A glass rod becomes positively charged when rubbed with silk, while the silk becomes negatively charged. (a) The glass rod is attracted to the silk because their charges are opposite. (b) Two similarly charged glass rods repel. (c) Two similarly charged silk cloths repel.

More sophisticated questions arise. Where do these charges come from? Can you create or destroy charge? Is there a smallest unit of charge? Exactly how does the force depend on the amount of charge and the distance between charges? Such questions obviously occurred to Benjamin Franklin and other early researchers, and they interest us even today.

## Charge Carried By Electrons and Protons

Franklin wrote in his letters and books that he could see the effects of electric charge but did not understand what caused the phenomenon. Today we have the advantage of knowing that normal matter is made of atoms, and that atoms contain positive and negative charges, usually in equal amounts.

Figure 9.5 shows a simple model of an atom with negative **electrons** orbiting its positive nucleus. The nucleus

is positive due to the presence of positively charged **protons**. Nearly all charge in nature is due to electrons and protons, which are two of the three building blocks of most matter. (The third is the **neutron**, which is neutral, carrying no charge.) Other charge-carrying particles are observed in cosmic rays and nuclear decay, and are created in particle accelerators. All but the electron and proton survive only a short time and are quite rare by comparison.

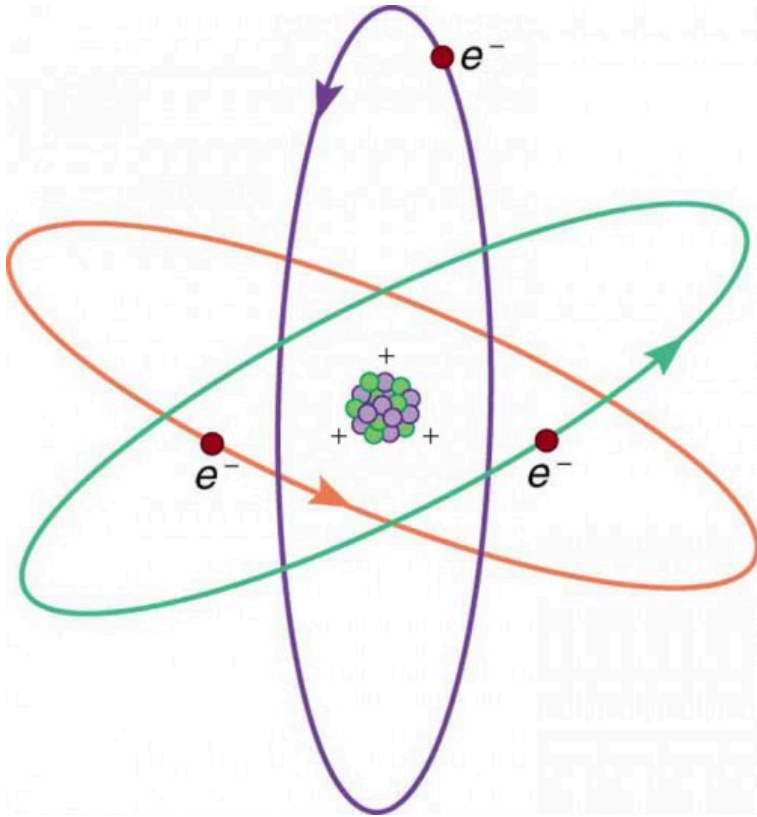


Fig 9.5 This simplified (and not to scale) view of an atom is called the planetary model of the atom. Negative electrons orbit a much heavier positive nucleus, as the planets orbit the much heavier sun. There the similarity ends, because forces in the atom are electromagnetic, whereas those in the planetary system are gravitational. Normal macroscopic amounts of matter contain immense numbers of atoms and molecules and, hence, even greater numbers of individual negative and positive charges.

The charges of electrons and protons are identical in magnitude but opposite in sign. Furthermore, all charged

objects in nature are integral multiples of this basic quantity of charge, meaning that all charges are made of combinations of a basic unit of charge. Usually, charges are formed by combinations of electrons and protons. The magnitude of this basic charge is

$$|q_e| = 1.60 \times 10^{-19} \text{C}.$$

The symbol  $q$  is commonly used for charge and the subscript  $e$  indicates the charge of a single electron (or proton).

The SI unit of charge is the coulomb (C). The number of protons needed to make a charge of 1.00 C is

$$1.00 \text{ C} \times (1 \text{ proton} / 1.60 \times 10^{-19} \text{C}) = 6.25 \times 10^{18} \text{ protons}.$$

Similarly,  $6.25 \times 10^{18}$  electrons have a combined charge of  $-1.00$  coulomb. Just as there is a smallest bit of an element (an **atom**), there is a smallest bit of charge. There is no directly observed charge smaller than  $|q_e|$ , and all observed charges are integral multiples of  $|q_e|$ .

Figure 9.6 shows a person touching a Van de Graaff generator and receiving excess positive charge. The expanded view of a hair shows the existence of both types of charges but an excess of positive. The repulsion of these positive like charges causes the strands of hair to repel other strands of hair and to stand up. The further blowup shows an artist's conception of an electron and a proton perhaps found in an atom in a strand of hair.



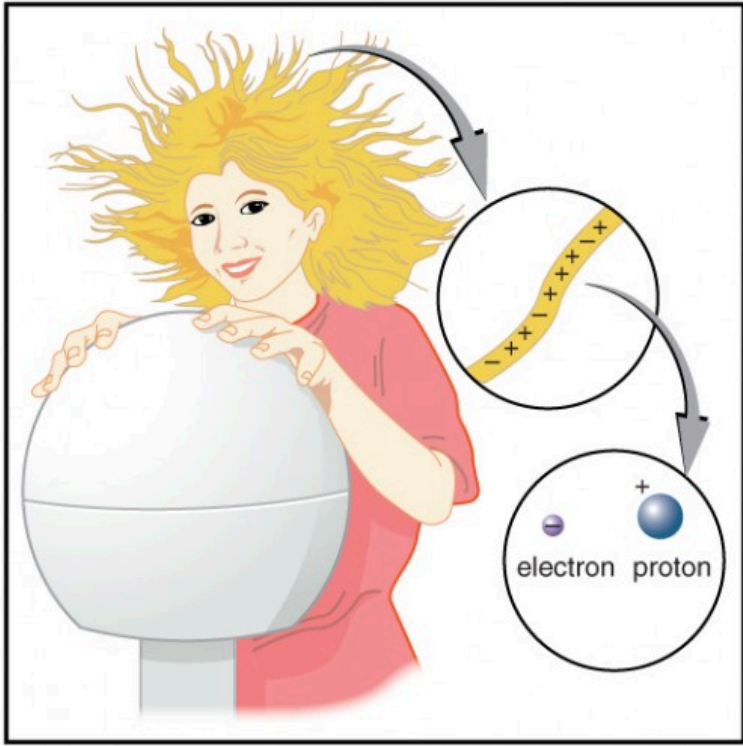


Fig. 9.6 When this person touches a Van de Graaff generator, some electrons are attracted to the generator, resulting in an excess of positive charge, causing her hair to stand on end. The charges in one hair are shown. An artist's conception of an electron and a proton illustrate the particles carrying the negative and positive charges. We cannot really see these particles with visible light because they are so small (the electron seems to be an infinitesimal point), but we know a great deal about their measurable properties, such as the charges they carry.

The electron seems to have no substructure; in contrast, when the substructure of protons is explored by scattering extremely

energetic electrons from them, it appears that there are point-like particles inside the proton. These sub-particles, named quarks, have never been directly observed, but they are believed to carry fractional charges as seen in Figure 9.7. Charges on electrons and protons and all other directly observable particles are unitary, but these quark substructures carry charges of either  $-1/3$  or  $+2/3$ . There are continuing attempts to observe fractional charge directly and to learn of the properties of quarks, which are perhaps the ultimate substructure of matter.

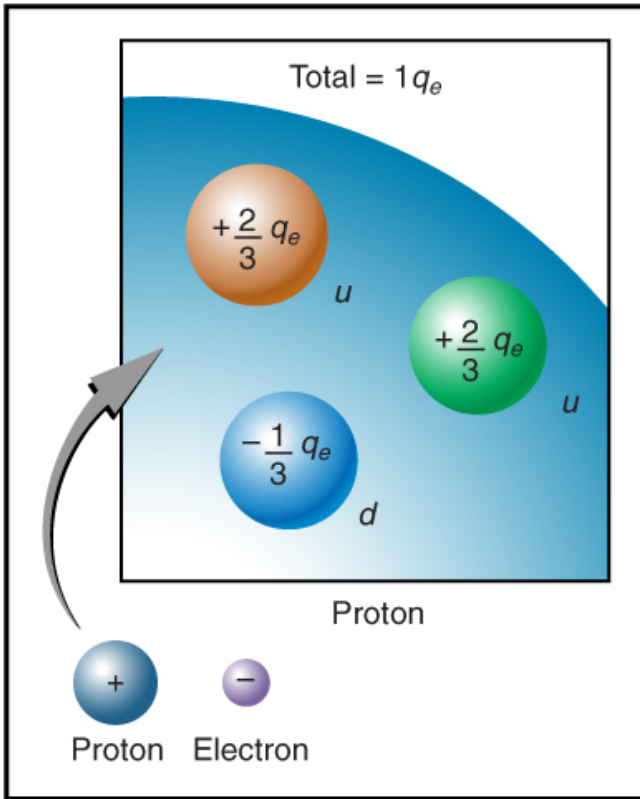


Fig. 9.7 Artist's conception of fractional quark charges inside a proton. A group of three quark charges add up to the single positive charge on the proton:  $-\frac{1}{3}q_e + \frac{2}{3}q_e + \frac{2}{3}q_e = +1q_e$ .

## Separation of Charges in Atoms

Charges in **atoms** and **molecules** can be separated—for example, by rubbing materials together. Some atoms and molecules have a greater affinity for electrons than others and

will become negatively charged by close contact in rubbing, leaving the other material positively charged. (See Figure 9.8.) Positive charge can similarly be induced by rubbing. Methods other than rubbing can also separate charges. Batteries, for example, use combinations of substances that interact in such a way as to separate charges. Chemical interactions may transfer negative charge from one substance to the other, making one battery terminal negative and leaving the first one positive.

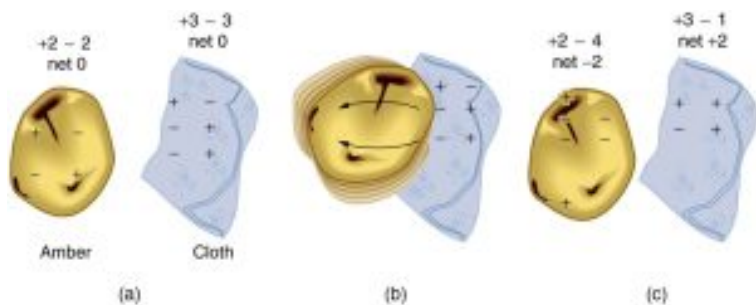


Fig 9.8 When materials are rubbed together, charges can be separated, particularly if one material has a greater affinity for electrons than another. (a) Both the amber and cloth are originally neutral, with equal positive and negative charges. Only a tiny fraction of the charges are involved, and only a few of them are shown here. (b) When rubbed together, some negative charge is transferred to the amber, leaving the cloth with a net positive charge. (c) When separated, the amber and cloth now have net charges, but the absolute value of the net positive and negative charges will be equal.

No charge is actually created or destroyed when charges are

separated as we have been discussing. Rather, existing charges are moved about. In fact, in all situations the total amount of charge is always constant. This universally obeyed law of nature is called **the law of conservation of charge**, which states that total charge is constant in any process.

## Making Connections: Conservation Laws

Only a limited number of physical quantities are universally conserved. Charge is one—energy, momentum, and angular momentum are others. Because they are conserved, these physical quantities are used to explain more phenomena and form more connections than other, less basic quantities. We find that conserved quantities give us great insight into the rules followed by nature and hints to the organization of nature. Discoveries of conservation laws have led to further discoveries, such as the weak nuclear force and the quark substructure of protons and other particles.

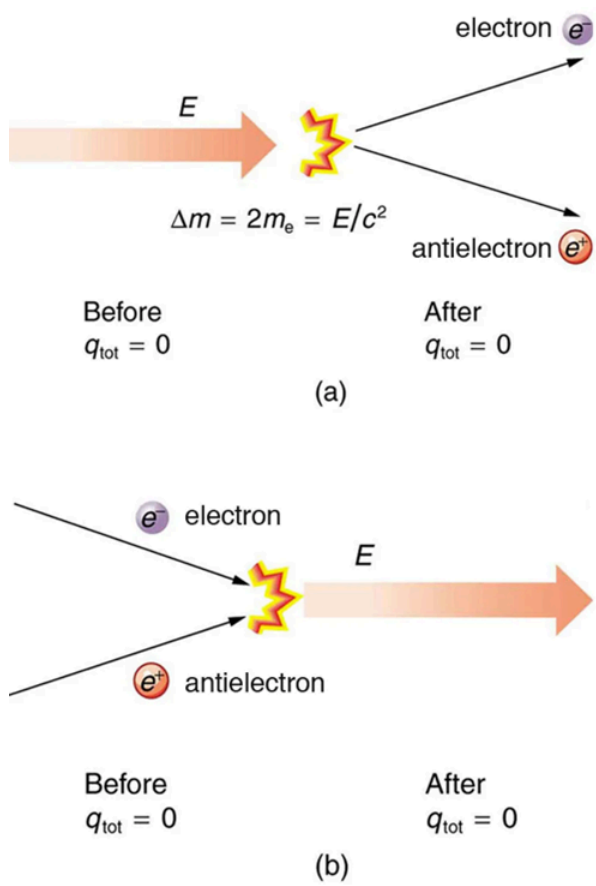


Fig. 9.9 9 (a) When enough energy is present, it can be converted into matter. Here the matter created is an electron–antielectron pair. ( $m_e$  is the electron’s mass.) The total charge before and after this event is zero. (b) When matter and antimatter collide, they annihilate each other; the total charge is conserved at zero before and after the annihilation.

The law of conservation of charge is absolute—it has never

been observed to be violated. Charge, then, is a special physical quantity, joining a very short list of other quantities in nature that are always conserved. Other conserved quantities include energy, momentum, and angular momentum.

## Coulomb's Law

Through the work of scientists in the late 18th century, the main features of the **electrostatic force**—the existence of two types of charge, the observation that like charges repel, unlike charges attract, and the decrease of force with distance—were eventually refined and expressed as a mathematical formula. The mathematical formula for the electrostatic force is called **Coulomb's law** after the French physicist Charles Coulomb (1736–1806), who performed experiments and first proposed a formula to calculate it.

Although the law determining the magnitude of the Coulomb electric force has the same form as the law of gravity, the electric constant is 20 orders of magnitude greater than the gravitational constant. That is why electricity normally dominates gravity at the atomic and molecular level. Since there is only one type of mass but two opposite types of electric charge, gravity will dominate in large bodies unless there is a separation of charge.

Coulomb's law states that any two charged particles ( $q_1, q_2$ ) — with charge measured in units of Coulombs — at a distance

$r$  from each other will experience a force of repulsion or attraction along the line joining them equal to:

$$F = k \frac{q_1 q_2}{r^2}$$

Although the formula for **Coulomb's law** is simple, it was no mean task to prove it. The experiments Coulomb did, with the primitive equipment then available, were difficult. Modern experiments have verified Coulomb's law to great precision. For example, it has been shown that the force is inversely proportional to distance between two objects squared ( $F \propto 1/r^2$ ) to an accuracy of 1 part in  $10^{16}$ . No exceptions have ever been found, even at the small distances within the atom.

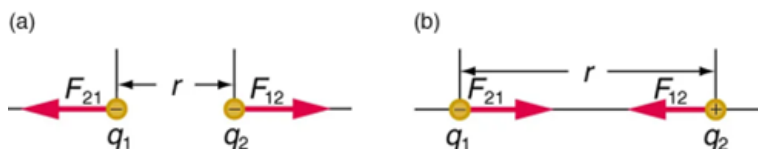


Fig 9.10 The magnitude of the electrostatic force  $F$  between point charges  $q_1$  and  $q_2$  separated by a distance  $r$  is given by Coulomb's law. Note that Newton's third law (every force exerted creates an equal and opposite force) applies as usual—the force on  $q_1$  is equal in magnitude and opposite in direction to the force it exerts on  $q_2$ . (a) Like charges. (b) Unlike charges.

## Electric Field and Electric Forces

Contact forces, such as between a baseball and a bat, are



explained on the small scale by the interaction of the charges in atoms and molecules in close proximity. They interact through forces that include the **Coulomb force**. Action at a distance is a force between objects that are not close enough for their atoms to “touch.” That is, they are separated by more than a few atomic diameters.

For example, a charged rubber comb attracts neutral bits of paper from a distance via the Coulomb force. It is very useful to think of an object being surrounded in space by a **force field**. The force field carries the force to another object (called a test object) some distance away.

## Electric Current and Electric Circuitry

**Electric current** is defined to be the rate at which charge flows. **Voltage** is the potential difference in energy between two points. It is this potential difference that drives electric current. We can think of various devices—such as batteries, generators, wall outlets, and so on—which are necessary to maintain a current. All such devices create a potential difference and are loosely referred to as voltage sources. When a voltage source is connected to a conductor, it applies a potential difference  $V$  that creates an **electric field**. The electric field in turn exerts force on charges, causing current.

## Ohm's Law

The current that flows through most substances is directly proportional to the voltage  $V$  applied to it. The German physicist Georg Simon Ohm (1787–1854) was the first to demonstrate experimentally that the current in a metal wire is *directly proportional to the voltage applied*:

$$I \propto V$$

This important relationship is known as **Ohm's law**. It can be viewed as a cause-and-effect relationship, with voltage the cause and current the effect.

## Resistance and Simple Circuits

If voltage drives current, what impedes it? The electric property that impedes current (crudely similar to friction and air resistance) is called **resistance** ( $R$ ). Collisions of moving charges with atoms and molecules in a substance transfer energy to the substance and limit current. Resistance is defined as inversely proportional to current, or

$$I \propto (1/R)$$

Thus, for example, current is cut in half if resistance doubles. Combining the relationships of current to voltage and current to resistance gives

$$I = (V/R)$$

This relationship is also called **Ohm's law**. Ohm's law in this form really defines resistance for certain materials. Ohm's

law (like Hooke's law) is not universally valid. The many substances for which Ohm's law holds are called ohmic. These include good conductors like copper and aluminum, and some poor conductors under certain circumstances. Ohmic materials have a resistance  $R$  that is independent of voltage  $V$  and current  $I$ . An object that has simple resistance is called a **resistor**, even if its resistance is small. The unit for resistance is an ohm and is given the symbol  $\Omega$  (upper case Greek omega).

Figure 9.11 shows the schematic for a simple **circuit**. A simple circuit has a single voltage source and a single resistor. The wires connecting the voltage source to the resistor can be assumed to have negligible resistance, or their resistance can be included in  $R$ .

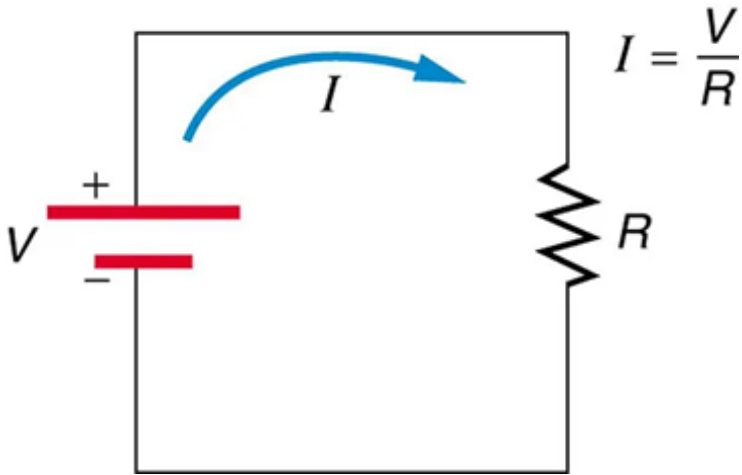


Fig. 9.11 A simple electric circuit in which a closed path for current to flow is supplied by conductors (usually metal wires) connecting a load to the terminals of a battery, represented by the red parallel lines. The zigzag symbol represents the single resistor and includes any resistance in the connections to the voltage source.

## Resistors in Parallel and In Series

Most circuits have more than one component, called a **resistor** that limits the flow of charge in the circuit. A measure of this limit on charge flow is called **resistance**. The simplest combinations of **resistors** are the series and parallel connections illustrated in Figure 9.12. The total resistance of a combination of resistors depends on both their individual values and how they are connected.

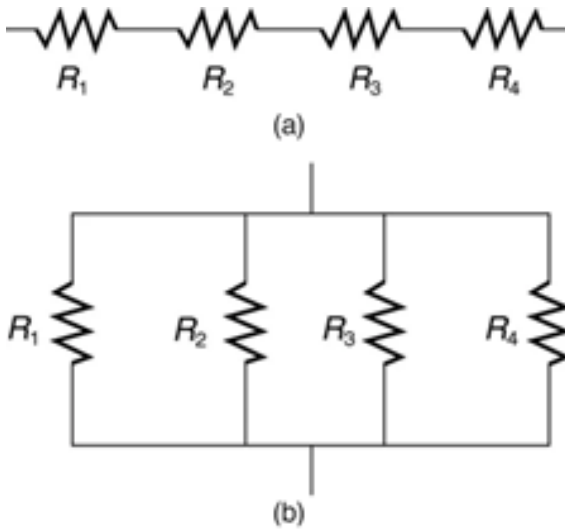


Fig. 9.12 (a) A series connection of resistors.  
(b) A parallel connection of resistors.

When are resistors in series? Resistors are in series whenever the flow of charge, called the current, must flow through devices sequentially. For example, if current flows through a person holding a screwdriver and into the Earth, then  $R_1$  in Figure 9.13(a) could be the resistance of the screwdriver's shaft,  $R_2$  the resistance of its handle,  $R_3$  the person's body resistance, and  $R_4$  the resistance of her shoes.

Figure 9.13 shows resistors in series connected to a voltage source. It seems reasonable that the total resistance is the sum of the individual resistances, considering that the current has to pass through each resistor in sequence. (This fact would be an advantage to a person wishing to avoid an electrical shock, who could reduce the current by wearing high-resistance

rubber-soled shoes. It could be a disadvantage if one of the resistances were a faulty high-resistance cord to an appliance that would reduce the operating current.)

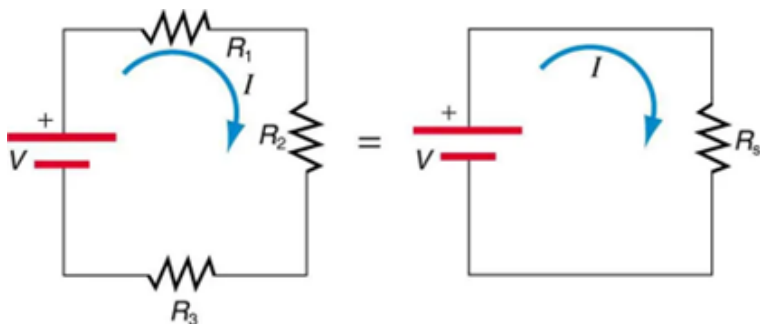


Fig. 9.13 Three resistors connected in series to a battery (left) and the equivalent single or series resistance (right).

Figure 9.14 shows resistors in parallel, wired to a voltage source. Resistors are in parallel when each resistor is connected directly to the voltage source by connecting wires having negligible resistance. Each resistor thus has the full voltage of the source applied to it.

Each resistor draws the same current it would if it alone were connected to the voltage source (provided the voltage source is not overloaded). For example, an automobile's headlights, radio, and so on, are wired in parallel, so that they utilize the full voltage of the source and can operate completely independently. The same is true in your house, or any building. (See Figure 9.14b.)

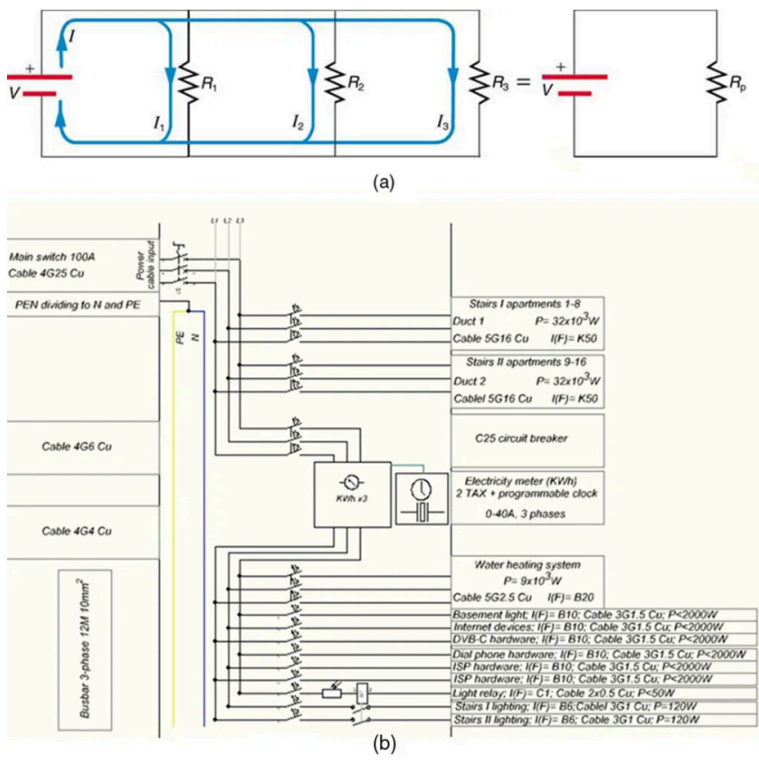


Fig. 9.14 (a) Three resistors connected in parallel to a battery and the equivalent single or parallel resistance. (b) Electrical power setup in a house.

# Direct Current and Alternating Current

Most of the examples dealt with so far, and particularly those utilizing batteries, have constant voltage sources. Once the current is established, it is thus also a constant. **Direct current** (DC) is the flow of electric charge in only one

direction. It is the steady state of a constant-voltage circuit. Most well-known applications, however, use a time-varying voltage source. **Alternating current** (AC) is the flow of electric charge that periodically reverses direction. If the source varies periodically, particularly sinusoidally, the circuit is known as an alternating current circuit. Examples include the commercial and residential power that serves so many of our needs. Fig 9.15 shows graphs of voltage and current versus time for typical DC and AC power. The AC voltages and frequencies commonly used in homes and businesses vary around the world.



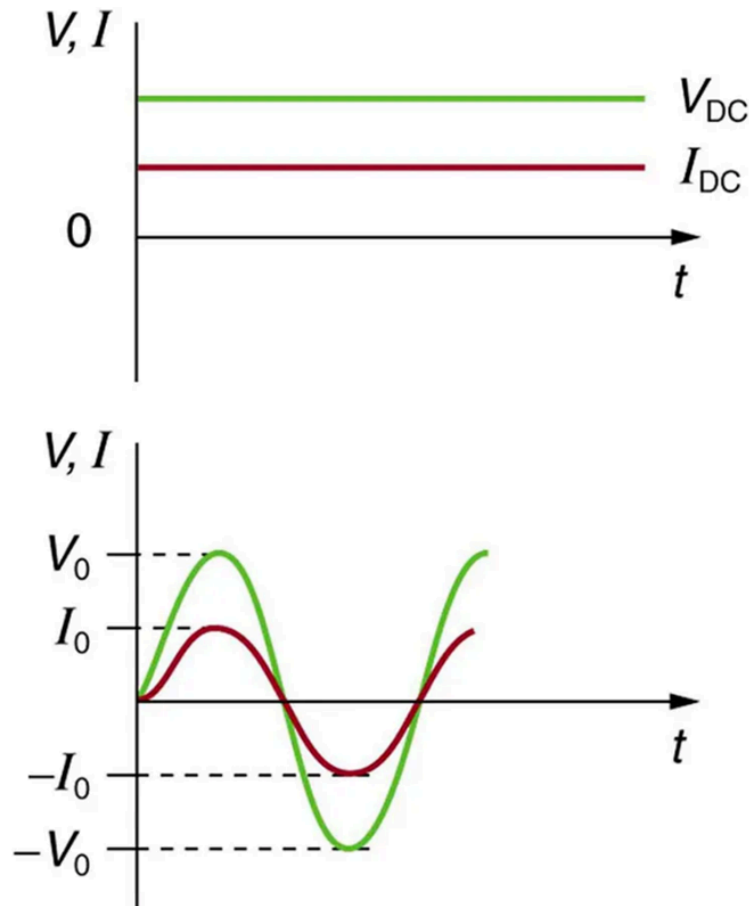


Fig. 9.15 (a) DC voltage and current are constant in time, once the current is established. (b) A graph of voltage and current versus time for 60-Hz AC power. The voltage and current are sinusoidal and are in phase for a simple resistance circuit. The frequencies and peak voltages of AC sources differ greatly.

## Voltmeters and Ammeters

Voltmeters measure **voltage**, whereas ammeters measure **current**. Some of the meters in automobile dashboards, digital cameras, cell phones, and tuner-amplifiers are voltmeters or ammeters. (See Figure 9.16.) The internal construction of the simplest of these meters and how they are connected to the system they monitor give further insight into applications of series and parallel connections.



Fig. 9.16 The fuel and temperature gauges (far right and far left, respectively) in this 1996 Volkswagen are voltmeters that register the voltage output of “sender” units, which are hopefully proportional to the amount of gasoline in the tank and the engine temperature.

Voltmeters are connected in parallel with whatever device's voltage is to be measured. A parallel connection is used because objects in parallel experience the same potential difference. (See Figure 9.17 where the **voltmeter** is represented by the symbol V.)

Ammeters are connected in series with whatever device's current is to be measured. A series connection is used because objects in series have the same current passing through them. (See Figure 9.18, where the **ammeter** is represented by the symbol A.)

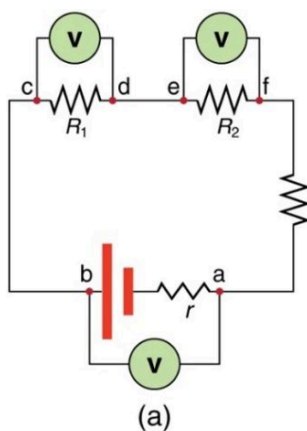


Fig 9.17 (a) To measure potential differences in this series circuit, the voltmeter (V) is placed in parallel with the voltage source or either of the resistors. Note that terminal voltage is measured between points  $a$  and  $b$ . It is not possible to connect the voltmeter directly across the emf without including its internal resistance,  $r$ . (b) A digital voltmeter in use.

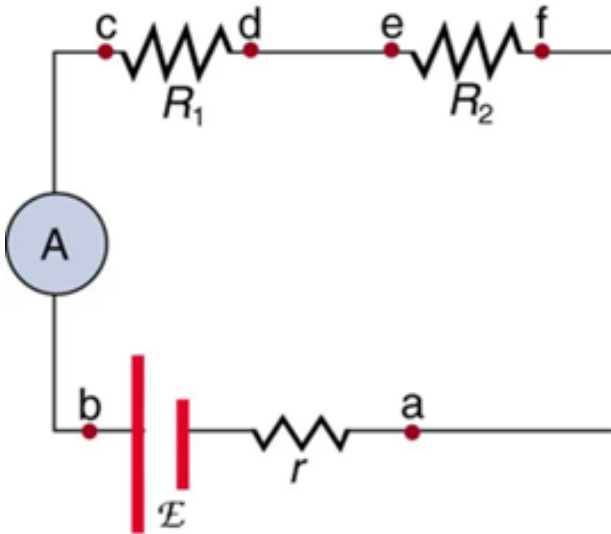


Fig. 9.18 An ammeter (A) is placed in series to measure current. All of the current in this circuit flows through the meter. The ammeter would have the same reading if located between points d and e or between points f and a as it does in the position shown. (Note that the script capital E stands for emf, and r stands for the internal resistance of the source of potential difference.)

## Check Your Understanding



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10.

# MAGNETISM

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## Learning Objectives

After successful completion of this section, you will be able to demonstrate competency in the following areas:

- Define: magnetic force (CLO1)(CLO2)(CLO3)
- Explain the effects of magnetic fields on moving charge (CLO2)(CLO3)
- Describe how electric motors work and the role of magnetism in their function. (CLO2)(CLO3)



Fig. 10.1 The magnificent spectacle of the Aurora Borealis, or northern lights, glows in the northern sky above Bear Lake near Eielson Air Force Base, Alaska. Shaped by the Earth's magnetic field, this light is produced by radiation spewed from solar storms.

## Introduction to Magnetism

One evening, an Alaskan sticks a note to his refrigerator with a small magnet. Through the kitchen window, the Aurora Borealis glows in the night sky. This grand spectacle is shaped by the same force that holds the note to the refrigerator.

People have been aware of magnets and magnetism for thousands of years. The earliest records date to well before the time of Christ, particularly in a region of Asia Minor called Magnesia (the name of this region is the source of words like *magnetic*). Magnetic rocks found in Magnesia, which is

now part of western Turkey, stimulated interest during ancient times. A practical application for magnets was found later, when they were employed as navigational compasses. The use of magnets in compasses resulted not only in improved long-distance sailing, but also in the names of “north” and “south” being given to the two types of magnetic poles.

Today **magnetism** plays many important roles in our lives. Physicists’ understanding of magnetism has enabled the development of technologies that affect our everyday lives. The iPod in your purse or backpack, for example, wouldn’t have been possible without the applications of magnetism and electricity on a small scale.

All electric motors, with uses as diverse as powering refrigerators, starting cars, and moving elevators, contain magnets. Generators, whether producing hydroelectric power or running bicycle lights, use magnetic fields. Recycling facilities employ magnets to separate iron from other refuse. Hundreds of millions of dollars are spent annually on magnetic containment of fusion as a future energy source. Magnetic resonance imaging (MRI) has become an important diagnostic tool in the field of medicine, and the use of magnetism to explore brain activity is a subject of contemporary research and development. The list of applications also includes computer hard drives, tape recording, detection of inhaled asbestos, and levitation of high-speed trains. Magnetism is used to explain atomic energy

levels, cosmic rays, and charged particles trapped in the Van Allen belts. Once again, we will find all these disparate phenomena are linked by a small number of underlying physical principles.

## Magnets

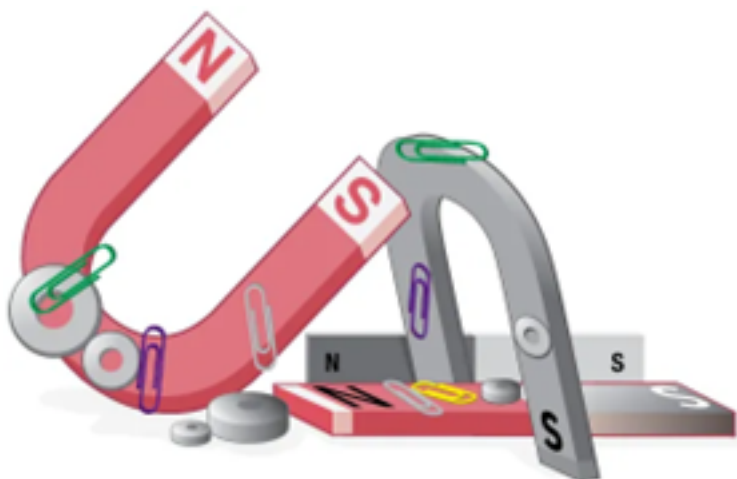


Fig 10.2 Magnets come in various shapes, sizes, and strengths. All have both a north pole and a south pole. There is never an isolated pole (a monopole).

All magnets attract iron, such as that in a refrigerator door. However, magnets may attract or repel other magnets. Experimentation shows that all magnets have two poles. If freely suspended, one pole will point toward the north. The two poles are thus named the **north magnetic pole** and

the **south magnetic pole** (or more properly, north-seeking and south-seeking poles, for the attractions in those directions).

It is a universal characteristic of all magnets that *like poles repel and unlike poles attract*. (Note the similarity with electrostatics: unlike charges attract and like charges repel.) Further experimentation shows that it is *impossible to separate north and south poles* in the manner that + and – charges can be separated.

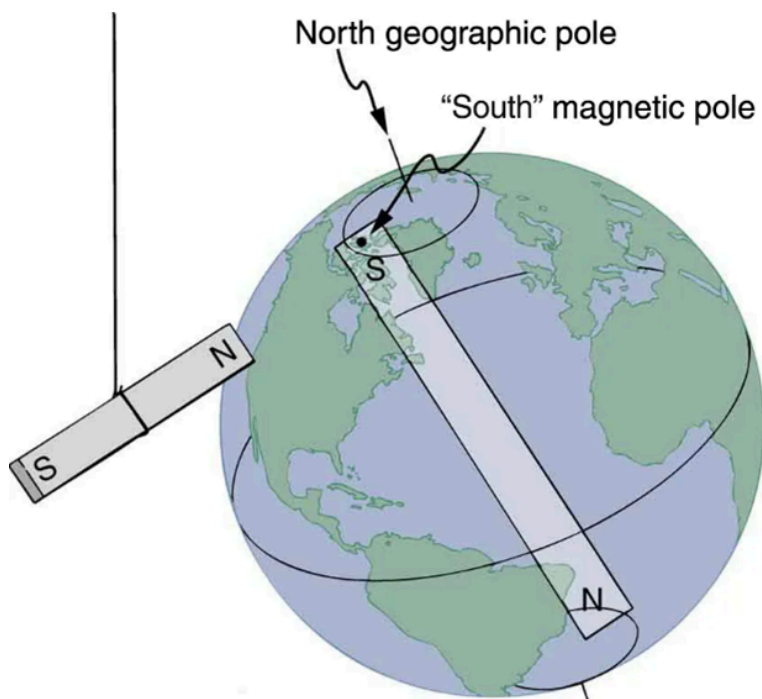


Fig. 10.3 One end of a bar magnet is suspended from a thread that points toward north. The magnet's two poles are labeled N and S for north-seeking and south-seeking poles, respectively.

### Misconception Alert

Earth acts like a very large bar magnet with its

south-seeking pole near the geographic North Pole. That is why the north pole of your compass is attracted toward the geographic north pole of Earth—because the magnetic pole that is near the geographic North Pole is actually a south magnetic pole! Confusion arises because the geographic term “North Pole” has come to be used (incorrectly) for the magnetic pole that is near the North Pole. Thus, “north magnetic pole” is actually a misnomer—it should be called the south magnetic pole.

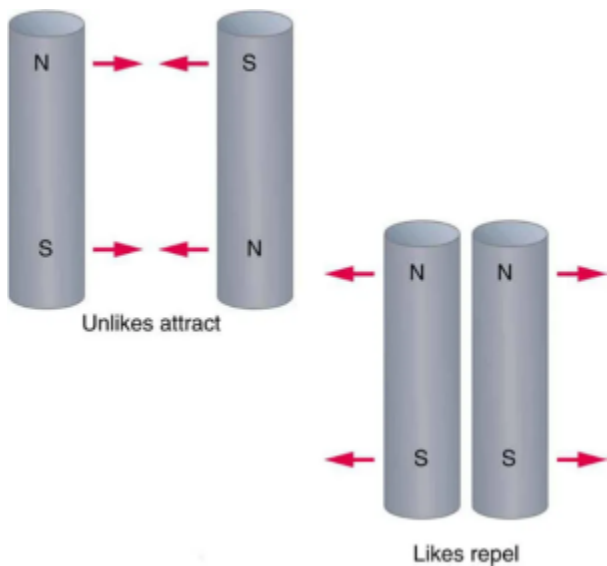


Fig. 10.4 Unlike poles attract, whereas like poles repel.



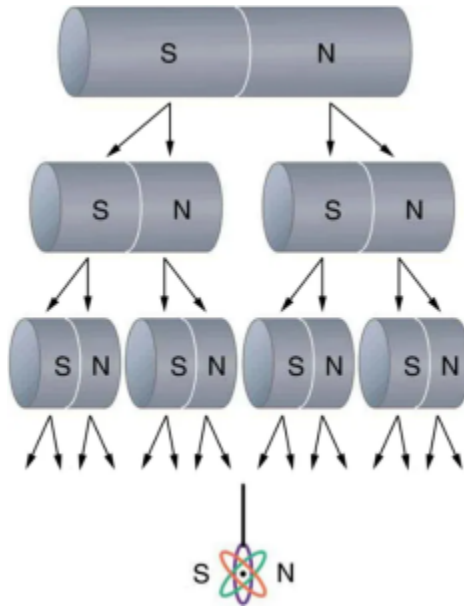


Fig. 10.5 North and south poles always occur in pairs. Attempts to separate them result in more pairs of poles. If we continue to split the magnet, we will eventually get down to an iron atom with a north pole and a south pole—these, too, cannot be separated.

The fact that magnetic poles always occur in pairs of north and south is true from the very large scale—for example, sunspots always occur in pairs that are north and south magnetic poles—all the way down to the very small scale. Magnetic atoms have both a north pole and a south pole, as do

many types of subatomic particles, such as electrons, protons, and neutrons.

### Take Home Experiment: Refrigerator Magnets

We know that like magnetic poles repel and unlike poles attract. See if you can show this for two refrigerator magnets. Will the magnets stick if you turn them over? Why do they stick to the door anyway? What can you say about the magnetic properties of the door next to the magnet? Do refrigerator magnets stick to metal or plastic spoons? Do they stick to all types of metal?

## Ferromagnets

Only certain materials, such as iron, cobalt, nickel, and gadolinium, exhibit strong magnetic effects. Such materials are called ferromagnetic, after the Latin word for iron, *ferrum*. A group of materials made from the alloys of the rare earth

elements are also used as strong and permanent magnets; a popular one is neodymium. Other materials exhibit weak magnetic effects, which are detectable only with sensitive instruments. Not only do **ferromagnets** respond strongly to magnets (the way iron is attracted to magnets), they can also be **magnetized** themselves—that is, they can be induced to be magnetic or made into permanent magnets.

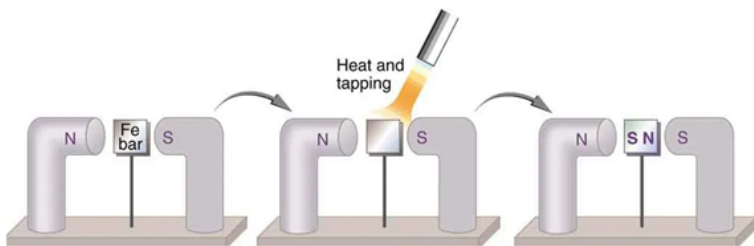


Fig 10.6 An unmagnetized piece of iron is placed between two magnets, heated, and then cooled, or simply tapped when cold. The iron becomes a permanent magnet with the poles aligned as shown: its south pole is adjacent to the north pole of the original magnet, and its north pole is adjacent to the south pole of the original magnet. Note that there are attractive forces between the magnets.

## Electromagnets

Early in the 19th century, it was discovered that electrical currents cause magnetic effects. The first significant

observation was by the Danish scientist Hans Christian Oersted (1777–1851), who found that a compass needle was deflected by a current-carrying wire. This was the first significant evidence that the movement of charges had any connection with magnets. **Electromagnetism** is the use of electric current to make magnets. Such a temporarily induced magnet is called an **electromagnet**. Electromagnets are employed for everything from a wrecking yard crane that lifts scrapped cars to controlling the beam of a 90-km-circumference particle accelerator to the magnets in medical imaging machines. (See Figure 10.7.)



Fig. 10.7 Instrument for magnetic resonance imaging (MRI). The device uses a superconducting cylindrical coil for the main magnetic field. The patient goes into this “tunnel” on the gurney.

Figure 10.8 shows the response of iron filings to a current-carrying coil and to a permanent bar magnet. The patterns are similar. In fact, electromagnets and ferromagnets have the same basic characteristics—for example, they have north and south poles that cannot be separated and for which like poles repel and unlike poles attract.

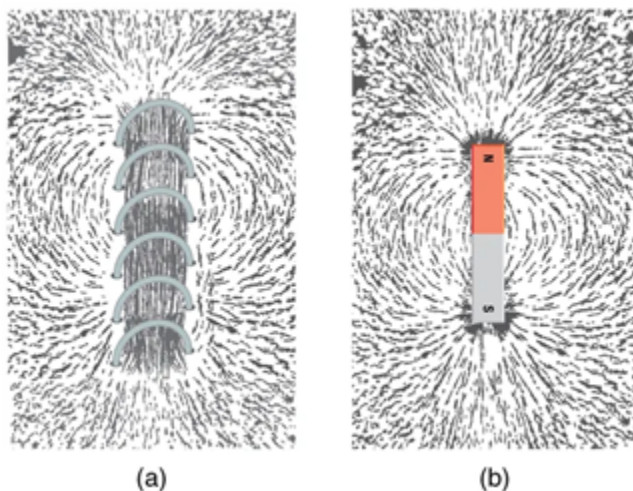


Fig. 10.8 Iron filings near (a) a current-carrying coil and (b) a magnet act like tiny compass needles, showing the shape of their fields. Their response to a current-carrying coil and a permanent magnet is seen to be very similar, especially near the ends of the coil and the magnet.

Combining a ferromagnet with an electromagnet can produce particularly strong magnetic effects. (See Figure 10.9.) Whenever strong magnetic effects are needed, such as lifting scrap metal, or in particle accelerators, electromagnets are enhanced by ferromagnetic materials.

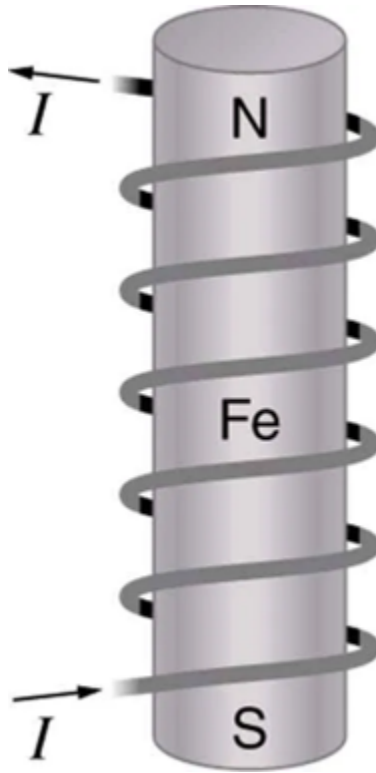


Fig. 10.9 An electromagnet with a ferromagnetic core can produce very strong magnetic effects. Alignment of domains in the core produces a magnet, the poles of which are aligned with the electromagnet.

Figure 10.10 shows a few uses of combinations of electromagnets and ferromagnets. Ferromagnetic materials can

act as memory devices, because the orientation of the magnetic fields of small domains can be reversed or erased. Magnetic information storage on videotapes and computer hard drives are among the most common applications. This property is vital in our digital world.



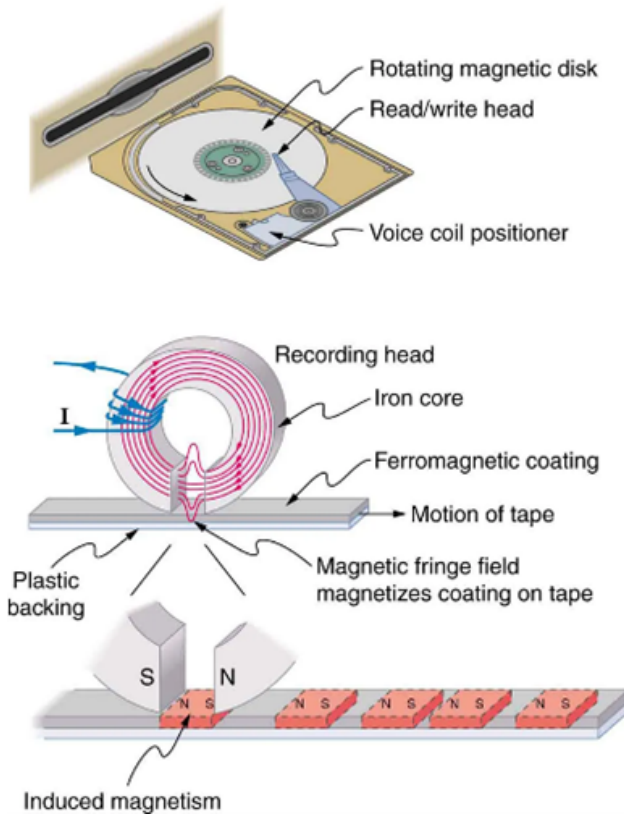


Fig. 10.10 An electromagnet induces regions of permanent magnetism on a floppy disk coated with a ferromagnetic material. The information stored here is digital (a region is either magnetic or not); in other applications, it can be analog (with a varying strength), such as on audiotapes.

## Current: Source of All Magnetism

An electromagnet creates **magnetism** with an electric **current**. But what about ferromagnets? Figure 10.12 shows models of how electric currents create magnetism at the submicroscopic level. (Note that we cannot directly observe the paths of individual **electrons** about atoms, and so a model or visual image, consistent with all direct observations, is made. We can directly observe the electron's orbital angular momentum, its spin momentum, and subsequent magnetic moments, all of which are explained with electric-current-creating subatomic magnetism.) Currents, including those associated with other submicroscopic particles like **protons**, allow us to explain ferromagnetism and all other magnetic effects. Ferromagnetism, for example, results from an internal cooperative alignment of electron spins, possible in some materials but not in others.

Crucial to the statement that electric current is the source of all **magnetism** is the fact that it is impossible to separate north and south magnetic poles. (This is far different from the case of positive and negative charges, which are easily separated.) A current loop always produces a magnetic dipole—that is, a magnetic field that acts like a north pole and south pole pair. Since isolated north and south magnetic poles, called **magnetic monopoles**, are not observed, currents are

used to explain all magnetic effects. If magnetic monopoles did exist, then we would have to modify this underlying connection that all magnetism is due to electrical current. There is no known reason that magnetic monopoles should not exist—they are simply never observed—and so searches at the subnuclear level continue. If they do *not* exist, we would like to find out why not. If they *do* exist, we would like to see evidence of them.

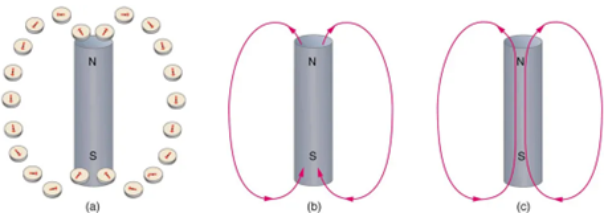


Figure 10.12  
Magnetic field lines are defined to have the direction that a small compass will point when placed at a location. (a) If small compasses are used to map the magnetic field around a bar magnet, they will point in the directions shown: away from the north pole of

## Magnetic Fields and Magnetic Field Lines

the magnet, toward the south pole of the magnet.

(Recall that the Earth's north magnetic pole is really a south pole in terms of definition s of poles on a bar magnet.) (b)

Connecting the arrows gives continuous magnetic field

lines. The strength of the field is proportional to the closeness

Einstein is said to have been fascinated by a compass as a child, perhaps musing on how the needle felt a force without direct physical contact. His ability to think deeply and clearly about action at a distance, particularly for gravitational, electric, and magnetic forces, later enabled him to create his revolutionary theory of relativity. Since magnetic forces act at a distance, we define a **magnetic**

### Link to Learning

Explore the interactions between a compass and bar magnet. Discover how you can use a battery and wire to make a magnet! Can you make it a stronger magnet? Can you make the magnetic field reverse? (Opens in a new window.)

[Click to view content.](#)

(or density) of the lines. (c) If the interior of the magnet could be probed, the field lines would be found to form continuous closed loops.

**field** to represent magnetic forces. The pictorial representation of **magnetic field lines** is very useful in visualizing the strength and direction of the magnetic field. As shown in Figure 10.12, the direction of magnetic field lines is defined to be the direction in which the north end of a compass needle points. The magnetic field is traditionally called the *B*-field.

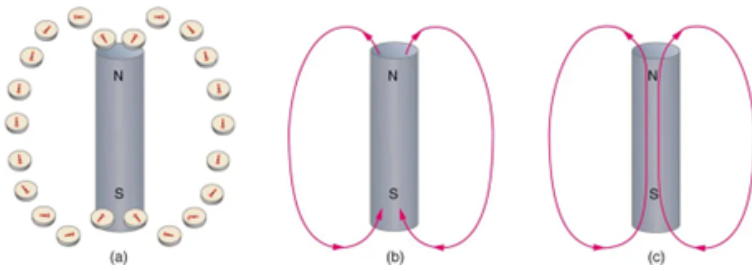


Figure 10.12 Magnetic field lines are defined to have the direction that a small compass will point when placed at a location. (a) If small compasses are used to map the magnetic field around a bar magnet, they will point in the directions shown: away from the north pole of the magnet, toward the south pole of the magnet. (Recall that the Earth's north magnetic pole is really a south pole in terms of definitions of poles on a bar magnet.) (b) Connecting the arrows gives continuous magnetic field lines. The strength of the field is proportional to the closeness (or density) of the lines. (c) If the interior of the magnet could be probed, the field lines would be found to form continuous closed loops.

Small compasses used to test a magnetic field will not disturb it. (This is analogous to the way we tested electric fields with a small test charge. In both cases, the fields represent only the object creating them and not the probe testing them.) Figure 10.13 shows how the magnetic field appears for a current loop and a long straight wire, as could be explored with small compasses. A small compass placed in these fields will align itself parallel to the field line at its location, with its north pole pointing in the direction of  $B$ . Note the symbols used for field into and out of the paper.

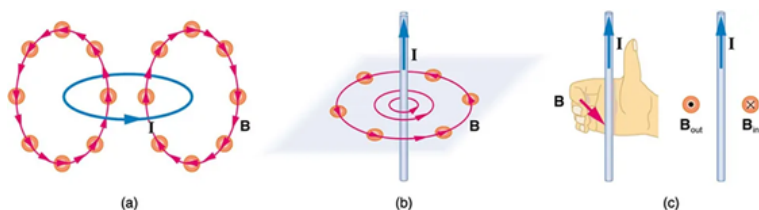


Fig. 10.13 Small compasses could be used to map the fields shown here. (a) The magnetic field of a circular current loop is similar to that of a bar magnet. (b) A long and straight wire creates a field with magnetic field lines forming circular loops. (c) When the wire is in the plane of the paper, the field is perpendicular to the paper. Note that the symbols used for the field pointing inward (like the tail of an arrow) and the field pointing outward (like the tip of an arrow).

## Making Connections: Concept of a Field

A field is a way of mapping forces surrounding any object that can act on another object at a distance without apparent physical connection. The field represents the object generating it. Gravitational fields map gravitational forces, electric fields map electrical forces, and magnetic fields map magnetic forces.

## Magnetic Field Rules

Extensive exploration of **magnetic fields** has revealed a number of hard-and-fast rules. We use **magnetic field lines** to represent the field (the lines are a pictorial tool, not a physical



entity in and of themselves). The properties of magnetic field lines can be summarized by these rules:

1. The direction of the magnetic field is tangent to the field line at any point in space. A small compass will point in the direction of the field line.
2. The strength of the field is proportional to the closeness of the lines. It is exactly proportional to the number of lines per unit area perpendicular to the lines (called the areal density).
3. Magnetic field lines can never cross, meaning that the field is unique at any point in space.
4. Magnetic field lines are continuous, forming closed loops without beginning or end. They go from the north pole to the south pole.

The last property is related to the fact that the north and south poles cannot be separated. It is a distinct difference from **electric field** lines, which begin and end on the positive and negative charges. If **magnetic monopoles** existed, then magnetic field lines would begin and end on them.

## Check Your Understanding

—



*An interactive H5P element has been excluded from this version of the text. You can view it online here:*

<https://louis.pressbooks.pub/physicalscience1/?p=683#h5p-26>



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# GLOSSARY TERMS

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## absolute zero

technically known as zero kelvins, equals  $-273.15$  degrees Celsius, or  $-459.67$  Fahrenheit, and marks the spot on the thermometer where a system reaches its lowest possible energy, or thermal motion

## absorption

transfer of the energy of a wave to matter as the wave passes through it

## acceleration

rate at which an object changes its velocity

## accuracy

closeness of a measured value to a standard or known value

## acid

substance that will dissolve in water to yield hydronium ions,  $\text{H}_3\text{O}^+$ .

## acid-base reaction

reaction in which a hydrogen ion,  $H^+$ , is transferred from one chemical species to another

## alternating current

flow of electric charge that periodically reverses direction

## ammeter

instrument for measuring electric current in amperes

## amplitude

distance between the resting position and the maximum displacement of the wave

## angle of incidence

angle formed between the normal and the incident ray at the point of incidence

## angle of reflection

angle between reflected ray and the normal at the point of incidence to a reflecting surface

## anion

negative ion formed by the loss of electrons

## atom

smallest particle of an element that has the properties of that element and can enter into a chemical combination

## atomic charge

number of protons in an atom minus the number of electrons

## atomic mass

mass of an atom of a chemical element expressed in atomic mass units. It is approximately equivalent to the number of protons and neutrons in the atom (the mass number) or to the average number allowing for the relative abundances of different isotopes

## atomic mass unit

12th of the mass of one carbon-12 atom;  $1 \text{ amu} = 1.6605 \times 10^{-24} \text{ g}$

## atomic number

number of protons in the nucleus of an atom

## Aufbau principle

procedure in which the electron configuration of the

elements is determined by “building” them in order of atomic numbers, adding one proton to the nucleus and one electron to the proper subshell at a time

### average acceleration

rate at which velocity changes

### average velocity

displacement divided by the time of travel

### balanced chemical equation

equation in which equal numbers of atoms for each element involved in the reaction are represented on the reactant (left) and product (right) sides.

### base

substance that will dissolve in water to yield hydroxide ions,  $\text{OH}^-$

### base unit

one of a set of simple units in a system of measurement that is based on a natural phenomenon or established standard and from which other units may be derived



## binary ionic compounds

ionic compounds composed of just two elements, a metal which forms the cations and a nonmetal which forms the anions.

## cation

positive ion formed by the loss of electrons

## Celsius

temperature scale, also called centigrade, according to which water freezes at zero degrees and boils at one hundred degrees.

## chemical bond

attraction between atoms or ions that holds them together in a molecule or a crystal

## chemical change

process in which one or more substances are altered into one or more new and different substances with different physical and chemical properties

## chemical energy

energy that is produced by breaking or forming chemical bonds between atoms and molecules

## chemical equation

symbolic representations of chemical reactions in which the reactants and the products are expressed in terms of their respective chemical formulae

## chemical reactions

common patterns of reactivity between elements

## chemical symbol

abbreviation used to indicate an element or an atom of an element

## circuit

closed loop through which charges can continuously move

## compounds

pure substances that are comprised of two or more elements

## conduction

heat transfer through contact between atoms

## conductivity

measure of the ease at which an electric charge or heat can pass through a material

## controls

standards to which comparisons are made in an experiment

## convection

heat transfer through the movement of gases or liquids.

## converging lens

lens that focuses the light rays to a specific point called the focal point

## conversion

change from one set of units to another, by multiplying or dividing

## convex lens

lens shaped so that all light rays that enter it parallel to its

axis cross one another at a single point on the opposite side of the lens

## Coulomb force

another term for the electrostatic force

## Coulomb's law

mathematical equation calculating the electrostatic force vector between two charged particles

## covalent bonds

chemical bonds that involve the sharing of electrons to form electron pairs between atoms

## crest

point on the medium (wave) that exhibits the maximum amount of positive or upward displacement from the rest position

## current

## deceleration

slowing down or decrease in the magnitude of the velocity of an object; also known as negative acceleration

## density

mass of a unit volume of a material substance.

## dependent variable

variable that changes as a result of the independent variable manipulation

## derived unit

unit that results from a mathematical combination of SI base units

## dimensional analysis

process that uses conversion factors to convert from one set of units to another

## direct current

flow of electric charge in only one direction

## displacement

change in position of an object

## distance

magnitude or size of displacement between two positions

## diverging lens

lens that causes a beam of parallel rays to diverge after refraction, as from a virtual image; a lens that has a negative focal length

## double bond

when two pairs of electrons are shared between a pair of atoms

## dynamics

study of forces that cause objects and systems to move

## electric charge

a physical property of an object that causes it to be attracted toward or repelled from another charged object; each charged object generates and is influenced by a force called an electromagnetic force

## electric current

rate at which charge flows

## electric field

region around a charged particle or object within which

a force would be exerted on other charged particles or objects

### electrical energy

energy (both kinetic and potential) in the charged particles of an atom that can be used to apply force and/or do work

### electromagnet

object that is temporarily magnetic when an electrical current is passed through it

### electromagnetic force

one of the four fundamental forces of nature; the electromagnetic force consists of static electricity, moving electricity and magnetism

### electromagnetic spectrum

range of wavelengths or frequencies over which electromagnetic radiation extends

### electromagnetic waves

form of radiation that travels through the universe; they are created because of vibrations between an electric field and a magnetic field

## electromagnetism

use of electrical currents to induce magnetism

## electron configuration

arrangement of electrons in the orbitals of an atom

## electronegativity

measure of the tendency of an atom to attract electrons

## electrons

negatively charged sub-atomic particles; together with protons and neutrons they compose all atoms

## electrostatic force

amount and direction of attraction or repulsion between two charged bodies

## elements

pure substances that cannot be broken down into simpler substances by chemical changes

## empirical formula

a formula for a compound that indicates the types of



atoms present and the simplest whole-number ratio of the number of atoms or ions in the compound

### endothermic process

reaction or change that absorbs heat

### energy

capacity of a physical system to perform work

### exothermic process

change that releases heat

### experiment

a scientific procedure undertaken to make a discovery, test a hypothesis, or demonstrate a known fact

### experiments

scientific procedures undertaken to make discoveries, test hypotheses, or demonstrate known facts

### Fahrenheit

scale for measuring temperature, in which water freezes at 32 degrees and boils at 212 degrees

## ferromagnets

materials, such as iron, cobalt, nickel, and gadolinium, that exhibit strong magnetic effects

## final velocity

vector quantity that measures the speed and direction of a moving body after it has reached its maximum acceleration

## first law of thermodynamics

total energy of a system remains constant, though energy may transform into another form

## focal length

distance from a lens or mirror to the focal point (F)

## focal point

point at which all radiation coming from a single direction and passing through a lens or striking a mirror converges, also called focus

## force

push or pull upon an object resulting from the object's interaction with another object

## force field

region in which a test particle will experience a force

## frequency

number of waves that pass a fixed point in unit time

## friction

force that resists motion when the surface of one object comes in contact with the surface of another

## fundamental unit of charge

equals the magnitude of the charge of an electron (e) with  
 $e = 1.602 \times 10^{-19} \text{ C}$

## gravitational energy

potential energy associated with gravitational force

## gravity

universal force of attraction acting between all matter

## Heat

energy from the movement of atoms or molecules. It may be considered as energy relating to temperature. Also known as thermal energy

## heat flow

movement of thermal energy that increases the energy of one body and decreases the energy of the other body

## heterogeneous mixture

mixture with a composition that varies from point to point

## homogeneous mixture

a mixture that exhibits a uniform composition and appears visually the same throughout; also called a solution

## hypothesis

a testable statement about the relationship between two or more variables or a proposed explanation for some observed phenomenon

## independent variable

variable that stands alone and isn't changed by the other variables you are trying to measure but manipulated by the researcher

## index of refraction

value calculated from the ratio of the speed of light in a vacuum to that in a second medium of greater density

## International System of Units

a set of base units, derived units, and a set of decimal-based multipliers that are used as prefixes

## ion

atom in which the number of subatomic particles is not equal

## ionic bonds

electrostatic forces of attraction between oppositely charged cations and anions

## ionic compound

compound that contains ions and is held together by ionic bonds

## isomers

compounds with the same chemical formula but different molecular structures

## isotope

each of two or more forms of the same element that contain equal numbers of protons but different numbers of neutrons in their nuclei, and hence differ in relative atomic mass but not in chemical properties; in particular, a radioactive form of an element

## joule

The SI derived unit used to measure energy or work. One joule is equal to the energy used to accelerate a body with a mass of one kilogram using one newton of force over a distance of one meter

## kelvin

base unit of thermodynamic temperature measurement in the International System of Units (SI). It is represented by the symbol K

## kilogram

base unit of mass in the International System of Units (SI)

## kinematics

study of motion without considering its causes

## kinetic energy

energy of motion

## law

statements, based on repeated experiments or observations, that describe or predict a range of natural phenomena

## law of conservation of charge

total charge is constant in any process

## law of conservation of energy

total energy of a system remains constant, though energy may transform into another form

## law of conservation of matter

in a closed or isolated system, matter cannot be created or destroyed. It can change forms but is conserved

## law of reflection

angle of reflection equals the angle of incidence;  $\theta_r = \theta_i$

## law of refraction

light-ray always deviates more towards the normal in the optically denser medium

## length

the measurement of the extent of something along its greatest dimension; a measurement of the physical quantity of distance

## Lewis structures

drawings that describe the bonding in molecules and polyatomic ions

## Lewis symbol

elemental symbol surrounded by one dot for each of its valence electrons

## liter

derived unit for the measure of volume for a substance using the metric system

## magnetic energy

energy that results from a magnetic field



## magnetic field

region around a magnetic material or a moving electric charge within which the force of magnetism acts

## magnetic field lines

pictorial representations of the strength and the direction of magnetic fields

## magnetic monopoles

isolated magnetic poles; a south pole without a north pole, or vice versa (no magnetic monopole has ever been observed)

## magnetism

phenomenon associated with magnetic fields, which arise from the motion of electric charges

## magnetized

to be turned into a magnet; to be induced to be magnetic

## magnitude

quantity or distance

## mass

a dimensionless quantity representing the amount of matter in a particle or object

## mass defect

the difference between the value of an atom's mass in atomic mass units and the atoms actual mass

## mass number

total number of protons and neutrons in an atom

## mass spectrometer

scientific instrument that analyzes and helps identify the substances in a sample of material

## matter

anything that takes up space and can be weighed

## measurement

the process of obtaining the magnitude of a quantity relative to an agreed standard

## mechanical energy

sum of the kinetic and potential energy of a body

## metallic bonds

bonds that result from the electrostatic attraction between metal cations and delocalized electrons

## metalloids

elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals; shaded purple on the periodic table

## metals

elements that are shiny, malleable, and good conductors of heat and electricity; shaded yellow on the periodic table

## meter

the base unit of length in the International System of Units (SI)

## metric system

the decimal measuring system based on the meter, liter, and gram as units of length, capacity, and weight or mass

## mixture

substance composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes

## molecular compounds

compounds that result when atoms share rather than transfer electrons; also called covalent compounds

## molecular formula

representation of a molecule that uses chemical symbols to indicate the types of atoms and subscripts to show the number of each atom of each type in the molecule

## molecular mass

sum of atomic masses for all atoms composing the molecule

## molecule

two or more atoms joined by strong forces called chemical bonds

## net external force

vector sum of all the forces acting on a system from the outside of the system

## neutrons

particles found in the nucleus of an atom; almost identical in mass to a protons, but carry no electric charge

## newton

force needed to accelerate a 1 kg system at the rate 1 m/s squared

## Newton's first law of motion

a body at rest remains at rest, or, if in motion, remains in motion at a constant velocity unless acted on by a net external force

## Newton's second law of motion

acceleration of a system is directly proportional to and in the same direction as the net external force acting on the system, and inversely proportional to its mass.

## Newton's third law of motion

whenever one body exerts a force on a second body, the

first body experiences a force that is equal in magnitude and opposite in direction to the force that it exerts

## nonmetals

elements that appear dull and are poor conductors of heat and electricity; shaded green on the periodic table

## north magnetic pole

end or the side of a magnet that is attracted toward Earth's geographic north pole

## nuclear energy

form of energy released from the nucleus, the core of atoms, made up of protons and neutrons

## nucleus

center of an atom composed of protons and neutrons, surrounded by orbiting electrons

## observations

: the collecting and recording of data, which enables scientists to construct and then test hypotheses and theories

## octet rule

tendency of main group atoms to form enough bonds to obtain eight valence electrons is known as the octet rule

## Ohm's law

an empirical relation stating that the current  $I$  is proportional to the potential difference  $V$ ,  $\propto V$ ; it is often written as  $I = V/R$ , where  $R$  is the resistance

## opaque

substance allowing no light to pass through

## oxidation

process that occurs when an atom, molecule, or ion loses one or more electrons in a chemical reaction

## period

time it takes for two successive crests (one wavelength) to pass a specified point. The time it takes the wave to complete one full oscillation

## periodic table

table of the chemical elements arranged in order of atomic number, usually in rows, so that elements with

similar atomic structure (and hence similar chemical properties) appear in vertical columns

### physical change

change in the state or properties of matter without any accompanying change in the chemical identities of the substances contained in the matter

### physical property

characteristic of matter that is not associated with a change in its chemical composition

### polar covalent bond

bond between different atoms in which electrons are not shared equally

### position

location of an object at any given time

### potential energy

energy that is stored in an object due to its position relative to some zero position

### pound

unit of weight in the British engineering system equal



to 4.448 newtons that is commonly used in the United States

## power

the repeated multiplication of a factor and the number which is raised to that base factor is the exponent

## precipitation reaction

reaction in which dissolved substances react to form one or more solid products

## precision

closeness of two or more measurements to each other.

## prefix

a letter or series of letters attached to the beginning of a word, word base, or suffix to produce a derivative word with a new meaning

## properties

qualities and characteristics of individual substances that describe and identify them

## protons

positively charged particles that, together with the

electrically neutral particles called neutrons, make up the nucleus of an atom

### pure covalent bonds

covalent bonds in which electrons are shared equally between atoms within the compound

### pure substance

substance for which all specimens have exactly the same makeup and properties

### radiant energy

energy emitted by electromagnetic waves that travel through space without any medium

### radiation

heat transfer through electromagnetic waves

### ray

a beam of light or radiation

### real image

image in which light rays from one point on the object actually cross at the location of the image and can be

projected onto a screen, a piece of film, or the retina of an eye

## redox reactions

reactions that involves the transfer of electrons between chemical species; also known as oxidation-reduction reactions

## reduction

process in which a chemical species decreases its oxidation number, usually by gaining electrons

## reflection

change in direction of a wave front at an interface between two different media so that the wave front returns into the medium from which it originated

## refraction

change in direction of a wave passing from one medium to another caused by its change in speed

## resistance

electric property that impedes current

## resistor

object that has simple resistance

## scientific method

a collection of principles and procedures for the systematic pursuit of knowledge involving the recognition and formulation of a problem, the collection of data through observation and experiment, and the formulation and testing of hypotheses

## scientific notation

a way of writing very large or very small numbers

## second

SI Unit for time;  $1/86,400$  of the time that it takes the Earth to rotate once on its axis

## simple harmonic motion

repetitive movement back and forth through an equilibrium, or central, position, so that the maximum displacement on one side of this position is equal to the maximum displacement on the other side

## simple machine

mechanical device that allows you to do the same amount of work with reduced effort by increasing the distance traveled to decrease the force; the six basic simple machines are the wheel and axle, inclined plane (or ramp), pulley, lever, wedge and the screw

## single bond

when a pair of atoms shares one pair of electrons

## solution

a mixture that exhibits a uniform composition and appears visually the same throughout; also called a homogeneous mixture

## sound energy

form of kinetic energy caused by the physical vibration of air particles or molecules

## sound waves

pattern of disturbance caused by the movement of energy traveling through a medium (such as air, water or any other liquid or solid matter) as it propagates away from the source of the sound

## south magnetic pole

end or the side of a magnet that is attracted toward Earth's geographic south pole

## speed

rate of change of position of an object in any direction

## states of matter

distinct forms in which matter can exist

## static electricity

a buildup of electric charge on the surface of an object

## structural formula

formula for a compound that gives the types and numbers of atoms in the molecule and shows how the atoms are connected in the molecule

## temperature

a measure of the average kinetic energy of the particles in an object

## theories

carefully thought-out explanations for observations of

the natural world that have been constructed using the scientific method, and which brings together a myriad of facts and hypotheses.

### theory

a carefully thought-out explanation for observations of the natural world that has been constructed using the scientific method, and which brings together a myriad of facts and hypotheses

### thermal energy

energy from the movement of atoms or molecules. It may be considered as energy relating to temperature; also referred to as heat

### thin lens

lens whose thickness allows rays to refract but does not allow properties such as dispersion and aberrations

### time

interval over which change occurs

### transmission

passage of electromagnetic radiation through a medium

## transparent

substance allowing light to pass through without appreciable scattering of light

## triple bond

when three electron pairs are shared by a pair of atoms

## trough

point on the medium (wave) that exhibits the maximum amount of negative or downward displacement from the rest position

## uncertainty

a range of possible values within which the true value of the measurement lies

## unit

a particular physical quantity, defined and adopted by convention, with which other quantities of the same kind are compared to express their value.

## valence electrons

outer electrons of an atom that have the highest energy and are more easily lost or shared than the core electrons



## vector

quantity that has both magnitude and direction

## velocity

directional speed of an object in motion as an indication of its rate of change in position as observed from a particular frame of reference and as measured by a particular standard of time

## virtual image

image that is on the same side of the lens as the object and cannot be projected on a screen

## voltage

electrical potential energy per unit charge; electric pressure created by a power source, such as a battery

## voltmeter

instrument for measuring electric potential in volts

## volume

quantity of three-dimensional space occupied by a liquid, solid, or gas

## watt

measure of the rate of energy transfer over a unit of time, with one watt equal to one joule (J) per second:  $W = J/s$

## wave

disturbance in which energy is transferred from one location to another.

## wave velocity

distance traversed by a periodic, or cyclic, motion per unit time (in any direction)

## wavelength

length of a wave from one peak to the next; measured from peak to trough

## weight

the force acting on an object due to gravity

## work

force causing the movement—or displacement—of an object