

SURVEY OF CHEMISTRY AND PHYSICS



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Survey of Chemistry and Physics

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CHAPTER OVERVIEW

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- 1.1: Science and Technology
- 1.2: The Scientific Method
- 1.3: Measurement and Scientific Notation
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- 1.5: Measuring Matter
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1.1: Science and Technology

Learning Objectives

- Learn the definitions of science, chemistry, and technology.
- Know the principles of green chemistry.

Science is a practice of learning about the natural world. Natural sciences include [physics](#), [chemistry](#), [biology](#), [geology](#) and [astronomy](#). Science uses [mathematics](#) and [logic](#), which are sometimes called "formal sciences". Natural science makes [observations](#) and experiments. Science produces accurate [facts](#), scientific laws and theories. 'Science' also refers to the large amount of knowledge that has been found using this process.

People who study and research science are scientists. Scientists study things by looking at them very carefully, by measuring, and by doing [experiments](#) and tests. Scientists try to explain why things act the way that they do, and predict what will happen.

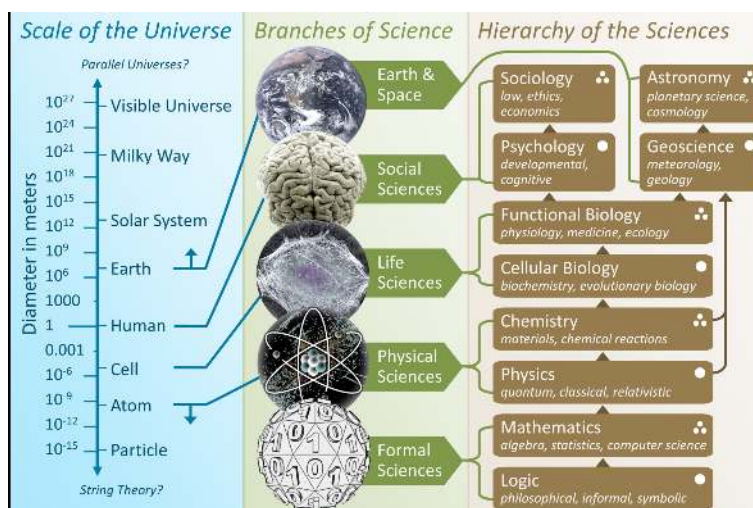


Figure 1.1.1: The scale of the universe mapped to the branches of science. (Source: Wikipedia)

Chemistry is the scientific discipline that involves the properties and behavior of matter, and the changes matter undergoes during a reaction with other substances.

In the scope of its subject, chemistry occupies an intermediate position between physics and biology. It is sometimes called *the central science*, because it provides a foundation for understanding both basic and applied scientific disciplines at a fundamental level (Figure 1.1.1). For example, chemistry explains aspects of plant chemistry (botany), the formation of igneous rocks (geology), how atmospheric ozone is formed and how environmental pollutants are degraded (ecology), the properties of the soil on the moon (astrophysics), how medications work (pharmacology), and how to collect DNA evidence at a crime scene (forensics).

Science is human understanding of how the universe and the things in it work. On the other hand, **technology** is the use of resources to solve a problem (such as knowledge, skills, processes, techniques, tools, and raw materials). Paleolithic technology is the oldest known, but people likely used technology long before then. The Paleolithic age began when hominids (early humans) started to use stones as tools for bashing, cutting, and scraping. The age ended when humans began to make small, fine tools (Mesolithic); it came to a close when they started to plant crops and develop other methods of agriculture (Neolithic).

Technology was once quite simple. It was passed on through word of mouth, until written word was invented. Writing allowed technology to develop much quicker. In the present day, people continue to understand more about the world and the universe. The use of the telescope by Galileo, Einstein's theory of relativity, lasers, and computing are all scientific discoveries. Technology is of great importance to science, medicine, and everyday life.

Green chemistry (also called **sustainable chemistry**) is a type of chemical research and engineering. It supports the design of products and processes that use the smallest amount of dangerous substances possible. In 1990, the Pollution Prevention Act was

passed in the United States. This law sought new and original ways to handle pollution. The Pollution Prevention Act aims to avoid problems before they happen.

Green chemistry applies organic chemistry, inorganic chemistry, biochemistry, analytical chemistry, and even physical chemistry. While green chemistry seems to focus on just industrial applications, it does apply to other scientific disciplines. Green chemists aim to reduce the hazards and increase the efficiency of any chemical choice. Green chemistry is distinct from environmental chemistry, which focuses on chemical phenomena in the environment.

The 12 Principles of Green Chemistry

1. **Prevent waste.**
 - Create products without waste, or with minimal wastes, so that the wastes do not need to be taken care of afterwards.
2. **Design safer chemicals and products.**
 - Design chemicals to have little or no toxicity, without altering effectiveness.
3. **Design less hazardous chemical syntheses.**
 - Design a way of synthesizing products, without them being toxic to humans or the environment.
4. **Use renewable raw materials.**
 - Use renewable raw materials, like plant materials; rather than depleting materials, such as fossil fuels.
5. **Use catalysts, not stoichiometric reagents.**
 - Use catalysts, because of their ability to be able to be reused. Catalysts are less harmful than reagents.
6. **Avoid chemical derivatives.**
 - Chemical derivatives generate wastes that can be avoided.
7. **Maximize atom economy.**
 - Maximize the percentage of reactant atoms that convert to usable product atoms in a chemical reaction, so as to minimize or eliminate atomic waste.
8. **Use safer solvents and reaction conditions.**
 - Avoid using harsh solvents; if that cannot be avoided, then use benign chemicals.
9. **Increase energy efficiency.**
 - Use the normal ambient temperature and pressure wherever possible.
10. **Design for degradation.**
 - Design materials to break down into benign substances, by bacterial or other environmentally sound ways.
11. **Analyze in real time to prevent pollution.**
 - Monitor and control the formation of by-products during a reaction.
12. **Minimize the potential for accidents.**
 - Design chemicals to minimize the potential for accidents.

Summary

- Science is the human method to make discoveries about the natural world.
- Chemistry is the branch of science involved with the properties and behavior of matter, and the changes it undergoes during a reaction with other substances.
- Green chemistry applies organic chemistry, inorganic chemistry, biochemistry, analytical chemistry, and physical chemistry to reduce and prevent pollution.
- There are 12 guiding principles in the practice of green chemistry.

Contributors and Attributions

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1.2: The Scientific Method

Learning Objectives

- To identify the components of the scientific method
- Classify measurements as being quantitative or qualitative.
- Evaluate science in the media.

The Scientific Method

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

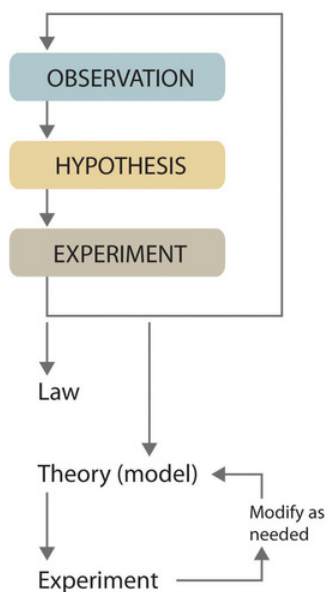


Figure 1.2.1: The Steps in the Scientific Method. (CC BY-SA-NC; anonymous).

Step 1: Make observations

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

Step 2: Formulate a hypothesis

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

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

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

Step 3: Design and perform experiments

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

Step 4: Accept or modify the hypothesis

A properly designed and executed experiment enables a scientist to determine whether the original hypothesis is valid. In which case he can proceed to step 5. In other cases, experiments often demonstrate that the hypothesis is incorrect or that it must be modified thus requiring further experimentation.

Step 5: Development into law and/or theory

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

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

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

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

A Real-World Application of the Scientific Method

In 2007, my husband and I journeyed to China to adopt our daughter. Upon arrival in Beijing, I became violently ill. Due to her visa paperwork, my husband, daughter, and I were required to stay in China for two weeks. Unfortunately, I was ill the entire time. Once the two-week period was up, the three of us flew back to the United States where I continued to be sick. For the next year, I remained ill and lost a total of 30 pounds. The picture below shows me holding my daughter eight months after we returned home from China.



Figure 1.2.2: Mother and child. (CC BY-SA-NC; Elizabeth R. Gordon).

I would like you to attempt to perform the scientific method on my situation described above. List the steps of the scientific method along with some plausible explanations. * **Please have this ready to discuss in class.***

✓ Exercise \(\PageIndex{2}\)

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

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

Answer a

experiment

Answer b

law

Answer c

theory

Answer d

hypothesis

Answer e

qualitative observation

Answer f

quantitative observation

Evaluating Science in the Media

A Rough Guide to SPOTTING BAD SCIENCE

Being able to evaluate the evidence behind a scientific claim is important. Being able to recognise bad science reporting, or faults in scientific studies, is equally important. These 12 points will help you separate the science from the pseudoscience.

1. SENSATIONALISED HEADLINES

Aa

Article headlines are commonly designed to entice viewers into clicking on and reading the article. At times, they can over-simplify the findings of scientific research. At worst, they sensationalise and misrepresent them.

7. UNREPRESENTATIVE SAMPLES USED



In human trials, subjects are selected that are representative of a larger population. If the sample is different from the population as a whole, then the conclusions from the trial may be biased towards a particular outcome.

2. MISINTERPRETED RESULTS



News articles can distort or misinterpret the findings of research for the sake of a good story, whether intentionally or otherwise. If possible, try to read the original research, rather than relying on the article based on it for information.

8. NO CONTROL GROUP USED



In clinical trials, results from test subjects should be compared to a 'control group' not given the substance being tested. Groups should also be allocated randomly. In general experiments, a control test should be used where all variables are controlled.

3. CONFLICTS OF INTEREST



Many companies will employ scientists to carry out and publish research - whilst this doesn't necessarily invalidate the research, it should be analysed with this in mind. Research can also be misrepresented for personal or financial gain.

9. NO BLIND TESTING USED



To try and prevent bias, subjects should not know if they are in the test or the control group. In 'double blind' testing, even researchers don't know which group subjects are in until after testing. Note, blind testing isn't always feasible, or ethical.

4. CORRELATION & CAUSATION



Be wary of any confusion of correlation and causation. A correlation between variables doesn't always mean one causes the other. Global warming increased since the 1800s, and pirate numbers decreased, but lack of pirates doesn't cause global warming.

10. SELECTIVE REPORTING OF DATA



Also known as 'cherry picking', this involves selecting data from results which supports the conclusion of the research, whilst ignoring those that do not. If a research paper draws conclusions from a selection of its results, not all, it may be guilty of this.

5. UNSUPPORTED CONCLUSIONS



Speculation can often help to drive science forward. However, studies should be clear on the facts their study proves, and which conclusions are as yet unsupported ones. A statement framed by speculative language may require further evidence to confirm.

11. UNREPLICABLE RESULTS



Results should be replicable by independent research, and tested over a wide range of conditions (where possible) to ensure they are consistent. Extraordinary claims require extraordinary evidence - that is, much more than one independent study!

6. PROBLEMS WITH SAMPLE SIZE



In trials, the smaller a sample size, the lower the confidence in the results from that sample. Conclusions drawn can still be valid, and in some cases small samples are unavoidable, but larger samples often give more representative results.

12. NON-PEER REVIEWED MATERIAL



Peer review is an important part of the scientific process. Other scientists appraise and critique studies, before publication in a journal. Research that has not gone through this process is not as reputable, and may be flawed.





Contributors and Attributions

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1.3: Measurement and Scientific Notation

Learning Outcomes

- Demonstrate use of scientific notation.
- Identify base and derived units for measurements.
- Memorize the metric prefixes, abbreviations, and values (Tables 1.1.1 and 1.1.3).
- Convert between temperatures in Celsius and Fahrenheit.

Many aspects of chemistry use quantitative measurements to describe behavior. In this section, we will look at how we deal with very large or very small numbers, units of measure and prefixes used with these units, and how numbers are reported.

Scientific Notation

Scientific notation is a way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power. It is a very useful tool for working with numbers that are either very large or very small. As an example, the distance from Earth to the Sun is about 150,000,000,000 meters - a very large distance indeed. In scientific notation, the distance is written as $1.5 \times 10^{11} \text{ m}$. The coefficient is the 1.5 and must be a number greater than or equal to 1 and less than 10. The power of 10, or exponent, is 11 because you would have to multiply 1.5 by 10^{11} to get the correct number. Scientific notation is sometimes referred to as exponential notation.

When working with small numbers, we use a negative exponent. So 0.1 meters is 1×10^{-1} meters, 0.01 is 1×10^{-2} and so forth. A red blood cell has a diameter of 0.000008 meters which is neither convenient to write nor say. It is much easier to report a diameter of $8 \times 10^{-6} \text{ m}$. Using scientific notation is one way to make writing very large or very small numbers more convenient. Note the use of a **leading zero** (the zero to the left of the decimal point) when writing very small numbers. The leading zero is there to help you see the decimal point more clearly. The figure 0.01 is less likely to be misunderstood than .01 where you may not see the decimal point.

Units

SI Base Units

All measurements depend on the use of units that are all well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. The **International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units, five of which are shown in the table below. While some of the base units have a calculation as their standard, the kilogram has a physical standard but there is a movement to change that. Read or listen to more about changes to the standard at <http://www.npr.org/templates/story/s...ryId=112003322>.

Table 1.3.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol

These five units are frequently encountered in chemistry and other units of measurement, such as volume, force, and energy, can be derived from the SI base units.



Figure 1.3.1: The standard for the unit kilogram is stored in France.

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



Figure 1.3.2: Countries which use the metric system.

Derived Units

Some units are combinations of SI base units. A **derived unit** is a unit that results from a mathematical combination of SI base units. These derived units are the standard for these quantities but they are not always the units used in normal practice. For example, the density of a solid or a liquid is often reported as g/cm^3 , or g/mL rather than kg/m^3 (the derived SI unit). Some common examples of derived units are listed in the table below.

Table 1.3.2: Derived SI Units

Quantity	Symbol	Unit	Unit Abbreviation	Derivation
Area	A	square meter	m^2	$\text{length} \times \text{width}$
Volume	V	cubic meter	m^3	$\text{length} \times \text{width} \times \text{height}$
Density	D	kilograms/cubic meter	kg/m^3	$\frac{\text{mass}}{\text{volume}}$
Concentration	c	moles/liter	mol/L	$\frac{\text{amount}}{\text{volume}}$
Speed (velocity)	v	meters/second	m/s	$\frac{\text{length}}{\text{time}}$
Acceleration	a	meters/second/second	m/s^2	$\frac{\text{speed}}{\text{time}}$
Force	F	newton	N	$\text{mass} \times \text{acceleration}$
Energy	E	joule	J	$\text{force} \times \text{length}$

Metric Prefixes

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter.

Metric prefixes are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word *μεγας*, meaning "great".

The table below lists the most common metric prefixes and their relationship to the base unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit. However, these prefixes can be used with *any* base unit.

Table 1.3.3: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter (Gm) = 10^9 m
mega	M	1,000,000	1 megameter (Mm) = 10^6 m
kilo	k	1,000	1 kilometer (km) = 1,000 m
		1	1 meter (m)
deci	d	1/10	1 decimeter (dm) = 0.1 m
centi	c	1/100	1 centimeter (cm) = 0.01 m
milli	m	1/1,000	1 millimeter (mm) = 0.001 m
micro	μ	1/1,000,000	1 micrometer (μm) = 10^{-6} m
nano	n	1/1,000,000,000	1 nanometer (nm) = 10^{-9} m

*Micro is often abbreviated "mc" in healthcare to avoid confusion between the μ and m which often look similar when written by hand.

There are more prefixes, some of them rarely used, that go far beyond what is given here. Have you ever heard of a zeptometer? You can learn more about prefixes at <http://www.bipm.org/en/measurement-units/>

Since the metric system is not the usual unit system we use, it can be challenging to understand the relative size of each unit. Go to <http://learn.genetics.utah.edu/content/cells/scale> and use the slider below the image to zoom in and see how the sizes of objects compare.

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

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs **6.5 kg**, this is easier than saying it weighs **6500 g**. Both are correct, but the **kg** units in this case make for a small and easily managed number.

Temperature Scales

There are three temperature scales that are commonly used in measurement. Their units are **°F** (degrees Fahrenheit), **°C** (degrees Celsius), and **K** (Kelvin). The **Fahrenheit scale**, which is the most commonly used scale in the United States, defines the normal freezing point and boiling point of water as **32°F** and **212°F**, respectively. The **Celsius scale** defines the normal freezing point and boiling point of water as **0°C** and **100°C**, respectively. The Celsius scale is commonly used in most countries across the globe. The **Kelvin scale**, which is also referred to as the absolute temperature scale, defines absolute zero as the lowest theoretically possible temperature, which means that temperatures expressed in Kelvin cannot be negative numbers.

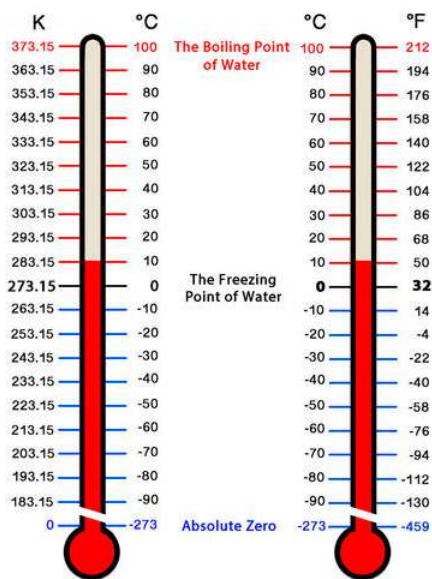


Figure 1.3.3: Comparison of Kelvin, Celsius on a Fahrenheit scale.

Converting Temperature Scales

Regardless of the temperature scale used, it is important to be able to convert from one scale to another. Here are the conversions we use.

°F to °C

$$T_{\text{°C}} = (T_{\text{°F}} - 32) \times \frac{5}{9} \quad (1.3.1)$$

°C to °F

$$T_{\text{°F}} = \frac{9}{5} \times (T_{\text{°C}}) + 32 \quad (1.3.2)$$

°C to K

$$T_{\text{K}} = T_{\text{°C}} + 273.15 \quad (1.3.3)$$

K to °C

$$T_{\text{°C}} = T_{\text{K}} - 273.15 \quad (1.3.4)$$

Example 1.3.1

The melting point of mercury is -38.84°C . Convert this value to degrees Fahrenheit and Kelvin.

Solution

The formulas above can be used to convert among temperature units. First, the given value can be used to convert from °C to °F.

$$\begin{aligned} T_{\text{°F}} &= \frac{9}{5} \times (-38.84^{\circ}\text{C}) + 32 \\ T_{\text{°F}} &= -37.91^{\circ}\text{F} \end{aligned}$$

Then, the same initial temperature can be used to find the temperature in Kelvin.

$$\begin{aligned} T_{\text{K}} &= -38.84^{\circ}\text{C} + 273.15 \\ T_{\text{K}} &= 234.75 \text{ K} \end{aligned}$$

Supplemental Resources

- Scientific notation practice: <http://www.mathisfun.com/numbers/sci...-notation.html>
- History of SI Units: <http://physics.nist.gov/cuu/Units/history.html>

Contributors and Attributions

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- [Allison Soult](#), Ph.D. (Department of Chemistry, University of Kentucky)

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1.4: Significant Figures and Scientific Notation

Learning Outcomes

- Distinguish between accuracy and precision.
- Explain the concept of significant figures.
- Report answers to calculations with the correct number of significant figures.

In chemistry, we are looking at measurements and not just numbers. Therefore, it is necessary to be able to count the number of significant figures and report measurements to the correct level of precision. The reported value should be as precise as possible without adding digits to a value that cannot be measured.

Accuracy and Precision

In everyday speech, the terms accuracy and precision are frequently used interchangeably. However, their scientific meanings are quite different. **Accuracy** is a measure of how close a measurement is to the correct or accepted value of the quantity being measured. **Precision** is a measure of how close a series of measurements are to one another. Precise measurements are highly reproducible, even if the measurements are not near the correct value. Darts thrown at a dartboard are helpful in illustrating accuracy and precision (see figure below).

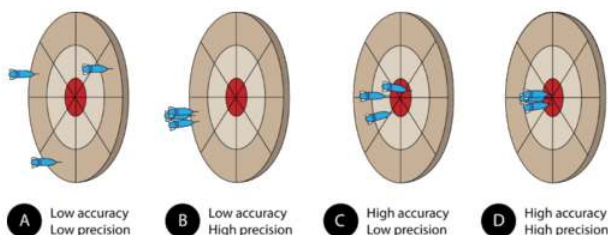


Figure 1.4.3: Distribution of darts on a dartboard to represent precision and accuracy.

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. Pictured above are the four possible outcomes.

(A) The darts have landed far from each other and far from the bulls-eye. This grouping demonstrates measurements that are neither accurate nor precise.

(B) The darts are close to one another, but far from the bulls-eye. This grouping demonstrates measurements that are precise, but not accurate. In a laboratory situation, high precision with low accuracy often results from a systematic error. Either the measurer makes the same mistake repeatedly or the measuring tool is somehow flawed. A poorly calibrated balance may give the same mass reading every time, but it will be far from the true mass of the object.

(C) The darts are not grouped very near to each other, but they are generally centered around the bulls-eye. This demonstrates poor precision but fairly high accuracy. This situation is not desirable because in a lab situation, we do not know where the "bulls-eye" actually is. Continuing with this analogy, measurements are taken in order to find the bulls-eye. If we could only see the locations of the darts and not the bulls-eye, the large spread would make it difficult to be confident about where the exact center was, even if we knew that the darts were thrown accurately (which would correspond to having equipment that is calibrated and operated correctly).

(D) The darts are grouped together and have hit the bulls-eye. This demonstrates high precision and high accuracy. Scientists always strive to maximize both in their measurements. Turning back to our laboratory situation, where we can see the darts but not the bulls-eye, we have a much narrower range of possibilities for the exact center than in the less precise situation depicted in part C.

Significant Figures in Measurements

Some error or uncertainty always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest **0.1 g**, other highly sensitive balances are capable of measuring to the nearest **0.001 g** or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Pictured below is an object (indicated by the blue arrow) whose length is being measured by two different rulers.

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

The top ruler contains marks for tenths of a centimeter (millimeters). Now, the same object may be measured as **2.55 cm**. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5.

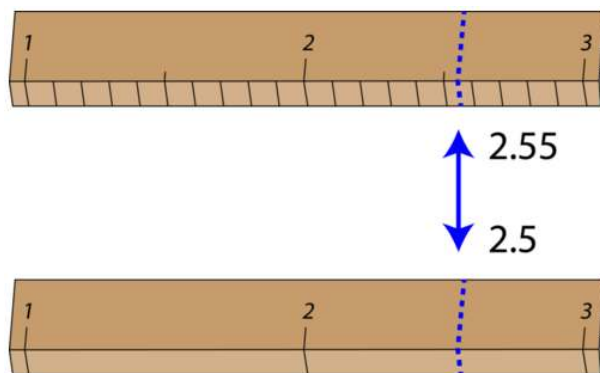


Figure 1.4.4: Measurements can be more or less precise depending on the markings on the ruler.

Again, another measurer may report the length to be **2.54 cm** or **2.66 cm**. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain (estimated) digit. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported.

Rules for Determination of Significant Figures

1. All nonzero digits are significant.

- 237 has three significant figures
- 1.897 has four significant figures

2. Zeroes between significant figures are significant.

- 39004 has five significant figures
- 2.03 has three significant figures

3. Zeroes that appear to the left of all of the nonzero digits are called placeholder zeroes and are not significant.

- 0.008 has one significant figure
- 0.0000416 has three significant figures

4. Zeroes that appear after the last nonzero digit may be significant.

a. If the number is greater than 1 and the zero(es) are not followed by a decimal point, then the number of significant figures is ambiguous.

- 140 is ambiguous

b. If the zero is followed by a decimal point, then the zero is significant.

- 140. has three significant figures (note the decimal point after the zero)

c. If the zero is after a decimal point, then the zero is significant.

- 141.0 has four significant figures

It needs to be emphasized that just because a certain digit is not significant does not mean that it is not important or that it can be left out. Though the zero in a measurement of 140 may not be significant, the value cannot simply be reported as 14. An insignificant zero functions as a placeholder for the decimal point. When numbers are written in scientific notation, this becomes more apparent. The measurement 140 can be written as 1.4×10^2 , with two significant figures in the coefficient or as 1.40×10^3 , with three significant figures. A number less than one, such as 0.000416, can be written in scientific notation as 4.16×10^{-4} , which has 3 significant figures. In some cases, scientific notation is the only way to correctly indicate the correct number of significant figures. In order to report a value of 15,000,00 with four significant figures, it would need to be written as 1.500×10^7 .

Exact Quantities

When numbers are known exactly, the significant figure rules do not apply. This occurs when objects are counted rather than measured. For example, a carton of eggs has 12 eggs. The actual value cannot be 11.8 eggs, since we count eggs in whole number quantities. So the 12 is an exact quantity. Exact quantities are considered to have an infinite number of significant figures; the importance of this concept will be seen later when we begin looking at how significant figures are dealt with during calculations. Numbers in many conversion factors, especially for simple unit conversions, are also exact quantities and have infinite significant figures. There are exactly 100 centimeters in 1 meter and exactly 60 seconds in 1 minute. Those values are definitions and are not the result of a measurement.

Calculations

Adding and Subtracting Significant Figures

The sum or difference is determined by the smallest number of digits to the right of the decimal point in any of the original numbers.

Example 1.4.1

What is the result of $89.332 + 1.1$ when answered to the correct number of significant figures?

Solution

The mathematical result is $89.332 + 1.1 = 90.432$. However, we need to round the answer to the correct number of significant figures. For addition and subtraction, the answer should have the same number of decimal places as the starting value with the *least* number of decimal places. Since 1.1 only has one digit after the decimal place, the answer will be rounded to one decimal place. The first digit to be dropped is less than five so we round down.

The correct answer is 90.4.

Multiplying and Dividing Significant Figures

The number of significant figures in the final product or quotient is equal to the number of significant figures in the starting value that has the fewest significant figures.

Example 1.4.2

What is the product of 2.8 and 4.5039 reported to the correct number of significant figures?

Solution

First, we find the mathematical answer to the calculation which is $2.8 \times 4.5039 = 12.61092$. Next, we need to round the answer to the correct number of significant figures. For multiplication and division, the answer should be rounded to that it has the same number of significant figures as the starting value with the least number of significant figures. The value 2.8 has two significant figures while 4.5039 has five. Therefore, the answer should have two significant figures so we should round the answer to 13 because the first digit to be dropped is greater than 5, so we need to round up.

Example 1.4.3

A bag of chocolate candies has a mass of 8.25 ounces. The candies were to be divided among 4 people. How many ounces should each person get? Report your answer to the correct number of significant figures.

Solution

First, we'll find the mathematical answer to this calculation: $8.25 \div 4 = 2.0625$. Now, we need to round to the correct number of significant figures so we look back at the starting values. There are three significant figures in 8.25 but 4 is an exact number so it is not considered when determining the number of significant figures in the answer. Therefore, the answer should be rounded to three significant figures or 2.06 ounces.

Supplemental Resources

Contributors and Attributions

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1.5: Measuring Matter

Learning Objectives

- Connect basic glassware and equipment with metric units.
- Provide SI units for mass, temperature, and length.
- Practice scientific notation and calculator skills (entering powers of ten).
- Memorize or place the bolded conversion factors on your reference card.

A variety of instruments are available for making direct measurements of the macroscopic properties of a chemical substance. For example, we usually measure the **volume** of a liquid sample with pipettes, burets, graduated cylinders, and volumetric flasks (Figure 1.5.1), whereas we usually measure the mass of a solid or liquid substance with a balance (Figure 1.5.2).

Instruments that Measure Volume

Many academic scientists report their volume measurements with the milliliter unit. On the other hand, analytical chemists tend to work with much smaller volumes, and they would use nanoliter and microliter samples in their laboratory. When considering the industrial realm, scientists work with larger volumes and would thus employ the liter unit of measurement. Scientists use beakers and flasks to contain liquids. For measuring this type of matter, graduated cylinders and pipettes are utilized.

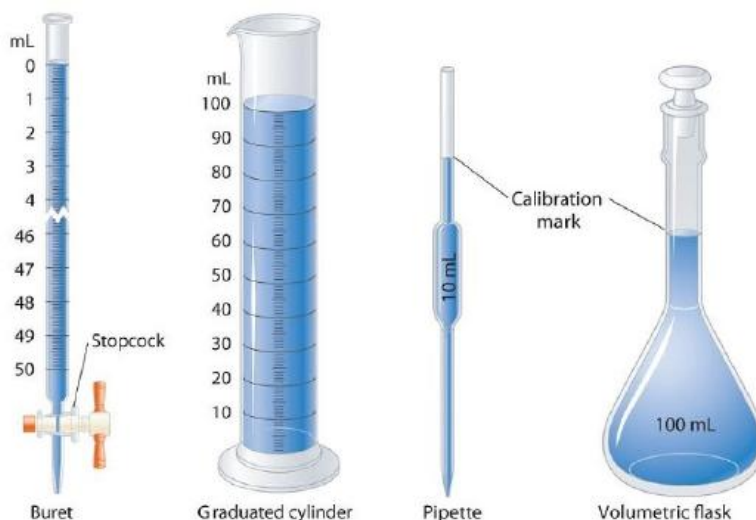


Figure 1.5.1: (left two) Graduated glassware is used to deliver variable volumes of liquid. (right two) Volumetric glassware is used to deliver (pipette) or contain (volumetric flask) a single volume accurately when filled to the mark. (CC BY-SA-NC 3.0; anonymous).

Instruments that Measure Mass

Balances are utilized to measure the **mass** of matter. Typically, high school and middle schools will use triple beam balances for laboratory activities. Colleges will rely on digital balances. For larger quantities of materials, top-loading digital balances are employed. For higher sensitivity and accuracy, some chemists use analytical balances.



Figure 1.5.2: Triple beam balance (image from: upload.wikimedia.org/wikipedia/m_balance2.JPG. Digital top-loading balance (upload.wikimedia.org/wikipedia/00-2M_-_01.jpg) and an analytical balance (upload.wikimedia.org/wikipedia/NaCl-1mol.jpg)

In the academic laboratory (like Furman University's), a chemist will commonly use the gram, or the even smaller unit, the milligram. Industrial chemists making larger quantities of material would work with kilograms of materials. Analytical chemists (industrial or academic), environmental scientists, and toxicologists are typically more concerned with the smaller units of the metric system. For example, many chemicals found in drinking water and air are regulated in concentrations like nanograms and or micrograms.

Tools that Measure Length

Length is measured using a metric stick or a metric ruler marked in millimeters and centimeters. Most devices used to measure length (in the United States) contain both English and metric markings.

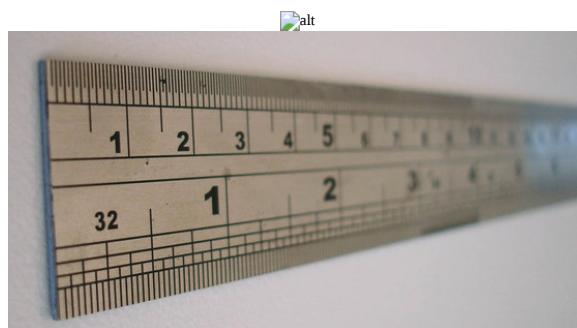


Figure 1.5.3: Metric and English measuring devices. (CC BY-SA 4.0; Ejay).

SI Units

All reported measurements must include an appropriate unit of measurement. To say that a substance has “a mass of 10” does not tell whether the mass was measured in grams, pounds, tons, or some other unit. For example, a nurse giving morphine to a patient needs to be specific in the unit dosages. A volume of liters would kill a patient quickly while a milliliter amount would provide sufficient pain relief. To establish worldwide standards for the consistent measurement of significant physical and chemical properties, an international body called the General Conference on Weights and Measures devised the *Système Internationale d’unités* (or SI). The International System of Units is based on metric units and requires that measurements be expressed in decimal form. Table 1.5.1 lists the seven base units of the SI system; all other SI units of measurement are derived from them.

Table 1.5.1: *SI Units*

Base Quantity	Unit Name	Abbreviation
mass	kilogram	kg
length	meter	m
time	second	s
temperature	Kelvin	K
amount of substance	mole	mol
intensity of light	candela	cd

The Metric System

Table 1.5.2: Prefixes Used with metric bases of meter, liter, and gram

Prefix	Symbol	Value	Power of 10	Meaning
tera	T	1,000,000,000,000	10^{12}	trillion
giga	G	1,000,000,000	10^9	billion
mega	M	1,000,000	10^6	million
kilo	k	1000	10^3	thousand
hecto	h	100	10^2	hundred
deca	da	10	10^1	ten
—	—	1 (this is base unit which could be gram, liter, or meter)	10^0	one
deci	d	0.1	10^{-1}	tenth
centi	c	0.01	10^{-2}	hundredth
milli	m	0.001	10^{-3}	thousandth
micro	μ (medical community uses mcg)	0.000001	10^{-6}	millionth
nano	n	0.000000001	10^{-9}	billionth
pico	p	0.000000000001	10^{-12}	trillionth
femto	f	0.000000000000001	10^{-15}	quadrillionth

Supplement Facts

Serving Size 1 Packet
Servings Per Container 10

	AM Packet		PM Packet	
Amount Per Serving	% Daily Value		% Daily Value	
Vitamin A	450 mcg	50%	450 mcg	50%
Vitamin C	90 mg	100%	90 mg	100%
Vitamin D	20 mcg (800 IU)	100%		
Vitamin E	15 mg	100%		
Thiamin	1.2 mg	100%	1.2 mg	100%
Riboflavin	1.3 mg	100%	1.3 mg	100%
Niacin	16 mg	100%	16 mg	100%
Vitamin B ₆	1.7 mg	100%	1.7 mg	100%
Folate	200 mcg DFE	50%	200 mcg DFE	50%
	(120 mcg folic acid)		(120 mcg folic acid)	
Vitamin B ₁₂	1.2 mcg	50%	1.2 mcg	50%
Biotin			3 mcg	10%
Pantothenic Acid	2.5 mg	50%	2.5 mg	50%

Ingredients: Sodium ascorbate, ascorbic acid, calcium pantothenate, niacinamide, dl-alpha tocopheryl acetate, microcrystalline cellulose, dextrin, starch, mono- and diglycerides, vitamin A acetate, magnesium stearate, gelatin, FD&C Blue #1, FD&C Red #3, artificial colors, thiamin mononitrate, pyridoxine hydrochloride, citric acid, lactose, sorbic acid (preservative), tricalcium phosphate, sodium benzoate (preservative), sodium caseinate, preservatives (methylparaben, potassium sorbate, BHA, BHT), ergocalciferol, cyanocobalamin, and artificial flavors.

Figure 1.5.4: Example of Supplement Facts by the FDA. (Public Domain, FDA.gov)

Two methods of Memorization for the Metric System

Using negative exponents (see above) is the traditional way to memorize the metric system. Many students who have been taught outside of the United States approach the metric system this way. For their conversion factors, students who have been taught the metric system in the United States tend to use positive exponents. Either method is fine, pick one method and use it consistently.

Two different ways to memorize the metric systems (be sure to know all bolded conversion factors listed in section 2.1 of this OER)

1 gigabase = 1×10^9 base

1 megabase = 1×10^6 base

1 kilobase = 1×10^3 base

• **1 centibase = 1×10^{-2} base**

• **1 millibase = 1×10^{-3} base**

• **1 microbase = 1×10^{-6} base**

• **1 nanobase = 1×10^{-9} base**

or 1×10^2 centibase = 1 base

or 1×10^3 millibase = 1 base

or 1×10^6 microbase = 1 base

or 1×10^9 nanobase = 1 base

Note: A base unit can be gram (mass), meter (length), volume (liter), and joules (energy).

Additional Units

When traveling to different regions of the globe, it is helpful to be able to convert from English to Metric units. **Please be sure to commit the bolded conversions below to memory.**

Unit Conversion
2.21 pounds = 1.00 kilograms
1 mile = 1.61 kilometers
1.00 inch = 2.54 centimeters
1.00 mL = 1.00 cc (cubic centimeters) = 1.00 cm³
1 teaspoon (tsp) = 5.00 mL
12 in = 1.00 ft.
16 oz. = 1.00 pounds

Scientific Notation

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

$$N \times 10^n$$

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

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

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

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases, and vice versa. The application of this rule is illustrated in Example 1.5.1.

✓ Example 1.5.1: scientific notation

Convert each number to scientific notation.

- 637.8
- 0.0479
- 7.86
- 12,378
- 0.00032
- 61.06700
- 2002.080
- 0.01020

Solution

Explanation	Solution
To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 6.378 Because the decimal point was moved two places to the left, $n = 2$.	$637.8 = 6.378 \times 10^2$
To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 4.79 Because the decimal point was moved two places to the right, $n = -2$.	$0.0479 = 4.79 \times 10^{-2}$
This is usually expressed simply as 7.86. (Recall that $10^0 = 1$).	7.86×10^0
Move the decimal point four places to the left and note a positive exponent.	1.2378×10^4
Move the decimal place four places to the right a note a negative exponent.	3.2×10^{-4}
Move the decimal place one space to the left and note a positive exponent.	6.106700×10^1
Move the decimal place three places to the left and note a positive exponent.	2.002080×10^3
Move the decimal place two places to the right and note a negative exponent.	1.020×10^{-2}

Using the Scientific Calculator

Please be sure to obtain a calculator that has an exponent button on it. It will either say "exp" or have an "ee" key indicating this function. This key represents the function of 10^x . You do not need to enter a 10 for an exponent when using this particular function. Read the steps below and practice on your calculator.

Use of Calculator

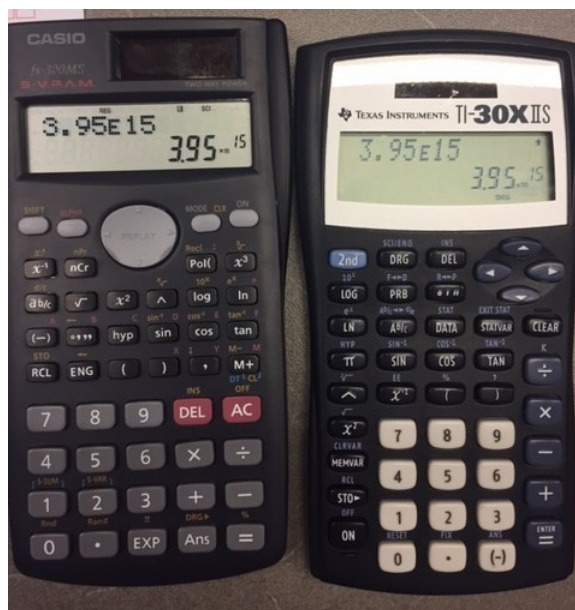
To input powers of ten:

1. Locate the ee or exp button on your calculator. Do not use the carrot (looks like the $\wedge$ button).
2. Put the coefficient number in and then hit "ee" or "exp."
3. If you need to have a negative "exp," then locate the +/- or small - button.
4. Lastly, input your power of ten. You do not need to put the actual number ten in for the exponent.

Table 1.5.3: Examples

Number	How to Input in a calculator
3.95×10^{15}	Put 3.95 in as the coefficient number. Hit the "ee" or "exp" button. Lastly, enter the 15 for the exponent. Your calculator should say $3.95e^{15}$.
9.26×10^{-4}	Put 9.26 in as the coefficient number. Hit the "ee" or "exp" button. Locate the small +/- button for negative (not minus function). Next, enter the number four. Your calculator should say $9.26e^{-4}$.

This is an image of suggested calculators for CHM101. You will need a non-programmable, scientific calculator that does not have metric conversion functions. In this picture, you will see the correct way to input one of the numbers shown in the example above. Please take time to practice with your calculator before tests and quizzes.



Contributors and Attributions

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1.6: Dimensional Analysis

Learning Objective

- Convert metric and English based units.
- Report converted values in correct scientific notation with the corresponding units.
- Master using your calculator.

Conversion Factors

We have an expression,

$$\frac{3 \text{ ft}}{1 \text{ yd}} = 1 \quad (1.6.1)$$

Equation 1.6.1 is a strange way to write 1, but it makes sense: 3 ft equal 1 yd, so the quantities in the numerator and denominator are the same quantity, just expressed with different units. The expression in Equation 1.6.1 is called a conversion factor and it is used to formally change the unit of a quantity into another unit. (The process of converting units in such a formal fashion is sometimes called *dimensional analysis* or the *factor label method*.)

To see how this happens, let us start with the original quantity: 4 yd and let us multiply this quantity by 1. Rather than multiplying by just 1, let us write 1 as

$$1 = \frac{3 \text{ ft}}{1 \text{ yd}}$$

When you multiply anything by 1, you don't change the value of the quantity.

$$\begin{aligned} 4 \text{ yd} &= 4 \text{ yd} \times 1 \\ &= 4 \text{ yd} \times \frac{3 \text{ ft}}{1 \text{ yd}} \end{aligned}$$

The 4 yd term can be thought of as yd/1; that is, it can be thought of as a fraction with 1 in the denominator. We are essentially multiplying fractions. If the same thing appears in the numerator and denominator of a fraction, they cancel. In this case, what cancels is the unit yard:

$$4 \cancel{\text{ yd}} \times \frac{3 \text{ ft}}{1 \cancel{\text{ yd}}}$$

That is all that we can cancel. Now, multiply and divide all the numbers to get the final answer:

$$\frac{4 \times 3 \text{ ft}}{1} = \frac{12 \text{ ft}}{1} = 12 \text{ ft}$$

Again, we get an answer of 12 ft, just as we did originally. But in this case, we used a more formal procedure that is applicable to a variety of problems.

How many millimeters are in 14.66 m? To answer this, we need to construct a conversion factor between millimeters and meters and apply it correctly to the original quantity. One method would entail using the conversion factor of $1 \text{ mm} = 1 \times 10^{-3} \text{ m}$

$$14.66 \cancel{\text{ m}} \times \frac{1 \text{ mm}}{10^{-3} \cancel{\text{ m}}} = 14660 \text{ mm} \quad (1.6.2)$$

$$= 1.46 \times 10^4 \text{ mm} \quad (1.6.3)$$

Which conversion factor do we use? The answer is based on *what unit you want to get rid of in your initial quantity*. The original unit of our quantity is meters, which we want to convert to millimeters. Because the original unit is assumed to be in the numerator, to get rid of it, we want the meter unit in the *denominator*; then they will cancel. Using the positive metric exponents, we have an alternative method to calculate the same answer as shown previously. Canceling units and performing the mathematics, we get:

$$14.66 \cancel{m} \times \frac{1000 \text{ mm}}{1 \cancel{m}} = 14660 \text{ mm} \quad (1.6.4)$$

$$= 1.46 \times 10^4 \text{ mm} \quad (1.6.5)$$

Note how m cancels, leaving mm , which is the unit of interest.

The ability to construct and apply proper conversion factors is a very powerful mathematical technique in chemistry. You need to master this technique if you are going to be successful in this and future science courses.

✓ Example 1.6.1

- Convert 35.9 kL to liters.
- Convert 555 nm to meters.

Solution

- We will use the fact that 1 kL = 1,000 L. Of the two conversion factors that can be defined, the one that will work is 1000L/1kL. Applying this conversion factor, we get:

$$35.9 \cancel{kL} \times \frac{1000 \text{ L}}{1 \cancel{kL}} = 35,900 \text{ L}$$

- We will use the fact that 1 nm = 1/1,000,000,000 m, which we will rewrite as 1,000,000,000 nm = 1 m, or $10^9 \text{ nm} = 1 \text{ m}$. Of the two possible conversion factors, the appropriate one has the nm unit in the denominator:

$$\frac{1 \text{ m}}{10^9 \text{ nm}}$$

Applying this conversion factor, we get

$$555 \cancel{nm} \times \frac{1 \text{ m}}{10^9 \cancel{nm}} = 0.000000555 \text{ m} \\ = 5.55 \times 10^{-7} \text{ m}$$

In the final step, we expressed the answer in scientific notation.

? Exercise 1.6.1

- Convert 67.08 μL to liters.
- Convert 56.8 m to kilometers.

Answer a

$$6.708 \times 10^{-5} \text{ L}$$

Answer b

$$5.68 \times 10^{-2} \text{ km}$$

? Exercise 1.6.2

How many milliliters are in 607.8 kL?

Answer

This problem involves multiple steps that require going through the metric base of liters. Think of a pathway of a conversion before setting up these types of problems. The suggested pathway would be $\text{mL} \Rightarrow \text{L} \Rightarrow \text{kL}$.

$$\begin{aligned}
 607.8 \text{ kL} \times \frac{1000 \text{ L}}{1 \text{ kL}} &= 6.08 \times 10^5 \text{ L} \\
 &= 6.08 \times 10^5 \text{ L} \frac{1000 \text{ mL}}{1 \text{ L}} \\
 &= 608000000 \text{ mL} \\
 &= 6.08 \times 10^8 \text{ mL}
 \end{aligned}$$

? Exercise 1.6.3

Kathleen Hsu Barron (Furman University, Class of 2011) is exploring Switzerland with her husband Patrick Barron (Furman University, Class of 2011). The sign she is pointing to is a typical speed limit in km/hr. Convert this distance to miles/hr and feet/hr.



Figure 1.6.1: Kathleen Hsu Barron standing by speed limit sign. (CC-BY-NC-ND, Patrick Barron)

Answer

Again, this problem is a multi-step problem. There are multiple ways that one could proceed to solve this problem. With the conversion factors in this course, the suggested round would be miles $\Rightarrow$ kilometers $\Rightarrow$ meters $\Rightarrow$ centimeters $\Rightarrow$ inches $\Rightarrow$ feet. Another method could utilize miles $\Rightarrow$ feet.

18.6 miles/hr, 9.8×10^4 ft/hr

In this course, we will not use significant figures. You will need to read the directions on your assignment, quiz, or test and round accordingly.

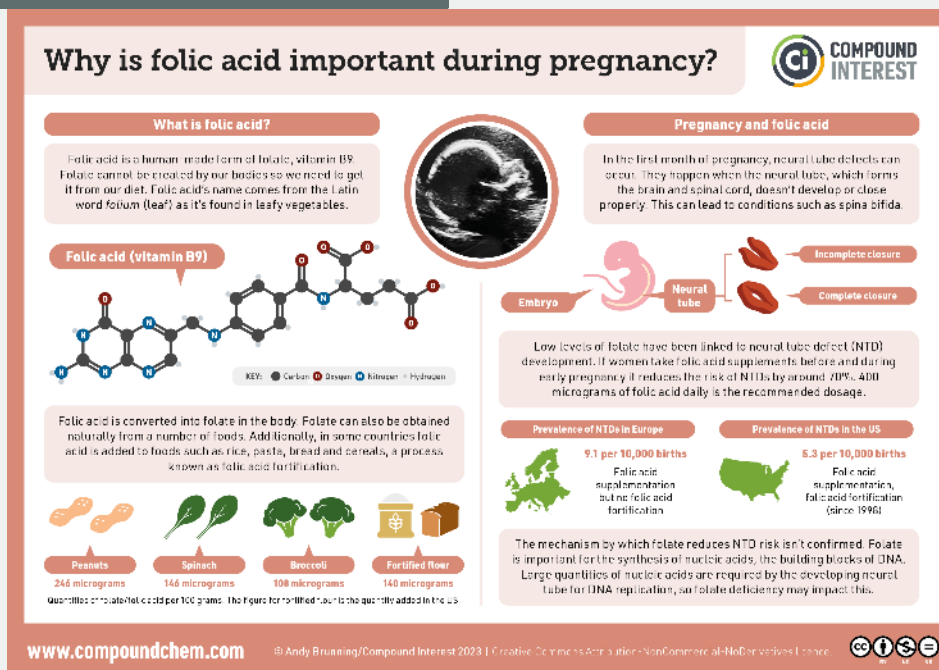


Figure 1.6.2: Why is folic acid important in pregnancy. (CC-BY-NC-ND, Andy Brunning).

Question: The FDA recommends taking 0.400 mg of folic acid a day before pregnancy. How many micrograms is this dosage? Hint- think of the pathway $\text{mg} \Rightarrow ? \Rightarrow \mu\text{g}$

Need More Practice?

- Turn to Section 2.E of this OER and work problems #1, #3, #5, #6, and #8.

Contributors and Attributions

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1.7: Density

Learning Objectives

- Calculate density, mass, or volume when given 2 of these three variables.
- Identify what units are required for the density equation.
- Review metric conversions.
- Compare densities of different chemical substances.
- Classify a substance as being a heterogeneous or homogeneous mixture if solubility data is provided.
- Identify where a chemical would appear in water if solubility and density data are provided.
- Compare any chemical substance's density to the density of water (please memorize this value).

Density is a physical property that is defined as a substance's mass divided by its volume:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (1.7.1)$$

$$d = \frac{m}{V} \quad (1.7.2)$$

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has *derived units*, just like velocity. Common units for density include g/mL, g/cm³, g/L, kg/L, and even kg/m³. Densities for some common substances are listed in Table 1.7.1. **Memorize the density of water with its appropriate units.**

Table 1.7.1: Densities of Some Common Substances

Substance	Density (g/mL or g/cm ³)
water	1.0
gold	19.3
mercury	13.6
air	0.0012
cork	0.22–0.26
aluminum	2.7
iron	7.87



Figure 1.7.1: Separatory Funnel containing oil and colored water to display density differences. (CC BY-SA 3.0; PRHaney via Wikipedia).

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3 . How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (Table 1.7.1), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

Start with Equation 1.7.1

$$\text{density} = \frac{m}{V}$$

and insert the relevant numbers

$$\frac{2.7 \text{ g}}{\text{cm}^3} = \frac{m}{7.88 \text{ cm}^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

$$7.88 \cancel{\text{cm}^3} \times \frac{2.7 \text{ g}}{\cancel{\text{cm}^3}} = 21 \text{ g of aluminum}$$

✓ Example 1.7.1: Mercury

What is the mass of 44.6 mL of mercury? If you need to review how to isolate variables, please click [here](#) for a short tutorial.

Solution

Use the density value for mercury from Table 1.7.1 and the definition of density (Equation 1.7.1)

$$\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow d = \frac{m}{V}$$

$$\frac{13.6 \text{ g}}{\text{mL}} = \frac{m}{44.6 \text{ mL}}$$

Remember to cross multiply here in order to isolate variable. Then, report answer with correct units.

$$44.6 \cancel{\text{mL}} \times \frac{13.6 \text{ g}}{\cancel{\text{mL}}} = 607 \text{ g}$$

The mass of the mercury is 607 g.

? Exercise 1.7.1

What is the mass of 25.0 cm^3 of iron?

Answer

Use the density value for iron from Table 1.7.1

$$\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow d = \frac{m}{V}$$

$$\frac{7.87 \text{ g}}{\text{cm}^3} = \frac{m}{25.0 \text{ cm}^3}$$

Cross multiplying both sides (right numerator $\times$ left denominator = left numerator $\times$ right denominator), we get the following expression with answer and appropriate unit.

$$25.0 \cancel{\text{cm}^3} \times \frac{7.87 \text{ g}}{\cancel{\text{cm}^3}} = 197 \text{ g of iron}$$

Another way of looking at density (some students choose to perform calculations using this method)

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

$$d = 19.3 \text{ g/mL} = \frac{19.3 \text{ g}}{\text{mL}}$$

Although this was not previously pointed out, it can be assumed that there is a 1 in the denominator:

$$d = 19.3 \text{ g/mL} = \frac{19.3 \text{ g}}{\text{mL}}$$

That is, the density value tells us that we have 19.3 grams for every 1 milliliter of volume, and the 1 is an exact number. When we want to use density to convert from mass to volume, the numerator and denominator of density need to be switched—that is, we must take the *reciprocal* of the density. In so doing, we move not only the units but also the numbers:

$$\frac{19.3 \cancel{\text{g}}}{\text{mL}} = \frac{45 \cancel{\text{g}}}{V}$$

Cross multiplying denominators with numerators, we obtain the following algebraic equation.

$$19.3V = 45.9 \text{ mL}$$

then you will need to isolate the variable (volume)

$$V = \frac{45.9 \text{ mL}}{19.3}$$

After multiplication, the answer would be

$$V = 2.38 \text{ mL}$$

✓ Example 1.7.2: Wine Cork

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/mL, what is the volume of the cork? *Regardless of the method that is used, you should still be able to obtain the same (and correct) answer.*

Solution

To use density as a conversion factor, we need to take the reciprocal so that the mass unit of density is in the denominator. Taking the reciprocal, we find

$$\frac{0.22 \text{ g}}{\text{mL}} = \frac{3.78 \text{ g}}{V}$$

Cross multiplying denominators with numerators, we obtain the following algebraic equation.

$$0.22V = 3.78 \text{ mL}$$

then you will need to isolate the variable (volume)

$$V = \frac{3.78 \text{ mL}}{0.22}$$

so, the volume of the cork is 17.2 mL.

? Exercise 1.7.2

What is the volume of 3.78 g of gold?

Answer

Before attempting this question, be sure to obtain the density of gold in the table above. If you were to need this value on a quiz or a test, then it would be provided. Once you have this value, plug it into the density equation. Next, you will need to isolate the volume variable (basic algebra). The final answer should be 0.196 cm³.

Care must be used with density as a conversion factor. Make sure the mass units are the same or the volume units are the same, before using density to convert to a different unit. Often, the unit of the given quantity must be first converted to the appropriate unit before applying density as a conversion factor.

Using Density in Environmental Applications

Along with solubility, density can help determine how a contaminant could affect an aquatic system. For example, imagine mercury has been spilled in Furman Lake. Looking at this element's density value and comparing it to liquid water, one could determine the location of the insoluble (you would be given solubility information) mercury layer. The more dense mercury layer would reside on the bottom of Furman Lake. If one were to take a cross-section of the lake, you could see that a heterogeneous mixture would result.

In contrast, spilling ethanol (density = 0.789g/mL) would result in the formation of a homogeneous mixture. Ethanol (grain alcohol) is soluble in water. This would make it miscible (mixable to form a solution) in water and one would not be able to denote separate layers. According to the density, an alcohol layer would remain on top, but would ultimately dissolve.

📌 Applications

What difficulties would arise from the separation and removal of contaminants?

1. Hg in Furman Lake
2. Ethanol in Furman Lake
3. oil (less dense, insoluble) in Furman Lake

Watch this [video](#) and record your observations.

1. What component was different in the two types of beverages (mass or volume)?
2. How does the above-mentioned difference affect the density equation?
3. Which beverage is denser than water?

Need More Practice?

- Turn to Section 2.E of this OER and work problems #2 and #9.

Contributors and Attributions

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CHAPTER OVERVIEW

2: Structure of Matter

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2.2: The Periodic Table of Elements

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2.3.1: Atomic Structure

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2.5: Molecules and Compounds

2.6: Solutions

2.6.1: Like Dissolves Like

2.6.2: Acids and Bases

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2.1: Classification and Phases of Matter

Learning Objectives

- Define matter, element, molecule, compound, mixture, homogeneous mixture and heterogeneous mixture.
- Describe the different states of matter.
- Be able to write and interpret symbols of elements, molecules, and compounds.

What is the matter?

Matter is anything that has mass and occupies space.

Matter is a natural material that makes up the universe. The matter comprises tiny particles called atoms held together by forces called bonds. The matter is classified as **pure** if it has a constant and a non-variable composition of the type of atoms. Pure matter is either an element or a compound.

Element

An element is pure matter that is composed of only one type of atom

For example, carbon, shown in Fig. 1.1.1, is an element. Elements can not convert to simpler matter by physical or chemical methods. There are around a hundred different elements known at this time. For example, hydrogen, oxygen, carbon, nitrogen, sodium, chlorine, iron, cobalt, gold, and silver are a few elements.

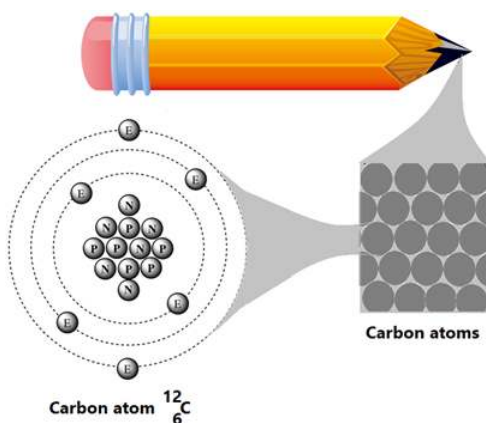


Figure 2.1.1: The pencil tip is carbon -an element composed of one type of atom, as shown in the model.

Symbol of an element

Symbols represent elements, the first alphabet of their English or non-English name, written in capital letters. For example, C for carbon, O for oxygen, and H for hydrogen. Usually, another alphabet is also chosen from the element's name and written as a small letter, e.g., He for helium, Co for cobalt. Some element symbols are derived from non-English names, e.g., Fe for iron is from its Latin name Ferrum, and Au for gold is from its Latin name Aurum.

Two consecutive capital letters do not represent an element; they may combine two elements.

CO is not a symbol of an element; it is a pure matter, a combination of carbon and oxygen bonded together in a 1:1 atom ratio.

Molecule

A molecule is a group of two or more atoms held together by forces called chemical bonds. The molecule is the smallest particle of matter that can exist freely. A single atom of some elements can exist freely and is also considered a molecule. For example, He, O₂, P₄, and S₈ are examples of elements having molecules composed of one, two, four, and eight atoms of the same element,

respectively, as illustrated in Fig. 1.1.2. Although metal elements exist as a vast number of atoms bonded together by a special type of bond called metallic bonds, their symbol is that of a single atom, e.g., Fe for iron and Au for gold.

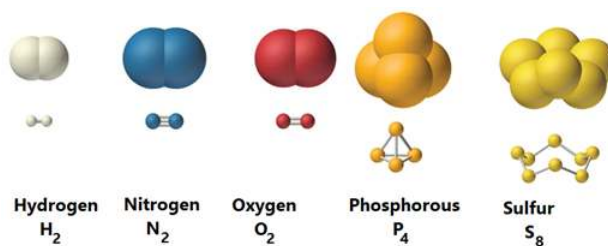


Figure 2.1.2: Models of molecules of some elements.

Compound

A compound is a pure matter composed of atoms of two or more different elements in a constant whole-number ratio held together by chemical bonds.

The symbol of a compound is a combination of its constituent elements with a subscript to the right of the element symbol representing the whole number ratio of the atoms of the element in the compound. For example, H_2O symbolizes a compound called water composed of hydrogen (H) and oxygen (O) atoms in a 2:1 ratio. Similarly, NaCl symbolizes a table salt compound composed of sodium (Na) and chlorine (Cl) atoms in a 1:1 ratio. If the symbol of a compound also represents a molecule of the compound, it is called a **molecular formula**. For example, H_2O is a molecule formula of water. On the other hand, table salt is another class of compound composed of a vast number of atoms of its constituent elements arranged in a specific arrangement in 3D space called a **crystal lattice**, as illustrated in Fig. 1.1.3. When the compound symbol does not represent a molecule, it only means the simple whole-number ratio of the constituent elements; it is called the **formula unit**. For example, NaCl is a chemical formula of a compound called table salt.

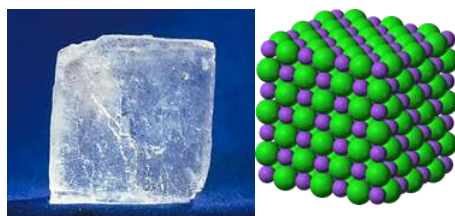


Figure 2.1.3: Sodium chloride (NaCl) – a compound, the crystal (left) and the model showing crystal lattice (right). Source: Halite(Salt)USGOV.jpg and Benjah-bmm27 (talk · contribs) / Public domain.

Law of definite proportion or law of constant composition

A chemical compound always contains the number of atoms of its component elements in a constant ratio, or the ratio of masses of the constituent elements is a constant in a given compound.

Pure matter, i.e., elements or compounds, has a fixed proportion of atoms of element/s independent of the source or method of their preparation. It is also called a **chemical** or a **substance**. The atoms in a compound are held together by attractive forces called chemical bonds. Constituent elements in a compound can be separated only by breaking chemical bonds and making new chemical bonds which is a **chemical reaction**.

Mixture

Two or more pure substances mixed such that ratio of the constituents is variable is called a mixture.

For example, table salt can be mixed with water, but the salt (NaCl) to water (H_2O) can be varied to give less salty or more salty water in a mixture. A mixture in which the components are thoroughly mixed, and the composition is constant throughout a given sample, is a **homogeneous mixture** or a **solution**. For example, table salt dissolved in water is an example of a homogeneous

mixture or a solution. Other homogeneous mixtures include sugar dissolved in water; the air, a mixture of nitrogen, oxygen, carbon dioxide, and other gases; and metal alloys like brass, a mix of copper and zinc metals, etc.



Figure 2.1.4: A heterogeneous mixture of iron and sulfur (left), separated by a magnet (right). Source: Asoult / CC BY (<https://creativecommons.org/licenses/by/4.0>)

It is a **heterogeneous** mixture if the constituents are not thoroughly mixed and the composition varies within different regions of a given sample. For example, sulfur (S) powder mixed with iron (Fe) filling is a heterogeneous mixture shown in Fig. 1.1.4. Other examples of heterogeneous mixtures include smoke which is a mixture of air and carbon particles; smog which is a mixture of liquid water droplets suspended in air; and orange juice which is a mixture of sugar, water, fiber particles, etc. Fig. 1.1.5 illustrates the classification of the matter described above.

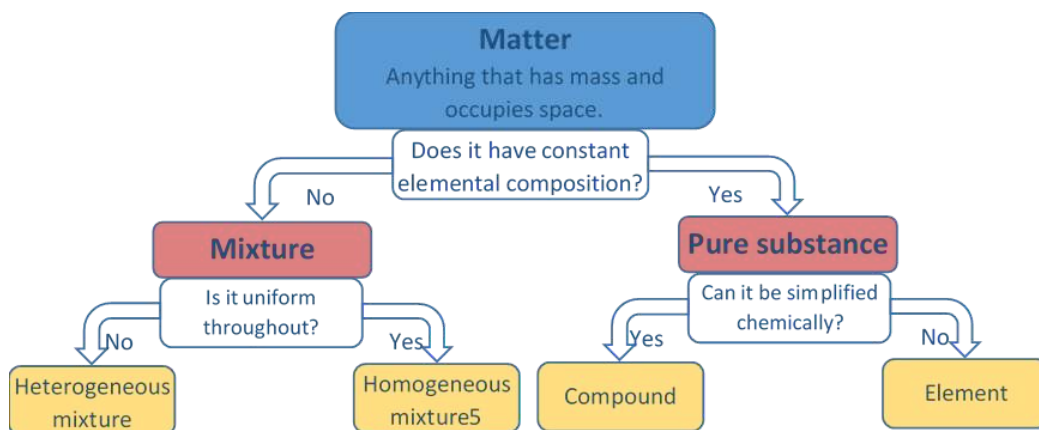


Figure 2.1.5: Classification of matter.

States of matter

Matter exists in one of the four physical states or phases, i.e., solid (s), liquid (l), gas (g), or plasma. Fig. 1.1.6 illustrates the four states of matter at the molecular level, and Fig. 1.1.7 shows examples of the four states of matter.

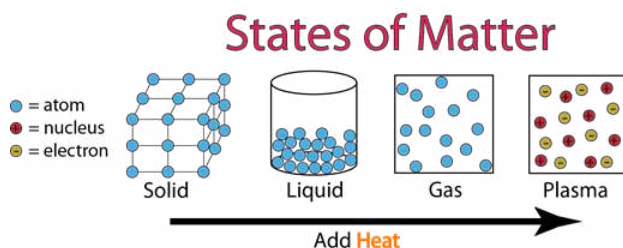


Figure 2.1.6: Four states of matter at the molecular level. Source: <https://www.hiclipart.com/free-trans...mwpjm/download>



Figure 2.1.7: Examples of the four states of matter. Clockwise from the top left are solid, liquid, plasma, and gas, represented by an ice sculpture, a drop of water, electrical arcing from a Tesla coil, and the air around clouds, respectively. Source: Spirit469 / CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0>)

Solid-state

In the solid state, the particles, i.e., atoms or molecules, are very close to each other and held strongly by intermolecular forces. The particles can vibrate around their mean positions, but they cannot slide past each other. Expansion and contraction in a solid state are negligible. The solid has a fixed shape and a fixed volume.

Liquid state

In the liquid state, the particles are close enough to experience strong intermolecular interactions that usually do not let the particles cross the liquid boundary. Still, the particles can move around within the liquid. Consequently, the particles in a liquid can flow and acquire the shape of the container but have a fixed volume. Expansion and contraction are negligible in the liquid state.

Gas state

In the gas state, the particles are far apart. The intermolecular interactions are negligible in the gas phase due to the large distances between the particles. The gas molecules move in straight lines in random directions until they collide with other molecules or the walls of the container. The collisions are elastic; that is, the molecules bounce off like elastic balls, and the total kinetic energy of the system is conserved. If exposed to space, the particles keep moving into space. In other words, the particles in a gas can flow, acquire the shape of the container, and expand or contract to fill up the available space. The gases do not have a fixed shape and do not have a fixed volume.

Plasma state

In the plasma state, the particles are far apart like gases, and a portion of the negative charge of the particles, i.e., electrons, are separated from the positive charge portion, i.e., the nucleus. In other words, the atoms in the plasma state are ionized. The plasma state is not common on Earth but is the universe's most common state of matter. For example, the matter in the sun and stars is in the plasma state. Examples of the plasma state on Earth include the matter in the lightning bolts and electrical sparks.

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2.2: The Periodic Table of Elements

Learning Objectives

- Number groups and periods of the periodic table
- Specify if an element is a metal, nonmetal or metalloids and its respective properties.
- Name the major groups of the periodic table.
- Identify the state of matter an element is in at ambient conditions.
- Define what an allotrope is.

Early developments

The discoveries of elements happened over a long time. As the list of known elements grew, scientists tried to arrange them based on their properties. **Mendeleev** arranged elements in a table based on atomic masses. It resulted in elements having similar properties placed next to each other in most cases. There were gaps intentionally left in the table for the elements that were predicted based on the knowledge from the periodic table but not yet discovered.

Few exceptions were there observed where the properties of the elements did not agree with the group in which they were placed based on their atomic masses. **Mosely** developed a method to measure atomic numbers based on X-ray spectroscopy. The arrangement of the elements based on the atomic number instead of atomic masses removed the discrepancies in Mendeleev's periodic table.

The modern periodic table of elements

- The modern periodic table of elements arranges the elements according to the increasing order of the atomic number starting from atomic number 1 for H and ending with atomic number 118 for Og, as shown in Fig. 2.4.1.
- The elements are arranged in horizontal rows called **periods** and vertical columns called **groups**.

Group→

Period↓

1

2

3

4

5

6

7

1

2

3

4

5

6

7

8

9

10

11

12

13

14

15

16

17

18

Metals

Metalloids

Nonmetals

Group
IA

Atomic Number

Symbol

Name

Atomic Mass

Font Black = solid Red = liquid Blue = gas

1

2

3

4

5

6

7

8

9

10

11

12

13

14

15

16

17

18

Alkali metals

Alkaline earth metals

Transition Metals

Post-Transition Metals

Halogens

Noble Gases

LANTHANIDES & ACTINIDES

Inner Transition metals

Figure 2.2.1: Periodic table of elements. Fort color of element symbols: blue are gases, red are liquids and black are solids at room temperature. Source: Modified from <https://www.exceltemplates.org/category/academic>)

Periods

The periodic table has seven horizontal rows called periods. The periods are numbered: 1 at the top to 7 at the bottom.

1. The 1st period has only two elements: hydrogen in group 1 and helium in group 18, with a gap from group 2 to group 17.
2. The 2nd and the 3rd periods have eight elements each, filling groups 1 and 2 followed by groups 13 to 18, leaving a gap from group 3 to group 12.
3. The periods 4th and 5th periods have eighteen elements that are filled successively from group 1 to group 18.
4. The periods 6th and 7th have 32 elements, each: the first two in groups 1 and 2, the next fourteen elements in separate rows below the table. These two rows of 14 elements each are called **Actinides** and **Lanthanoids**, respectively. Then the next sixteen elements fill the groups 3 to 18.

Groups

The periodic table has 18 vertical columns called groups or families. The groups are numbered starting from 1 on the leftmost and going through to 18 at the rightmost.

 Alkali metals

The 1st group is called alkali metals. The alkali metals include lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr), shown in Fig. 2.4.2. The alkali metals are the most reactive among the metals in the periodic table. They react vigorously with water, as shown in Fig. 2.4.3.



Figure 2.2.2: Alkali metals, from left to right: lithium, sodium, potassium, rubidium, and cesium. Source: Tomihahndorf at German Wikipedia, Dnn87 Contact email: Dnn87@yahoo.dk, and <http://images-of-elements.com/potassium.php>.

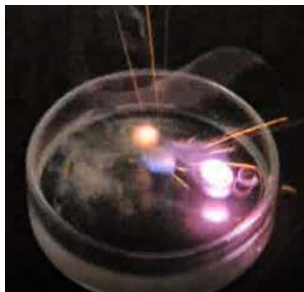


Figure 2.2.3: Potassium reacting with water. Source: Ozone aurora / Philip Evans / CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0>)

Hydrogen

Hydrogen is in group 1 but is not included in alkaline earth metals. Hydrogen is a nonmetal and has properties quite different from alkali metals or any other group of elements.

Alkaline earth metals

The 2nd group is called alkaline earth metals. It includes beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). They are reactive metals but less reactive than alkali metals. Alkaline earth metals impart characteristic color to a flame. Salts of alkali metals are used in firework formulation to give distinctive colors to the firework, as shown in Fig. 2.4.4.



Figure 2.2.4: New years eve fire works at Al Majaz waterfront, Sharjah UAE. Metal salts including alkali metals are used to give distinctive colors in fireworks. Source: Fariz Safarulla / CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0>)

Transition metals

Groups 3 to 12 are called transition metals. They include precious metals like gold, silver, platinum, and construction metals like iron. Some make catalysts and are found in enzymes and other bio-molecules, like hemoglobin and chlorophyll.

Group 13 to 16

Group 13 to group 16 does not have a unique name. They comprise nonmetals at the top and metals at the bottom of each group called **post-transition metals**. Important nonmetals include carbon, nitrogen, oxygen, phosphorous, and sulfur.

Halogens

Group 17 elements are called halogens. The halogens include fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). The halogens are highly reactive nonmetals. Chlorine is gas, bromine is liquid, and iodine is solid at room temperature, as shown in Fig. 2.4.5.



Figure 2.2.5: From left to right: chlorine, bromine, and iodine. Source: W. Oelen / CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0>)

Noble gases

Group 18 is called noble gases. They include helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are the least reactive of all the elements. Noble gases are used to create an inert atmosphere for chemical reactions. Noble gases are also used in the lighting system because of their chemically inert nature, as illustrated in Fig. 2.4.6.



Figure 2.2.6: Glowing noble gases in glass vials with low pressure inside. Power supply: 5 kV, 20 mA, 25 kHz. Source: New work Alchemist-hp (talk) www.pse-mendelejew.de; original single images: Jurii, <http://images-of-elements.com>. / CC BY (<https://creativecommons.org/licenses/by/3.0>)

Metals, metalloids, and nonmetals

Metals

The elements towards the right-bottom corner of the periodic table are metals except for hydrogen, which is a nonmetal. Metals have common characteristics, including:

1. they are shiny,
2. solid at room temperature (except mercury which is liquid),
3. malleable (can be hammered into sheets) and ductile (can be drawn into wires),
4. good conductors of heat and electricity, and
5. tend to lose electrons and form ionic compounds when they react with nonmetals.

The metals of group 1 are called alkali metals; group 2 are called alkaline earth metals, group 3 to group 13 are called transition metals, the two-row below transition metals in the periodic table are called **lanthanoids**, and **actinides** or **inner transition metals**, and metals in the group 13 to group 16 are called **post-transition metals**. Elements other than transition or inner-transition metals, i.e., the elements of groups 1 and 2 and groups 13 to 18, are collectively called **main group elements** or **representative elements**.

Metalloids

The dividing line between metals and nonmetals is a staircase line starting from $_{5}\text{B}$ and ending at $_{85}\text{At}$. The elements on the staircase line are metalloids except for aluminum and polonium, which are considered metals. Metalloids have properties in-between metals and nonmetals; e.g., they have moderate heat and electrical conductivity.

Nonmetals

Elements towards the top-right corner of the periodic table and hydrogen are called nonmetals.

1. The nonmetals usually have properties opposite to metals, e.g., they are not typically shiny, brittle if solid, and poor conductors of heat and electricity.
2. Nonmetals tend to make ionic compounds by accepting electrons from metals and making molecular compounds by reacting with each other.

Two groups in nonmetals also have unique names, i.e., group 17 is called halogens, and group 18 is called Noble gases.

The most reactive and nonreactive elements

Generally, alkali metals are the most reactive, followed by alkaline earth metals, and halogens are the most reactive nonmetals. Noble gases are the least reactive nonmetals, also called inert gases.

Fig. 2.4.7 shows examples of a metal, a metalloid, and a nonmetal.



Figure 2.2.7: Gold nugget –a metal, pure silicon –a metalloid, and sulfur –a nonmetal (from left to right). Source: Gold: Rob Lavinsky, iRocks.com – CC-BY-SA-3.0 / CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0>), Silicon: Enricoros at English Wikipedia / Public domain, Sulfur: Benjah-bmm27 / Public domain

Elements essential for life

The elements that are the main constituents of humans and other living organisms are oxygen, carbon, hydrogen, nitrogen, and sulfur. Phosphorous is present in bone, teeth, and DNA. Calcium and magnesium are the main constituents of bones and teeth and perform some other body functions. Sodium and potassium cations are the main electrolytes in body fluids, and chloride anion balances the charge. Iron is present in hemoglobin that carries oxygen to the cells. These elements are essential to life, and they are macronutrients. Besides these, several other elements are needed in a small amount. They are essential for life and called micronutrients. Fig. 2.4.8 shows the macronutrients in pink and micronutrients in blue color in a periodic table.

[illegible]

Physical and chemical states of elements

The term molecule is generally used for an electrically neutral group of two or more atoms held together by chemical bonds. In the kinetic molecular theory of gases, a **molecule** is the smallest particle of an element or compound with a stable and independent existence. Atoms of noble gases exist as independent species in the gas phase at room temperature, i.e., as **monoatomic molecules** like He, Ne, Ar, Kr, Xe, and Rn. Other elements that are gases at room temperature or **diatomic molecules**, i.e., H_2 , N_2 , O_2 , F_2 , and Cl_2 , are called dihydrogen, dinitrogen, dioxygen, and dichlorine, respectively. Note that usually prefix di is not used for the name of the element, i.e., these are usually called hydrogen, nitrogen, oxygen, fluorine, and chlorine, respectively. Monoatomic species of these elements, i.e., H, N, O, F, and Cl, exist, but they are very reactive species called free radicals, and they do not survive for a long time.

Two elements exist as liquids at room temperature, i.e., mercury (Hg), a metal, and bromine (Br₂), a non-metal that exists as diatomic molecules. Four elements, francium, cesium, gallium, and rubidium, are solid metals at 25 °C, but become liquid when the temperature is slightly warmer.

All elements not mentioned in the previous two sections are solid at room temperature. They may be diatomic or polyatomic molecules, e.g., I_2 , O_3 , P_4 , S_8 , diiodine, ozone, tetraphosphorous, and octasulfur, with two or three, four, and eight atoms, respectively, in a molecule.

Different forms of the same element in the same physical state are called allotropes.

The allotropes are different structural modifications of the element. For example, carbon exists in several allotropic forms; two of them are shown in Fig. 2.4.9. Another example is O_2 and O_3 are gaseous forms of oxygen.

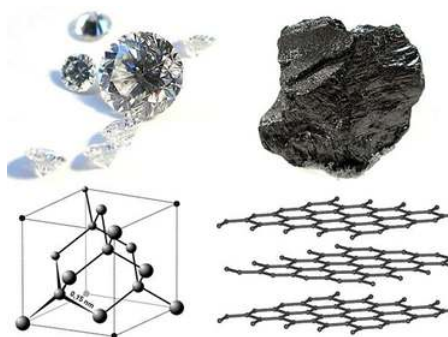


Figure 2.2.9: Diamond and graphite are two of the allotropes of carbon: pure forms of the same element that differ in structure.
Source: Diamond_and_graphite.jpg; User:Itubderivative work: Materialschemist / CC BY-SA
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Elements that exist as giant molecules

Some elements exist as **giant molecules**, i.e., a collection of many atoms bonded with each other through a 3D network of bonds. For example, carbon is a giant molecule in several allotropic forms, including diamond, graphite, carbon nanotubes, and fullerenes. Fig. 2.4.9 shows the bonding in diamond and graphite allotropes. The whole piece of diamond is one molecule with carbon atoms interconnected in a 3D network of bonds.

Metals also exist as a collection of many atoms bonded together by metallic bonds. The metal atoms exist as $+$ ions arranged in a well-defined 3D arrangement called crystal lattice with some of the outermost electrons roaming around in the whole piece of the metal, forming a sea of electrons around the metal atoms, as illustrated in Fig. 2.4.10.

The elements in giant molecules and metals are represented by element symbols without any subscript, e.g., C is carbon, Fe is iron, Au is gold, etc.

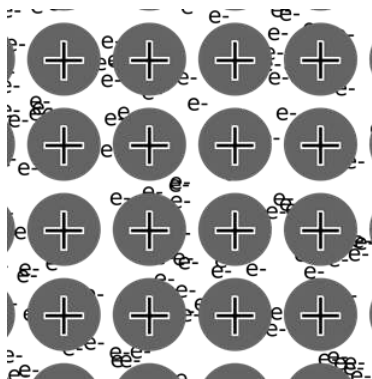


Figure 2.2.10: Metal ions in a sea of electrons. Source: Steven Legg (Sjlegg) / Public domain

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2.2.1: Chemical Formulas of Elements

Learning Objectives

- Explain how all matter is composed of atoms.
- Describe the modern atomic theory.
- Recognize which elements exist as diatomic molecules.

Take some aluminum foil. Cut it in half. Now you have two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil.

It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can you take this exercise, at least in theory? Can you continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? (Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century B.C.)

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 2.2.1.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

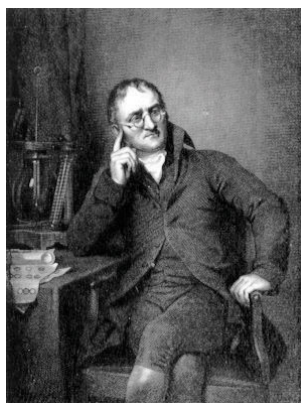


Figure 2.2.1.1 John Dalton was an English scientist who enunciated the modern atomic theory.

Most elements in their pure form exist as **individual atoms**. For example, a macroscopic chunk of iron (Fe) metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called **diatomic molecules**. In representing a diatomic molecule, we use the symbol of the element and include the **subscript "2"** to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H₂), nitrogen (N₂), fluorine (F₂), oxygen (O₂), iodine (I₂), chlorine (Cl₂) and bromine (Br₂). A few other elements can exist as 3-atom molecules like ozone (O₃) and 4-atom molecules like phosphorus (P₄). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is S₈ (Figure 2.2.1.2).

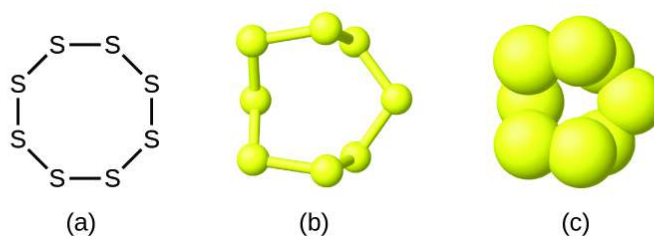


Figure 2.2.1.2: A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8 . It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

Figure A shows eight sulfur atoms, symbolized with the letter S, that are bonded to each other to form an octagon. Figure B shows a 3-D, ball-and-stick model of the arrangement of the sulfur atoms. The shape is clearly not octagonal as it is represented in the structural formula. Figure C is a space-filling model that shows each sulfur atom is partially embedded into the sulfur atom it bonds with.

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, H_2 and $2H$ represent distinctly different species. 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 $2H$, on the other hand, indicates two separate hydrogen atoms that are **not** combined as a unit. The expression $2H_2$ represents two molecules of diatomic hydrogen (Figure 2.2.1.3).

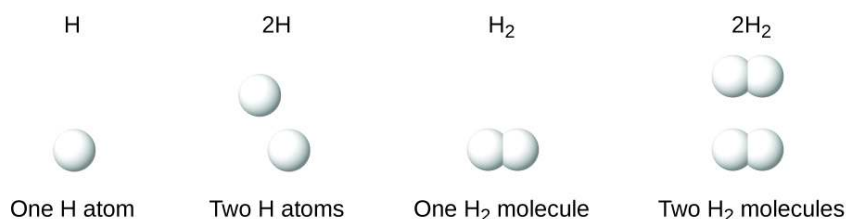


Figure 2.2.1.3: The symbols H , $2H$, H_2 , and $2H_2$ represent very different entities.

This figure shows four diagrams. The diagram for H shows a single, white sphere and is labeled one H atom. The diagram for $2H$ shows two white spheres that are not bonded together. It is labeled 2 H atoms. The diagram for H_2 shows two white spheres bonded together. It is labeled one H subscript 2 molecule. The diagram for $2H_2$ shows two sets of bonded, white spheres. It is labeled 2 H subscript 2 molecules.

✓ Example 2.2.1.1

Write the chemical formula of the following elements:

- oxygen
- carbon
- potassium

Solution

Memorizing the diatomic molecules is worthwhile for our future endeavors. A mnemonic device to help in the memorization of the diatomic elements is as follows: **Have No Fear Of Ice Cold Beer**. (hydrogen, nitrogen, fluorine, oxygen, iodine, chlorine and bromine).

- Oxygen is one of the seven diatomic molecular elements. Its chemical formula is O_2 .
- Carbon is monatomic, hence its formula is C.
- Potassium is monatomic hence its formula is K.

2.2.1.1Exercise

Write the chemical formula of the following elements:

- hydrogen
- nitrogen
- neon

Answer a

Hydrogen is one of the seven diatomic molecular elements. Its chemical formula is H_2 .

Answer b

Nitrogen is one of the seven diatomic molecular elements. Its chemical formula is N_2 .

Answer c

Neon is a monatomic element, hence its formula is Ne .

 Looking Closer: Atomic Theory

Dalton's ideas are called the *modern* atomic theory because the concept of atoms is very old. The Greek philosophers Leucippus and Democritus originally introduced atomic concepts in the fifth century BC. (The word *atom* comes from the Greek word *atomos*, which means “indivisible” or “uncuttable.”) Dalton had something that the ancient Greek philosophers didn't have, however; he had experimental evidence, such as the formulas of simple chemicals and the behavior of gases. In the 150 years or so before Dalton, natural philosophy had been maturing into modern science, and the scientific method was being used to study nature. So when Dalton announced a modern atomic theory, he was proposing a fundamental theory to describe many previous observations of the natural world; he was not just participating in a philosophical discussion.

Concept Review Exercises

1. What is the modern atomic theory?
2. What are atoms?

Answers

1. The modern atomic theory states that all matter is composed of atoms.
2. Atoms are the smallest parts of an element that maintain the identity of that element.

Key Takeaways

- **Atoms** are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of **atoms** and how they compose matter.
- The **diatomic** elements are hydrogen, nitrogen, fluorine, oxygen, iodine, chlorine and bromine.

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2.3: Atomic Theory

Learning Objectives

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

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

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

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

These concepts form the basis of chemistry. Although the word *atom* comes from a Greek word that means "indivisible," we understand now that atoms themselves are composed of smaller parts called *subatomic particles*. The first part to be discovered was the electron, a tiny subatomic particle with a negative charge. It is often represented as e^{-} , with the right superscript showing the negative charge. Later, two larger particles were discovered. The proton is a more massive (but still tiny) subatomic particle with a positive charge, represented as p^{+} . The neutron is a subatomic particle with about the same mass as a proton, but no charge. It is represented as either n or n^0 . We now know that all atoms of all elements are composed of electrons, protons, and (with one exception) neutrons. Table 2.3.1 summarizes the properties of these three subatomic particles.

Table 2.3.1: Properties of the Three Subatomic Particles

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

How are these particles arranged in atoms? They are not arranged at random. Experiments by Ernest Rutherford in England in the 1910s pointed to a nuclear model with atoms that has the protons and neutrons in a central nucleus with the electrons in orbit about the nucleus. The relatively massive protons and neutrons are collected in the center of an atom, in a region called the nucleus of the atom (plural *nuclei*). The electrons are outside the nucleus and spend their time orbiting in space about the nucleus. (Figure 2.3.1).

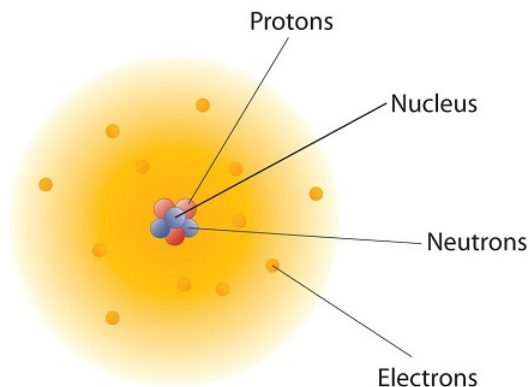


Figure 2.3.1: The Structure of the Atom. Atoms have protons and neutrons in the center, making the nucleus, while the electrons orbit the nucleus.

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

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

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

✓ Example 2.3.1

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

Solution

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

? Exercise 2.3.1

The number of protons in the nucleus of a tin atom is 50, while the number of neutrons in the nucleus is 68. What are the atomic number and the mass number of this isotope?

Answer

Atomic number = 50, mass number = 118

When referring to an atom, we simply use the element's name: the term *sodium* refers to the element as well as an atom of sodium. But it can be unwieldy to use the name of elements all the time. Instead, chemistry defines a symbol for each element. The atomic symbol is a one- or two-letter representation of the name of an element. By convention, the first letter of an element's symbol is always capitalized, while the second letter (if present) is lowercase. Thus, the symbol for hydrogen is H, the symbol for sodium is Na, and the symbol for nickel is Ni. Most symbols come from the English name of the element, although some symbols come from an element's Latin name. (The symbol for sodium, Na, comes from its Latin name, *natrium*.) Table 2.3.2 lists some common elements and their symbols. You should memorize the symbols in Table 2.3.2, as this is how we will be representing elements throughout chemistry.

Table 2.3.2: Names and Symbols of Common Elements

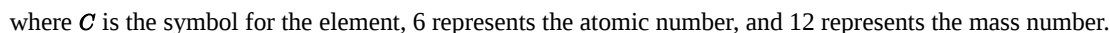
Element Name	Symbol	Element Name	Symbol
Aluminum	Al	Mercury	Hg
Argon	Ar	Molybdenum	Mo
Arsenic	As	Neon	Ne
Barium	Ba	Nickel	Ni
Beryllium	Be	Nitrogen	N
Bismuth	Bi	Oxygen	O
Boron	B	Palladium	Pd
Bromine	Br	Phosphorus	P
Calcium	Ca	Platinum	Pt
Carbon	C	Potassium	K
Chlorine	Cl	Radium	Ra
Chromium	Cr	Radon	Rn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

The elements are grouped together in a special chart called the periodic table of all the elements. A simple periodic table is shown in Figure 2.3.2, while one may view a more extensive periodic table from another source. The elements on the periodic table are listed in order of ascending atomic number. The periodic table has a special shape that will become important to us when we consider the organization of electrons in atoms (Chapter 8). One immediate use of the periodic table helps us identify metals and nonmetals. Nonmetals are in the upper right-hand corner of the periodic table, on one side of the heavy line splitting the right-side part of the chart. All other elements are metals.

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

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

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



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

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

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

11 protons

<https://chem.libretexts.org/@go/page/491413>

Key Takeaways

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

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2.3.1: Atomic Structure

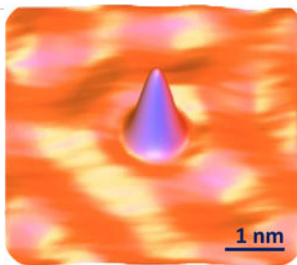


Figure 2.3.1.1 (Public Domain; NIST, Joseph Stroscio et. al. via [Commons Wikimedia, STM Image of Single Co Atom](https://commons.wikimedia.org/wiki/File:STM_Image_of_Single_Co_Atom) [commons.wikimedia.org])

What could this hilly blue surface possibly be? Do you have any idea? The answer is a single atom of the element cobalt. The picture was created using a scanning tunneling microscope. No other microscope can make images of things as small as atoms. How small are atoms? You will find out in this lesson.

What Are Atoms?

Atoms are the building blocks of matter. They are the smallest particles of an element that still have the element's properties. Elements, in turn, are pure substances—such as nickel, hydrogen, and helium—that make up all kinds of matter. All the atoms of a given element are identical in that they have the same number of protons, one of the building blocks of atoms (see below). They are also different from the atoms of all other elements, as atoms of different elements have different number of protons.

Size of Atoms

Unlike bricks, atoms are extremely small. The radius of an atom is well under 1 nanometer, which is one-billionth of a meter. If a size that small is hard to imagine, consider this: trillions of atoms would fit inside the period at the end of this sentence. Although all atoms are very small, elements vary in the size of their atoms. The figure below compares the sizes of atoms of more than 40 different elements. The elements in the figure are represented by chemical symbols, such as H for hydrogen and He for helium. Of course, real atoms are much smaller than the circles representing them in the figure below.

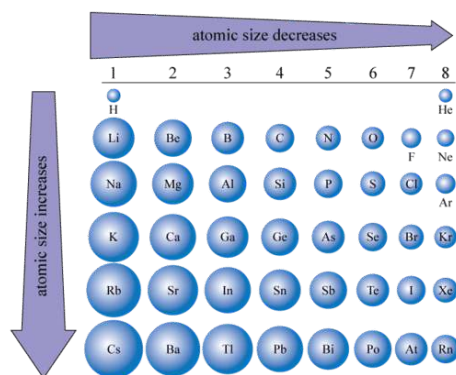


Figure 2.3.1.2 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: Which element in the figure above has the biggest atoms?

A: The element in the figure with the biggest atoms is cesium (Cs).

Subatomic Particles

Although atoms are very tiny, they consist of even smaller particles. Three main types of particles that make up all atoms are:

- protons, which have a positive electric charge.
- electrons, which have a negative electric charge.
- neutrons, which are neutral in electric charge.

The model in the figure below shows how these particles are arranged in an atom. The particular atom represented by the model is helium, but the particles of all atoms are arranged in the same way. At the center of the atom is a dense area called the nucleus, where all the protons and neutrons are clustered closely together. The electrons constantly move around the nucleus. Helium has two protons and two neutrons in its nucleus and two electrons moving around the nucleus. Atoms of other elements have different numbers of subatomic particles, but the number of protons always equals the number of electrons. This makes atoms neutral in charge because the positive and negative charges "cancel out."

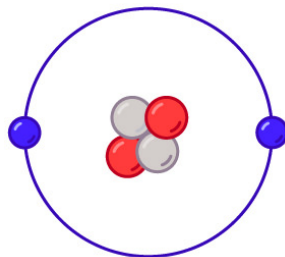


Figure 2.3.1.3: Model of a helium atom. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Q: Lithium has three protons, four neutrons, and three electrons. Sketch a model of a lithium atom, similar to the model above for helium.

A: Does your sketch resemble the model in the figure below? The model has three protons (blue) and four neutrons (gray) in the nucleus, with three electrons (red) moving around the nucleus.

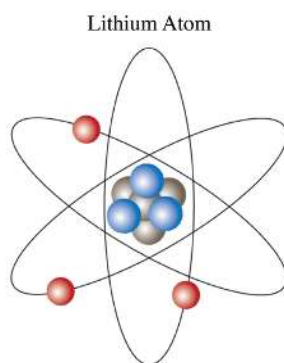


Figure 2.3.1.4: Copy and Paste Caption here. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: All atoms of carbon have six protons. How many electrons do carbon atoms have?

A: Carbon atoms must have six electrons to "cancel out" the positive charges of the six protons. That's because atoms are always neutral in electric charge.



Summary

- Atoms are the building blocks of matter. They are the smallest particles of an element that still have the element's properties.
- All atoms are very small, but atoms of different elements vary in size.
- Three main types of particles that make up all atoms are protons, neutrons, and electrons.

Review

1. What is an atom?
2. Which of the following statement(s) are true about the atoms of any element?
 - a. The number of protons in an atom of an element is unique to each element.
 - b. The number of protons and neutrons in an atom of an element is unique to each element
 - c. A proton in an atom of one element is identical to a proton in an atom of another element.
 - d. The number of protons in an atom of an element is the same for all elements.
3. Which of the following statements explains why atoms are always neutral in charge
 - a. They have the same number of protons as the atoms of all other elements.
 - b. They have protons that are identical to the protons of all other elements.
 - c. They have the same size as the atoms of all other elements.
 - d. They have the same number of protons as electrons.

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2.3.1.1: The Bohr Model of Atoms

Learning Objectives

- Describe the arrangement of electrons using the shell model.

Building Atoms: Main Shells

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

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

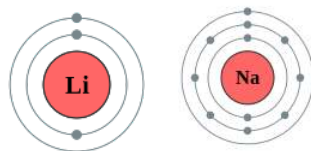


Figure 2.3.1.1.10 A shell diagram of lithium (left) and Sodium (right)

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

Drawing Shell Models



Video 2.3.1.1.1 How to draw the shell model for sulfur.

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

Table 2.3.1.1.1 shows the number of electrons that fill each shell for neutral atoms of several elements. As mentioned earlier, the innermost shell (corresponding to lowest energy) is filled first and only a fixed number of electrons is allowed in each shell. The only electron in hydrogen ($Z=1$) goes to the first shell. In lithium atom ($Z=3$), the two electrons fill the first shell, and the third electron goes to the second shell. An argon atom ($Z=18$) has 18 electrons. The 10 electrons fill the first and second shells, and the remaining 8 electrons go to the third shell. The electron configuration for elements pass argon are covered in more detail in section 3.7.

Table 2.3.1.1.1 Electron Configuraion (Arrangement) of Several Elements.

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

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

Summary

- The shell model is a good representation of electron arrangement only for elements 1-18.

Contributors and Attributions

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2.3.1.2: Valence Electrons

Learning Objectives

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

Valence Electrons

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

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

The number of valence electrons of an element can be determined by the periodic table group (vertical column) in which the element is categorized (Table 2.3.1.2.1). With the exception of groups 3–12 (the transition metals), the units digit of the group number identifies how many valence electrons are associated with a neutral atom of an element listed under that particular column.

Table 2.3.1.2.1: Number of Valence Electrons for Main Periodic Table Groups.

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

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

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

 Valence Electrons and the Periodic Table



Video 2.3.1.2.1 Finding the valence electron for an element.

Summary

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

References

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3. Petrucci, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

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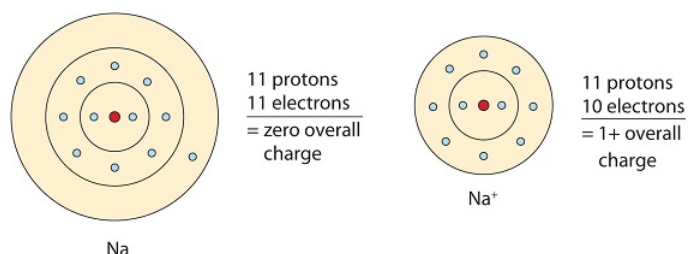
2.4: The Octet Rule

Learning Objective

- Define the octet rule.

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

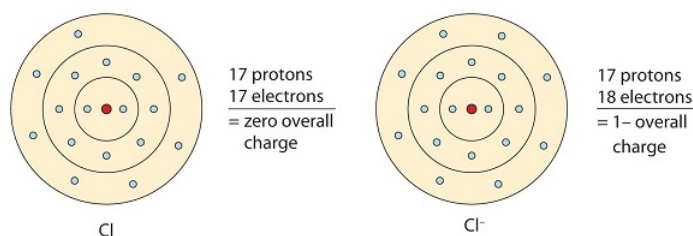
Referring to the octet rule, atoms attempt to get a noble gas electron configuration, which is eight valence electrons. Metals have few electrons in their outer-most orbitals. By losing those electrons, these metals can achieve noble gas configuration and satisfy the octet rule. Sodium has one valence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



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

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

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



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

Summary

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

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Organic Chemistry Supplemental_Modules_(Organic_Chemistry)/Fundamentals/Ionic_and_Covalent_Bonds

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2.4.1: Periodic Trends

Properties of elements generally show a periodic trend that correlates with their position in the periodic table. The properties and their periodic trends are described below.

Valence electrons

Electrons in the outermost shell are the valence electrons. Fig. 2.5.8 shows the valence electrons of the first twenty elements in red fonts.

Periodic trend of valence electrons

All elements in a group have the same number of valence electrons equal to the first digit of their group number. For example, 1st group of hydrogen and alkali metals has one valence electron, 2nd group of alkali metals has two valence electrons, halogens in 17th group have seven valence electrons, and noble gases in 18th group have eight valence electrons.

Caution

The transition metals in groups 3 to 12, and inner-transition metals, i.e., lanthanoids and actinoids -the two rows of elements placed below the periodic table, are the exception to the general trend of valence electrons described above. Valence electron configurations of transition metals and inner-transition metals are not described here; it is beyond the scope of this book.

The valence electrons mainly determine the chemical properties of the elements. The elements in the same group have similar chemical properties because they have the same valence shell electron configuration. The elements in a row show a gradual change in chemical properties because their valence shell electron configuration changes gradually along the row.

Lewis symbols

1. Lewis symbols show the valence electrons as dots around the symbol of an element. One dot represents one valence electron, e.g., $\text{H}^\bullet$.
2. The dots are shown on any of the four sides of the symbol.
3. A single dot on the top, bottom, left, or right is shown four valence electrons. Then start pairing the dots beyond four valence electrons, as shown in Fig. 2.6.1 for the first twenty elements.
4. Helium is an exception that has only two valence electrons, but they are shown paired.

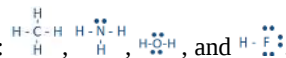
Group → Row ↓	1	2	13	14	15	16	17	18
1	$\text{H}^\bullet$							$\text{He}^\bullet$
2	$\text{Li}^\bullet$	$\text{Be}^\bullet$	$\text{B}^\bullet$	$\text{C}^\bullet$	$\text{N}^\bullet$	$\text{O}^\bullet$	$\text{F}^\bullet$	$\text{Ne}^\bullet$
3	$\text{Na}^\bullet$	$\text{Mg}^\bullet$	$\text{Al}^\bullet$	$\text{Si}^\bullet$	$\text{P}^\bullet$	$\text{S}^\bullet$	$\text{Cl}^\bullet$	$\text{Ar}^\bullet$
4	$\text{K}^\bullet$	$\text{Ca}^\bullet$						

Figure 2.4.1.1: Lewis symbols or electron-dot symbols of the first twenty elements in the periodic table.

The electron dots in the Lewis structure are a convenient way to determine how many bonds an atom of an element can make.

Generally, each unpaired dot can make one bond.

For example, a hydrogen atom with one unpaired dot can make one bond as in $\text{H}-\text{H}$. A bond is represented by a line between the bonded atoms. A bond is formed by sharing unpaired valence electrons. It is called a covalent bond. Carbon, nitrogen, oxygen, and fluorine with 4, 3, 2, and 1 unpaired dot can make 4, 3, 2, and 1 bond, e.g., in the following molecules:



- Each line in these molecules represents a **bonding electron pair**, and

- the pair of dots represent valence electrons that are not involved in bonding, called **lone pair of electrons**.

Atomic size

Electrons exist around the nucleus in a cloud-like appearance with no clearly defined boundaries. Therefore, the atomic size generally refers to the **covalent radius** of an atom that is one-half of the distance between the nuclei covalently bonded in a homonuclear molecule, like Cl_2 , I_2 , H_2 , as illustrated in Fig. 2.6.2.

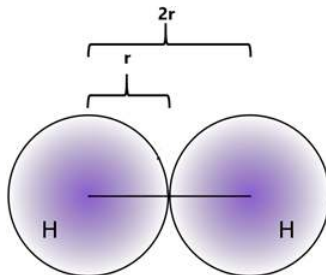


Figure 2.4.1.2: Diagram showing the atomic radius of H_2 . Source: modified from CK-12 Foundation / CC BY-SA (<https://creativecommons.org/licenses/by-sa/3.0>)

Periodic trend of atomic sizes

- The atomic size generally increases from top to bottom in a group because the valence electrons add in the higher shell in each consecutive member of a group down the column.
- The atomic size generally decreases from left to right in a row because the valence electrons are in the same shell while more protons add to the nucleus, increasing the pull on the valence electrons, as illustrated in Fig. 2.6.3.

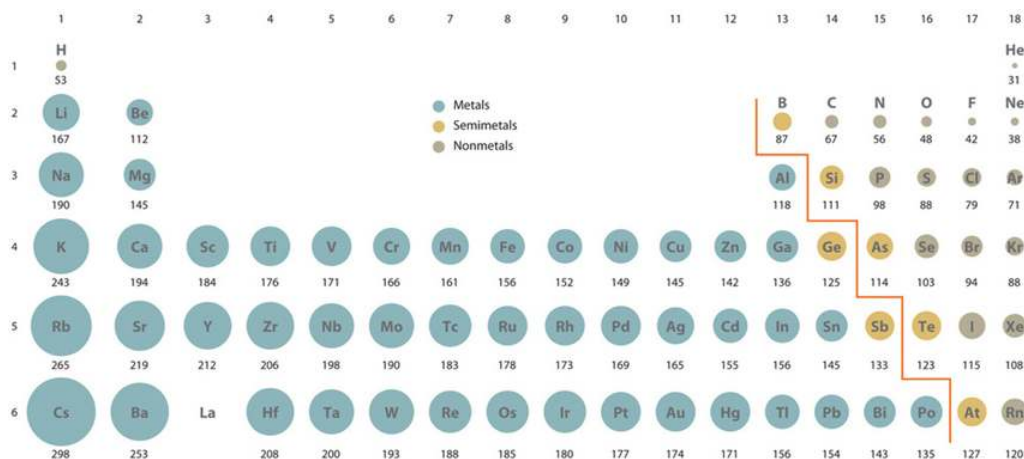


Figure 2.4.1.3: Illustration of the periodic trend of atomic sizes, Calculated Atomic Radii (in Picometers). Source: <https://chem.libretexts.org/@api/dek...jpg?revision=1>

Ionization energy

The positively charged nucleus attracts the negatively charged electrons. Therefore removal of an electron from the atom requires energy. The ionization produces a cation with fewer electrons than the parent neutral atom, i.e., cations.

Ionization energy

Ionization energy is the energy needed to remove an electron from a neutral atom, as in the following reaction.



Periodic trend in ionization energy

1. The ionization energy generally decreases from top to bottom in a column because the valence electrons are further away and experience less pull to the nucleus down the column.
2. The ionization energy generally increases from left to right in a row because the valence electrons are in the same shell while more protons add to the nucleus, which increases the pull on the valence electrons.

Fig. 2.6.4 illustrates the periodic trend in the ionization energy.

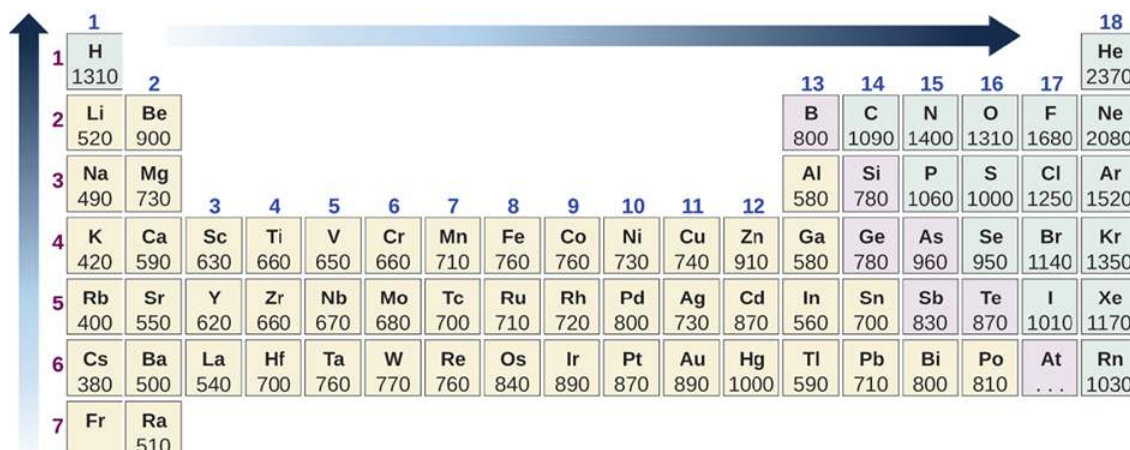


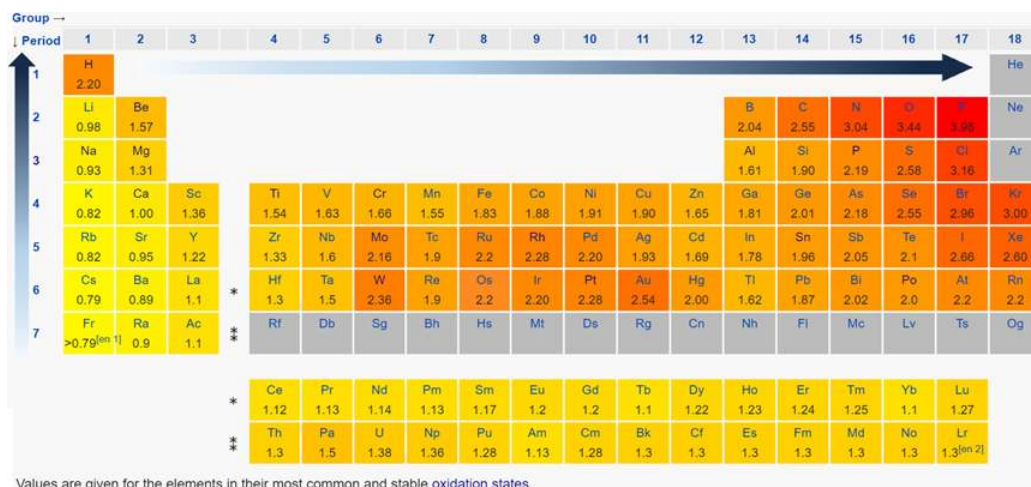
Figure 2.4.1.4: The 1st ionization energy in kJ/mol. Source: Download for free at <https://openstax.org/details/books/chemistry>

Electronegativity

Definition: Electronegativity

Electronegativity is the ability of an atom in a compound to attract the bonding electron pair to itself. Electronegativity is a property of an atom in a compound, i.e., a bonded atom, not a feature of an individual atom.

There are several electronegativity scales. The most commonly used is the Pauling electronegativity scale. Fig. 2.6.5 shows the electronegativities values on Pauling's electronegativity scale.



Values are given for the elements in their most common and stable oxidation states.

Figure 2.4.1.5: Periodic table of elements by Pauling scale. Source: <https://en.Wikipedia.org/wiki/Electronegativity>

Periodic trend in electronegativity

1. The electronegativity generally decreases from top to bottom in a column because the atomic size increases down the column making the nucleus less effective in pulling the bonding electrons.

- The electronegativity generally increases from left to right in a row because the atomic size decreases from left to right in a row, making the nucleus more effective in pulling the bonding electron pair to itself.

Summary of the periodic trends

The ionization energy and the electronegativity generally increase from left to right in a row and from bottom to top in a column. The atomic size and the metallic character are opposite, i.e., they increase from right to left in a row and from top to bottom in a column. Fig. 2.6.6 summarizes the periodic trend in the properties of the elements.

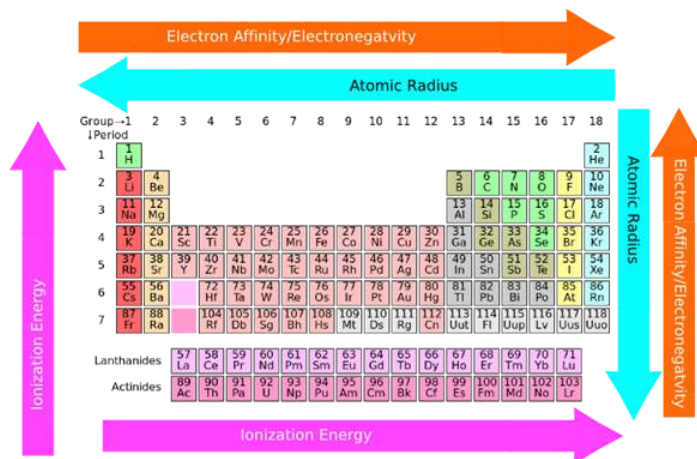


Figure 2.4.1.6: Summary of the periodic trend in the properties of elements. Source: <https://www.hiclipart.com/free-trans...ntzga/download>

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2.5: Molecules and Compounds

The tendency for two or more elements to combine and form a molecule that is stabilized by covalent bonds (a molecular compound) can be predicted simply by the location of the various elements on the periodic table. So far, we have divided the elements in the periodic table into (seemingly) arbitrary groupings; the metals, the non-metals, the semi-metals, and so on. These groupings are not arbitrary, but are largely based on physical properties and on the tendency of the various elements to bond with other elements by forming either an ionic or a covalent bond. As a general rule of thumb, compounds that involve a metal binding with either a non-metal or a semi-metal will display ionic bonding. Compounds that are composed of only non-metals or semi-metals with non-metals will display covalent bonding and will be classified as molecular compounds. Thus, the compound formed from sodium and chlorine will be ionic (a metal and a non-metal). Nitrogen monoxide (NO) will be a covalently bound molecule (two non-metals), silicon dioxide (SiO₂) will be a covalently bound molecule (a semi-metal and a non-metal) and MgCl₂ will be ionic (a metal and a non-metal). Later in this chapter we will see that many covalent compounds have bonds that are highly **polarized** with greater electron density around one atom than the other.

? Exercise 2.5.1

- a. Determine whether each of the following compounds is likely to exist as a molecule, or as an ionic compound:
 - i. Hydrogen fluoride; HF
 - ii. Silicon tetrachloride; SiCl₄
 - iii. Elemental sulfur as S₈
 - iv. Disodium dioxide; Na₂O
 - v. PF₃
 - vi. Be₃N₂
 - vii. AlP
 - viii. CBr₄

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2.6: Solutions

Learning Objectives

- Describe the difference between homogenous and heterogenous mixtures.
- Distinguish a homogenous mixture as a solution or colloid.

As introduced previously, *mixtures* are combinations of two or more substances that each retain their individual physical properties. A mixture can be classified as either *heterogenous* or *homogenous*. In a heterogeneous mixture, the composition is *not* uniform throughout the sample, and sometimes the individual substances that make up the mixture can be differentiated by eye. Trail mix, salad, and blood (which is also called a suspension) are examples of heterogenous mixtures. Homogenous mixtures are uniform and have the same composition throughout. Air, simple syrup, and seawater are examples of homogenous mixtures.

Homogenous mixtures can be further classified, based on the *size* of their particles, as solutions or colloids. **Solutions** are composed of particles the size of an ion or small molecule, $\sim 0.1\text{-}2.0\text{ nm}$. The examples provided above are all considered solutions; air is a solution of small gas molecules, simple syrup is a solution of sucrose in water, and seawater is a solution of ions and water. Homogenous mixtures with larger particles, $\sim 2.0\text{-}500\text{ nm}$, are classified as **colloids**. Milk, fog, and butter are all considered colloids. Mixtures with particles larger than 500 nm are called **suspensions** and are considered to be heterogenous mixtures in which the particles will settle upon standing. Many medications are classified as suspensions that need to be re-mixed before taking in order to redistribute the particles throughout the mixture. Some examples and distinguishing characteristics of solutions, colloids, and suspensions are listed in Table 2.6.1 below.

Table 2.6.1: Properties of Liquid Solutions, Colloids, and Suspensions

Type of Mixture	Approximate Size of Particles (nm)	Characteristic Properties	Examples
solution	< 2	not filterable; does not separate on standing; does not scatter visible light	air, white wine, gasoline, salt water
colloid	$2\text{-}500$	scatters visible light; translucent or opaque; not filterable; does not separate on standing	smoke, fog, ink, milk, butter, cheese
suspension	$500\text{-}1000$	cloudy or opaque; filterable; separates on standing	muddy water, hot cocoa, blood, paint

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

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table 2.6.2 lists some common types of solutions, with examples of each.

Table 2.6.2: Types of Solutions

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

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2.6.1: Like Dissolves Like

Learning Objectives

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

Like Dissolves Like

A simple way to predict which compounds will dissolve in other compounds is the phrase "like dissolves like". What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve in water but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, an nonpolar molecules, does dissolve in water and it is this oxygen that the fish take in through their gills. Or, one more example of a nonpolar compound that dissolves in water is the reason we can enjoy carbonated sodas. Pepsi-cola and all the other sodas have carbon dioxide gas, CO_2 , a nonpolar compound, dissolved in a sugar-water solution. In this case, to keep as much gas in solution as possible, the sodas are kept under pressure.

This general trend of "like dissolves like" is summarized in the following table:

Table 2.6.1.2 Summary of Solubilities

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

Note that every time charged particles (ionic compounds or polar substances) are mixed, a solution is formed. When particles with no charges (nonpolar compounds) are mixed, they will form a solution. However, if substances with charges are mixed with other substances without charges a solution does not form.

When an ionic compound is considered "insoluble", it doesn't necessarily mean the compound is completely untouched by water. All ionic compounds dissolve to some extent. An insoluble compound just doesn't dissolve in any noticeable or appreciable amount.

What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding.



Figure 2.6.1.1 Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water. *Liquids that dissolve in one another in all proportions are said to be **miscible**. Liquids that do not dissolve in one another are called **immiscible**.*

Figure 2.6.1.2 Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

Summary

- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.
- Liquids that dissolve in one another in all proportions are said to be **miscible**.
- Liquids that do not dissolve in one another are called **immiscible**.

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2.6.2: Acids and Bases

Learning Objectives

- Examine properties of acids and bases.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling. Figure 2.6.2.1 provides some examples of acids and bases.



Figure 2.6.2.1 Some examples of common household acids and bases.

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

- Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current.
- Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
- Acids change the color of certain acid-base indicators. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
- Acids react with active metals to yield hydrogen gas.
- Acids react with bases to produce a salt compound and water.

Common examples: Lemons, oranges, vinegar, urine, sulfuric acid, hydrochloric acid

Bases

Bases have properties that mostly contrast with those of acids.

- Aqueous solutions of bases are also electrolytes.
- Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
- Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink.
- Bases do not react with metals in the way that acids do.
- Bases react with acids to produce a salt and water.

Common Examples: Soap, toothpaste, bleach, cleaning agents, limewater, ammonia water, sodium hydroxide.



Figure 2.6.2.2 Phenolphthalein indicator in presence of base.

Warning

Tasting chemicals and touching them are NOT good lab practices and should be avoided - in other words, don't do this at home.

Summary

- A brief summary of properties of acids and bases are given.
- Properties of acids and bases mostly contrast each other.

Contributors and Attributions

-
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- Wikipedia
- [Marisa Alviar-Agnew](#) (Sacramento City College)
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2.6.2.1: The pH Scale

Learning Objectives

- Define **pH**.
- Determine the pH of acidic and basic solutions.

The **pH** scale ranges in values from 0 to 14 that describes the acidity or basicity of a solution. You can use **pH** to make a quick determination whether a given aqueous solution is acidic, basic, or neutral. The lower the pH, the more acidic it is.

pH is a *logarithmic* scale. A solution that has a pH of 1.0 is 10 times as acidic as a solution with a pH of 2.0, which in turn is 10 times as acidic as a solution with a pH of 3.0 and so forth. If you live in California or somewhere that is known for earthquakes, this may sound familiar to you! The Richter scale, which is used for judging the strength of an earthquake, is also a logarithmic scale! That's why a 5.0 earthquake is so much stronger than a 4.0!

Warning: The pH scale has no limits

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

- If $pH < 7$, then the solution is acidic.
- If $pH = 7$, then the solution is neutral.
- If $pH > 7$, then the solution is basic.

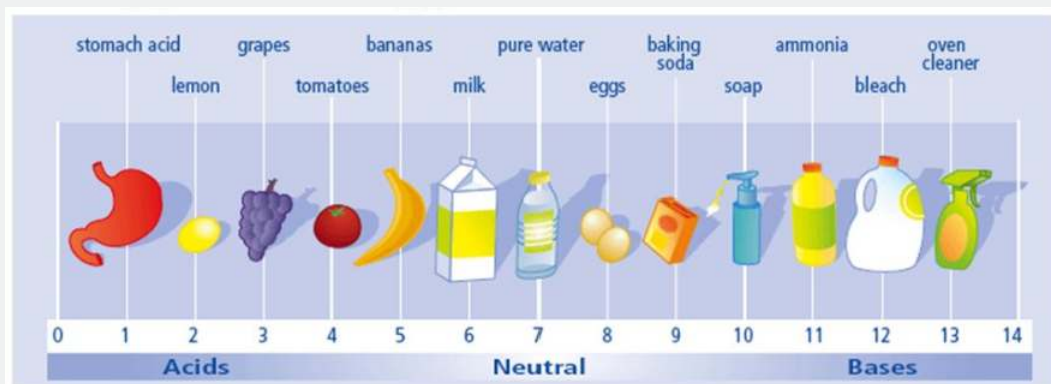


Figure 2.6.2.1.1 The pH scale.

Figure 2.6.2.1.2 illustrates the relationship between pH and the human body.

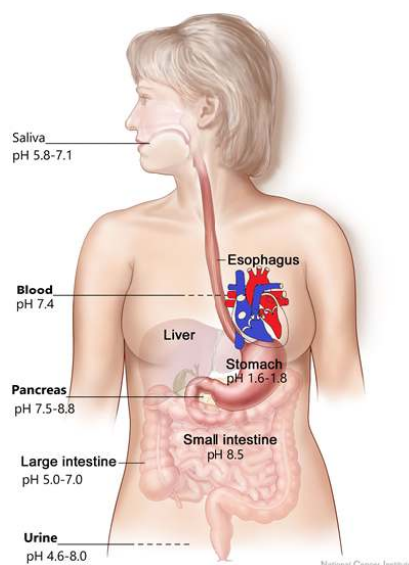


Figure 2.6.2.1.2 The pH values for different areas of the human body.

✓ Example 2.6.2.1.1

Label each solution as acidic, basic, or neutral based only on the stated pH .

- milk of magnesia, $pH = 10.5$
- pure water, $pH = 7$
- wine, $pH = 3.0$

Answer

- With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
- Pure water, with a pH of 7, is neutral.
- With a pH of less than 7, wine is acidic.

? Exercise 2.6.2.1.1

Identify each substance as acidic, basic, or neutral based only on the stated pH .

- human blood with $pH = 7.4$
- household ammonia with $pH = 11.0$
- cherries with $pH = 3.6$

Answer a

basic

Answer b

basic

Answer c

acidic

Table 2.6.2.1.1: gives the typical pH values of some common substances. Note that several food items are on the list, and most of them are acidic.

Table 2.6.2.1.1 Typical pH Values of Various Substances*

Substance	pH
stomach acid	1.7
lemon juice	2.2
vinegar	2.9
soda	3.0
wine	3.5
coffee, black	5.0
milk	6.9
pure water	7.0
blood	7.4
seawater	8.5
milk of magnesia	10.5
ammonia solution	12.5
1.0 M NaOH	14.0

*Actual values may vary depending on conditions

Summary

pH is usually (but not always) between 0 and 14.

- If $\text{pH} < 7$, then the solution is acidic.
- If $\text{pH} = 7$, then the solution is neutral.
- If $\text{pH} > 7$, then the solution is basic.

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CHAPTER OVERVIEW

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3.1: Matter and Energy

Learning Objectives

- Define *energy*.
- Know the units of energy.
- Define *heat*.
- Understand the law of conservation of energy.

Energy is the ability to do work. Think about it: when you have a lot of energy, you can do a lot of work; but if you're low on energy, you do not want to do much work. Work (w) itself is defined as a force (F) operating over a distance (Δx):

$$w = F \times \Delta x$$

In SI, force has units of newtons (N), while distance has units of meters. Therefore, work has units of N·m. This compound unit is redefined as a **Joule (J)**.

$$1 \text{ joule} = 1 \text{ newton} \cdot \text{meter} \quad 1 \text{ J} = 1 \text{ N} \cdot \text{m}$$

Because energy is the ability to do work, energy is also measured in joules. This is the primary unit of energy we will use here.

How much is 1 J? It is enough to warm up about one-fourth of a gram of water by 1°C. It takes about 12,000 J to warm a cup of coffee from room temperature to 50°C. So a joule is not a lot of energy. It will not be uncommon to measure energies in thousands of joules, so the kilojoule (kJ) is a common unit of energy, with 1 kJ equal to 1,000 J.

An older—but still common—unit of energy is the *calorie*. The calorie (cal) was originally defined in terms of warming up a given quantity of water. The modern definition of calorie equates it to joules:

$$1 \text{ cal} = 4.184 \text{ J}$$

The calorie is used when considering nutrition. Energy contents of foods are often expressed in calories. However, the calorie unit used for foods is actually the kilocalorie (kcal). Most foods indicate this by spelling the word with a capital C—Calorie. Figure 3.1.1 - Calories on Food Labels, shows one example.

Nutrition Facts			
Serving Size 3 oz. (85g)			
Amount Per Serving		As Served	
Calories 38		Calories from Fat 0	
		% Daily Value	
Total Fat	0g	0%	
Saturated Fat	0g	0%	
Cholesterol	0g	0%	
Sodium	0g	2%	
Total Carbohydrate	0g	3%	
Dietary Fiber	0g	8%	
Sugars	0g		
Protein	0g		
Vitamin A	270%	•	Vitamin C 10%
Calcium	2%	•	Iron 0%
Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs:			
	Calories	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	80g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

Figure 3.1.1 Calories on Food Labels © Thinkstock. This label expresses the energy content of the food, but in Calories (which are actually kilocalories).

✓ Example 3.1.1

The label in Figure 3.1.1 states that the serving has 38 Cal. How many joules is this?

Solution

We recognize that with a capital C, the Calories unit is actually kilocalories. To determine the number of joules, we convert first from kilocalories to calories (using the definition of the *kilo-* prefix) and then from calories to joules (using the relationship between calories and joules). So

$$38 \cancel{\text{kcal}} \times \frac{1000 \cancel{\text{cal}}}{1 \cancel{\text{kcal}}} \times \frac{4.184 \text{ J}}{1 \cancel{\text{cal}}} = 160,000 \text{ J}$$

? Exercise 3.1.1

A serving of breakfast cereal usually has 110 Cal. How many joules of energy is this?

Answer

460,000 J

Heat is another aspect of energy. **Heat** is the transfer of energy from one body to another due to a difference in temperature. For example, when we touch something with our hands, we interpret that object as either hot or cold depending on how energy is transferred: If energy is transferred into your hands, the object feels hot. If energy is transferred from your hands to the object, your hands feel cold. Because heat is a measure of energy transfer, heat is also measured in joules.

In the study of energy, we use the term **system** to describe the part of the universe under study: a beaker, a flask, or a container whose contents are being observed and measured. An **isolated system** is a system that does not allow a transfer of energy or matter into or out of the system. A good approximation of an isolated system is a closed, insulated thermos-type bottle. The fact that the thermos-type bottle is closed keeps matter from moving in or out, and the fact that it is insulated keeps energy from moving in or out.

One of the fundamental ideas about the total energy of an isolated system is that it does not increase or decrease. When this happens to a quantity, we say that the quantity is *conserved*. The **law of conservation of energy** states that the total energy of an isolated system does not change. As a scientific law, this concept occupies the highest level of understanding we have about the natural universe.

Key Takeaways

- Energy is the ability to do work and uses the unit joule.
- Heat is the transfer of energy due to temperature differences
- The law of conservation of energy states that the total energy of an isolated system does not increase or decrease.

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3.1.1: Phase Changes

Learning Objectives

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

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

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

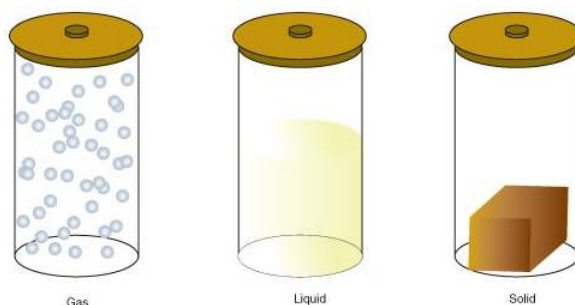


Figure 3.1.1.1: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

In Figure 3.1.1.2, all three phases are shown at the submicroscopic level in animations. Notice how the movement and freedom of molecules steadily increases as attractive forces decrease from solid to liquid to gas phase.

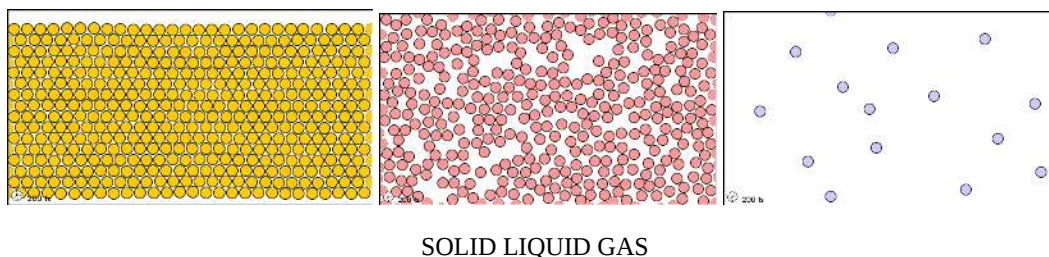


Figure 3.1.1.2 Animation of all three phases at the submicroscopic level.

The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Figures 3.1.1.3 and 3.1.1.4 show the differences among solids, liquids, and gases at the molecular level, while Table 3.1.1.1 lists the different characteristics of these states.

Table 3.1.1.1 Characteristics of the Three States of Matter

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

Phase Changes

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

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C . At 0°C , both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O . If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O . A similar process can occur at 100°C : adding heat increases the amount of gaseous H_2O , while removing heat increases the amount of liquid H_2O (Figure 3.1.1.3).

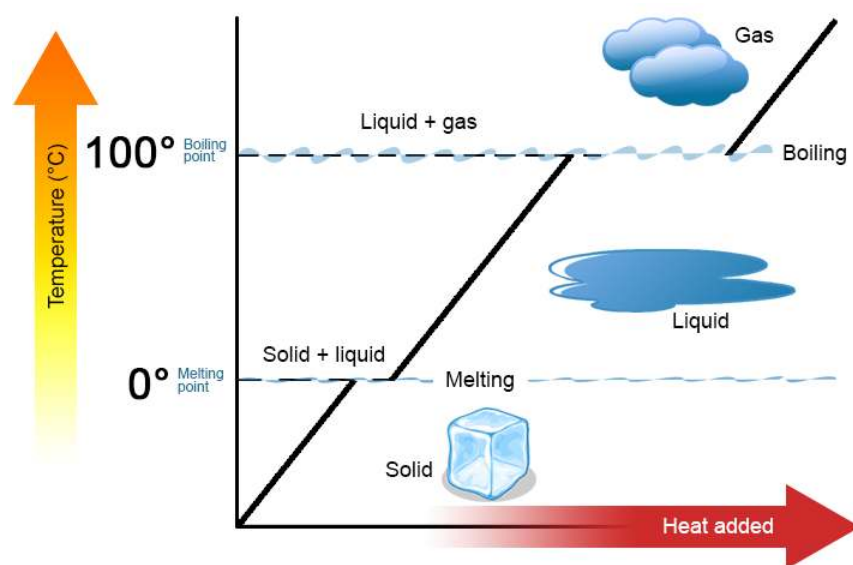


Figure 3.1.1.3: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0°C , the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).

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

Second, as shown in Figure 3.1.1.3, **the temperature of a substance does not change** as the substance goes **from one phase to another**. Again, consider H_2O as an example. Solid water (ice) can exist at 0°C . If heat is added to ice at 0°C , some of the solid changes phase to make liquid, which is also at 0°C . Remember, the solid and liquid phases of H_2O can coexist at 0°C . Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

Summary of Phase Changes

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

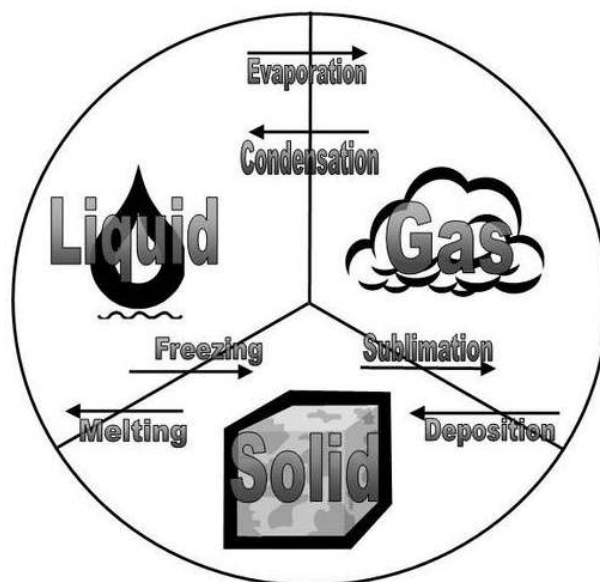


Figure 3.1.1.4: Solid, liquid, and gas states with the terms for each change of state that occurs between them.

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

Summary

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

Contributors and Attributions

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3.1.2: Physical and Chemical Changes

Learning Outcomes

- Distinguish between physical and chemical changes.
- Give examples of physical and chemical changes.

Physical Changes

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. **Melting** is an example of a **physical change**. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.



Figure 3.1.2.1: Melting ice in the Beaufort Sea.

When a piece of wood is ground into sawdust, that change is irreversible since the sawdust could not be reconstituted into the same piece of wood that it was before. Cutting the grass or pulverizing a rock would be other irreversible physical changes. Firewood also represents an irreversible physical change since the pieces cannot be put back together to form the tree.



Figure 3.1.2.2: Firewood being cut is a physical change because the composition doesn't change when being cut.

Chemical Changes

When exposed to air, an object made of iron will eventually begin to rust (see figure below).



Figure 3.1.2.3: Rust (iron oxide) forms on an unprotected iron surface.

As the rust forms on the surface of the iron, it flakes off to expose more iron, which will continue to rust. Rust is clearly a substance that is different from iron. Rusting is an example of a chemical change.

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

A chemical change is also called a chemical reaction. A **chemical reaction** is a process that occurs when one or more substances are changed into one or more new substances. Zinc (**Zn**) is a silver-gray element that can be ground into a powder. If zinc is mixed at room temperature with powdered sulfur (**S**), a bright yellow element, the result will simply be a mixture of zinc and sulfur. No chemical reaction occurs. However, if energy is provided to the mixture in the form of heat, the zinc will chemically react with the sulfur to form the compound zinc sulfide (**ZnS**). Pictured below are the substances involved in this reaction.

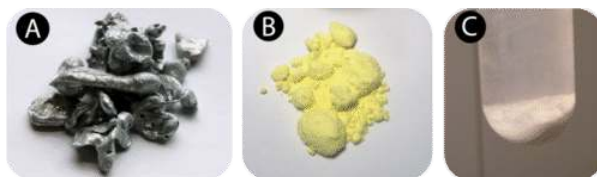


Figure 3.1.2.4: Zinc (A) and sulfur (B) are two elements that undergo a chemical reaction when heated to form the compound zinc sulfide (C).

The reaction between zinc and sulfur can be depicted in something called a chemical equation. In words, we could write the reaction as:



A more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved:



The substance(s) to the left of the arrow in a chemical equation are called reactants. A **reactant** is a substance that is present at the start of a chemical reaction. The substance(s) to the right of the arrow are called products. A **product** is a substance that is present at the end of a chemical reaction. In the equation above, zinc and sulfur are the reactants that chemically combine to form zinc sulfide as a product.

Recognizing Chemical Reactions

How can you tell if a chemical reaction is taking place? Certain visual clues indicate that a chemical reaction is likely (but not necessarily) occurring, including the following examples:

1. A change of color occurs during the reaction.
2. A gas is produced during the reaction.
3. A solid product, called a precipitate, is produced in the reaction.
4. A visible transfer of energy occurs in the form of light as a result of the reaction.

When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced (see figure below). The production of a gas is also an indication that a chemical reaction may be occurring.



Figure 3.1.2.5: Zinc reacts with hydrochloric acid to produce bubbles of hydrogen gas.

When a colorless solution of lead (II) nitrate is added to a colorless solution of potassium iodide, a yellow solid called a precipitate is instantly produced (see figure below). A **precipitate** is a solid product that forms from a reaction and settles out of a liquid mixture. The formation of a precipitate may also indicate the occurrence of a chemical reaction.

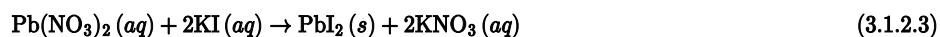


Figure 3.1.2.6: A yellow precipitate of solid lead (II) iodide forms immediately when solutions of lead (II) nitrate and potassium iodide are mixed.

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3.1.3: Chemical Potential Energy

Gunpowder was originally developed by the Chinese in the ninth century AD, primarily for rockets. This material is composed of charcoal, sulfur, and saltpeter (potassium nitrate). The explosive reaction that occurs involves the conversion of the charcoal to carbon dioxide, with the potassium nitrate providing the extra oxygen needed for a rapid reaction. Although sulfur was included to stabilize the product, gunpowder is still highly explosive.

Types of Energy

Two basic types of energy exist: potential energy and kinetic energy. **Potential energy** is stored energy. It has not yet been released, but is ready to go. **Kinetic energy** is the energy of motion. It causes work to be done through movement.

Chemical Potential Energy

Energy is the capacity for doing work or supplying heat. When you fill your car with gasoline, you are providing it with potential energy. **Chemical potential energy** is the energy stored in the chemical bonds of a substance. The various chemicals that make up gasoline contain a large amount of chemical potential energy that is released when the gasoline is burned in a controlled way in the engine of the car. The release of that energy does two things: some of the potential energy is transformed into work, which is used to move the car; at the same time, some of the potential energy is converted to heat and makes the car's engine very hot. The energy changes of a system occur as either heat or work, or some combination of both.



Figure 3.1.3.1: A dragster is able to accelerate because of the chemical potential energy of its fuel. The burning of the fuel also produces large amounts of heat. (CC BY-NC; CK-12)

Dynamite is another example of chemical potential energy. The major component of dynamite is nitroglycerin, a very unstable material. By mixing it with diatomaceous earth, the stability is increased and it is less likely to explode if it receives a physical shock. When ignited, the nitroglycerin explodes rapidly, releasing large amounts of nitrogen and other gases along with a massive amount of heat.



Figure 3.1.3.2: Dynamite explosion. (CC BY-NC; CK-12)

Summary

- Potential energy is stored energy.
- Kinetic energy is the energy of motion.
- Chemical potential energy is energy available in the chemical bonds of a compound.

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3.1.4: Heat

Blacksmiths heat solid iron in order to shape it into a variety of different objects. Iron is a rigid, solid metal. At room temperature, it is extremely difficult to bend iron. However, when heated to a high enough temperature, iron can be easily worked. The heat energy in the forge is transferred to the metal, making the iron atoms vibrate more and move around more readily.

Heat is energy that is transferred from one object or substance to another because of a difference in temperature between the two. Heat always flows from an object at a higher temperature to an object at a lower temperature (see figure below). The flow of heat will continue until the two objects are at the same temperature.

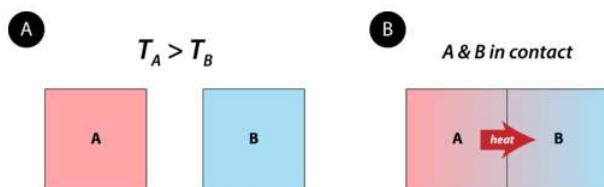


Figure 3.1.4.1: Object A starts with a higher temperature than object B. No heat flows when the objects are isolated from each other. When brought into contact, heat flows from A to B until the temperatures of the two objects are the same.

Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state. When chemical reactions occur, some chemical bonds are broken, while new chemical bonds form. As a result of the rearrangement of atoms, the total chemical potential energy of the system either increases or decreases.

Summary

- Heat is transferred energy from a site of higher energy to a site of lower energy.
- Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state.

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3.1.4.1: Exothermic and Endothermic Processes

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

Exothermic and Endothermic Processes

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

A chemical reaction or physical change is **endothermic** if heat is absorbed by the system from the surroundings. In the course of an endothermic process, the system gains heat from the surroundings, and so the temperature of the surroundings decreases. The quantity of heat for a process is represented by the letter **q**. The sign of **q** for an endothermic process is positive because the system is gaining heat. A chemical reaction or physical change is **exothermic** if heat is released by the system into the surroundings. Because the surroundings are gaining heat from the system, the temperature of the surroundings increases. The sign of **q** for an exothermic process is negative because the system is losing heat.

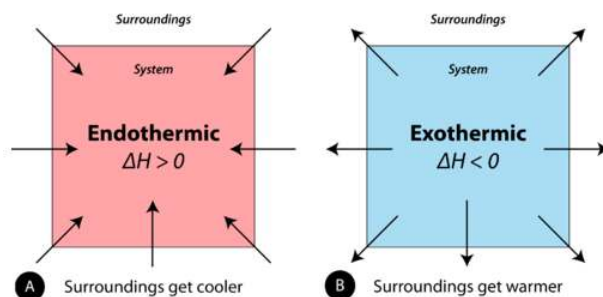


Figure 3.1.4.1.1: (A) Endothermic reaction. (B) Exothermic reaction. (CC BY-NC; CK-12)

Units of Heat

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

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

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

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

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

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

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

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

Summary

- The law of conservation of energy states that in any physical or chemical process, energy is neither created nor destroyed.
- A specific portion of matter in a given space that is being studied during an experiment or an observation is the *system*.
- The *surroundings* is everything in the universe that is not part of the system.
- A chemical reaction or physical change is endothermic if heat is absorbed by the system from the surroundings.
- A reaction or change is exothermic if heat is released by the system into the surroundings.

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3.2: Chemical Bonding and Reactivity



Figure 3.2.1 (CC by 2.0; Flickr: Amelia Wells via Flickr)

Have you ever made cupcakes from scratch?

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

What Is a Chemical Bond?

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

Q: Why do you think that chemical bonds form?

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

Why Bonds Form

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

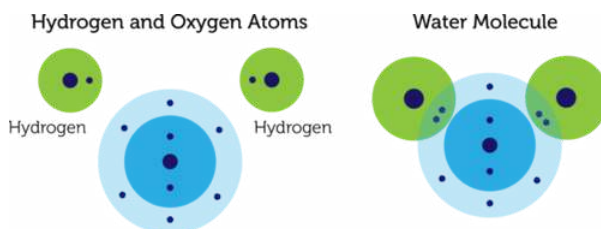


Figure 3.2.2: Copy and Paste Caption here. (CC by-NC 3.0; Christopher AuYeung via CK-12 Foundation)

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

Types of Chemical Bonds

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



- A covalent bond is the force of attraction that holds together two nonmetal atoms that share a pair of electrons. One electron is provided by each atom, and the pair of electrons is attracted to the positive nuclei of both atoms. The water molecule represented in Figure 3.2.2 contains covalent bonds.
- An ionic bond is the force of attraction that holds together oppositely charged ions. Ionic bonds form crystals instead of molecules. Table salt contains ionic bonds.
- A metallic bond is the force of attraction between a positive metal ion and the valence electrons that surround it—both its own valence electrons and those of other ions of the same metal. The ions and electrons form a lattice-like structure. Only metals, such as the copper pictured in Figure 3.2.3, form metallic bonds.



Figure 3.2.3: Metallic bonds explain many of the properties of metals. This coil of wire is made of the metal copper. Like other metals, copper is shiny, can be formed into wires, and conducts electricity. (CC BY 2.0; Flavio~ - "Industrial Photography: Copper Wires-003" via Flickr)

Summary

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

Review

1. What is a chemical bond?
2. Explain why hydrogen and oxygen atoms are more stable when they form bonds in a water molecule.
3. How do ionic bonds and covalent bonds differ?

Explore More

Watch this video about covalent bonds, and then answer the questions below.



1. Which types of elements can form covalent bonds?
2. How can you tell the number of covalent bonds an element can form?
3. Why does one atom of nitrogen form bonds with three atoms of hydrogen?

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3.2.1: Ions and Ionic Compounds

Learning Objectives

- Know how ions form.
- Learn the characteristic charges that ions have.
- Construct a proper formula for an ionic compound.
- Generate a proper name for an ionic compound.

So far, we have discussed elements and compounds that are electrically neutral. They have the same number of electrons as protons, so the negative charges of the electrons are balanced by the positive charges of the protons. However, this is not always the case. Electrons can move from one atom to another; when they do, species with overall electric charges are formed. Such species are called ions. Species with overall positive charges are termed **cations**, while species with overall negative charges are called **anions**. Remember that ions are formed only when *electrons* move from one atom to another; a proton never moves from one atom to another. Compounds formed from positive and negative ions are **ionic compounds**.

Individual atoms can gain or lose electrons. When they do, they become *monatomic* ions. When atoms gain or lose electrons, they usually gain or lose a characteristic number of electrons and so take on a characteristic overall charge. Table 3.2.1.1 lists some common ions in terms of how many electrons they lose (making cations) or gain (making anions). There are several things to notice about the ions in Table 3.2.1.1. First, each element that forms a cation is a metal, except for one (hydrogen), while each element that forms an anion is a nonmetal. This is actually one of the chemical properties of metals and nonmetals: metals tend to form cations, while nonmetals tend to form anions. Second, most atoms form ions of a single characteristic charge. When sodium atoms form ions, they always form a 1+ charge, never a 2+ or 3+ or even 1– charge. Thus, if you commit the information in Table 3.2.1.1 to memory, you will always know what charges most atoms form.

Table 3.2.1.1: Monatomic Ions of Various Charges

Ions formed by losing a single electron	H^+
Na^+	
K^+	
Rb^+	
Ag^+	
Au^+	
Ions formed by losing two electrons	Mg^{2+}
	Ca^{2+}
	Sr^{2+}
	Fe^{2+}
	Co^{2+}
	Ni^{2+}
	Cu^{2+}
	Zn^{2+}
	Sn^{2+}
	Hg^{2+}
Ions formed by losing three electrons	Pb^{2+}
	Sc^{3+}
	Fe^{3+}
	Co^{3+}
	Ni^{3+}
	Au^{3+}
	Al^{3+}
Ions formed by losing four electrons	Cr^{3+}
	Ti^{4+}
	Sn^{4+}
Ions formed by gaining a single electron	Pb^{4+}
	F^-
	Cl^-
Ions formed by gaining two electrons	Br^-
	I^-
	O^{2-}
Ions formed by gaining three electrons	S^{2-}
	Se^{2-}
	N^{3-}
	P^{3-}

Third, there are some exceptions to the previous point. A few elements, all of which are metals, can form more than one possible charge. For example, iron atoms can form $2+$ cations or $3+$ cations. Cobalt is another element that can form more than one possible charged ion ($2+$ and $3+$), while lead can form $2+$ or $4+$ cations. Unfortunately, there is little understanding which two charges a metal atom may take, so it is best to just memorize the possible charges a particular element can have.

Note the convention for indicating an ion. The magnitude of the charge is listed as a right superscript next to the symbol of the element. If the charge is a single positive or negative one, the number 1 is not written; if the magnitude of the charge is greater than 1, then the number is written *before* the $+$ or $-$ sign. An element symbol without a charge written next to it is assumed to be the uncharged atom.

Naming an ion is straightforward. For a cation, simply use the name of the element and add the word *ion* (or if you want to be more specific, add *cation*) after the element's name. So Na^+ is the sodium ion; Ca^{2+} is the calcium ion. If the element has more than one possible charge, the value of the charge comes after the element name and before the word *ion*. Thus, Fe^{2+} is the iron two ion, while Fe^{3+} is the iron three ion. In print, we use roman numerals in parentheses to represent the charge on the ion; so these two iron ions would be represented as the iron(II) cation and the iron(III) cation, respectively.

For a monatomic anion, use the stem of the element name and append the suffix *-ide* to it, and then add *ion*. This is similar to how we named molecular compounds. Thus, Cl^- is the chloride ion, and N^{3-} is the nitride ion.

✓ Example 3.2.1.1

Name each species.

- a. O^{2-}
- b. Co
- c. Co^{2+}

Solution

- a. This species has a $2-$ charge on it, so it is an anion. Anions are named using the stem of the element name with the suffix *-ide* added. This is the oxide anion.
- b. Because this species has no charge, it is an atom in its elemental form. This is cobalt.
- c. In this case, there is a $2+$ charge on the atom, so it is a cation. We note from Table 3.2.1.1 that cobalt cations can have two possible charges, so the name of the ion must specify which charge the ion has. This is the cobalt(II) cation.

? Exercise 3.2.1.1

Name each species.

- a. P^{3-}
- b. Sr^{2+}

Answers

- a. the phosphide anion
- b. the strontium cation

Ionic Formulas

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula has a cation and an anion in it; an ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is simple: the total positive charge must balance the total negative charge. Because the charges on the ions are characteristic, sometimes we have to have more than one of a cation or an anion to balance the overall positive and negative charges. It is conventional to use the lowest ratio of ions that are needed to balance the charges.

For example, consider the ionic compound between Na^+ and Cl^- . Each ion has a single charge, one positive and one negative, so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions: (1) write the formula for the cation first and the formula for the anion second, but (2) do not write the charges on the ions. Thus, for

the compound between Na^+ and Cl^- , we have the ionic formula NaCl (Figure 3.2.1.1). The formula Na_2Cl_2 also has balanced charges, but the convention is to use the lowest ratio of ions, which would be one of each. (Remember from our conventions for writing formulas that we do not write a 1 subscript if there is only one atom of a particular element present.) For the ionic compound between magnesium cations (Mg^{2+}) and oxide anions (O^{2-}), again we need only one of each ion to balance the charges. By convention, the formula is MgO .



Figure 3.2.1.1: The ionic compound **NaCl** is table salt and is very common. (Public Domain; Pexels.)

For the ionic compound between Mg^{2+} ions and Cl^- ions, we now consider the fact that the charges have different magnitudes: 2+ on the magnesium ion and 1- on the chloride ion. To balance the charges with the lowest number of ions possible, we need to have two chloride ions to balance the charge on the one magnesium ion. Rather than write the formula MgClCl , we combine the two chloride ions and write it with a 2 subscript: MgCl_2 .

What is the formula MgCl_2 telling us? There are two chloride ions in the formula. Although chlorine as an element is a diatomic molecule, Cl_2 , elemental chlorine is not part of this ionic compound. The chlorine is in the form of a negatively charged *ion*, not the neutral *element*. The 2 subscript is in the ionic formula because we need two Cl^- ions to balance the charge on one Mg^{2+} ion.

✓ Example 3.2.1.2

Write the proper ionic formula for each of the two given ions.

- Ca^{2+} and Cl^-
- Al^{3+} and F^-
- Al^{3+} and O^{2-}

Solution

- We need two Cl^- ions to balance the charge on one Ca^{2+} ion, so the proper ionic formula is CaCl_2 .
- We need three F^- ions to balance the charge on the Al^{3+} ion, so the proper ionic formula is AlF_3 .
- With Al^{3+} and O^{2-} , note that neither charge is a perfect multiple of the other. This means we have to go to a least common multiple, which in this case will be six. To get a total of 6+, we need two Al^{3+} ions; to get 6-, we need three O^{2-} ions. Hence the proper ionic formula is Al_2O_3 .

? Exercise 3.2.1.2

Write the proper ionic formulas for each of the two given ions.

- Fe^{2+} and S^{2-}
- Fe^{3+} and S^{2-}

Answers

- FeS
- Fe_2S_3

Naming ionic compounds is simple: combine the name of the cation and the name of the anion, in both cases omitting the word *ion*. *Do not use numerical prefixes if there is more than one ion necessary to balance the charges.* NaCl is sodium chloride, a combination of the name of the cation (sodium) and the anion (chloride). MgO is magnesium oxide. MgCl₂ is magnesium chloride—not magnesium dichloride.

When naming ionic compounds whose cations can have more than one possible charge, we must also include the charge, in parentheses and in roman numerals, as part of the name. Hence FeS is iron(II) sulfide, while Fe₂S₃ is iron(III) sulfide. Again, no numerical prefixes appear in the name. The number of ions in the formula is dictated by the need to balance the positive and negative charges.

✓ Example 3.2.1.3

Name each ionic compound.

- a. CaCl₂
- b. AlF₃
- c. Co₂O₃

Solution

- a. Using the names of the ions, this ionic compound is named calcium chloride. It is *not* calcium(II) chloride, because calcium forms only one cation when it forms an ion, and it has a characteristic charge of 2+.
- b. The name of this ionic compound is aluminum fluoride.
- c. We know that cobalt can have more than one possible charge; we just need to determine what it is. Oxide always has a 2- charge, so with three oxide ions, we have a total negative charge of 6-. This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. Therefore, the proper name for this ionic compound is cobalt(III) oxide.

? Exercise 3.2.1.3

Name each ionic compound.

- a. Sc₂O₃
- b. AgCl

Answers

- a. scandium oxide
- b. silver chloride

How do you know whether a formula—and by extension, a name—is for a molecular compound or for an ionic compound? Molecular compounds form between nonmetals and nonmetals, while ionic compounds form between metals and nonmetals. The periodic table can be used to determine which elements are metals and nonmetals.

There also exists a group of ions that contain more than one atom. These are called **polyatomic ions**. Table 3.2.1.2 lists the formulas, charges, and names of some common polyatomic ions. Only one of them, the ammonium ion, is a cation; the rest are anions. Most of them also contain oxygen atoms, so sometimes they are referred to as **oxyanions**. Some of them, such as nitrate and nitrite, and sulfate and sulfite, have very similar formulas and names, so care must be taken to get the formulas and names correct. Note that the -ite polyatomic ion has one less oxygen atom in its formula than the -ate ion but with the same ionic charge.

Table 3.2.1.2: Common Polyatomic Ions

Name	Formula and Charge	Name	Formula and Charge
ammonium	NH_4^+	hydroxide	OH^-
acetate	$\text{C}_2\text{H}_3\text{O}_2^-$, or CH_3COO^-	nitrate	NO_3^-
bicarbonate (hydrogen carbonate)	HCO_3^-	nitrite	NO_2^-
bisulfate (hydrogen sulfate)	HSO_4^-	peroxide	O_2^{2-}
carbonate	CO_3^{2-}	perchlorate	ClO_4^-
chlorate	ClO_3^-	phosphate	PO_4^{3-}
chromate	CrO_4^{2-}	sulfate	SO_4^{2-}
cyanide	CN^-	sulfite	SO_3^{2-}
dichromate	$\text{Cr}_2\text{O}_7^{2-}$	triiodide	I_3^-

The naming of ionic compounds that contain polyatomic ions follows the same rules as the naming for other ionic compounds: simply combine the name of the cation and the name of the anion. Do not use numerical prefixes in the name if there is more than one polyatomic ion; the only exception to this is if the name of the ion itself contains a numerical prefix, such as dichromate or triiodide.

Writing the formulas of ionic compounds has one important difference. If more than one polyatomic ion is needed to balance the overall charge in the formula, enclose the formula of the polyatomic ion in parentheses and write the proper numerical subscript to the right and *outside* of the parentheses. Thus, the formula between calcium ions, Ca^{2+} , and nitrate ions, NO_3^- , is properly written $\text{Ca}(\text{NO}_3)_2$, not CaNO_3_2 or CaN_2O_6 . Use parentheses where required. The name of this ionic compound is simply calcium nitrate.

✓ Example 3.2.1.4

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

- NH_4^+ and S^{2-}
- Al^{3+} and PO_4^{3-}
- Fe^{2+} and PO_4^{3-}

Solution

- Because the ammonium ion has a 1+ charge and the sulfide ion has a 2- charge, we need two ammonium ions to balance the charge on a single sulfide ion. Enclosing the formula for the ammonium ion in parentheses, we have $(\text{NH}_4)_2\text{S}$. The compound's name is ammonium sulfide.
- Because the ions have the same magnitude of charge, we need only one of each to balance the charges. The formula is AlPO_4 , and the name of the compound is aluminum phosphate.
- Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, we need three iron(II) ions, and to get 6-, we need two phosphate ions. The proper formula is $\text{Fe}_3(\text{PO}_4)_2$, and the compound's name is iron(II) phosphate.

? Exercise 3.2.1.4

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

- NH_4^+ and PO_4^{3-}
- Co^{3+} and NO_2^-

Answers

- a. $(\text{NH}_4)_3\text{PO}_4$, ammonium phosphate
- b. $\text{Co}(\text{NO}_2)_3$, cobalt(III) nitrite

📌 Food and Drink Application: Sodium in Your Food

The element sodium, at least in its ionic form as Na^+ , is a necessary nutrient for humans to live. In fact, the human body is approximately 0.15% sodium, with the average person having one-twentieth to one-tenth of a kilogram in their body at any given time, mostly in fluids outside cells and in other bodily fluids.

Sodium is also present in our diet. The common table salt we use on our foods is an ionic sodium compound. Many processed foods also contain significant amounts of sodium added to them as a variety of ionic compounds. Why are sodium compounds used so much? Usually sodium compounds are inexpensive, but, more importantly, most ionic sodium compounds dissolve easily. This allows processed food manufacturers to add sodium-containing substances to food mixtures and know that the compound will dissolve and distribute evenly throughout the food. Simple ionic compounds such as sodium nitrite (NaNO_2) are added to cured meats, such as bacon and deli-style meats, while a compound called sodium benzoate is added to many packaged foods as a preservative. Table 3.2.1.3 is a partial list of some sodium additives used in food. Some of them you may recognize after reading this chapter. Others you may not recognize, but they are all ionic sodium compounds with some negatively charged ion also present.

Table 3.2.1.3: Some Sodium Compounds Added to Food

Sodium Compound	Use in Food
Sodium acetate	preservative, acidity regulator
Sodium adipate	food acid
Sodium alginate	thickener, vegetable gum, stabilizer, gelling agent, emulsifier
Sodium aluminum phosphate	acidity regulator, emulsifier
Sodium aluminosilicate	anti-caking agent
Sodium ascorbate	antioxidant
Sodium benzoate	preservative
Sodium bicarbonate	mineral salt
Sodium bisulfite	preservative, antioxidant
Sodium carbonate	mineral salt
Sodium carboxymethylcellulose	emulsifier
Sodium citrates	food acid
Sodium dehydroacetate	preservative
Sodium erythorbate	antioxidant
Sodium erythorbin	antioxidant
Sodium ethyl para-hydroxybenzoate	preservative
Sodium ferrocyanide	anti-caking agent
Sodium formate	preservative
Sodium fumarate	food acid
Sodium gluconate	stabilizer
Sodium hydrogen acetate	preservative, acidity regulator
Sodium hydroxide	mineral salt
Sodium lactate	food acid
Sodium malate	food acid
Sodium metabisulfite	preservative, antioxidant, bleaching agent
Sodium methyl para-hydroxybenzoate	preservative
Sodium nitrate	preservative, color fixative
Sodium nitrite	preservative, color fixative
Sodium orthophenyl phenol	preservative
Sodium propionate	preservative
Sodium propyl para-hydroxybenzoate	preservative
Sodium sorbate	preservative
Sodium stearyl lactylate	emulsifier

Sodium Compound	Use in Food
Sodium succinates	acidity regulator, flavor enhancer
Sodium salts of fatty acids	emulsifier, stabilizer, anti-caking agent
Sodium sulfite	mineral salt, preservative, antioxidant
Sodium sulfite	preservative, antioxidant
Sodium tartrate	food acid
Sodium tetraborate	preservative

The use of so many sodium compounds in prepared and processed foods has alarmed some physicians and nutritionists. They argue that the average person consumes too much sodium from his or her diet. The average person needs only about 500 mg of sodium every day; most people consume more than this—up to 10 times as much. Some studies have implicated increased sodium intake with high blood pressure; newer studies suggest that the link is questionable. However, there has been a push to reduce the amount of sodium most people ingest every day: avoid processed and manufactured foods, read labels on packaged foods (which include an indication of the sodium content), avoid oversalting foods, and use other herbs and spices besides salt in cooking.

Nutrition Facts		
Serving Size 8 oz (227 g/8 oz)		
Servings Per Container About 3		
Amount Per Serving		
Calories	180	Calories from Fat 60
% Daily Value*		
Total Fat	6g	10%
Saturated Fat	1g	5%
Trans Fat	0g	
Cholesterol	5mg	2%
Sodium	75mg	3%
Total Carbohydrate	26g	9%
Dietary Fiber	5g	19%
Sugars	11g	
Protein	8g	
Vitamin A 60% • Vitamin C 70%		
Calcium 8% • Iron 10%		
* Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs.		
	Calories	2,000 2,500
Total Fat	Less than	65g 80g
Sat Fat	Less than	20g 25g
Cholesterol	Less than	300mg 300mg
Sodium	Less than	2,400mg 2,400mg
Total Carbohydrate		300g 375g
Dietary Fiber		25g 30g
Calories per gram:		
	Fat 9 • Carbohydrate 4 • Protein 4	

Figure 3.2.1.2: Food labels include the amount of sodium per serving. This particular label shows that there are 75 mg of sodium in one serving of this particular food item.

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions and negative ions.
- Ionic formulas balance the total positive and negative charges.
- Ionic compounds have a simple system of naming.
- Groups of atoms can have an overall charge and make ionic compounds.

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3.2.1.1: Ionic Bonding



Figure 3.2.1.1.1 (Public Domain; Leon Brooks via [Commons Wikimedia, Waneroo Beach](https://commons.wikimedia.org/wiki/File:Waneroo_Beach) [commons.wikimedia.org])

Does the sea really have salt in it?

We can get common table salt from several sources. It can be mined in the solid form in salt mines, or found as a solid in deposits. We can also get salt from the ocean, but it really does not exist as a salt when in solution. The sodium ions and chloride ions are dissolved, but not combined into a structure until all the water is removed.

Most of the rocks and minerals that make up the Earth's crust are composed of positive and negative ions held together by ionic bonding. An ionic compound is an electrically neutral compound consisting of positive and negative ions. You are very familiar with some ionic compounds, such as sodium chloride (NaCl). A sodium chloride crystal consists of equal numbers of positive sodium ions (Na^+) and negative chloride ions (Cl^-).

Ionic Bonds

Oppositely charged particles attract each other. This attractive force is often referred to as an **electrostatic force**. An **ionic bond** is the electrostatic force that holds ions together in an **ionic compound**. The strength of the ionic bond is directly dependent upon the quantity of the charges and inversely dependent on the distance between the charged particles. A cation with a $2+$ charge will make a stronger ionic bond than a cation with a $1+$ charge. A larger ion makes a weaker ionic bond because of the greater distance between its electrons and the nucleus of the oppositely charged ion.



Electron Dot Diagrams

We will use sodium chloride as an example to demonstrate the nature of the ionic bond and how it forms. As you know, sodium is a metal and loses its one valence electron to become a cation. Chlorine is a nonmetal and gains one electron in becoming an anion. Both achieve a noble-gas electron configuration. However, electrons cannot be simply "lost" to nowhere in particular. A more accurate way to describe what is happening is that a single electron is transferred from the sodium atom to the chlorine atom, as shown below.

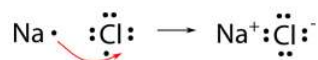


Figure 3.2.1.1.2 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The ionic bond is the attraction of the Na^+ ion for the Cl^- ion. It is conventional to show the cation without dots around the symbol to emphasize that the original energy level that contained the valence electron is now empty. The anion is now shown with a complete octet of electrons.

For a compound such as magnesium chloride, it is not quite as simple. Because magnesium has two valence electrons, it needs to lose both to achieve the noble-gas configuration. Therefore, two chlorine atoms will be needed.

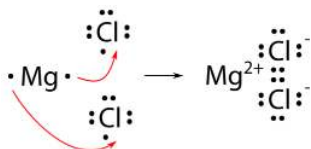


Figure 3.2.1.1.3 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The final formula for magnesium chloride is MgCl_2 .

Summary

- An ionic compound contains positive and negative ions.
- An ionic bond is electrostatic in nature.
- Electron dot diagrams can be used to illustrate electron movements and ion formation.

Review

1. What is an ionic compound?
2. What is an ionic bond?
3. Which cation (Na^+ or Ca^{2+}) would form a stronger ionic bond with Cl^- ?

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3.2.2: Molecules and Molecular Compounds

Learning Objectives

- Define *molecule*.
- Name simple molecules based on their formulas.
- Determine a formula of a molecule based on its name.

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multiatom combinations are called molecules. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

Table 3.2.2.1: Elements That Exist as Diatomic Molecules

Hydrogen (H_2)	Oxygen (O_2)	Nitrogen (N_2)	Fluorine (F_2)
Chlorine (Cl_2)	Bromine (Br_2)	Iodine (I_2)	

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules (Table 3.2.2.1). As with any molecule, these elements are labeled with a molecular formula, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H_2 , with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O_2 , N_2 , and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S_8 , while phosphorus exists as a four-atom molecule, P_4 (Figure 3.2.2.1). Otherwise, we will assume that elements exist as individual atoms, rather than molecules. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element's symbol.

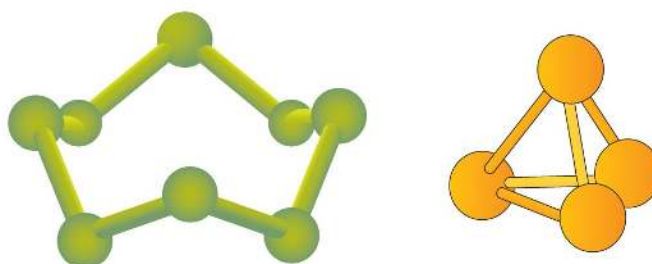


Figure 3.2.2.1: Molecular Art of S_8 and P_4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S_8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 3.2.2.1 shows two examples of how we will be representing molecules in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond and is the connection between two atoms in a molecule.

Many compounds exist as molecules. In particular, when nonmetals connect with other nonmetals, the compounds typically exist as molecules. (Compounds between a metal and a nonmetal are different and will be considered in Section 3.4.) In some cases, there are many different kinds of molecules that can be formed between any given elements, with all the different molecules having different chemical and physical properties. How do we tell them apart?

The answer is a very specific system of naming compounds, called *chemical nomenclature*. By following the rules of nomenclature, each and every compound has its own unique name, and each name refers to one and only one compound. Here, we will start with relatively simple molecules that have only two elements in them, the so-called *binary compounds*:

1. Identify the elements in the molecule from its formula.

2. Begin the name with the element name of the first element. If there is more than one atom of this element in the molecular formula, use a numerical prefix to indicate the number of atoms, as listed in Table 3.2.2.2. *Do not use the prefix mono- if there is only one atom of the first element.*
3. Name the second element by using three pieces:
 - a numerical prefix indicating the number of atoms of the second element, plus
 - the stem of the element name (e.g., *ox* for oxygen, *chlor* for chlorine, etc.), plus
 - the suffix *-ide*
4. Combine the two words, leaving a space between them.

Table 3.2.2.2: Numerical Prefixes Used in Naming Molecular Compounds

The Number of Atoms of an Element	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Let us see how these steps work for a molecule whose molecular formula is SO_2 , which has one sulfur atom and two oxygen atoms—this completes step 1. According to step 2, we start with the name of the first element—sulfur. Remember, we do not use the *mono-* prefix for the first element. Now for step 3, we combine the numerical prefix *di-* (see Table 3.2.2.2) with the stem *ox-* and the suffix *-ide*, to make *dioxide*. Bringing these two words together, we have the unique name for this compound—sulfur dioxide.

Why all this trouble? There is another common compound consisting of sulfur and oxygen whose molecular formula is SO_3 , so the compounds need to be distinguished. SO_3 has three oxygen atoms in it, so it is a different compound with different chemical and physical properties. The system of chemical nomenclature is designed to *give this compound its own unique name*. Its name, if you go through all the steps, is sulfur trioxide. Different compounds have different names.

In some cases, when a prefix ends in *a* or *o*; and the element name begins with *o*, we drop the *a* or *o* on the prefix. So we see *monoxide* or *pentoxide*, rather than *monooxide* or *pentaoxide* in molecule names.

✓ Example 3.2.2.1

Name each molecule.

- PF_3
- CO
- Se_2Br_2

Solution

- A molecule with a single phosphorus atom and three fluorine atoms is called phosphorus trifluoride.
- A compound with one carbon atom and one oxygen atom is properly called carbon monoxide, not carbon monooxide.
- There are two atoms of each element, selenium and bromine. According to the rules, the proper name here is *diselenium dibromide*.

? Exercise 3.2.2.1

Name each molecule.

- a. SF₄
- b. P₂S₅

Answer a

sulfur tetrafluoride

Answer b

diphosphorus pentasulfide

One great thing about this system is that it works both ways. From the name of a compound, you should be able to determine its molecular formula. Simply list the element symbols, with a numerical subscript if there is more than one atom of that element, in the order of the name (we do not use the subscript 1 if there is only one atom of the element present; 1 is implied). From the name *nitrogen trichloride*, you should be able to get NCl₃ as the formula for this molecule. From the name *diphosphorus pentoxide*, you should be able to get the formula P₂O₅ (note the numerical prefix on the first element, indicating there is more than one atom of phosphorus in the formula).

✓ Example 3.2.2.2

Give the formula for each molecule.

- a. carbon tetrachloride
- b. silicon dioxide
- c. trisilicon tetranitride

Solution

- a. The name *carbon tetrachloride* implies one carbon atom and four chlorine atoms, so the formula is CCl₄.
- b. The name *silicon dioxide* implies one silicon atom and two oxygen atoms, so the formula is SiO₂.
- c. We have a name that has numerical prefixes on both elements. *Tri-* means three, and *tetra-* means four, so the formula of this compound is Si₃N₄.

? Exercise 3.2.2.2

Give the formula for each molecule.

- a. disulfur difluoride
- b. iodine pentabromide

Answer a

S₂F₂

Answer b

IBr₅

Some simple molecules have common names that we use as part of the formal system of chemical nomenclature. For example, H₂O is given the name *water*, not *dihydrogen monoxide*. NH₃ is called *ammonia*, while CH₄ is called *methane*. We will occasionally see other molecules that have common names; we will point them out as they occur.

Key Takeaways

- Molecules are groups of atoms that behave as a single unit.
- Some elements exist as molecules: hydrogen, oxygen, sulfur, and so forth.
- There are rules that can express a unique name for any given molecule, and a unique formula for any given name.

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3.2.2.1: Covalent Bonds



Figure 3.2.2.1.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

In a tennis match, two players keep hitting the ball back and forth. The ball bounces from one player to the other, over and over again. The ball keeps the players moving together on the court. What if the two players represented the nuclei of two atoms and the ball represented valence electrons? What would the back and forth movement of the ball represent? The answer is a covalent bond.

Sharing Electrons

A covalent bond is the force of attraction that holds together two atoms that share a pair of valence electrons. The shared electrons are attracted to the nuclei of both atoms. This forms a molecule consisting of two or more atoms. Covalent bonds form only between atoms of nonmetals.

Covalent Compounds and Diatomic Elements

The two atoms that are held together by a covalent bond may be atoms of the same element or different elements. When atoms of different elements form covalent bonds, a new substance, called a covalent compound, results. Water is an example of a covalent compound. A water molecule is modeled in the figure below. A molecule is the smallest particle of a covalent compound that still has the properties of the compound.

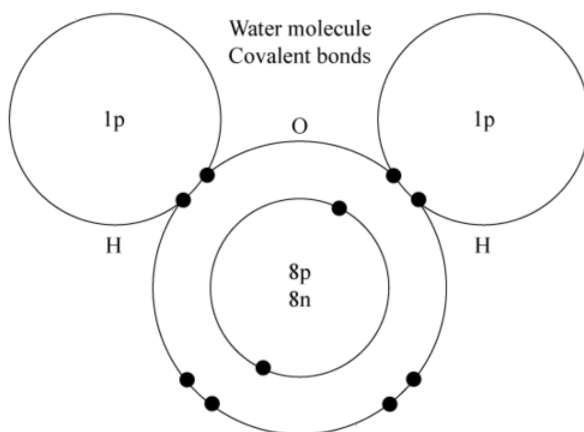


Figure 3.2.2.1.2 (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Q: How many valence electrons does the oxygen atom (O) share with each hydrogen atom (H)? How many covalent bonds hold the water molecule together?

A: The oxygen atom shares one pair of valence electrons with each hydrogen atom. Each pair of shared electrons represents one covalent bond, so two covalent bonds hold the water molecule together.

The diagram in the figure below shows an example of covalent bonds between two atoms of the same element, in this case two atoms of oxygen. The diagram represents an oxygen molecule, so it's not a new compound. Oxygen normally occurs in diatomic ("two-atom") molecules. Several other elements also occur as diatomic molecules: hydrogen, nitrogen, and all but one of the halogens (fluorine, chlorine, bromine, and iodine).

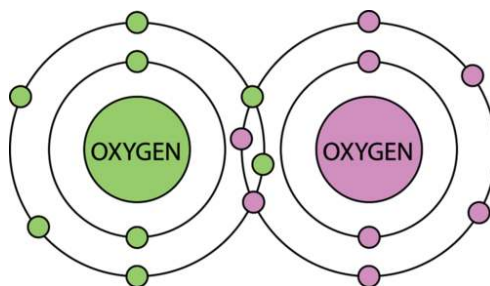


Figure 3.2.2.1.3 (CC BY-NC 3.0; Christopher AuYeung via CK-12 Foundation)

Q: How many electrons do these two oxygen atoms share? How many covalent bonds hold the oxygen molecule together?

A: The two oxygen atoms share two pairs of electrons, so two covalent bonds hold the oxygen molecule together.

Why Covalent Bonds Form

Covalent bonds form because they give atoms a more stable arrangement of electrons. Look at the oxygen atoms in the figure above. Alone, each oxygen atom has six valence electrons. By sharing two pairs of valence electrons, each oxygen atom has a total of eight valence electrons. This fills its outer energy level, giving it the most stable arrangement of electrons. The shared electrons are attracted to both oxygen nuclei, and this force of attraction holds the two atoms together in the oxygen molecule.

Summary

- A covalent bond is the force of attraction that holds together two atoms that share a pair of valence electrons. Covalent bonds form only between atoms of nonmetals.
- The two atoms that are held together in a covalent bond may be atoms of the same element or different elements. When atoms of different elements bond together, it forms a covalent compound.
- Covalent bonds form because the shared electrons fill each atom's outer energy level and this is the most stable arrangement of electrons.

Review

1. What is a covalent bond?
2. Nitrogen is a diatomic element with five valence electrons. Create a model of a molecule of nitrogen.
3. Which of the following represents a covalent compound?
 - a. O_2
 - b. CO_2
 - c. Cl_2
 - d. $NaCl$

Explore More

Watch the video about covalent bonding and then answer the questions below.



1. How can you tell the number covalent bonds the atoms of an element can form?
2. How many covalent bonds can nitrogen form? How many covalent bonds can chlorine form?
3. Why are covalent bonds stronger than ionic bonds?
4. Why can't molecules of a covalent compound conduct electricity?

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3.2.2.2: Energy and Covalent Bonds



Figure 3.2.2.2.1 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Why do BeCl_2 and LiCl bond differently?

We have learned that halide salts of elements in group 1 are typically ionic compounds. We would expect LiCl to exist as Li^+ cations and Cl^- anions (and it does). However, if we move one column to the right, lithium's neighbor beryllium forms a different type of bond altogether. This bond consists of shared electrons between the **Be** and **Cl** atoms, not electrostatic attraction among ions.

Energy and Covalent Bond Formation

Molecular compounds are those that take the form of an individual molecule. Molecular compounds are generally comprised of two or more nonmetal atoms. Familiar examples include water (H_2O), carbon dioxide (CO_2), and ammonia (NH_3). Recall that the molecular formula shows the number of each atom that occurs in a molecule of that compound. One molecule of water contains two hydrogen atoms and one oxygen atom. Hydrogen (H_2) is an example of an element that exists naturally as a diatomic molecule. A **diatomic molecule** is a molecule containing two atoms.

Most atoms attain a lower potential energy when they are bonded to other atoms than when they are separated. Consider two isolated hydrogen atoms that are separated by a distance large enough to prevent any interaction between them. At this distance, the potential energy of the system is said to be equal to zero (see figure below).

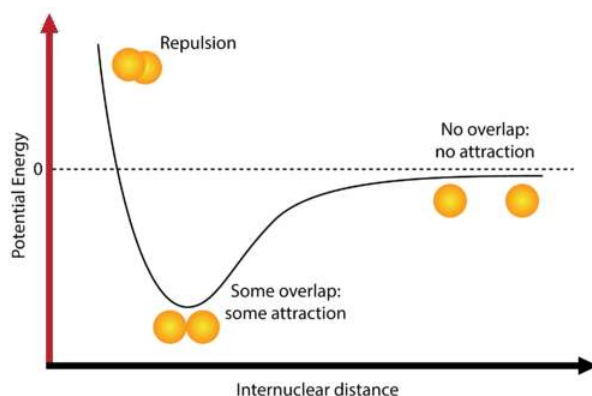


Figure 3.2.2.2.2: The graph shows how the potential energy of two hydrogen atoms changes as a function of their separation distance. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

As the atoms approach one another, their electron clouds gradually begin to overlap. Now there are several interactions which begin to occur. One is that the single electrons that each hydrogen atom possesses begin to repel each other. This repulsive force would tend to make the potential energy of the system increase. However, the electron of each atom begins to be attracted to the nucleus of the other atom. This attractive force tends to make the potential energy of the system decrease.

As the atoms first begin to interact, the attractive force is stronger than the repulsive force and so the potential energy of the system decreases, as seen in the diagram. Remember that the lower potential energy increases the stability of the system. As the two hydrogen atoms move closer and closer together, the potential energy is at its lowest possible point. If the hydrogen atoms move any closer together, a third interaction begins to dominate, and that is the repulsive force between the two positively-charged nuclei. This repulsive force is very strong, as can be seen by the sharp rise in energy at the far left of the diagram.

The point at which the potential energy reached its minimum represents the ideal distance between hydrogen atoms for a stable chemical bond to occur. This type of chemical bond is called a covalent bond. A **covalent bond** is a bond in which two atoms share one or more pairs of electrons. The single electrons from each of the two hydrogen atoms are shared when the atoms come together to form a hydrogen molecule (H_2).



Summary

- Covalent bonds are formed when atoms share electrons between them.

Review

1. What is a diatomic molecule?
2. What does lower potential energy do to a system?
3. What is the ideal distance for two atoms?

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3.2.2.3: Bond Energy and Reactivity

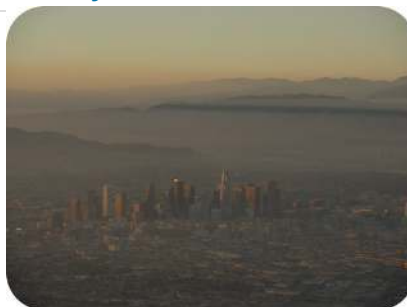


Figure 3.2.2.3.1 (CC by 2.0; Warner Hocker (Flickr: TravelingOtter) via Flickr)

How does smog form?

We know that nitrogen makes up about **79%** of the gases in the air, and that this gas is chemically very inert. However, we also know that a major contributor to the production of smog is a mixture of nitrogen compounds referred to as NO_x . Due to the high combustion temperatures in modern car engines (brought about by the need for better fuel efficiency), we make more NO_x that will react with other materials to create smog. So, our inert nitrogen gas can be converted to other compounds if enough energy is available to break the triple bonds in the N_2 molecule.

Bond Energy

The formation of a chemical bond results in a decrease in potential energy. Consequently, breaking a chemical bond requires an input of energy. **Bond energy** is the energy required to break a covalent bond between two atoms. A high bond energy means that a bond is strong and the molecule that contains that bond is likely to be stable and less reactive. More reactive compounds will contain bonds that have generally lower bond energies. Some bond energies are listed in the table below.

Table 3.2.2.3.1: Bond Energies

Bond	Bond Energy (kJ/mol)
H—H	436
C—H	414
C—C	347
C=C	620
C $\equiv$ C	812
F—F	157
Cl—Cl	243
Br—Br	193
I—I	151
N $\equiv$ N	941

The halogen elements all exist naturally as diatomic molecules (F_2 , Cl_2 , Br_2 , and I_2). They are generally very reactive and thus have relatively low bond energies.

As can be seen by a comparison of the bond energies for the various carbon-carbon bonds, double bonds are substantially stronger than single bonds. Likewise, triple bonds are even stronger. The triple bond that exists between the nitrogen atoms in nitrogen gas (N_2) makes it very unreactive. All plants and animals require the element nitrogen, but it cannot be obtained from the direct absorption of nitrogen gas from the atmosphere because of its strong, unreactive triple bond. Instead, bacteria convert the nitrogen

to a more usable form such as ammonium and nitrate ions, which are then absorbed by plants from the soil. Animals only obtain nitrogen by eating those plants.



Summary

- Bond energy is an indication of the amount of energy needed for a chemical reaction.
- The higher the bond energy, the less reactive the bond is.

Review

1. What does a high bond energy mean?
2. What kind of bonds do more reactive compounds have?
3. Which will react more readily: a C-H bond or a Cl-Cl bond?

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CHAPTER OVERVIEW

4: Motion, Forces, and Energy

4.1: Motion

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4.1.3: Acceleration

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4.1: Motion



Figure 4.1.1: The motion of an

American kestrel through the air can be described by the bird's displacement, speed, velocity, and acceleration. When it flies in a straight line without any change in direction, its motion is said to be one dimensional. (credit: Vince Maidens, Wikimedia Commons)

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?* 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.

Our 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 one-dimensional kinematics we will study only the *motion* of a football, for example, without worrying about what forces cause or change its motion. Such considerations come in other chapters. In this chapter, we examine the simplest type of motion—namely, motion along a straight line, or one-dimensional motion.

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4.1.1: Position, Displacement, and Distance

Learning Objectives

- Define position, displacement, distance, and distance traveled.
- Explain the relationship between position and displacement.
- Distinguish between displacement and distance traveled.
- Calculate displacement and distance given initial position, final position, and the path between the two.



Figure 4.1.1.1: 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. (credit: Suzan Black, Fotopedia)

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 4.1.1.2.) 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 4.1.1.3.)

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.

Definition: DISPLACEMENT

Displacement is the *change in position* of an object:

$$\Delta x = x_f - x_0,$$

where Δx is displacement, x_f is the final position, and x_0 is the initial position.

In this text the upper case Greek letter Δ always means “change in” whatever quantity follows it; thus, Δx means *change in position*. Always solve for displacement by subtracting initial position x_0 from final position x_f .

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.

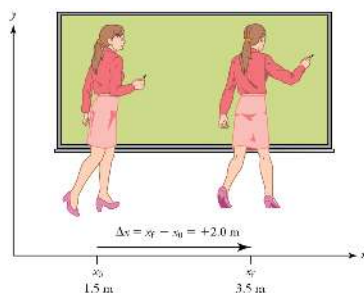


Figure 4.1.1.2: A professor paces left and right while lecturing. Her position relative to Earth is given by x . The $+2.0\text{ m}$ displacement of the professor relative to Earth is represented by an arrow pointing to the right.

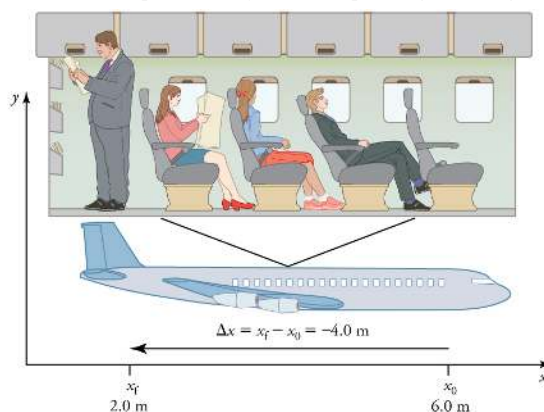


Figure 4.1.1.3: A passenger moves from his seat to the back of the plane. His location relative to the airplane is given by x . The -4.0-m displacement of the passenger relative to the plane is represented by an arrow toward the rear of the plane. Notice that the arrow representing his displacement is twice as long as the arrow representing the displacement of the professor (he moves twice as far) in Figure 4.1.1.2.

Note that displacement has a direction as well as a magnitude. The professor's displacement is 2.0 m to the right, and the airline passenger's displacement is 4.0 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 $x_f = 2.0\text{ m}$, so his displacement is

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

His displacement is negative because his motion is toward the rear of the plane, or in the negative x 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 .



MISCONCEPTION ALERT: DISTANCE TRAVELED VS. MAGNITUDE OF DISPLACEMENT

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.

? Exercise 4.1.1.1

A cyclist rides 3 km west and then turns around and rides 2 km east. (a) What is her displacement? (b) What distance does she ride? (c) What is the magnitude of her displacement?

Answer

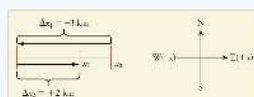


Figure 4.1.1.4

- (a) The rider's displacement is $\Delta x = x_f - x_0 = -1 \text{ km}$. (The displacement is negative because we take east to be positive and west to be negative.)
- (b) The distance traveled is $3 \text{ km} + 2 \text{ km} = 5 \text{ km}$.
- (c) The magnitude of the displacement is 1 km .

Section Summary

- Kinematics is the study of motion without considering its causes. In this chapter, it is limited to motion along a straight line, called one-dimensional motion.
- Displacement is the change in position of an object.
- In symbols, displacement Δx is defined to be

$$\Delta x = x_f - x_0,$$

where x_0 is the initial position and x_f is the final position. In this text, the Greek letter Δ (delta) always means "change in" whatever quantity follows it. The SI unit for displacement is the meter (m). Displacement has a direction as well as a magnitude.

- When you start a problem, assign which direction will be positive.
- Distance is the magnitude of displacement between two positions.
- Distance traveled is the total length of the path traveled between two positions.

Glossary

kinematics

the study of motion without considering its causes

position

the location of an object at a particular time

displacement

the change in position of an object

distance

the magnitude of displacement between two positions

distance traveled

the total length of the path traveled between two positions

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4.1.2: Time, Velocity, and Speed

Learning Objectives

- Explain the relationships between instantaneous velocity, average velocity, instantaneous speed, average speed, displacement, and time.
- Calculate velocity and speed given initial position, initial time, final position, and final time.
- Interpret graphs of position vs. time and velocity vs. time.



Figure 4.1.2.1: The motion of these racing snails can be described by their speeds and their velocities. (credit: tobiasflickr, Flickr)

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.

Time

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, abbreviated 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, of course, 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 Δt** is the difference between the ending time and beginning time,

$$\Delta t = t_f - t_0,$$

where Δt 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, Δ , means the change in the quantity that follows it.)

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.

Definition: AVERAGE VELOCITY

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

$$\bar{v} = \frac{\Delta x}{\Delta t} = \frac{x_f - x_0}{t_f - t_0},$$

where $\bar{v}$ is the *average* (indicated by the bar over the v) velocity, Δx is the change in position (or displacement), and x_f and x_0 are the final and beginning positions at times t_f and t_0 , respectively.

Velocity has both magnitude (size) 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

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

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.

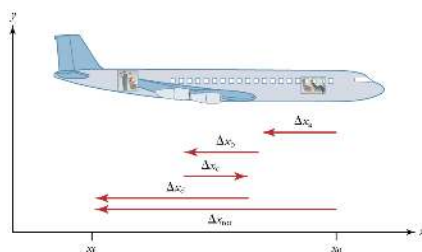


Figure 4.1.2.2: A more detailed record of an airplane passenger heading toward the back of the plane, showing smaller segments of his trip.

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.) **Instantaneous velocity v** is the average velocity at a specific instant in time (or over an infinitesimally small time interval).

Mathematically, finding instantaneous velocity, v , at a precise instant t can involve taking a limit, a calculus operation beyond the scope of this text. However, under many circumstances, we can find precise values for instantaneous velocity without calculus.

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. 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 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.

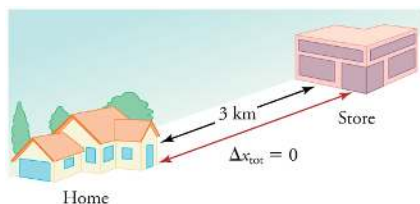


Figure 4.1.2.3: During a 30-minute round trip to the store, the total distance traveled is 6 km. The average speed is 12 km/h. The displacement for the round trip is zero, since there was no net change in position. Thus the average velocity is zero.

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 4.1.2.4. (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.)

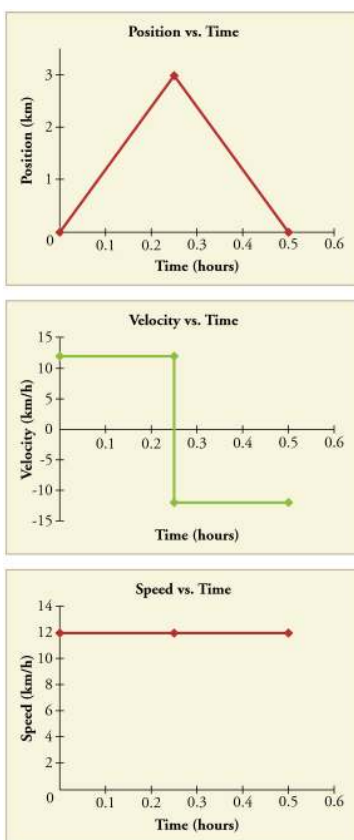


Figure 4.1.2.4: Position vs. time, velocity vs. time, and speed vs. time on a trip. Note that the velocity for the return trip is negative.

✚ MAKING CONNECTIONS: TAKE-HOME INVESTIGATION—GETTING A SENSE OF SPEED

If you have spent much time driving, you probably have a good sense of speeds between about 10 and 70 miles per hour. But what are these in meters per second? What do we mean when we say that something is moving at 10 m/s? To get a better sense of what these values really mean, do some observations and calculations on your own:

- calculate typical car speeds in meters per second
- estimate jogging and walking speed by timing yourself; convert the measurements into both m/s and mi/h
- determine the speed of an ant, snail, or falling leaf

? Exercise 4.1.2.1

A commuter train travels from Baltimore to Washington, DC, and back in 1 hour and 45 minutes. The distance between the two stations is approximately 40 miles. What is (a) the average velocity of the train, and (b) the average speed of the train in m/s?

Answer

(a) The average velocity of the train is zero because $x_f = x_0$; the train ends up at the same place it starts.

(b) The average speed of the train is calculated below. Note that the train travels 40 miles one way and 40 miles back, for a total distance of 80 miles.

$$\frac{\text{distance}}{\text{time}} = \frac{80 \text{ miles}}{105 \text{ minutes}} \quad (4.1.2.1)$$

$$\frac{80 \text{ miles}}{105 \text{ minutes}} \times \frac{5280 \text{ feet}}{1 \text{ mile}} \times \frac{1 \text{ meter}}{3.28 \text{ feet}} \times \frac{1 \text{ minute}}{60 \text{ seconds}} = 20 \text{ m/s} \quad (4.1.2.2)$$

Section Summary

- Time is measured in terms of change, and its SI unit is the second (s). Elapsed time for an event is

$$\Delta t = t_f - t_0,$$

where t_f is the final time and t_0 is the initial time.

- Average velocity $\bar{v}$ is defined as displacement divided by the travel time. In symbols, average velocity is

$$\bar{v} = \frac{\Delta x}{\Delta t} = \frac{x_f - x_0}{t_f - t_0}.$$

- The SI unit for velocity is m/s.
- Velocity has a direction.
- Instantaneous velocity v is the velocity at a specific instant or the average velocity for an infinitesimal interval.
- Instantaneous speed is the magnitude of the instantaneous velocity.
- Instantaneous speed has no direction specified.
- Average speed is the total distance traveled divided by the elapsed time. (Average speed is *not* the magnitude of the average velocity.) Speed has no direction associated with it.

Glossary

average speed

distance traveled divided by time during which motion occurs

average velocity

displacement divided by time over which displacement occurs

instantaneous velocity

velocity at a specific instant, or the average velocity over an infinitesimal time interval

instantaneous speed

magnitude of the instantaneous velocity

time

change, or the interval over which change occurs

model

simplified description that contains only those elements necessary to describe the physics of a physical situation

elapsed time

the difference between the ending time and beginning time

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4.1.3: Acceleration

Learning Objectives

- Define and distinguish between velocity and acceleration, and between instantaneous and average acceleration.
- Calculate acceleration given initial time, initial velocity, final time, and final velocity.



Figure 4.1.3.1: A plane slows down as it comes in for landing in St. Maarten. It is accelerating in a direction opposite to its velocity.
(credit: Steve Conry, Flickr)

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.

Definition: AVERAGE ACCELERATION

Average Acceleration is the rate at which velocity changes,

$$\bar{a} = \frac{\Delta v}{\Delta t} = \frac{v_f - v_0}{t_f - t_0},$$

where $\bar{a}$ is average acceleration, v is velocity, and t is time. (The bar over the a means *average* acceleration.)

Because acceleration is velocity in m/s divided by time in s, the SI units for acceleration are **m/s²**, 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 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 in the same direction as the *change* in velocity, Δv . Since velocity can change either in magnitude or in direction, acceleration is therefore a change in either speed or 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*. If acceleration is in a direction opposite to the direction of motion, the object slows down.



Figure 4.1.3.2: A subway train in Sao Paulo, Brazil, slows down as it comes into a station. It is accelerating in a direction opposite to its direction of motion. (credit: Yusuke Kawasaki, Flickr)

✓ Example 4.1.3.1: Calculating Acceleration: A Racehorse Leaves the Gate

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



Figure 4.1.3.3

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.

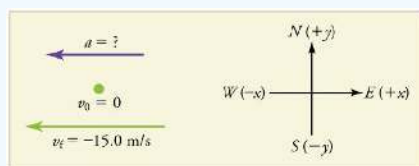


Figure 4.1.3.4

We can solve this problem by identifying Δv and Δt from the given information and then calculating the average acceleration directly from the equation $\bar{a} = \frac{\Delta v}{\Delta t} = \frac{v_f - v_0}{t_f - t_0}$.

Solution

1. Identify the knowns. $v_0 = 0$, $v_f = -15.0 \text{ m/s}$ (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 m/s , its change in velocity equals its final velocity: $\Delta v = v_f = -15.0 \text{ m/s}$.
3. Plug in the known values (Δv and Δt) and solve for the unknown $\bar{a}$.

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

Discussion

The negative sign for acceleration indicates that acceleration is toward the west. An acceleration of 8.33 m/s^2 due west means that the horse increases its velocity by 8.33 m/s due west each second, that is, 8.33 meters per second per second, which we write as 8.33 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.

Instantaneous Acceleration

Instantaneous acceleration a , or the *acceleration at a specific instant in time*, is obtained by the same process as discussed for instantaneous velocity—that is, by considering an infinitesimally 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 4.1.3.5 shows graphs of instantaneous acceleration versus time for two very different motions. In Figure 4.1.3.5 (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 m/s^2). In Figure 4.1.3.5 (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 to 1.0 s and from 1.0 to 3.0 s as separate motions with accelerations of $+3.0 \text{ m/s}^2$ and -2.0 m/s^2 , respectively.

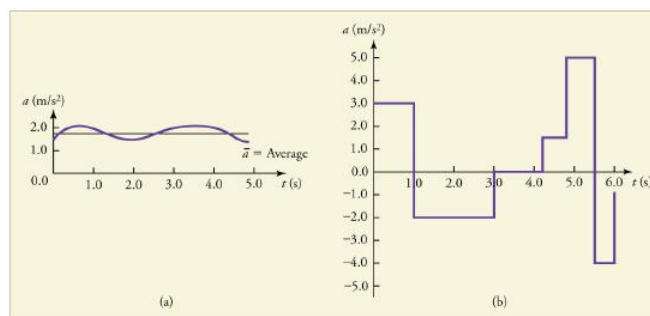


Figure 4.1.3.5: Graphs of instantaneous acceleration versus time for two different one-dimensional motions. (a) Here acceleration varies only slightly and is always in the same direction, since it is positive. The average over the interval is nearly the same as the acceleration at any given time. (b) 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 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 4.1.3.6. 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.

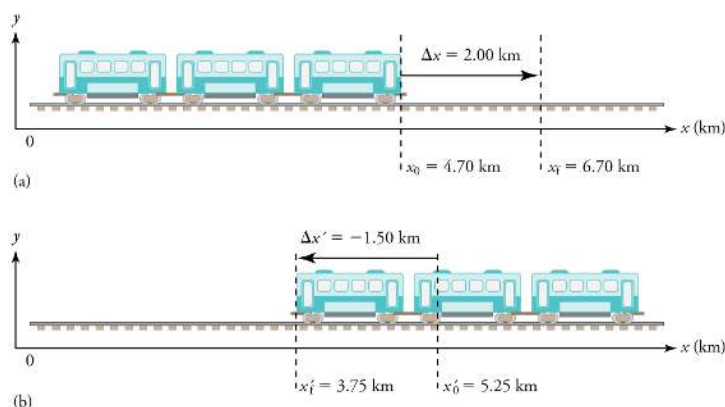


Figure 4.1.3.6: One-dimensional motion of a subway train considered in Example 4.1.3.2, Example 4.1.3.3, Example 4.1.3.4, Example 4.1.3.5, Example 4.1.3.6, and Example 4.1.3.7. Here we have chosen the x -axis so that + means to the right and - means to the left for displacements, velocities, and accelerations. (a) The subway train moves to the right from x_0 to x_f . Its displacement Δx is +2.0 km. (b) The train moves to the left from x'_0 to x'_f . Its displacement $\Delta x'$ is -1.5 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 4.1.3.2: 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 4.1.3.6?

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 4.1.3.3: 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 4.1.3.6?

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 4.1.3.2. Distance traveled is the total length of the path traveled between the two positions. In the case of the subway train shown in Figure 4.1.3.6, 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 $+2.00 \text{ km}$. 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 .

Discussion

Distance has magnitude but no sign to indicate direction.

✓ Example 4.1.3.4: Calculating Acceleration: A Subway Train Speeding Up

Suppose the train in Figure 4.1.3.6 (a) accelerates from rest to 30.0 km/h in the first 20.0 s of its motion. What is its average acceleration during that time interval?

Strategy

It is worth it at this point to make a simple sketch:

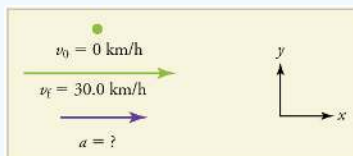


Figure 4.1.3.7

This problem involves three steps. First we must determine the change in velocity, then we must determine the change in time, and finally we use these values to calculate the acceleration.

Solution

1. Identify the knowns. $v_0 = 0$ (the train starts at rest), $v_f = 30.0 \text{ km/h}$, and $\Delta t = 20.0 \text{ s}$.
2. Calculate Δv . Since the train starts from rest, its change in velocity is $\Delta v = +30.0 \text{ km/h}$, where the plus sign means velocity to the right.
3. Plug in known values and solve for the unknown, $\bar{a}$.

$$\bar{a} = \frac{\Delta v}{\Delta t} = \frac{+30.0 \text{ km/h}}{20.0 \text{ s}}$$

4. Since the units are mixed (we have both hours and seconds for time), we need to convert everything into SI units of meters and seconds.

$$\bar{a} = \left(\frac{+30 \text{ km/h}}{20.0 \text{ s}} \right) \left(\frac{10^3 \text{ m}}{1 \text{ km}} \right) \left(\frac{1 \text{ h}}{3600 \text{ s}} \right) = 0.417 \text{ m/s}^2$$

Discussion

The plus sign means that acceleration is to the right. This is reasonable because the train starts from rest and ends up with a velocity to the right (also positive). So acceleration is in the same direction as the *change* in velocity, as is always the case.

✓ Example 4.1.3.5: Calculate Acceleration: A Subway Train Slowing Down

Now suppose that at the end of its trip, the train in Figure 4.1.3.6 (a) slows to a stop from a speed of 30.0 km/h in 8.00 s. What is its average acceleration while stopping?

Strategy

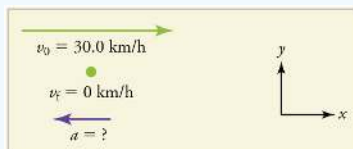


Figure 4.1.3.8

In this case, the train is slowing down and its acceleration is negative because it is toward the left. As in the previous example, we must find the change in velocity and the change in time and then solve for acceleration.

Solution

1. Identify the knowns. $v_0 = 30.0 \text{ km/h}$, $v_f = 0 \text{ km/h}$ (the train is stopped, so its velocity is 0), and $\Delta t = 8.00 \text{ s}$.
2. Solve for the change in velocity, Δv .

$$\Delta v = v_f - v_0 = 0 - 30.0 \text{ km/h} = -30.0 \text{ km/h}$$

3. Plug in the knowns, Δv and Δt , and solve for $\bar{a}$.

$$\bar{a} = \frac{\Delta v}{\Delta t} = \frac{-30.0 \text{ km/h}}{8.00 \text{ s}}$$

4. Convert the units to meters and seconds.

$$\bar{a} = \frac{\Delta v}{\Delta t} = \left(\frac{-30.0 \text{ km/h}}{8.00 \text{ s}} \right) \left(\frac{10^3 \text{ m}}{1 \text{ km}} \right) \left(\frac{1 \text{ h}}{3600 \text{ s}} \right) = -1.04 \text{ m/s}^2.$$

Discussion

The minus sign indicates that acceleration is to the left. This sign is reasonable because the train initially has a positive velocity in this problem, and a negative acceleration would oppose the motion.

The graphs of position, velocity, and acceleration vs. time for the trains in Example 4.1.3.4 and Example 4.1.3.5 are displayed in Figure 4.1.3.9. (We have taken the velocity to remain constant from 20 to 40 s, after which the train slows down.)

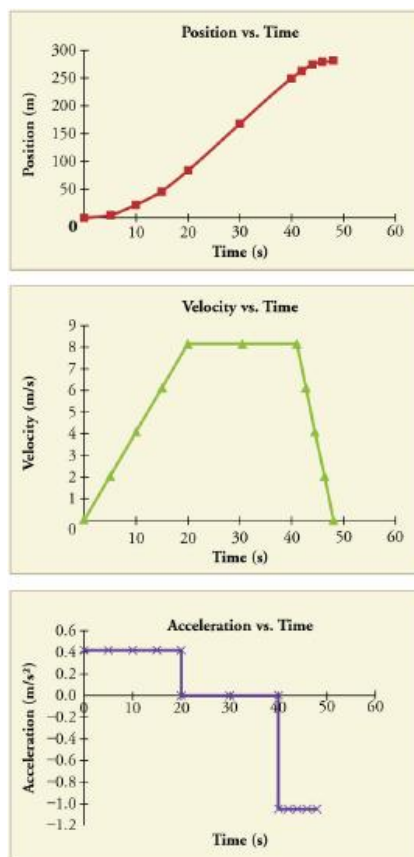


Figure 4.1.3.9: (a) Position of the train over time. Notice that the train's position changes slowly at the beginning of the journey, then more and more quickly as it picks up speed. Its position then changes more slowly as it slows down at the end of the journey. In the middle of the journey, while the velocity remains constant, the position changes at a constant rate. (b) Velocity of the train over time. The train's velocity increases as it accelerates at the beginning of the journey. It remains the same in the middle of the journey (where there is no acceleration). It decreases as the train slows down at the end of the journey. (c) The acceleration of the train over time. The train has positive acceleration as it speeds up at the beginning of the journey. It has no acceleration as it travels at constant velocity in the middle of the journey. Its acceleration is negative as it slows down at the end of the journey.

✓ Example 4.1.3.6: Calculating Average Velocity: The Subway Train

What is the average velocity of the train in part b of Example 4.1.3.2, and shown again below, if it takes 5.00 min to make its trip?

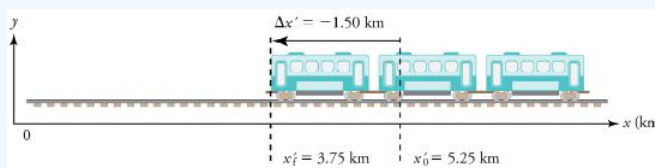


Figure 4.1.3.10

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. $x_f = 3.75 \text{ km}$, $x_i = 5.25 \text{ km}$, $\Delta t = 5.00 \text{ min}$.
2. Determine displacement, $\Delta x'$. We found $\Delta x'$ to be -1.5 km in Example 4.1.3.2.
3. Solve for average velocity.

$$\bar{v} = \frac{\Delta x'}{\Delta t} = \frac{-1.50 \text{ km}}{5.00 \text{ min}}$$

4. Convert units.

$$\bar{v} = \frac{\Delta x'}{\Delta t} = \left(\frac{-1.50 \text{ km}}{5.00 \text{ min}} \right) \left(\frac{60 \text{ min}}{1 \text{ h}} \right) = -18.0 \text{ km/h}$$

Discussion

The negative velocity indicates motion to the left.

✓ Example 4.1.3.7: Calculating Deceleration: The Subway Train

Finally, suppose the train in Figure 4.1.3.10 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:

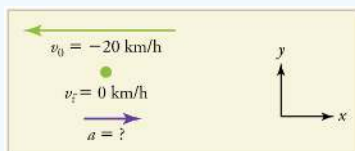


Figure 4.1.3.11

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 Δv . 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 $\bar{a}$.

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

4. Convert units.

$$\bar{a} = \left(\frac{+20.0 \text{ km/h}}{10.0 \text{ s}} \right) \left(\frac{10^3 \text{ m}}{1 \text{ km}} \right) \left(\frac{1 \text{ h}}{3600 \text{ s}} \right) = +0.556 \text{ m/s}^2$$

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).

Sign and Direction

Perhaps the most important thing to note about these examples is the signs of the answers. In our chosen coordinate system, plus means the quantity is to the right and minus means it is to the left. This is easy to imagine for displacement and velocity. But it is a little less obvious for acceleration. Most people interpret negative acceleration as the slowing of an object. This was not the case in Example 4.1.3.7, where a positive acceleration slowed a negative velocity. The crucial distinction was that the acceleration was in the opposite direction from the velocity. In fact, a negative acceleration will *increase* a negative velocity. For example, the train moving to the left in Figure 4.1.3.10 is sped up by an acceleration to the left. In that case, both v and a are negative. The plus and minus signs give the directions of the accelerations. If acceleration has the same sign as the change in velocity, the object is speeding up. If acceleration has the opposite sign of the change in velocity, the object is slowing down.

? Exercise 4.1.3.1

An airplane lands on a runway traveling east. Describe its acceleration.

Answer

If we take east to be positive, then the airplane has negative acceleration, as it is accelerating toward the west. It is also slowing down: its acceleration is opposite in direction to its velocity.

Section Summary

- Acceleration is the rate at which velocity changes. In symbols, **average acceleration** $\bar{a}$ is

$$\bar{a} = \frac{\Delta v}{\Delta t} = \frac{v_f - v_0}{t_f - t_0}$$

- The SI unit for acceleration is m/s^2 .
- Acceleration has both a magnitude and direction.
- Acceleration can be caused by either a change in the magnitude or the direction of the velocity.
- Instantaneous acceleration a is the acceleration at a specific instant in time.
- When an acceleration is in a direction opposite to that of the velocity of an object, the object slows down.

Glossary

acceleration

the rate of change in velocity; the change in velocity over time

average acceleration

the change in velocity divided by the time over which it changes

instantaneous acceleration

acceleration at a specific point in time

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4.2: Force and Newton's Laws of Motion

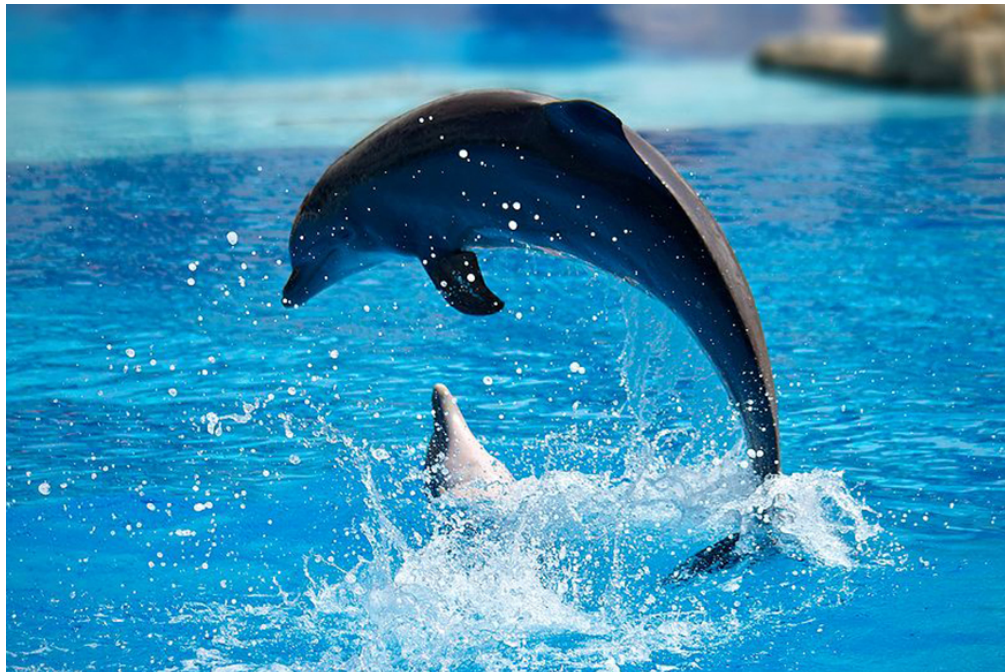


Figure 4.2.1: Newton's laws of

motion describe the motion of the dolphin's path. (credit: Jin Jang)

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.

Issac Newton's (1642–1727) laws of motion were just one part of the monumental work that has made him legendary. The development of Newton's laws marks the transition from the Renaissance into the modern era. This transition was characterized by a revolutionary change in the way people thought about the physical universe. For many centuries natural philosophers had debated the nature of the universe based largely on certain rules of logic with great weight given to the thoughts of earlier classical philosophers such as Aristotle (384–322 BC). Among the many great thinkers who contributed to this change were Newton and Galileo.

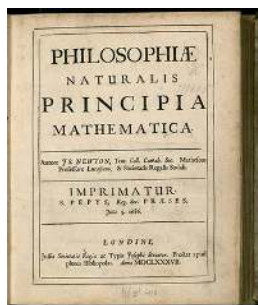


Figure 4.2.2: Issac Newton's monumental work, *Philosophiæ Naturalis Principia Mathematica*, was published in 1687. It proposed scientific laws that are still used today to describe the motion of objects. (credit: Service commun de la documentation de l'Université de Strasbourg)

Galileo was instrumental in establishing *observation* as the absolute determinant of truth, rather than “logical” argument. Galileo's use of the telescope was his most notable achievement in demonstrating the importance of observation. He discovered moons orbiting Jupiter and made other observations that were inconsistent with certain ancient ideas and religious dogma. For this reason, and because of the manner in which he dealt with those in authority, Galileo was tried by the Inquisition and punished. He spent the

final years of his life under a form of house arrest. Because others before Galileo had also made discoveries by *observing* the nature of the universe, and because repeated observations verified those of Galileo, his work could not be suppressed or denied. After his death, his work was verified by others, and his ideas were eventually accepted by the church and scientific communities.

Galileo also contributed to the formation of what is now called Newton's first law of motion. Newton made use of the work of his predecessors, which enabled him to develop laws of motion, discover the law of gravity, invent calculus, and make great contributions to the theories of light and color. It is amazing that many of these developments were made with Newton working alone, without the benefit of the usual interactions that take place among scientists today.

It was not until the advent of modern physics early in the 20th century that it was discovered that Newton's laws of motion produce a good approximation to motion only when the objects are moving at speeds much, much less than the speed of light and when those objects are larger than the size of most molecules (about 10^{-9} m in diameter). These constraints define the realm of classical mechanics. At the beginning of the 20th century, Albert Einstein (1879–1955) developed the theory of relativity and, along with many other scientists, developed quantum theory. This theory does not have the constraints present in classical physics. All of the situations we consider in this chapter are in the realm of classical physics.

MAKING CONNECTIONS: PAST AND PRESENT PHILOSOPHY

The importance of observation and the concept of *cause and effect* were not always so entrenched in human thinking. This realization was a part of the evolution of modern physics from natural philosophy. The achievements of Galileo, Newton, Einstein, and others were key milestones in the history of scientific thought. Most of the scientific theories that are described in this book descended from the work of these scientists.

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4.2.1: Dynamics and Force

Learning Objectives

- Understand the definition of force.

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 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.

A more quantitative definition of force can be based on some standard force, just as distance is measured in units relative to a standard distance. One possibility is to stretch a spring a certain fixed distance, as illustrated in Figure 4.2.1.1, and use the force it exerts to pull itself back to its relaxed shape—called a *restoring force*—as a standard. The magnitude of all other forces can be stated as multiples of this standard unit of force. Many other possibilities exist for standard forces. Some alternative definitions of force will be given later in this chapter.

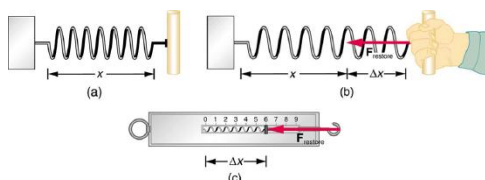


Figure 4.2.1.1: The force exerted by a stretched spring can be used as a standard unit of force. (a) This spring has a length x when undistorted. (b) When stretched a distance Δx , the spring exerts a restoring force, $\mathbf{F}_{\text{restore}}$, which is reproducible. (c) A spring scale is one device that uses a spring to measure force. The force $\mathbf{F}_{\text{restore}}$ is exerted on whatever is attached to the hook. Here $\mathbf{F}_{\text{restore}}$ has a magnitude of 6 units in the force standard being employed.

TAKE-HOME EXPERIMENT: FORCE STANDARDS

To investigate force standards and cause and effect, get two identical rubber bands. Hang one rubber band vertically on a hook. Find a small household item that could be attached to the rubber band using a paper clip, and use this item as a weight to investigate the stretch of the rubber band. Measure the amount of stretch produced in the rubber band with one, two, and four of these (identical) items suspended from the rubber band. What is the relationship between the number of items and the amount of stretch? How large a stretch would you expect for the same number of items suspended from two rubber bands? What happens to the amount of stretch of the rubber band (with the weights attached) if the weights are also pushed to the side with a pencil?

Section Summary

- Dynamics** is the study of how forces affect the motion of objects.
- Force** is a push or pull that can be defined in terms of various standards, and it is a vector having both magnitude and direction.

Glossary

dynamics

the study of how forces affect the motion of objects and systems

force

a push or pull on an object with a specific magnitude and direction; can be expressed as a multiple of a standard force

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4.2.2: Newton's First Law of Motion

Learning Objectives

- Define mass and inertia.
- Understand Newton's first law of motion.

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

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.

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 farther yet. 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.

Mass

The property of a body to remain at rest or to remain in motion with constant velocity is called **inertia**. Newton's first law is often called the **law of inertia**. As we know from experience, some objects have more inertia than others. It is obviously more difficult to change the motion of a large boulder than that of a basketball, for example. The inertia of an object is measured by its **mass**. Roughly speaking, mass is a measure of the amount of “stuff” (or matter) in something. The quantity or amount of matter in an object is determined by the numbers of atoms and molecules of various types it contains. Unlike weight, mass does not vary with location. The mass of an object is the same on Earth, in orbit, or on the surface of the Moon. In practice, it is very difficult to count and identify all of the atoms and molecules in an object, so masses are not often determined in this manner. Operationally, the masses of objects are determined by comparison with the standard kilogram.

? Exercise 4.2.2.1

Which has more mass: a kilogram of cotton balls or a kilogram of gold?

Answer

They are equal. A kilogram of one substance is equal in mass to a kilogram of another substance. The quantities that might differ between them are volume and density.

Section Summary

- **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 known as the **law of inertia**.
- **Inertia** is the tendency of an object to remain at rest or remain in motion. Inertia is related to an object's mass.
- **Mass** is the quantity of matter in a substance.

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4.2.3: Newton's Second Law of Motion

Learning Objectives

- Define net force, external force, and system.
- Understand Newton's second law of motion.
- Apply Newton's second law to determine the weight of an object.

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. 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** of interest. For example, in Figure 4.2.3.1(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 4.2.3.1(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. (The internal forces actually cancel, as we shall see in the next section.) *You must define the boundaries of the system before you can determine which forces are external.* Sometimes the system is obvious, whereas other times identifying the boundaries of a system is more subtle. The concept of a system is fundamental to many areas of physics, as is the correct application of Newton's laws. This concept will be revisited many times on our journey through physics.

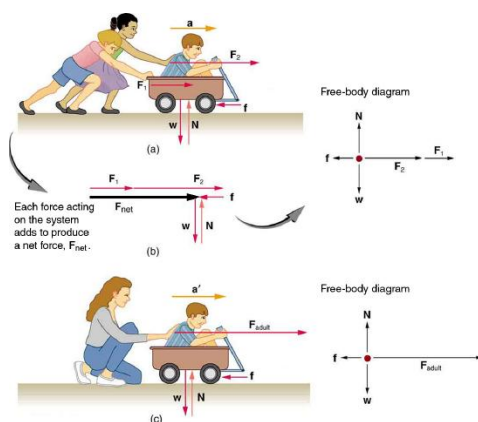


Figure 4.2.3.1: Different forces exerted on the same mass produce different accelerations. (a) Two children push a wagon with a child in it. Arrows representing all external forces are shown. The system of interest is the wagon and its rider. The weight w of the system and the support of the ground N are also shown for completeness and are assumed to cancel. 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 extends from this dot. Because there are two forces acting to the right, we draw the lines collinearly. (c) A larger net external force produces a larger acceleration ($|a'| > |a|$) when an adult pushes the child.

Figure 4.2.3.1 is our first example of a **free-body diagram**, which is a technique used to illustrate all the external forces acting on a body. The body is represented by a single isolated point (or free body), and only those forces acting on the body from the outside (external forces) are shown. Free-body diagrams are very useful in analyzing forces acting on a system and are employed extensively in the study and application of Newton's laws of motion.

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 4.2.3.1. 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 w and the support of the ground N , and the horizontal force f represents the force of friction. These will be discussed in more detail in later sections. For now, we will say that friction is a force that opposes the motion past each other of objects that are touching. Figure 4.2.3.1(b) shows how external forces add together to produce a net force, F_{net} .

To obtain an equation for Newton's second law, we first write the relationship of acceleration and net external force as the proportionality

$$a \propto F_{\text{net}}$$

where the symbol $\propto$ means "proportional to," and F_{net} is the **net external force**, the sum of all external forces. This proportionality states what we have said in words—*acceleration is directly proportional to the net external force*. Once the system of interest is chosen, it is important to identify the external forces, so that the internal forces can be ignored (again, as we will see later, they exactly cancel out, allowing for a tremendous simplification).

Now, hopefully it also seems reasonable that acceleration should be inversely proportional to the mass of the system. That is, if the system has more "stuff", then for the same net external force applied, the acceleration is less. And indeed, as illustrated in Figure 4.2.3.2, the same net external force applied to a car produces a much smaller acceleration than when applied to a basketball. The proportionality is written as

$$a \propto \frac{1}{m}$$

where m is the mass of the system. In fact, this is how we will *define* mass. That is, for an object B to have double the mass of an object A, the same net external force applied on object B produces half the acceleration it does on object A.

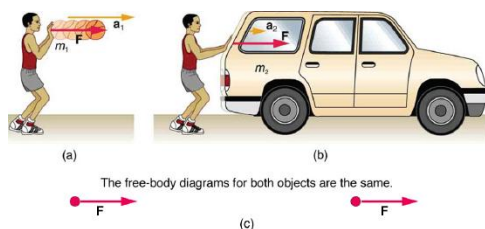


Figure 4.2.3.2: The same force exerted on systems of different masses produces different accelerations. (a) A basketball player pushes on a basketball to make a pass. (The effect of gravity on the ball is ignored.) (b) The same player exerts an identical force on a stalled SUV and produces a far smaller acceleration (even if friction is negligible). (c) The free-body diagrams are identical, permitting direct comparison of the two situations. A series of patterns for the free-body diagram will emerge as you do more problems.

As it turns out, the acceleration of an object depends only on the net external force and the mass of the object. Combining the two proportionalities just given yields Newton's second law of motion.

NEWTON'S SECOND LAW OF MOTION

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

$$a = \frac{F_{\text{net}}}{m}$$

This is often written in the more familiar form

$$F_{\text{net}} = ma,$$

with the notation indicating that the net external force is in the same direction as acceleration. When only the magnitude of force and acceleration are considered, this equation is simply

$$F_{\text{net}} = ma.$$

Although these 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, acceleration is *caused* by a net external force, not vice versa, as the second equation might mistakenly imply.

Units of Force

$F_{\text{net}} = ma$ 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 1 N is the force needed to accelerate a system of mass 1 kg at the rate of 1 m/s^2 . Putting these together,

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

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 $1 \text{ N} = 0.225 \text{ lb}$.

Weight and the Gravitational Force

When an object is dropped, it accelerates toward the center of Earth. Newton's second law states that a net force on an object is responsible for its acceleration. If air resistance is negligible, the net force on a falling object is the gravitational force, commonly called its **weight**, w . Weight can be denoted as w because it has a direction; *down* is, by definition, the direction of gravity, so w is directed downward. The magnitude of weight is denoted as w . Galileo was instrumental in showing that, in the absence of air resistance, all objects fall with the same acceleration g . Using Galileo's result and Newton's second law, we can derive an equation for the magnitude of weight.

Consider an object with mass m falling downward toward Earth. It experiences only the downward force of gravity, which has magnitude w . Newton's second law states that the magnitude of the net external force on an object is $F_{\text{net}} = ma$.

Since the object experiences only the downward force of gravity, $F_{\text{net}} = w$. We know that the acceleration of an object due to gravity is g , or $a = g$. Substituting these into Newton's second law gives

Definition: WEIGHT

This is the equation for *weight*—the gravitational force on a mass m :

$$w = mg.$$

Since $g = 9.80 \text{ m/s}^2$ on Earth, the weight of a 1.0 kg object on Earth is 9.8 N, as we see:

$$w = mg = (1.0 \text{ kg})(9.80 \text{ m/s}^2) = 9.8 \text{ N}.$$

When the net external force on an object is its weight, we say that it is in **free-fall**. That is, the only force acting on the object is the force of gravity. In the real world, when objects fall downward toward Earth, they are never truly in free-fall because there is always some upward force from the air acting on the object.

The acceleration due to gravity g varies slightly over the surface of Earth, so that the weight of an object depends on location and is not an intrinsic property of the object. Weight varies dramatically if one leaves Earth. On the Moon, for example, the acceleration due to gravity is only 1.67 m/s^2 . A 1.0-kg mass thus has a weight of 9.8 N on Earth and only about 1.7 N on the Moon.

The broadest definition of weight in this sense is that *the weight of an object is the gravitational force on it from the nearest large body*, such as Earth, the Moon, the Sun, and so on. This is the most common and useful definition of weight in physics. It differs dramatically, however, from the definition of weight used by NASA and the popular media in relation to space travel and exploration. When they speak of “weightlessness” and “microgravity,” they are really referring to the phenomenon we call “free-fall” in physics. We shall use the above definition of weight, and we will make careful distinctions between free-fall and actual weightlessness.

It is important to be aware that weight and mass are very different physical quantities, although they are closely related. Mass is the quantity of matter (how much “stuff”) and does not vary, whereas weight is the gravitational force and does vary depending on

gravity. It is tempting to equate the two, since most of our examples take place on Earth, where the weight of an object only varies a little with the location of the object. Furthermore, the terms *mass* and *weight* are used interchangeably in everyday language; for example, our medical records often show our “weight” in kilograms, but never in the correct units of newtons.

COMMON MISCONCEPTIONS: MASS VS. WEIGHT

Mass and weight are often used interchangeably in everyday language. However, in science, these terms are distinctly different from one another. Mass is a measure of how much matter is in an object. The typical measure of mass is the kilogram (or the “slug” in English units). Weight, on the other hand, is a measure of the force of gravity acting on an object. Weight is equal to the mass of an object (m) multiplied by the acceleration due to gravity (g). Like any other force, weight is measured in terms of newtons (or pounds in English units). This interchangeable use is why you might have heard an expression such as, “1 kilogram is 2.2 pounds.” The *correct* statement is “1 kilogram mass weighs 2.2 pounds on Earth.”

Assuming the mass of an object is kept intact, it will remain the same, regardless of its location. However, because weight depends on the acceleration due to gravity, the weight of an object *can change* when the object enters into a region with stronger or weaker gravity. For example, the acceleration due to gravity on the Moon is 1.67 m/s^2 (which is much less than the acceleration due to gravity on Earth, 9.80 m/s^2). If you measured your weight on Earth and then measured your weight on the Moon, you would find that you “weigh” much less, even though you do not look any skinnier. This is because the force of gravity is weaker on the Moon. In fact, when people say that they are “losing weight,” they really mean that they are losing “mass” (which in turn causes them to weigh less).

TAKE-HOME EXPERIMENT: MASS AND WEIGHT

What do bathroom scales measure? When you stand on a bathroom scale, what happens to the scale? It depresses slightly. The scale contains springs that compress in proportion to your weight—similar to rubber bands expanding when pulled. The springs provide a measure of your weight (for an object which is not accelerating). This is a force in newtons (or pounds). In most countries, the measurement in newtons is divided by 9.80 to give a reading in mass units of kilograms. The scale measures weight but is calibrated to provide information about mass. While standing on a bathroom scale, push down on a table next to you. What happens to the reading? Why? Would your scale measure the same “mass” on Earth as on the Moon?

✓ Example 4.2.3.1: 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?



Figure 4.2.3.3: 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 F_{net} and m are given, the acceleration can be calculated directly from Newton’s second law as stated in $F_{\text{net}} = ma$.

Solution

The magnitude of the acceleration a is $a = \frac{F_{\text{net}}}{m}$. Entering known values gives

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

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

$$a = \frac{51 \text{ kg} \cdot \text{m/s}^2}{24 \text{ kg}} = 2.1 \text{ m/s}^2.$$

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 forward must be greater than the friction opposing the motion (directed backward) to result in a forward net force, and the vertical forces must cancel if there is to be no acceleration in the vertical direction. The acceleration found is small enough to be reasonable for a person pushing a mower. When the person reaches the top speed, the acceleration will be zero (no change of velocity) and the force exerted by the person pushing the mower will be equal to the friction opposing the motion.

Section Summary

- Acceleration, a , is defined as a rate of change in velocity, resulting from a change in the magnitude and/or the direction of the velocity.
- An external force is one acting on a system from outside the system, as opposed to internal forces, which act between components within the system.
- Newton's second law of motion states 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 $a = \frac{F_{\text{net}}}{m}$, often written in the more familiar form: $F_{\text{net}} = ma$.
- The weight w of an object is defined as the force of gravity acting on an object of mass m . Given acceleration due to gravity g , the magnitude of weight is:

$$w = mg$$

Glossary

acceleration

the rate at which an object's velocity changes over a period of time

free-fall

a situation in which the only force acting on an object is the force due to gravity

external force

a force acting on an object or system that originates outside of the object or system

net external force

the vector sum of all external forces acting on an object or system; causes a mass to accelerate

free-body diagram

a sketch showing all of the external forces acting on an object or system; the system is represented by a dot, and the forces are represented by vectors extending outward from the dot

Newton's second law of motion

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.

system

the object or the group of objects under consideration

weight

the force due to gravity; $w = mg$ for objects on Earth

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4.2.4: Newton's Third Law of Motion

Learning Objectives

- Understand Newton's third law of motion.
- Apply Newton's third law to define systems and solve problems of motion.

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 by the second body that is equal in magnitude and opposite in direction to the force that it itself exerts.

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 4.2.4.1. 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 bodies*. In this case, there are two bodies to consider: the swimmer and the wall. If we define the swimmer as the system of interest, as in the figure, then $\mathbf{F}_{\text{wall on feet}}$ is an external force on this system and affects its motion. The swimmer moves in the direction of $\mathbf{F}_{\text{wall on feet}}$. In contrast, the force $\mathbf{F}_{\text{feet on wall}}$ acts on the wall and not on our system of interest. Thus $\mathbf{F}_{\text{feet on wall}}$ does not affect the motion of the system and does not counteract $\mathbf{F}_{\text{wall on feet}}$. Note that the swimmer pushes the wall in the direction opposite to that in which she wishes to move. The reaction force by the wall on her pushes her in the desired direction.

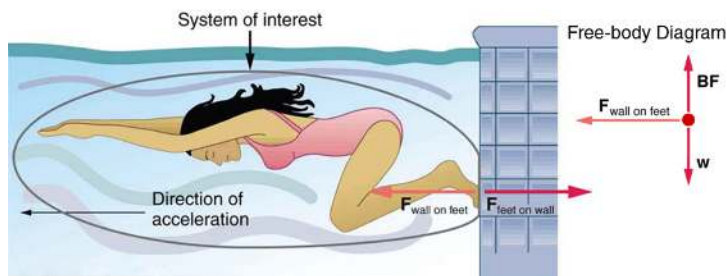


Figure 4.2.4.1: When the swimmer exerts a force $\mathbf{F}_{\text{feet on wall}}$ on the wall, she accelerates in the direction opposite to that of her push. This means the net external force on her is in the direction opposite to $\mathbf{F}_{\text{feet on wall}}$. This opposition occurs because, in accordance with Newton's third law of motion, the wall exerts a force $\mathbf{F}_{\text{wall on feet}}$ 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 $\mathbf{F}_{\text{feet on wall}}$ does not act on this system (the swimmer) and, thus, does not cancel $\mathbf{F}_{\text{wall on feet}}$. Thus the free-body diagram shows only $\mathbf{F}_{\text{wall on feet}}$, w , the gravitational force, and $\mathbf{BF}$, the buoyant force of the water supporting the swimmer's weight. The vertical forces w and $\mathbf{BF}$ 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 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. The rockets propel themselves by pushing their own exhaust gases, and that is how they work even in a vacuum. Helicopters 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.

COMMON MISCONCEPTION: REACTION VS. RESULT

The phrases "action" and "reaction" in reference to the pair of forces described by Newton's third law often leads to unfortunate misunderstandings regarding the nature of the reaction force. Because of the common English usage of the word "reaction" (root word: react)—for example, in film, "reaction shot" means a shot of actors' faces as they react to an event that just happened—many students intuitively guess that there is a cause-and-effect relationship between "action force" and "reaction force."

There is no such relationship. Newton's third law describes *symmetry* in forces. Action force and reaction force are symmetric—they occur simultaneously; they are two sides of one and the same interaction. Action force does not cause reaction force, no more than the head of a coin causes the tail of the coin (they merely occur together, always). Here's a quick check to see if you understood the pairing of action force and reaction force correctly: try swapping the labels. That is, if you originally called A "action force" and B "reaction force", try calling B "action force" and A "reaction force." Does this re-labeling make some sense? If not, chances are, you have mis-identified the pair of action force and reaction force. (Here, a little weirdness is O.K., but if it feels like complete nonsense to swap the labels—like reversal of a cause-and-effect relationship—that tells you you used "reaction" in the common English usage, not how the word is used in Newton's third law.)

Section Summary

- **Newton's third law of motion** represents a basic symmetry in nature. It states: 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 the first body exerts.
- A **thrust** is a reaction force that pushes a body forward in response to a backward force. Rockets, airplanes, and cars are pushed forward by a thrust reaction force.

Glossary

Newton's third law of motion

Whenever one body exerts a force on a second body, the first body experiences a force by the second body that is equal in magnitude and opposite in direction to the force that it itself exerts.

thrust

a reaction force that pushes a body forward; rockets, airplanes, and cars are pushed forward by a thrust, a reaction force to propellants pushed backward

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4.2.5: The Four Basic Forces

Learning Objectives

By the end of this section, you will be able to:

- Understand the four basic forces that underlie the processes in nature.

One of the most remarkable simplifications in physics is that only four distinct forces account for all known phenomena. In fact, nearly all of the forces we experience directly are due to only one basic force, called the electromagnetic force. (The gravitational force is the only force we experience directly that is not electromagnetic.) This is a tremendous simplification of the myriad of *apparently* different forces we can list, only a few of which were discussed in the previous section. As we will see, the basic forces are all thought to act through the exchange of microscopic carrier particles, and the characteristics of the basic forces are determined by the types of particles exchanged. Action at a distance, such as the gravitational force of Earth on the Moon, is explained by the existence of a force field rather than by “physical contact.”

The *four basic forces* are the gravitational force, the electromagnetic force, the weak nuclear force, and the strong nuclear force. Their properties are summarized in Table. Since the weak and strong nuclear forces act over an extremely short range, the size of a nucleus or less, we do not experience them directly, although they are crucial to the very structure of matter. These forces determine which nuclei are stable and which decay, and they are the basis of the release of energy in certain nuclear reactions. Nuclear forces determine not only the stability of nuclei, but also the relative abundance of elements in nature. The properties of the nucleus of an atom determine the number of electrons it has and, thus, indirectly determine the chemistry of the atom. More will be said of all of these topics in later chapters.

CONCEPT CONNECTIONS: THE FOUR BASIC FORCES

The four basic forces will be encountered in more detail as you progress through the text. The gravitational force is defined in [Uniform Circular Motion and Gravitation](#), electric force in [Electric Charge and Electric Field](#), magnetic force in [Magnetism](#), and nuclear forces in [Radioactivity and Nuclear Physics](#). On a macroscopic scale, electromagnetism and gravity are the basis for all forces. The nuclear forces are vital to the substructure of matter, but they are not directly experienced on the macroscopic scale.

Force	Approximate relative strengths	Range	Attraction/Repulsion	Carrier Particle
Gravitational	10^{-38}	∞	attractive only	Graviton
Electromagnetic	10^{-2}	∞	attractive and repulsive	Photon
Weak Nuclear	10^{-13}	$< 10^{-18} \text{ m}$	attractive and repulsive	W^+, W^-, Z^0
Strong Nuclear	1	$< 10^{-15} \text{ m}$	attractive and repulsive	gluons

The gravitational force is surprisingly weak—it is only because gravity is always attractive that we notice it at all. Our weight is the gravitational force due to the *entire* Earth acting on us. On the very large scale, as in astronomical systems, the gravitational force is the dominant force determining the motions of moons, planets, stars, and galaxies. The gravitational force also affects the nature of space and time. As we shall see later in the study of general relativity, space is curved in the vicinity of very massive bodies, such as the Sun, and time actually slows down near massive bodies.

Electromagnetic forces can be either attractive or repulsive. They are long-range forces, which act over extremely large distances, and they nearly cancel for macroscopic objects. (Remember that it is the *net* external force that is important.) If they did not cancel, electromagnetic forces would completely overwhelm the gravitational force. The electromagnetic force is a combination of electrical forces (such as those that cause static electricity) and magnetic forces (such as those that affect a compass needle). These two forces were thought to be quite distinct until early in the 19th century, when scientists began to discover that they are different manifestations of the same force. This discovery is a classical case of the *unification of forces*. Similarly, friction, tension, and all of

the other classes of forces we experience directly (except gravity, of course) are due to electromagnetic interactions of atoms and molecules. It is still convenient to consider these forces separately in specific applications, however, because of the ways they manifest themselves.

CONCEPT CONNECTIONS: UNIFYING FORCES

Attempts to unify the four basic forces are discussed in relation to elementary particles later in this text. By “unify” we mean finding connections between the forces that show that they are different manifestations of a single force. Even if such unification is achieved, the forces will retain their separate characteristics on the macroscopic scale and may be identical only under extreme conditions such as those existing in the early universe.

Physicists are now exploring whether the four basic forces are in some way related. Attempts to unify all forces into one come under the rubric of Grand Unified Theories (GUTs), with which there has been some success in recent years. It is now known that under conditions of extremely high density and temperature, such as existed in the early universe, the electromagnetic and weak nuclear forces are indistinguishable. They can now be considered to be different manifestations of one force, called the *electroweak* force. So the list of four has been reduced in a sense to only three. Further progress in unifying all forces is proving difficult—especially the inclusion of the gravitational force, which has the special characteristics of affecting the space and time in which the other forces exist.

While the unification of forces will not affect how we discuss forces in this text, it is fascinating that such underlying simplicity exists in the face of the overt complexity of the universe. There is no reason that nature must be simple—it simply is.

Action at a Distance: Concept of a Field

All forces act at a distance. This is obvious for the gravitational force. Earth and the Moon, for example, interact without coming into contact. It is also true for all other forces. Friction, for example, is an electromagnetic force between atoms that may not actually touch. What is it that carries forces between objects? One way to answer this question is to imagine that a **force field** surrounds whatever object creates the force. A second object (often called a *test object*) placed in this field will experience a force that is a function of location and other variables. The field itself is the “thing” that carries the force from one object to another. The field is defined so as to be a characteristic of the object creating it; the field does not depend on the test object placed in it. Earth’s gravitational field, for example, is a function of the mass of Earth and the distance from its center, independent of the presence of other masses. The concept of a field is useful because equations can be written for force fields surrounding objects (for gravity, this yields $w = mg$ at Earth’s surface), and motions can be calculated from these equations. (See Figure.)

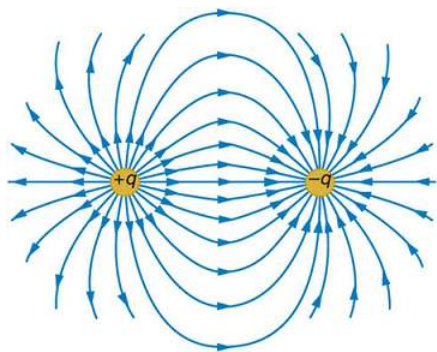


Figure 4.2.5.1: The electric force field between a positively charged particle and a negatively charged particle. When a positive test charge is placed in the field, the charge will experience a force in the direction of the force field lines.

CONCEPT CONNECTIONS: FORCE FIELDS

The concept of a *force field* is also used in connection with electric charge and is presented in [Electric Charge and Electric Field](#). It is also a useful idea for all the basic forces, as will be seen in [Particle Physics](#). Fields help us to visualize forces and how they are transmitted, as well as to describe them with precision and to link forces with subatomic carrier particles.

The field concept has been applied very successfully; we can calculate motions and describe nature to high precision using field equations. As useful as the field concept is, however, it leaves unanswered the question of what carries the force. It has been proposed in recent decades, starting in 1935 with Hideki Yukawa's (1907–1981) work on the strong nuclear force, that all forces are transmitted by the exchange of elementary particles. We can visualize particle exchange as analogous to macroscopic phenomena such as two people passing a basketball back and forth, thereby exerting a repulsive force without touching one another. (See Figure.)

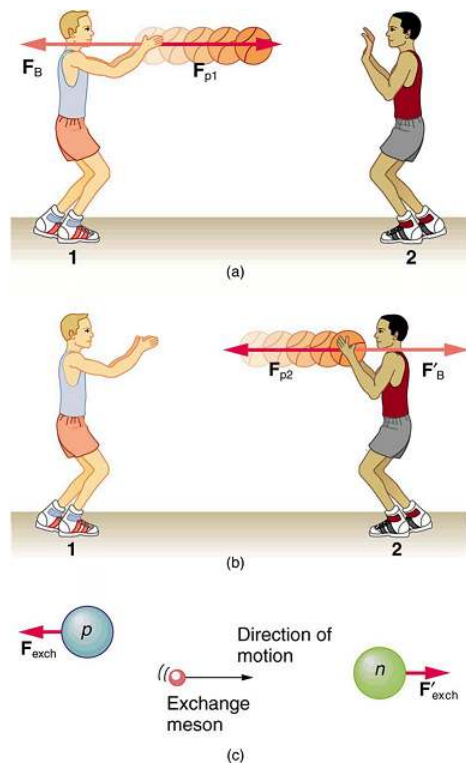


Figure 4.2.5.2: The exchange of masses resulting in repulsive forces. (a) The person throwing the basketball exerts a force F_{p1} on it toward the other person and feels a reaction force F_B away from the second person. (b) The person catching the basketball exerts a force F_{p2} on it to stop the ball and feels a reaction force F'_B away from the first person. (c) The analogous exchange of a meson between a proton and a neutron carries the strong nuclear forces F_{exch} and F'_{exch} . An attractive force can also be exerted by the exchange of a mass—if person 2 pulled the basketball away from the first person as he tried to retain it, then the force between them would be attractive

This idea of particle exchange deepens rather than contradicts field concepts. It is more satisfying philosophically to think of something physical actually moving between objects acting at a distance. Table lists the exchange or **carrier particles**, both observed and proposed, that carry the four forces. But the real fruit of the particle-exchange proposal is that searches for Yukawa's proposed particle found it *and* a number of others that were completely unexpected, stimulating yet more research. All of this research eventually led to the proposal of quarks as the underlying substructure of matter, which is a basic tenet of GUTs. If successful, these theories will explain not only forces, but also the structure of matter itself. Yet physics is an experimental science, so the test of these theories must lie in the domain of the real world. As of this writing, scientists at the CERN laboratory in Switzerland are starting to test these theories using the world's largest particle accelerator: the Large Hadron Collider. This accelerator (27 km in circumference) allows two high-energy proton beams, traveling in opposite directions, to collide. An energy of 14 trillion electron volts will be available. It is anticipated that some new particles, possibly force carrier particles, will be found. (See Figure.) One of the force carriers of high interest that researchers hope to detect is the Higgs boson. The observation of its properties might tell us why different particles have different masses.



Figure 4.2.5.3: The world's largest particle accelerator spans the border between Switzerland and France. Two beams, traveling in opposite directions close to the speed of light, collide in a tube similar to the

central tube shown here. External magnets determine the beam's path. Special detectors will analyze particles created in these collisions. Questions as broad as what is the origin of mass and what was matter like the first few seconds of our universe will be explored. This accelerator began preliminary operation in 2008. (credit: Frank Hommes)

Tiny particles also have wave-like behavior, something we will explore more in a later chapter. To better understand force-carrier particles from another perspective, let us consider gravity. The search for gravitational waves has been going on for a number of years. Almost 100 years ago, Einstein predicted the existence of these waves as part of his general theory of relativity. Gravitational waves are created during the collision of massive stars, in black holes, or in supernova explosions—like shock waves. These gravitational waves will travel through space from such sites much like a pebble dropped into a pond sends out ripples—except these waves move at the speed of light. A detector apparatus has been built in the U.S., consisting of two large installations nearly 3000 km apart—one in Washington state and one in Louisiana! The facility is called the Laser Interferometer Gravitational-Wave Observatory (LIGO). Each installation is designed to use optical lasers to examine any slight shift in the relative positions of two masses due to the effect of gravity waves. The two sites allow simultaneous measurements of these small effects to be separated from other natural phenomena, such as earthquakes. Initial operation of the detectors began in 2002, and work is proceeding on increasing their sensitivity. Similar installations have been built in Italy (VIRGO), Germany (GEO600), and Japan (TAMA300) to provide a worldwide network of gravitational wave detectors.

International collaboration in this area is moving into space with the joint EU/US project LISA (Laser Interferometer Space Antenna). Earthquakes and other Earthly noises will be no problem for these monitoring spacecraft. LISA will complement LIGO by looking at much more massive black holes through the observation of gravitational-wave sources emitting much larger wavelengths. Three satellites will be placed in space above Earth in an equilateral triangle (with 5,000,000-km sides) (Figure). The system will measure the relative positions of each satellite to detect passing gravitational waves. Accuracy to within 10% of the size of an atom will be needed to detect any waves. The launch of this project might be as early as 2018.

“I’m sure LIGO will tell us something about the universe that we didn’t know before. The history of science tells us that any time you go where you haven’t been before, you usually find something that really shakes the scientific paradigms of the day. Whether gravitational wave astrophysics will do that, only time will tell.” —David Reitze, LIGO Input Optics Manager, University of Florida.

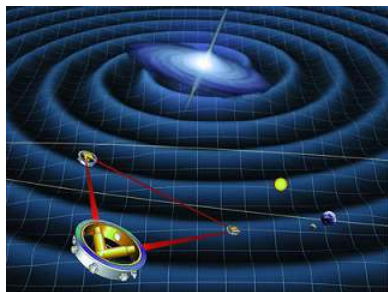


Figure 4.2.5.4: Space-based future experiments for the measurement of gravitational waves.

Shown here is a drawing of LISA's orbit. Each satellite of LISA will consist of a laser source and a mass. The lasers will transmit a signal to measure the distance between each satellite's test mass. The relative motion of these masses will provide information about passing gravitational waves. (credit: NASA)

The ideas presented in this section are but a glimpse into topics of modern physics that will be covered in much greater depth in later chapters.

Summary

- The various types of forces that are categorized for use in many applications are all manifestations of the *four basic forces* in nature.
- The properties of these forces are summarized in [Table](#).
- Everything we experience directly without sensitive instruments is due to either electromagnetic forces or gravitational forces. The nuclear forces are responsible for the submicroscopic structure of matter, but they are not directly sensed because of their short ranges. Attempts are being made to show all four forces are different manifestations of a single unified force.
- A force field surrounds an object creating a force and is the carrier of that force.

Footnotes

The graviton is a proposed particle, though it has not yet been observed by scientists. See the discussion of gravitational waves later in this section. The particles W^+ , W^- , and Z^0 are called vector bosons; these were predicted by theory and first observed in 1983. There are eight types of gluons proposed by scientists, and their existence is indicated by meson exchange in the nuclei of atoms.

Glossary

carrier particle

a fundamental particle of nature that is surrounded by a characteristic force field; photons are carrier particles of the electromagnetic force

force field

a region in which a test particle will experience a force

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4.3: Work and Energy



Figure 4.3.1: How many forms of energy can you identify in this photograph of a wind farm in Iowa? (credit: Jürgen from Sandesneben, Germany, Wikimedia Commons)

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. Not only does energy have 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. We will briefly examine the world's energy use patterns at the end of this chapter.

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. Because of the association of energy with work, we begin the chapter with a discussion of work. Work is intimately related to energy and how energy moves from one system to another or changes form.

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4.3.1: What is Work?

Learning Objectives

- Explain how an object must be displaced for a force on it to do work.
- Explain how relative directions of force and displacement determine whether the work done is positive, negative, or zero.

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 motion or displacement in the direction of the force.

Formally, the **work** done on a system by a constant force is defined to be *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 dimension, this is expressed in equation form as

$$W = Fd$$

as long as one keeps in mind that the force is in the same direction as the distance.

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 dimension, this is expressed in equation form as

$$W = Fd$$

Where W is work, d is the distance the force acts, F is the force parallel to the direction of motion.

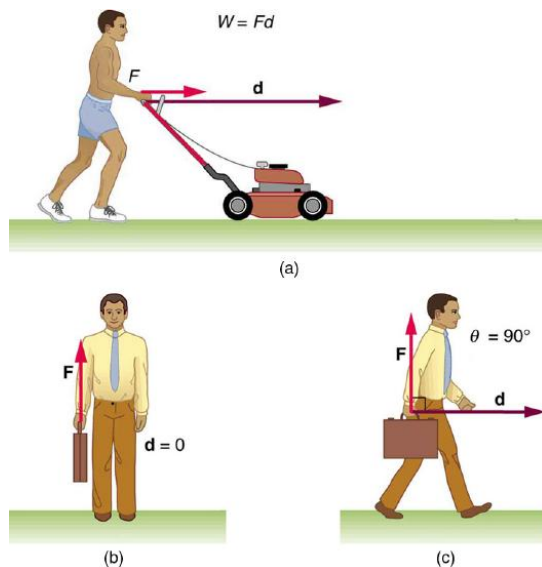


Figure 4.3.1.1: Examples of work. (a) Note that the force is in the direction of motion. (b) A person holding a briefcase does no work on it, because there is no motion. No energy is transferred to or from the briefcase. (c) The person moving the briefcase horizontally at a constant speed does no work on it, and transfers no energy to it. The force acting on the briefcase (against gravity) is perpendicular to the direction of motion.

To examine what the definition of work means, let us consider the other situations shown in Figure 4.3.1.1. The person holding the briefcase in Figure 4.3.1.1(b) does no work, for example. Here $d = 0$, so $W = 0$. Why is it you get tired just holding a load? The

answer is that your muscles are doing work against one another, *but they are doing no work on the system of interest* (the “briefcase-Earth system”). There must be motion for work to be done, and there must be a component of the force in the direction of the motion. For example, the person carrying the briefcase on level ground in Figure 4.3.1.1(c) does no work on it, because the force is perpendicular to the motion.

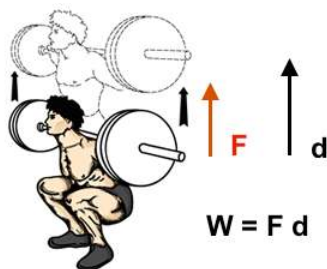


Figure 4.3.1.2: Work done when lifting a mass. Work is positive on the way up because force and displacement point in the same direction. On the way down work is negative due to force and displacement pointing in opposite directions (force up and displacement down).

Work can be positive or negative. In Figure 4.3.1.2 work done lifting the mass is positive because both force and displacement are in the same direction. Likewise, when the mass is lowered the work done is negative because the force and displacement are in *opposite* directions. We will soon see that positive work adds energy to the system and negative work removes energy from a system.

Calculating Work

Work and energy have the same units. From the definition of work, we see that those units are force times distance. Thus, in SI units, work and energy are measured in **newton-meters**. A newton-meter is given the special name **joule (J)**, and $1 \text{ J} = 1 \text{ N} \cdot \text{m} = 1 \text{ kg} \cdot \text{m}^2/\text{s}^2$. One joule is not a large amount of energy; it would lift a small 100-gram apple a distance of about 1 meter.

✓ Example 4.3.1.1: Calculating the Work You Do to Push a Lawn Mower Across a Large Lawn

How much work is done on the lawn mower by the person in Figure 4.3.1.1(a) if he exerts a constant force of 75.0 N and pushes the mower 25.0 m on level ground? Compare it with this person’s average daily intake of 10,000 kJ (about 2400 kcal) of food energy.

Strategy

We can solve this problem by substituting the given values into the definition of work done on a system, stated in the equation $W = Fd$. The force and displacement are given, so that only the work W is unknown. Note that force and displacement are in the same direction.

Solution

The equation for the work is

$$W = Fd.$$

Substituting the known values gives

$$W = (75.0 \text{ N})(25.0 \text{ m}) = 1875 \text{ J} = 1.88 \times 10^3 \text{ J}.$$

The ratio of the work done to the daily consumption is

$$\frac{W}{1.00 \times 10^4 \text{ kJ}} = \frac{1.88 \times 10^3 \text{ J}}{1.00 \times 10^7 \text{ J}} = 1.88 \times 10^{-4}.$$

Discussion

This ratio is a tiny fraction of what the person consumes, but it is typical. Very little of the energy released in the consumption of food is used to do work. Even when we “work” all day long, less than 10% of our food energy intake is used to do work and more than 90% is converted to thermal energy or stored as chemical energy in fat.

✓ Example 4.3.1.2: Calculating the Work Done Lifting a Mass

The 60.0 kg mass shown in [Figure 4.3.1.2](#) is raised at a constant speed through a vertical distance of 0.800 m. What is the force that must be exerted? How much work is done lifting the mass?

Strategy

Constant speed tells us that the upward force must have the same magnitude as the downward force of gravity $F = mg$ (to have a zero net force). Using that information, along with the given values for mass and displacement, we can determine work using $W = Fd$.

Solution

The force is given by

$$F = mg = (60.0 \text{ kg})(9.80 \text{ m/s}^2) = 588 \text{ N}.$$

The work done is

$$W = Fd = (588 \text{ N})(0.800 \text{ m}) = 470 \text{ J}.$$

Discussion

Note that the work done lowering the mass back to its starting position would be -470 J because force and displacement point in opposite directions. The total work done raising and lowering the mass is $470 \text{ J} - 470 \text{ J} = 0 \text{ J}$.

Section Summary

- Work is the transfer of energy by a force acting on an object as it is displaced.
- The work W that a force F does on an object is the product of the magnitude F of the force parallel to the motion, times the magnitude d of the displacement. In symbols,

$$W = Fd.$$

- The SI unit for work and energy is the joule (J), where $1 \text{ J} = 1 \text{ N} \cdot \text{m} = 1 \text{ kg} \cdot \text{m}^2/\text{s}^2$.
- The work done by a force is zero if the displacement is either zero or perpendicular to the force.
- The work done is positive if the force and displacement have the same direction, and negative if they have opposite direction.

Glossary

energy

the ability to do work

work

the transfer of energy by a force that causes an object to be displaced; the product of the component of the force in the direction of the displacement and the magnitude of the displacement

joule

SI unit of work and energy, equal to one newton-meter

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4.3.2: Net Force, Work, and Kinetic Energy

Learning Objectives

- Explain work as a transfer of energy and net work as the work done by the net force.
- Explain and apply the work-energy theorem.

Work Transfers Energy

What happens to the work done on a system? Energy is transferred into the system, but in what form? Does it remain in the system or move on? The answers depend on the situation. For example, if a lawn mower is pushed just hard enough to keep it going at a constant speed, then energy put into the mower by the person is removed continuously by friction, and eventually leaves the system in the form of heat transfer. In contrast, work done on a briefcase by a person carrying it up stairs is stored in the briefcase-Earth system and can be recovered at any time. In fact, the building of the pyramids in ancient Egypt is an example of storing energy in a system by doing work on the system. Some of the energy imparted to the stone blocks in lifting them during construction of the pyramids remains in the stone-Earth system and has the potential to do work.

In this section we begin the study of various types of work and forms of energy. We will find that some types of work leave the energy of a system constant, for example, whereas others change the system in some way, such as making it move. We will also develop definitions of important forms of energy, such as the energy of motion.

Net Work and the Work-Energy Theorem

We know from the study of Newton's laws that net force causes acceleration. We will see in this section that work done by the net force gives a system energy of motion, and in the process we will also find an expression for the energy of motion.

Let us start by considering the total, or net, work done on a system. **Net work** is defined to be the sum of work done by all external forces—that is, net work is the work done by the net external force F_{net} over a distance, d , and is given by the equation

$$W_{\text{net}} = F_{\text{net}}d.$$

By using Newton's second law, and doing some algebra, we can reach an interesting conclusion. Substituting $F=ma$ from Newton's second law gives

$$W_{\text{net}} = mad.$$

After doing more algebraic substitution, we eventually derive the work-energy theorem, which tells us that the net work on a system depends on the mass and velocity of an object.

THE WORK-ENERGY THEOREM

The net work on a system equals the change in the quantity $\frac{1}{2}mv^2$.

$$W_{\text{net}} = \frac{1}{2}mv^2 - \frac{1}{2}mv_0^2$$

The quantity $\frac{1}{2}mv^2$ in the work-energy theorem is defined to be the translational **kinetic energy** (KE) of a mass m moving at a speed v . In equation form, the translational kinetic energy,

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

is the energy associated with translational motion. Kinetic energy is a form of energy associated with the motion of a particle, single body, or system of objects moving together. The work-energy theorem can be compactly written as,

$$W_{\text{net}} = \Delta\text{KE}$$

where ΔKE is understood to mean "change in KE."

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 speed squared. This proportionality means, for example, that a car traveling at 40 m/s has four times the kinetic energy it has at 20 m/s, helping to explain why high-speed collisions are so devastating. We will now consider a series of examples to illustrate various aspects of work and energy.

✓ Example 4.3.2.1: Calculating the Kinetic Energy of a Package

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 v are given, the kinetic energy can be calculated from its definition as given in the equation $\text{KE} = \frac{1}{2}mv^2$.

Solution

The kinetic energy is given by

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

Entering known values gives

$$\text{KE} = 0.5(30.0 \text{ kg})(0.500 \text{ m/s})^2,$$

which yields

$$\text{KE} = 3.75 \text{ kg} \cdot \text{m}^2/\text{s}^2 = 3.75 \text{ J}.$$

Discussion

Note that the unit of kinetic energy is the joule, the same as the unit of work, as mentioned when work was first defined. It is also 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.

Some of the examples in this section can be solved without considering energy, but at the expense of missing out on gaining insights about what work and energy are doing in this situation. On the whole, solutions involving energy are generally shorter and easier than those using kinematics and dynamics alone.

Section Summary

- The net work W_{net} is the work done by the net force acting on an object.
- Work done on an object transfers energy to the object.
- The translational kinetic energy of an object of mass m moving at speed v is $\text{KE} = \frac{1}{2}mv^2$.
- The work-energy theorem states that the net work W_{net} on a system changes its kinetic energy, $W_{\text{net}} = \Delta\text{KE} = \frac{1}{2}mv^2 - \frac{1}{2}mv_0^2$.

Glossary

net work

work done by the net force, or vector sum of all the forces, acting on an object

work-energy theorem

the result, based on Newton's laws, that the net work done on an object is equal to its change in kinetic energy

kinetic energy

the energy an object has by reason of its motion, equal to $\frac{1}{2}mv^2$ for the translational (i.e., non-rotational) motion of an object of mass m moving at speed v

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4.3.3: Forms of Energy

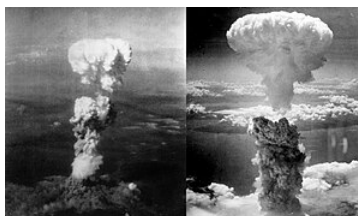
learning objectives

- Compare the different forms of energy interrelate to one another.
- Summarize the consequence of the first law of thermodynamics on the total energy of a system.

Other Forms of Energy

Thermal, chemical, electric, radiant, nuclear, magnetic, elastic, sound, mechanical, luminous, and mass are forms that energy can exist in. Energy can come in a variety of forms. These forms include:

- **Thermal Energy:** This is energy associated with the microscopic random motion of particles in the media under consideration. An example of something that stores thermal energy is warm bath water.
- **Chemical Energy:** This is energy due to the way that atoms are arranged in molecules and various other collections of matter. An example of something that stores chemical energy is food. When your body digests and metabolizes food it utilizes its chemical energy.
- **Electric Energy:** This is energy that is from electrical potential energy, a result of Coulombic forces. Electrical potential energy is associated with the way that point charges in a system are arranged. An example of something that stores electric energy is a capacitor. A capacitor collects positive charge on one plate and negative charge on the other plate. Energy is thus stored in the resulting electrostatic field.
- **Radiant Energy:** This is any kind of electromagnetic radiation (see key term). An example of an electromagnetic wave is light.
- **Nuclear Energy:** This type of energy is liberated during the nuclear reactions of fusion and fission. Examples of things that utilize nuclear energy include nuclear power plants and nuclear weapons.
- **Magnetic Energy:** Technically magnetic energy is electric energy; the two are related by Maxwell's equations. An example of something that stores magnetic energy is a superconducting magnet used in an MRI.
- **Elastic Energy:** This is potential mechanical energy that is stored in the configuration of a material or physical system as work is performed to distort its volume or shape. An example of something that stores elastic energy is a stretched rubber band.
- **Sound Energy:** This is energy that is associated with the vibration or disturbance of matter. An example of something that creates sound energy is your voice box (larynx).
- **Mechanical Energy:** This is energy that is associated with the motion and position of an object. It is the sum of all of the kinetic and potential energy that the object has. An example of something that utilizes mechanical energy is a pendulum.
- **Luminous Energy:** This is energy that can be seen because it is visible light. An example of luminous energy is light from a flashlight.
- **Mass:** Can be converted to energy via: $E = mc^2$. For example, mass is converted into energy when a nuclear bomb explodes.



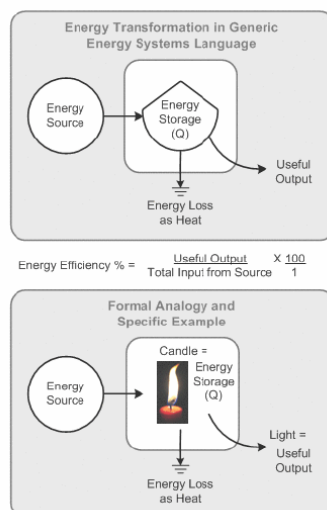
Atomic bomb explosion: The mushroom clouds of the atomic bombings of Hiroshima (left) and Nagasaki (right), Japan

In each of the aforementioned forms, energy exists as either kinetic energy, potential energy, or a combination of both. It is important to note that the above list is not necessarily complete as we may discover new forms of energy in the future such as “dark energy.” Also, each of the forms that energy can take on (as listed above) are not necessarily mutually exclusive. For example, luminous energy is radiant energy.

Energy Transformations

Energy transformation occurs when energy is changed from one form to another, and is a consequence of the first law of thermodynamics.

Energy transformation occurs when energy is changed from one form to another. It is a consequence of the first law of thermodynamics that the total energy of a given system can only be changed when energy is added or subtracted from the system. Often it appears that energy has been lost from a system when it simply has been transformed. For example, an internal combustion engine converts the potential chemical energy in gasoline and oxygen into heat energy. This heat energy is then converted to kinetic energy, which is then used to propel the vehicle that is utilizing the engine. The technical term for a device that converts energy from one form to another is a *transducer*.



Energy Transformation: These figures illustrate the concepts of energy loss and useful energy output.

When analyzing energy transformations, it is important to consider the efficiency of conversion. The efficiency of conversions describes the ratio between the useful output and input of an energy conversion machine. Some energy transformations can occur with an efficiency of essentially 100%. For example, imagine a pendulum in a vacuum. When the pendulum's mass reaches its maximum height, all of its energy exists in the form of potential energy. However, when the pendulum is at its lowest point, all of its energy exists in the form of kinetic energy.

Other energy transformations occur with a much lower efficiency of conversion. For example, the theoretical limit of the energy efficiency of a wind turbine (converting the kinetic energy of the wind to mechanical energy) is 59%. The process of photosynthesis is able to transform the light energy of the sun into chemical energy that can be used by a plant with an efficiency of conversion of a mere 6%.

Key Points

- Thermal, chemical, electric, radiant, nuclear, magnetic, elastic, sound, mechanical, luminous, and mass are forms that energy can exist in.
- Energy exists as either kinetic energy, potential energy, or a combination of both.
- We may discover new forms of energy (like "dark energy") in the future.
- The total energy of a given system can only be changed when energy is added or subtracted from the system.
- Often it appears that energy has been lost from a system when it simply has been transformed.

Key Terms

- **fusion:** A nuclear reaction in which nuclei combine to form more massive nuclei with the concomitant release of energy.
- **electromagnetic radiation:** radiation (quantized as photons) consisting of oscillating electric and magnetic fields oriented perpendicularly to each other, moving through space

- **fission:** The process of splitting the nucleus of an atom into smaller particles; nuclear fission.
- **pendulum:** A body suspended from a fixed support so that it swings freely back and forth under the influence of gravity; it is commonly used to regulate various devices such as clocks.
- **first law of thermodynamics:** A version of the law of conservation of energy, specialized for thermodynamical systems. It is usually formulated by stating that the change in the internal energy of a closed system is equal to the amount of heat supplied to the system, minus the amount of work done by the system on its surroundings.
- **potential energy:** The energy an object has because of its position (in a gravitational or electric field) or its condition (as a stretched or compressed spring, as a chemical reactant, or by having rest mass)
- **kinetic energy:** The energy possessed by an object because of its motion, equal to one half the mass of the body times the square of its velocity.

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4.3.3.1: Transfer of Energy

Learning Objectives

- Define conservative force, potential energy, and mechanical energy.
- Apply conservation of mechanical energy to simple physical situations.
- Explain the law of the conservation of energy.
- Describe some of the many forms of energy.

Potential Energy and Conservative Forces

Work is done by a force, and some forces, such as weight, have special characteristics. A **conservative force** is one, like the gravitational force, for which work done by or against it depends only on the starting and ending points of a motion and not on the path taken. This happens if the force is a function of position alone and not time, velocity, or other parameters. We can define a **potential energy (PE)** for any conservative force, just as we did for the gravitational force. For example, when you wind up a toy, an egg timer, or an old-fashioned watch, you do work against its spring and store energy in it. (We treat these springs as ideal, in that we assume there is no friction and no production of thermal energy.) This stored energy is recoverable as work, and it is useful to think of it as potential energy contained in the spring. Indeed, the reason that the spring has this characteristic is that its force is *conservative*. That is, a conservative force results in stored or potential energy. Gravitational potential energy is one example, as is the energy stored in a spring. We will also see how conservative forces are related to the conservation of energy.

POTENTIAL ENERGY AND CONSERVATIVE FORCES

Potential energy is the energy a system has due to position, shape, or configuration. It is stored energy that is completely recoverable.

A conservative force is a force that is a function of position alone, with the result that the work done by the force depends only on the starting and ending points of a motion and not on the particular path taken.

Change in potential energy comes from the work done against a conservative force. Or flipping this around, potential energy increases as a conservative force does a negative work. That is,

$$\Delta PE = -W_c,$$

where W_c is work done by a conservative force.

Conservation of Mechanical Energy

Consider an object in a system. If it gains any kinetic energy, this is a result of a net work done on the object, according to the work-kinetic energy theorem. If only conservative forces, such as the gravitational force or a spring force, do work in this system, then as kinetic energy increases with the net work done by the conservative forces, the system loses potential energy. That is, $\Delta KE = W_c = -\Delta PE$. In other words,

$$\Delta KE + \Delta PE = 0.$$

This equation means that the total kinetic and potential energy is constant for any process involving only conservative forces. That is,

$$\left. \begin{array}{l} KE + PE = \text{constant} \\ \text{or} \\ KE_i + PE_i = KE_f + PE_f \end{array} \right\} \text{(conservative forces only),}$$

where i and f denote initial and final values. This equation is a form of the work-energy theorem for conservative forces; it is known as the **conservation of mechanical energy** principle. The total kinetic plus potential energy of a system is defined to be its **mechanical energy, (KE+PE)**. In a system that experiences only conservative forces, there is a potential energy associated with each force, and the energy only changes form between **KE** and the various types of **PE**, with the total energy remaining constant.

Conservation of Total 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.

We have explored some forms of energy and some ways it can be transferred from one system to another, in the case of work done by conservative forces transforming between potential energy and kinetic energy. Together, these make up the mechanical energy ($KE+PE$), and the mechanical energy is not always conserved, because it can be transformed into other forms of energy through work done by non-conservative forces. Here, we list some of the many forms energy can take.

Some of the Many Forms of Energy

Here are some of the many forms of energy. You probably have heard of some of these before; many of these will be covered in later chapters, but let us detail a few here. **Electrical energy** is a common form that is converted to many other forms and does work in a wide range of practical situations. Fuels, such as gasoline and food, carry **chemical energy** that can be transferred to a system through oxidation. Chemical fuel can also produce electrical energy, such as in batteries. Batteries can in turn produce light, which is a very pure form of energy. Most energy sources on Earth are in fact stored energy from the energy we receive from the Sun. We sometimes refer to this as **radiant energy**, or electromagnetic radiation, which includes visible light, infrared, and ultraviolet radiation. **Nuclear energy** comes from processes that convert measurable amounts of mass into energy. Nuclear energy is transformed into the energy of sunlight, into electrical energy in power plants, and into the energy of the heat transfer and blast in weapons. Atoms and molecules inside all objects are in random motion. This internal mechanical energy from the random motions is called **thermal energy**, because it is related to the temperature of the object. These and all other forms of energy can be converted into one another and can do work.

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. This important point is discussed later in this section.)

Another example of energy conversion occurs in a solar cell. Sunlight impinging on a solar cell (see [Figure 4.3.3.1.2](#)) 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.



Figure 4.3.3.1.2: Solar energy is converted into electrical energy by solar cells, which is used to run a motor in this solar-power aircraft. (credit: NASA)

Section Summary

- A conservative force is one for which work depends only on the starting and ending points of a motion, not on the path taken.
- We can define potential energy (**PE**) for any conservative force, just as we defined PE_g for the gravitational force.
- Mechanical energy is defined to be $KE+PE$ for a conservative force.
- When only conservative forces act on and within a system, the total mechanical energy is conserved, $\Delta KE + \Delta PE = 0$.
- The law of conservation of energy states that the total energy, including the mechanical energy and other forms of energy, is constant in any process. Energy may change in form or be transferred from one system to another, but the total remains the

same.

- Commonly encountered forms of energy include electric energy, chemical energy, radiant energy, nuclear energy, and thermal energy.

Glossary

conservative force

a force that is a function of position alone, with the result that the work done by the force depends only on the starting and ending points of a motion and not on the particular path taken

potential energy

energy due to position, shape, or configuration

conservation of mechanical energy

the rule that the sum of the kinetic energies and potential energies remains constant if only conservative forces act on and within a system

mechanical energy

the sum of kinetic energy and potential energy

law of conservation of energy

the general law that total energy is constant in any process; energy may change in form or be transferred from one system to another, but the total remains the same

electrical energy

the energy carried by a flow of charge

chemical energy

the energy in a substance stored in the bonds between atoms and molecules that can be released in a chemical reaction

radiant energy

the energy carried by electromagnetic waves

nuclear energy

energy released by changes within atomic nuclei, such as the fusion of two light nuclei or the fission of a heavy nucleus

thermal energy

the energy within an object due to the random motion of its atoms and molecules that accounts for the object's temperature

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4.3.3.2: Heat as Energy

Learning Objectives

- Define heat as transfer of energy.

We have defined work as force times distance and learned that work done on an object changes its kinetic energy. We have also seen that temperature is proportional to the (average) kinetic energy of atoms and molecules. We say that a thermal system has a certain internal energy: its internal energy is higher if the temperature is higher. If two objects at different temperatures are brought in contact with each other, energy is transferred from the hotter to the colder object until equilibrium is reached and the bodies reach thermal equilibrium (i.e., they are at the same temperature). No work is done by either object, because no force acts through a distance. The transfer of energy is caused by the temperature difference, and ceases once the temperatures are equal. These observations lead to the following definition of **heat**: Heat is the spontaneous transfer of energy due to a temperature difference.

Heat is often confused with temperature. For example, we may say the heat was unbearable, when we actually mean that the temperature was high. Heat is a form of energy, whereas temperature is not. The misconception arises because we are sensitive to the flow of heat, rather than the temperature.

Owing to the fact that heat is a form of energy, it has the SI unit of *joule* (J). The *calorie* (cal) is a common unit of energy, defined as the energy needed to change the temperature of 1.00 g of water by 1.00°C —specifically, between 14.5°C and 15.5°C, since there is a slight temperature dependence. Perhaps the most common unit of heat is the **kilocalorie** (kcal), which is the energy needed to change the temperature of 1.00 kg of water by 1.00°C. Since mass is most often specified in kilograms, kilocalorie is commonly used. Food calories (given the notation Cal, and sometimes called “big calorie”) are actually kilocalories (1kilocalorie=1000 calories), a fact not easily determined from package labeling.

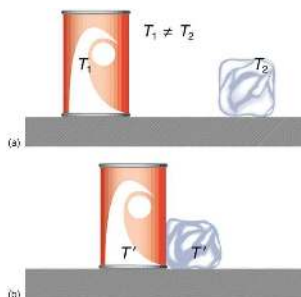


Figure 4.3.3.2.1: In figure (a) the soft drink and the ice have different temperatures, T_1 and T_2 , and are not in thermal equilibrium. In figure (b), when the soft drink and ice are allowed to interact, energy is transferred until they reach the same temperature T' , achieving equilibrium. Heat transfer occurs due to the difference in temperatures. In fact, since the soft drink and ice are both in contact with the surrounding air and bench, the equilibrium temperature will be the same for both.

Summary

- Heat and work are the two distinct methods of energy transfer.
- Heat is energy transferred solely due to a temperature difference.
- Any energy unit can be used for heat transfer, and the most common are kilocalorie (kcal) and joule (J).
- Kilocalorie is defined to be the energy needed to change the temperature of 1.00 kg of water between 14.5°C and 15.5°C.

Glossary

heat

the spontaneous transfer of energy due to a temperature difference

kilocalorie

$1 \text{ kilocalorie} = 1000 \text{ calories}$

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4.3.3.2.1: Heat Transfer

Learning Objectives

- Discuss the three different methods of heat transfer.

Equally as interesting as the effects of heat transfer on a system are the methods by which this occurs. Whenever there is a temperature difference, heat transfer occurs. Heat transfer may occur rapidly, such as through a cooking pan, or slowly, such as through the walls of a picnic ice chest. We can control rates of heat transfer by choosing materials (such as thick wool clothing for the winter), controlling air movement (such as the use of weather stripping around doors), or by choice of color (such as a white roof to reflect summer sunlight). So many processes involve heat transfer, so that it is hard to imagine a situation where no heat transfer occurs. Yet every process involving heat transfer takes place by only three methods:

1. **Conduction** is heat transfer through stationary matter by physical contact. (The matter is stationary on a macroscopic scale—we know there is thermal motion of the atoms and molecules at any temperature above absolute zero.) Heat transferred between the electric burner of a stove and the bottom of a pan is transferred by conduction.
2. **Convection** is the heat transfer by the macroscopic movement of a fluid. This type of transfer takes place in a forced-air furnace and in weather systems, for example.
3. Heat transfer by **radiation** occurs when microwaves, infrared radiation, visible light, or another form of electromagnetic radiation is emitted or absorbed. An obvious example is the warming of the Earth by the Sun. A less obvious example is thermal radiation from the human body.

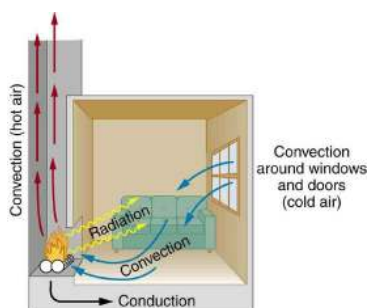


Figure 4.3.3.2.1.1: In a fireplace, heat transfer occurs by all three methods: conduction, convection, and radiation. Radiation is responsible for most of the heat transferred into the room. Heat transfer also occurs through conduction into the room, but at a much slower rate. Heat transfer by convection also occurs through cold air entering the room around windows and hot air leaving the room by rising up the chimney.

Figure 4.3.3.2.1.1 illustrates these three methods of heat transfer occurring in one system. While we skip detailed discussion of methods of heat transfer, the one thing all three methods—conduction, convection, and radiation—share is they are driven by temperature difference.

Exercise 4.3.3.2.1.1

Name an example from daily life (different from the text) for each mechanism of heat transfer.

Answer

Conduction: Heat transfers into your hands as you hold a hot cup of coffee.

Convection: Heat transfers as the barista “steams” cold milk to make hot *cocoa*.

Radiation: Reheating a cold cup of coffee in a microwave oven.

Summary

- Heat is transferred by three different methods: conduction, convection, and radiation.

conduction

heat transfer through stationary matter by physical contact

convection

heat transfer by the macroscopic movement of fluid

radiation

heat transfer which occurs when microwaves, infrared radiation, visible light, or other electromagnetic radiation is emitted or absorbed

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CHAPTER OVERVIEW

5: Electricity and Magnetism

5.1: Electricity

5.1.1: Conservation of Charge

5.1.1.1: Current

5.1.1.1.1: Ohm's Law

5.1.1.1.1.1: Circuits and Resistors

5.2: Magnetism

5.2.1: Magnets

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5.1: Electricity



Figure 5.1.1: 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. (credit: Ken Bosma/Wikimedia Commons)

The image of American politician and scientist Benjamin Franklin (1706–1790) flying a kite in a thunderstorm is familiar to every schoolchild. (See [Figure.](#)) 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.

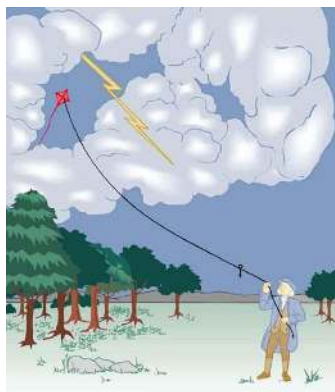


Figure 5.1.2: When Benjamin Franklin demonstrated that lightning was related to static electricity,

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.

Our understanding of electricity and the range of phenomena that are electrical in nature has vastly expanded since those early days. For example, 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. Furthermore, 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.)

In this chapter, we will develop the concepts of electric fields and electric potential (also known as *voltage*), which are used to describe electric force and electrical energy. Consider, for example, great amounts of electrical energy stored in batteries or transmitted cross-country through power lines, or consider electrical signals sent in our nervous systems at molecular levels, with *ions* crossing cell membranes and transferring information, or the role of electricity in our household appliances and devices,

including your laptop and smartphones. We will look at the laws that help us describe electric circuits and begin to explore some of the many applications of electricity.

Glossary

static electricity

a buildup of electric charge on the surface of an object

electromagnetic force

one of the four fundamental forces of nature; the electromagnetic force consists of static electricity, moving electricity and magnetism

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5.1.1: Conservation of Charge

Learning Objectives

- Define electric charge, and describe how the two types of charge interact.
- Describe three common situations that generate static electricity.
- State the law of conservation of charge.



Figure 5.1.1.1: 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. (credit: Sebakoamber, Wikimedia Commons)

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 5.1.1.1](#)). 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 aluminum foil on the bottoms to avoid creating sparks which may ignite 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 5.1.1.2](#) shows how these simple materials can be used to explore the nature of the force between charges.



Figure 5.1.1.2: 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 5.1.1.3 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.

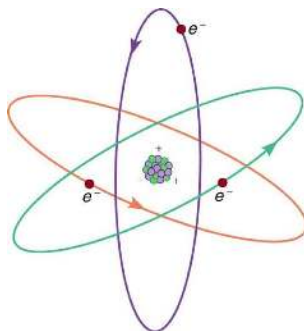


Figure 5.1.1.3: This simplified (and not to scale) view of an atom is called the planetary model of the atom. Negative electrons orbit a much heavier positive nucleus, as the planets orbit the much heavier sun. There the similarity ends, because forces in the atom are electromagnetic, whereas those in 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

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

Similarly, 6.25×10^{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 $|q_e|$.

THINGS GREAT AND SMALL: THE SUBMICROSCOPIC ORIGIN OF CHARGE

With the exception of exotic, short-lived particles, all charge in nature is carried by electrons and protons. Electrons carry the charge we have named negative. Protons carry an equal-magnitude charge that we call positive. (See Figure 5.1.1.4.) Electron and proton charges are considered fundamental building blocks, since all other charges are integral multiples of those carried by electrons and protons. Electrons and protons are also two of the three fundamental building blocks of ordinary matter. The neutron is the third and has zero total charge.

Figure 5.1.1.4 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.

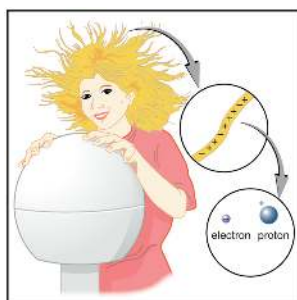


Figure 5.1.1.4: When this person touches a Van de Graaff generator, she receives an excess of positive charge, causing her hair to stand on end. The charges in one hair are shown. An artist's conception of an electron and a proton illustrate the particles carrying the negative and positive charges. We 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.

Separation of Charge 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 5.1.1.6.) 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.

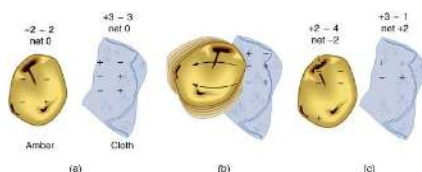


Figure 5.1.1.6: When materials are rubbed together, charges can be separated, particularly if one material has a greater affinity for electrons than another. (a) Both the amber and cloth are originally neutral, with equal positive and negative charges. Only a tiny fraction of the charges are involved, 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**.

LAW OF CONSERVATION OF CHARGE

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.

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.

Section Summary

- There are only two types of charge, which we call positive and negative.
- Like charges repel, unlike charges attract, and the force between charges decreases with the square of the distance.
- The vast majority of positive charge in nature is carried by protons, while the vast majority of negative charge is carried by electrons.
- The electric charge of one electron is equal in magnitude and opposite in sign to the charge of one proton.
- An ion is an atom or molecule that has nonzero total charge due to having unequal numbers of electrons and protons.
- The SI unit for charge is the coulomb (C), with protons and electrons having charges of opposite sign but equal magnitude; the magnitude of this basic charge $|q_e|$ is

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

- Whenever charge is created or destroyed, equal amounts of positive and negative are involved.
- Most often, existing charges are separated from neutral objects to obtain some net charge.
- Both positive and negative charges exist in neutral objects and can be separated by rubbing one object with another. For macroscopic objects, negatively charged means an excess of electrons and positively charged means a depletion of electrons.
- The law of conservation of charge ensures that whenever a charge is created, an equal charge of the opposite sign is created at the same time.

Glossary

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

law of conservation of charge

states that whenever a charge is created, an equal amount of charge with the opposite sign is created simultaneously

electron

a particle orbiting the nucleus of an atom and carrying the smallest unit of negative charge

proton

a particle in the nucleus of an atom and carrying a positive charge equal in magnitude and opposite in sign to the amount of negative charge carried by an electron

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5.1.1.1: Current

Learning Objectives

- Define electric current and ampere.
- Describe the direction of charge flow in conventional current.

Electric Current

Electric current is defined to be the rate at which charge flows. A large current, such as that used to start a truck engine, moves a large amount of charge in a small time, whereas a small current, such as that used to operate a hand-held calculator, moves a small amount of charge over a long period of time. In equation form, **electric current** I is defined to be

$$I = \frac{\Delta Q}{\Delta t},$$

where ΔQ is the amount of charge passing through a given area in time Δt . (As in previous chapters, initial time is often taken to be zero, in which case $\Delta t = t$.) (See Figure 5.1.1.1.1.) The SI unit for current is the **ampere** (A), named for the French physicist André-Marie Ampère (1775–1836). Since $I = \Delta Q/\Delta t$, we see that an ampere is one coulomb per second:

$$1 \text{ A} = 1 \text{ C/s}$$

Not only are fuses and circuit breakers rated in amperes (or amps), so are many electrical appliances.

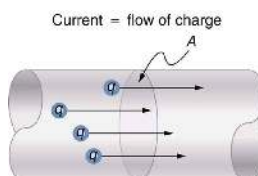


Figure 5.1.1.1.1: The rate of flow of charge is current. An ampere is the flow of one coulomb through an area in one second.

✓ Example 5.1.1.1.1: Calculating Currents: Current in a Truck Battery and a Handheld Calculator

- (a) What is the current involved when a truck battery sets in motion 720 C of charge in 4.00 s while starting an engine? (b) How long does it take 1.00 C of charge to flow through a handheld calculator if a 0.300-mA current is flowing?

Strategy

We can use the definition of current in the equation $I = \Delta Q/\Delta t$ to find the current in part (a), since charge and time are given. In part (b), we rearrange the definition of current and use the given values of charge and current to find the time required.

Solution for (a)

Entering the given values for charge and time into the definition of current gives

$$I = \frac{\Delta Q}{\Delta t} = \frac{720 \text{ C}}{4.00 \text{ s}} = 180 \text{ C/s} \\ = 180 \text{ A.}$$

Discussion for (a)

This large value for current illustrates the fact that a large charge is moved in a small amount of time. The currents in these “starter motors” are fairly large because large frictional forces need to be overcome when setting something in motion.

Solution for (b)

Solving the relationship $I = \Delta Q/\Delta t$ for time Δt , and entering the known values for charge and current gives

$$\Delta t = \frac{\Delta Q}{I} = \frac{1.00 \text{ C}}{0.300 \times 10^{-3} \text{ C/s}} \\ = 3.33 \times 10^3 \text{ s.}$$

Discussion for (b)

This time is slightly less than an hour. The small current used by the hand-held calculator takes a much longer time to move a smaller charge than the large current of the truck starter. So why can we operate our calculators only seconds after turning them on? It's because calculators require very little energy. Such small current and energy demands allow handheld calculators to operate from solar cells or to get many hours of use out of small batteries. Remember, calculators do not have moving parts in the same way that a truck engine has with cylinders and pistons, so the technology requires smaller currents.

Figure 5.1.1.1.2 shows a simple circuit and the standard schematic representation of a battery, conducting path, and load (a resistor). Schematics are very useful in visualizing the main features of a circuit. A single schematic can represent a wide variety of situations. The schematic in Figure 5.1.1.1.2(b), for example, can represent anything from a truck battery connected to a headlight lighting the street in front of the truck to a small battery connected to a penlight lighting a keyhole in a door. Such schematics are useful because the analysis is the same for a wide variety of situations. We need to understand a few schematics to apply the concepts and analysis to many more situations.

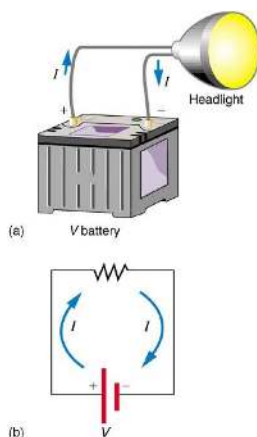


Figure 5.1.1.1.2: (a) A simple electric circuit. A closed path for current to flow through is supplied by conducting wires connecting a load to the terminals of a battery. (b) In this schematic, the battery is represented by the two parallel red lines, conducting wires are shown as straight lines, and the zigzag represents the load. The schematic represents a wide variety of similar circuits.

Note that the direction of current flow in Figure 5.1.1.1.2 is from positive to negative. *The direction of conventional current is the direction that positive charge would flow.* Depending on the situation, positive charges, negative charges, or both may move. In metal wires, for example, current is carried by electrons—that is, negative charges move. In ionic solutions, such as salt water, both positive and negative charges move. This is also true in nerve cells. A Van de Graaff generator used for nuclear research can produce a current of pure positive charges, such as protons. Figure 5.1.1.1.3 illustrates the movement of charged particles that compose a current. The fact that conventional current is taken to be in the direction that positive charge would flow can be traced back to American politician and scientist Benjamin Franklin in the 1700s. He named the type of charge associated with electrons negative, long before they were known to carry current in so many situations. Franklin, in fact, was totally unaware of the small-scale structure of electricity.

It is important to realize that there is an electric field in conductors responsible for producing the current, as illustrated in Figure 5.1.1.1.3. Unlike static electricity, where a conductor in equilibrium cannot have an electric field in it, conductors carrying a current have an electric field and are not in static equilibrium. An electric field is needed to supply energy to move the charges.

✚ MAKING CONNECTIONS: TAKE-HOME INVESTIGATION—ELECTRIC CURRENT ILLUSTRATION

Find a straw and little peas that can move freely in the straw. Place the straw flat on a table and fill the straw with peas. When you pop one pea in at one end, a different pea should pop out the other end. This demonstration is an analogy for an electric current. Identify what compares to the electrons and what compares to the supply of energy. What other analogies can you find for an electric current?

Note that the flow of peas is based on the peas physically bumping into each other; electrons flow due to mutually repulsive electrostatic forces.

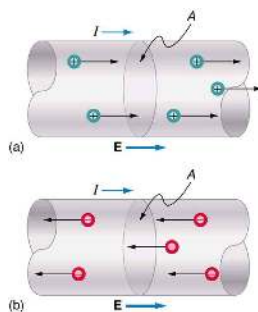


Figure 5.1.1.1.3: Current I is the rate at which charge moves through an area A , such as the cross-section of a wire. Conventional current is defined to move in the direction of the electric field. (a) Positive charges move in the direction of the electric field and the same direction as conventional current. (b) Negative charges move in the direction opposite to the electric field. Conventional current is in the direction opposite to the movement of negative charge. The flow of electrons is sometimes referred to as electronic flow.

✓ Example 5.1.1.1.2: Calculating the Number of Electrons that Move through a Calculator

If the 0.300-mA current through the calculator mentioned in the [Example 5.1.1.1.1](#) is carried by electrons, how many electrons per second pass through it?

Strategy

The current calculated in the previous example was defined for the flow of positive charge. For electrons, the magnitude is the same, but the sign is opposite, $I_{\text{electrons}} = -0.300 \times 10^{-3} \text{C/s}$. Since each electron (e^-) has a charge of $-1.60 \times 10^{-19} \text{C}$, we can convert the current in coulombs per second to electrons per second.

Solution

Starting with the definition of current, we have

$$I_{\text{electrons}} = \frac{\Delta Q_{\text{electrons}}}{\Delta t} = \frac{-0.300 \times 10^{-3} \text{C}}{\text{s}}$$

We divide this by the charge per electron, so that

$$\begin{aligned} \frac{e^-}{\text{s}} &= \frac{-0.300 \times 10^{-3} \text{C}}{\text{s}} \times \frac{1e^-}{-1.60 \times 10^{-19} \text{C}} \\ &= 1.88 \times 10^{15} \frac{e^-}{\text{s}} \end{aligned}$$

Discussion

There are so many charged particles moving, even in small currents, that individual charges are not noticed, just as individual water molecules are not noticed in water flow. Even more amazing is that they do not always keep moving forward like soldiers in a parade. Rather they are like a crowd of people with movement in different directions but a general trend to move forward. There are lots of collisions with atoms in the metal wire and, of course, with other electrons.

Section Summary

- Electric current I is the rate at which charge flows, given by

$$I = \frac{\Delta Q}{\Delta t},$$

where ΔQ is the amount of charge passing through an area in time Δt .

- The direction of conventional current is taken as the direction in which positive charge moves.
- The SI unit for current is the ampere (A), where $1 \text{ A} = 1 \text{ C/s}$.
- Current is the flow of free charges, such as electrons and ions.

electric current

the rate at which charge flows, $I = \Delta Q / \Delta t$

ampere

(amp) the SI unit for current; $1 \text{ A} = 1 \text{ C/s}$

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5.1.1.1.1: Ohm's Law

Learning Objectives

- Explain the origin of Ohm's law.
- Calculate voltages, currents, or resistances with Ohm's law.
- Explain what an ohmic material is.
- Describe a simple circuit.

What drives 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. This is an empirical law like that for friction—an experimentally observed phenomenon. Such a linear relationship doesn't always occur.

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

$$I \propto \frac{1}{R}.$$

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

$$I = \frac{V}{R}.$$

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 Ω (upper case Greek omega). Rearranging $I = V/R$ gives $R = V/I$, and so the units of resistance are $1 \text{ ohm} = 1 \text{ volt per ampere}$:

$$1 \Omega = 1 \frac{V}{A}.$$

Figure 5.1.1.1.1 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 .

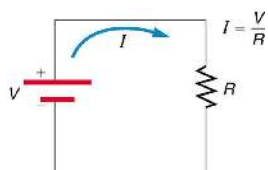


Figure 5.1.1.1.1.1: A simple electric circuit in which a closed path for current to flow is supplied by conductors (usually metal wires) connecting a load to the terminals of a battery, represented by the red parallel lines. The zigzag symbol represents the single resistor and includes any resistance in the connections to the voltage source.

Example 5.1.1.1.1.1: Calculating Resistance: An Automobile Headlight

What is the resistance of an automobile headlight through which 2.50 A flows when 12.0 V is applied to it?

Strategy

We can rearrange Ohm's law as stated by $I = V/R$ and use it to find the resistance.

Solution

Rearranging $I = V/R$ and substituting known values gives

$$R = \frac{V}{I} = \frac{12.0 \text{ V}}{2.50 \text{ A}} = 4.80 \Omega.$$

Discussion

This is a relatively small resistance, but it is larger than the cold resistance of the headlight. Resistance usually increases with temperature, and so the bulb has a lower resistance when it is first switched on and will draw considerably more current during its brief warm-up period.

Resistances range over many orders of magnitude. Some ceramic insulators, such as those used to support power lines, have resistances of $10^{12} \Omega$ or more. A dry person may have a hand-to-foot resistance of $10^5 \Omega$, whereas the resistance of the human heart is about $10^3 \Omega$. A meter-long piece of large-diameter copper wire may have a resistance of $10^{-5} \Omega$, and superconductors have no resistance at all (they are non-ohmic). Resistance is related to the shape of an object and the material of which it is composed.

Additional insight is gained by solving $I = V/R$ for V yielding

$$V = IR.$$

This expression for V can be interpreted as the *voltage drop across a resistor produced by the flow of current I* . The phrase *IR drop* is often used for this voltage. For instance, the headlight in [Example 5.1.1.1.1.1](#) has an *IR* drop of 12.0 V. If voltage is measured at various points in a circuit, it will be seen to increase at the voltage source and decrease at the resistor. Voltage is similar to fluid pressure. The voltage source is like a pump, creating a pressure difference, causing current—the flow of charge. The resistor is like a pipe that reduces pressure and limits flow because of its resistance. Conservation of energy has important consequences here. The voltage source supplies energy (causing an electric field and a current), and the resistor converts it to another form (such as thermal energy). In a simple circuit (one with a single simple resistor), the voltage supplied by the source equals the voltage drop across the resistor, since $\mathbf{PE} = q\Delta V$, and the same q flows through each. Thus the energy supplied by the voltage source and the energy converted by the resistor are equal. (See [Figure 5.1.1.1.1.2](#).)

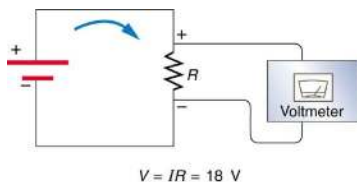


Figure 5.1.1.1.1.2: The voltage drop across a resistor in a simple circuit equals the voltage output of the battery.

MAKING CONNECTIONS: CONSERVATION OF ENERGY

In a simple electrical circuit, the sole resistor converts energy supplied by the source into another form. Conservation of energy is evidenced here by the fact that all of the energy supplied by the source is converted to another form by the resistor alone. We will find that conservation of energy has other important applications in circuits and is a powerful tool in circuit analysis.

Section Summary

- A simple circuit is one in which there is a single voltage source and a single resistance.
- One statement of Ohm's law gives the relationship between current I , voltage V , and resistance R in a simple circuit to be $I = \frac{V}{R}$.
- Resistance has units of ohms (Ω), related to volts and amperes by $1 \Omega = 1 \text{ V/A}$.
- There is a voltage or IR drop across a resistor, caused by the current flowing through it, given by $V = IR$.

Glossary

Ohm's law

an empirical relation stating that the current I is proportional to the potential difference V . It is often written as $I = V/R$, where R is the resistance

resistance

the electric property that impedes current; for ohmic materials, it is the ratio of voltage to current, $R = V/I$

ohm

the unit of resistance, given by $1 \Omega = 1 \text{ V/A}$

ohmic

a type of a material for which Ohm's law is valid

simple circuit

a circuit with a single voltage source and a single resistor

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5.1.1.1.1.1: Circuits and Resistors

Learning Objectives

- Draw a circuit with resistors in parallel and in series.
- Calculate the voltage drop of a current across a resistor using Ohm's law.
- Contrast the way total resistance is calculated for resistors in series and in parallel.
- Explain why total resistance of a parallel circuit is less than the smallest resistance of any of the resistors in that circuit.
- Calculate total resistance of a circuit that contains a mixture of resistors connected in series and in parallel.

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 5.1.1.1.1.1](#). The total resistance of a combination of resistors depends on both their individual values and how they are connected.

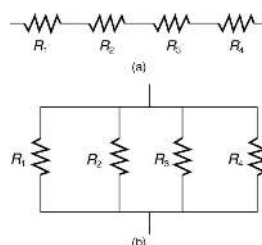


Figure 5.1.1.1.1.1: (a) A series connection of resistors. (b) A parallel connection of resistors.

Resistors in Series

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 5.1.1.1.1.1](#) (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 5.1.1.1.1.2](#) 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.)

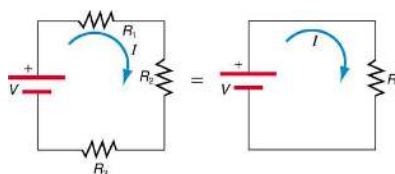


Figure 5.1.1.1.1.2: Three resistors connected in series to a battery (left) and the equivalent single or series resistance (right).

To verify that resistances in series do indeed add, let us consider the loss of electrical power, called a **voltage drop**, in each resistor in [Figure 5.1.1.1.1.2](#).

According to **Ohm's law**, the voltage drop, V , across a resistor when a current flows through it is calculated using the equation $V = IR$, where I equals the current in amps (A) and R is the resistance in ohms (Ω). Another way to think of this is that V is the voltage necessary to make a current I flow through a resistance R .

So the voltage drop across R_1 is $V_1 = IR_1$, that across R_2 is $V_2 = IR_2$, and that across R_3 is $V_3 = IR_3$. The sum of these voltages equals the voltage output of the source; that is,

$$V = V_1 + V_2 + V_3.$$

This equation is based on the conservation of energy and conservation of charge. Electrical potential energy can be described by the equation $PE = qV$, where q is the electric charge and V is the voltage. Thus the energy supplied by the source is qV , while that dissipated by the resistors is

$$qV_1 + qV_2 + qV_3.$$

📌 CONNECTIONS: CONSERVATION LAWS

The derivations of the expressions for series and parallel resistance are based on the laws of conservation of energy and conservation of charge, which state that total charge and total energy are constant in any process. These two laws are directly involved in all electrical phenomena and will be invoked repeatedly to explain both specific effects and the general behavior of electricity.

These energies must be equal, because there is no other source and no other destination for energy in the circuit. Thus, $qV = qV_1 + qV_2 + qV_3$. The charge q cancels, yielding $V = V_1 + V_2 + V_3$, as stated. (Note that the same amount of charge passes through the battery and each resistor in a given amount of time, since there is no capacitance to store charge, there is no place for charge to leak, and charge is conserved.)

Now substituting the values for the individual voltages gives

$$V = IR_1 + IR_2 + IR_3 = I(R_1 + R_2 + R_3).$$

Note that for the equivalent single series resistance R_s , we have

$$V = IR_s.$$

This implies that the total or equivalent series resistance R_s of three resistors is $R_s = R_1 + R_2 + R_3$.

This logic is valid in general for any number of resistors in series; thus, the total resistance R_s of a series connection is

$$R_s = R_1 + R_2 + R_3 + \dots$$

as proposed. Since all of the current must pass through each resistor, it experiences the resistance of each, and resistances in series simply add up.

✓ Example 5.1.1.1.1.1: Calculating Resistance, Current, Voltage Drop, and Power Dissipation: Analysis of a Series Circuit

Suppose the voltage output of the battery in [Figure 5.1.1.1.1.2](#) is 12.0 V , and the resistances are $R_1 = 1.00 \text{ } \Omega$, $R_2 = 6.00 \text{ } \Omega$, and $R_3 = 13.0 \text{ } \Omega$. (a) What is the total resistance? (b) Find the current. (c) Calculate the voltage drop in each resistor, and show these add to equal the voltage output of the source.

Strategy and Solution for (a)

The total resistance is simply the sum of the individual resistances, as given by this equation:

$$\begin{aligned} R_s &= R_1 + R_2 + R_3 \\ &= 1.00 \text{ } \Omega + 6.00 \text{ } \Omega + 13.0 \text{ } \Omega \\ &= 20.0 \text{ } \Omega. \end{aligned}$$

Strategy and Solution for (b)

The current is found using Ohm's law, $V = IR$. Entering the value of the applied voltage and the total resistance yields the current for the circuit:

$$I = \frac{V}{R_s} = \frac{12.0 \text{ V}}{20.0 \text{ } \Omega} = 0.600 \text{ A}.$$

Strategy and Solution for (c)

The voltage—or IR drop—in a resistor is given by Ohm's law. Entering the current and the value of the first resistance yields

$$V_1 = IR_1 = (0.600 \text{ A})(1.0 \text{ } \Omega) = 0.600 \text{ V}.$$

Similarly,

$$V_2 = IR_2 = (0.600 \text{ A})(6.0 \Omega) = 3.60 \text{ V}$$

and

$$V_3 = IR_3 = (0.600 \text{ A})(13.0 \Omega) = 7.80 \text{ V}.$$

Discussion for (c)

The three IR drops add to 12.0 V , as predicted:

$$V_1 + V_2 + V_3 = (0.600 + 3.60 + 7.80)\text{V} = 12.0 \text{ V}.$$

MAJOR FEATURES OF RESISTORS IN SERIES

1. Series resistances add: $R_s = R_1 + R_2 + R_3 + \dots$
2. The same current flows through each resistor in series.
3. Individual resistors in series do not get the total source voltage, but divide it.

Resistors in Parallel

Figure 5.1.1.1.1.3 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 5.1.1.1.1.3(b).)

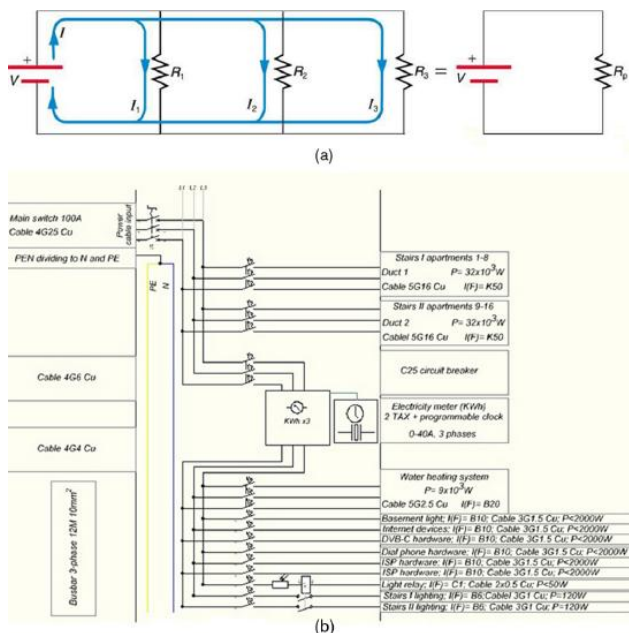


Figure 5.1.1.1.1.3: (a) Three resistors connected in parallel to a battery and the equivalent single or parallel resistance. (b) Electrical power setup in a house. (credit: Dmitry G, Wikimedia Commons)

To find an expression for the equivalent parallel resistance R_p , let us consider the currents that flow and how they are related to resistance. Since each resistor in the circuit has the full voltage, the currents flowing through the individual resistors are $I_1 = \frac{V}{R_1}$, $I_2 = \frac{V}{R_2}$, and $I_3 = \frac{V}{R_3}$. Conservation of charge implies that the total current I produced by the source is the sum of these currents:

$$I = I_1 + I_2 + I_3.$$

Substituting the expressions for the individual currents gives

$$I = \frac{V}{R_1} + \frac{V}{R_2} + \frac{V}{R_3} = V \left(\frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} \right).$$

Note that Ohm's law for the equivalent single resistance gives

$$I = \frac{V}{R_p} = V \left(\frac{1}{R_p} \right).$$

The terms inside the parentheses in the last two equations must be equal. Generalizing to any number of resistors, the total resistance R_p of a parallel connection is related to the individual resistances by

$$\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} + \dots$$

This relationship results in a total resistance R_p that is less than the smallest of the individual resistances. (This is seen in the next example.) When resistors are connected in parallel, more current flows from the source than would flow for any of them individually, and so the total resistance is lower.

✓ Example 5.1.1.1.1.2: Calculating Resistance, Current, Power Dissipation, and Power Output: Analysis of a Parallel Circuit

Let the voltage output of the battery and resistances in the parallel connection in [Figure 5.1.1.1.1.3](#) be the same as the previously considered series connection: $V = 12.0 \text{ V}$, $R_1 = 1.00 \text{ } \Omega$, $R_2 = 6.00 \text{ } \Omega$, and $R_3 = 13.0 \text{ } \Omega$. (a) What is the total resistance? (b) Find the total current. (c) Calculate the currents in each resistor, and show these add to equal the total current output of the source.

Strategy and Solution for (a)

The total resistance for a parallel combination of resistors is found using the equation below. Entering known values gives

$$\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} = \frac{1}{1.00 \text{ } \Omega} + \frac{1}{6.00 \text{ } \Omega} + \frac{1}{13.0 \text{ } \Omega}.$$

Thus,

$$\frac{1}{R_p} = \frac{1.00}{\Omega} + \frac{0.1667}{\Omega} + \frac{0.07692}{\Omega} = \frac{1.2436}{\Omega}.$$

(Note that in these calculations, each intermediate answer is shown with an extra digit.)

We must invert this to find the total resistance R_p . This yields

$$R_p = \frac{1}{1.2436} \text{ } \Omega = 0.8041 \text{ } \Omega.$$

The total resistance with the correct number of significant digits is $R_p = 0.804 \text{ } \Omega$

Discussion for (a)

R_p is, as predicted, less than the smallest individual resistance.

Strategy and Solution for (b)

The total current can be found from Ohm's law, substituting R_p for the total resistance. This gives

$$I = \frac{V}{R_p} = \frac{12.0 \text{ V}}{0.8041 \text{ } \Omega} = 14.92 \text{ A}.$$

Discussion for (b)

Current I for each device is much larger than for the same devices connected in series (see the previous example). A circuit with parallel connections has a smaller total resistance than the resistors connected in series.

Strategy and Solution for (c)

The individual currents are easily calculated from Ohm's law, since each resistor gets the full voltage. Thus,

$$I_1 = \frac{V}{R_1} = \frac{12.0 \text{ V}}{1.00 \, \Omega} = 12.0 \text{ A.}$$

Similarly,

$$I_2 = \frac{V}{R_2} = \frac{12.0 \text{ V}}{6.00 \, \Omega} = 2.00 \text{ A}$$

and

$$I_3 = \frac{V}{R_3} = \frac{12.0 \text{ V}}{13.0 \, \Omega} = 0.92 \text{ A}$$

Discussion for (c)

The total current is the sum of the individual currents:

$$I_1 + I_2 + I_3 = 14.92 \text{ A.}$$

This is consistent with conservation of charge.

MAJOR FEATURES OF RESISTORS IN PARALLEL

1. Parallel resistance is found from $\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} + \dots$, and it is smaller than any individual resistance in the combination.
2. Each resistor in parallel has the same full voltage of the source applied to it. (Power distribution systems most often use parallel connections to supply the myriad devices served with the same voltage and to allow them to operate independently.)
3. Parallel resistors do not each get the total current; they divide it.

Combinations of Series and Parallel

More complex connections of resistors are sometimes just combinations of series and parallel. These are commonly encountered, especially when wire resistance is considered. In that case, wire resistance is in series with other resistances that are in parallel.

Combinations of series and parallel can be reduced to a single equivalent resistance using the technique illustrated in [Figure 5.1.1.1.1.4](#). Various parts are identified as either series or parallel, reduced to their equivalents, and further reduced until a single resistance is left. The process is more time consuming than difficult.

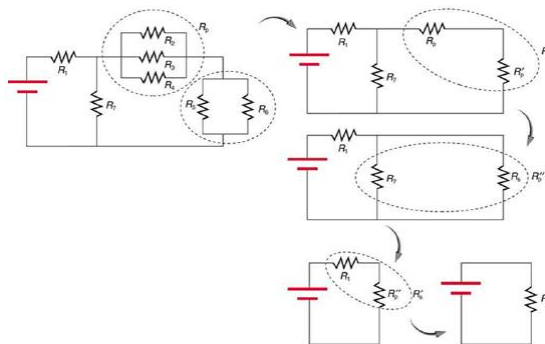


Figure 5.1.1.1.1.4: This combination of seven resistors has both series and parallel parts. Each is identified and reduced to an equivalent resistance, and these are further reduced until a single equivalent resistance is reached.

The simplest combination of series and parallel resistance, shown in [Figure 5.1.1.1.1.5](#), is also the most instructive, since it is found in many applications. For example, R_1 could be the resistance of wires from a car battery to its electrical devices, which are in parallel. R_2 and R_3 could be the starter motor and a passenger compartment light. We have previously assumed that wire resistance is negligible, but, when it is not, it has important effects, as the next example indicates.

Practical Implications

One implication of this last example is that resistance in wires reduces the current and power delivered to a resistor. If wire resistance is relatively large, as in a worn (or a very long) extension cord, then this loss can be significant. If a large current is drawn, the IR drop in the wires can also be significant.

For example, when you are rummaging in the refrigerator and the motor comes on, the refrigerator light dims momentarily. Similarly, you can see the passenger compartment light dim when you start the engine of your car (although this may be due to resistance inside the battery itself).

What is happening in these high-current situations is illustrated in Figure 5.1.1.1.1.6. The device represented by R_3 has a very low resistance, and so when it is switched on, a large current flows. This increased current causes a larger IR drop in the wires represented by R_1 , reducing the voltage across the light bulb (which is R_2), which then dims noticeably.

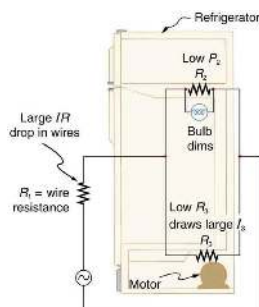


Figure 5.1.1.1.1.6: Why do lights dim when a large appliance is switched on? The answer is that the large current the appliance motor draws causes a significant IR drop in the wires and reduces the voltage across the light.

PROBLEM-SOLVING STRATEGIES FOR SERIES AND PARALLEL RESISTORS

1. Draw a clear circuit diagram, labeling all resistors and voltage sources. This step includes a list of the knowns for the problem, since they are labeled in your circuit diagram.
2. Identify exactly what needs to be determined in the problem (identify the unknowns). A written list is useful.
3. Determine whether resistors are in series, parallel, or a combination of both series and parallel. Examine the circuit diagram to make this assessment. Resistors are in series if the same current must pass sequentially through them.
4. Use the appropriate list of major features for series or parallel connections to solve for the unknowns. There is one list for series and another for parallel. If your problem has a combination of series and parallel, reduce it in steps by considering individual groups of series or parallel connections, as done in this module and the examples. Special note: When finding R_p , the reciprocal must be taken with care.
5. Check to see whether the answers are reasonable and consistent. Units and numerical results must be reasonable. Total series resistance should be greater, whereas total parallel resistance should be smaller, for example. Power should be greater for the same devices in parallel compared with series, and so on.

Section Summary

- The total resistance of an electrical circuit with resistors wired in a series is the sum of the individual resistances: $R_s = R_1 + R_2 + R_3 + \dots$
- Each resistor in a series circuit has the same amount of current flowing through it.
- The voltage drop, or power dissipation, across each individual resistor in a series is different, and their combined total adds up to the power source input.
- The total resistance of an electrical circuit with resistors wired in parallel is less than the lowest resistance of any of the components and can be determined using the formula:

$$\frac{1}{R_p} = \frac{1}{R_1} + \frac{1}{R_2} + \frac{1}{R_3} + \dots$$

- Each resistor in a parallel circuit has the same full voltage of the source applied to it.
- The current flowing through each resistor in a parallel circuit is different, depending on the resistance.

- If a more complex connection of resistors is a combination of series and parallel, it can be reduced to a single equivalent resistance by identifying its various parts as series or parallel, reducing each to its equivalent, and continuing until a single resistance is eventually reached.

Glossary

series

a sequence of resistors or other components wired into a circuit one after the other

resistor

a component that provides resistance to the current flowing through an electrical circuit

resistance

causing a loss of electrical power in a circuit

Ohm's law

the relationship between current, voltage, and resistance within an electrical circuit: $V = IR$

voltage

the electrical potential energy per unit charge; electric pressure created by a power source, such as a battery

voltage drop

the loss of electrical power as a current travels through a resistor, wire or other component

current

the flow of charge through an electric circuit past a given point of measurement

parallel

the wiring of resistors or other components in an electrical circuit such that each component receives an equal voltage from the power source; often pictured in a ladder-shaped diagram, with each component on a rung of the ladder

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5.2: Magnetism

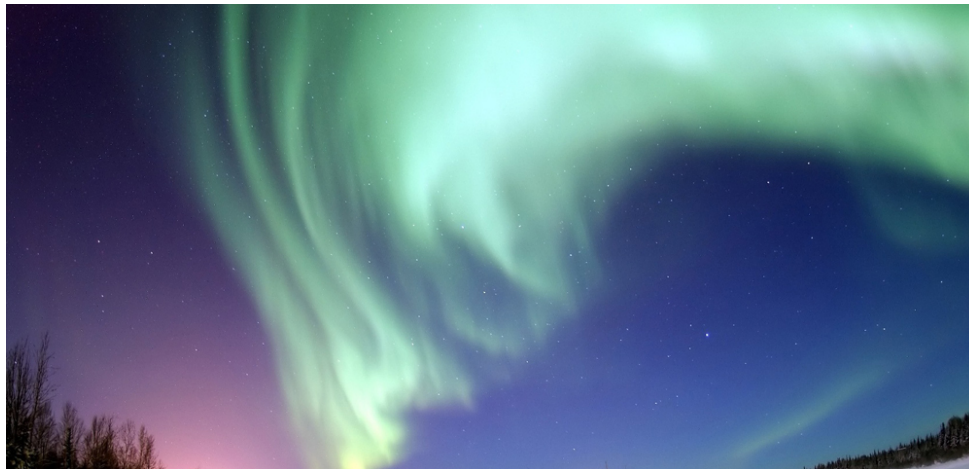


Figure 5.2.1: The magnificent

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. (credit: Senior Airman Joshua Strang, via Flickr)

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.

The discovery that weak changes in a magnetic field in a thin film of iron and chromium could bring about much larger changes in electrical resistance was one of the first large successes of nanotechnology. The 2007 Nobel Prize in Physics went to Albert Fert from France and Peter Grunberg from Germany for this discovery of *giant magnetoresistance* and its applications to computer memory.

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.



Figure 5.2.2: Engineering of technology like iPods would not be possible without a deep understanding magnetism. (credit: Jesse! S?, Flickr)

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5.2.1: Magnets

Learning Objectives

- Describe the difference between the north and south poles of a magnet.
- Describe how magnetic poles interact with each other.



Figure 5.2.1.1: 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).

UNIVERSAL CHARACTERISTICS OF MAGNETS AND MAGNETIC POLES

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.

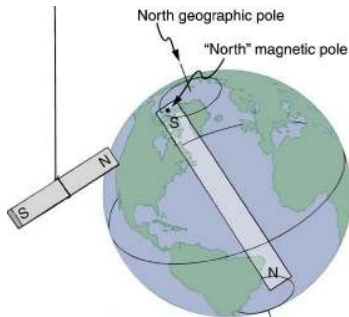


Figure 5.2.1.2: One end of a bar magnet is suspended from a thread that points toward north. The magnet's two poles are labeled N and S for north-seeking and south-seeking poles, respectively.

MISCONCEPTION ALERT: EARTH'S GEOGRAPHIC NORTH POLE HIDES AN S

The 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 the 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.

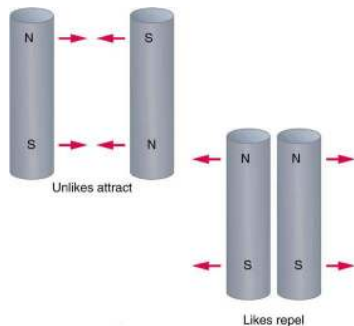
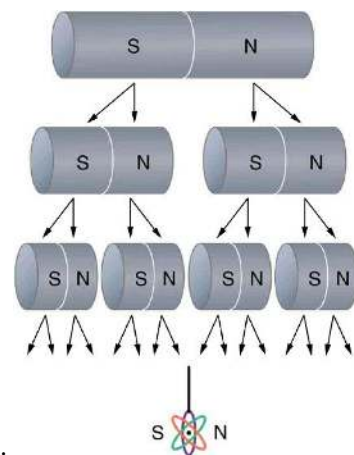


Figure 5.2.1.3: Unlike poles attract, whereas like poles repel.

Figure 5.2.1.4: 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.

MAKING CONNECTIONS: 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?

Section Summary

- Magnetism is a subject that includes the properties of magnets, the effect of the magnetic force on moving charges and currents, and the creation of magnetic fields by currents.
- There are two types of magnetic poles, called the north magnetic pole and south magnetic pole.
- North magnetic poles are those that are attracted toward the Earth's geographic north pole.
- Like poles repel and unlike poles attract.
- Magnetic poles always occur in pairs of north and south—it is not possible to isolate north and south poles.

Glossary

north magnetic pole

the end or the side of a magnet that is attracted toward Earth's geographic north pole

south magnetic pole

the end or the side of a magnet that is attracted toward Earth's geographic south pole

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CHAPTER OVERVIEW

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6.1: Waves



Figure 6.1.1: There are at least four

types of waves in this picture—only the water waves are evident. There are also sound waves, light waves, and waves on the guitar strings. (credit: John Norton)

What do an ocean buoy, a child in a swing, the cone inside a speaker, a guitar, atoms in a crystal, the motion of chest cavities, and the beating of hearts all have in common? They all **oscillate**—that is, they move back and forth between two points. Many systems oscillate, and they have certain characteristics in common. All oscillations involve force and energy. You push a child in a swing to get the motion started. The energy of atoms vibrating in a crystal can be increased with heat. You put energy into a guitar string when you pluck it.

Some oscillations create **waves**. A guitar creates sound waves. You can make water waves in a swimming pool by slapping the water with your hand. You can no doubt think of other types of waves. Some, such as water waves, are visible. Some, such as sound waves, are not. But *every wave is a disturbance that moves from its source and carries energy*. Other examples of waves include earthquakes and visible light. Even subatomic particles, such as electrons, can behave like waves.

By studying oscillatory motion and waves, we shall find that a small number of underlying principles describe all of them and that wave phenomena are more common than you have ever imagined.

Glossary

oscillate

moving back and forth regularly between two points

wave

a disturbance that moves from its source and carries energy

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6.1.1: Period and Frequency

Learning Objectives

- Understand the relationship between the frequency and the period of oscillations.
- Determine the frequency of oscillations.



Figure 6.1.1.1: The strings on this guitar vibrate at regular time intervals. (credit: JAR)

When you pluck a guitar string, the resulting sound has a steady tone and lasts a long time. Each successive vibration of the string takes the same time as the previous one. We define **periodic motion** to be a motion that repeats itself at regular time intervals, such as exhibited by the guitar string or by an object on a spring moving up and down. The time to complete one oscillation remains constant and is called the **period T** . Its units are usually seconds, but may be any convenient unit of time. The word period refers to the time for some event whether repetitive or not; but we shall be primarily interested in periodic motion, which is by definition repetitive. A concept closely related to period is the frequency of an event. For example, if you get a paycheck twice a month, the frequency of payment is two per month and the period between checks is half a month. **Frequency f** is defined to be the number of events per unit time. For periodic motion, frequency is the number of oscillations per unit time. The relationship between frequency and period is

$$f = \frac{1}{T}.$$

The SI unit for frequency is the *cycle per second*, which is defined to be a *hertz* (Hz):

$$1 \text{ Hz} = 1 \frac{\text{cycle}}{\text{sec}} \text{ or } 1 \text{ Hz} = \frac{1}{\text{s}}$$

A cycle is one complete oscillation. Note that a vibration can be a single or multiple event, whereas oscillations are usually repetitive for a significant number of cycles.

Example 6.1.1.1: Determine the Frequency of Two Oscillations: Medical Ultrasound and the Period of Middle C

We can use the formulas presented in this module to determine both the frequency based on known oscillations and the oscillation based on a known frequency. Let's try one example of each. (a) A medical imaging device produces ultrasound by oscillating with a period of $0.400 \mu\text{s}$. What is the frequency of this oscillation? (b) The frequency of middle C on a typical musical instrument is 264 Hz. What is the time for one complete oscillation?

Strategy

Both questions (a) and (b) can be answered using the relationship between period and frequency. In question (a), the period T is given and we are asked to find frequency f . In question (b), the frequency f is given and we are asked to find the period T .

Solution a

1. Substitute $0.400 \mu\text{s}$ for T in $f = \frac{1}{T}$:

$$f = \frac{1}{T} = \frac{1}{0.400 \times 10^{-6} \text{ s}}.$$

Solve to find

$$f = 2.50 \times 10^6 \text{ Hz}.$$

Discussion a

The frequency of sound found in (a) is much higher than the highest frequency that humans can hear and, therefore, is called ultrasound. Appropriate oscillations at this frequency generate ultrasound used for noninvasive medical diagnoses, such as observations of a fetus in the womb.

Solution b

1. Identify the known values:

The time for one complete oscillation is the period T :

$$f = \frac{1}{T}.$$

2. Solve for T :

$$T = \frac{1}{f}.$$

3. Substitute the given value for the frequency into the resulting expression:

$$T = \frac{1}{f} = \frac{1}{264 \text{ Hz}} = \frac{1}{264 \text{ cycles/s}} = 3.79 \times 10^{-3} \text{ s} = 3.79 \text{ ms}.$$

Discussion

The period found in (b) is the time per cycle, but this value is often quoted as simply the time in convenient units (ms or milliseconds in this case).

Exercise 6.1.1.1

Identify an event in your life (such as receiving a paycheck) that occurs regularly. Identify both the period and frequency of this event.

Answer

I visit my parents for dinner every other Sunday. The frequency of my visits is 26 per calendar year. The period is two weeks.

Section Summary

- Periodic motion is a repetitious oscillation.
- The time for one oscillation is the period T .
- The number of oscillations per unit time is the frequency f .
- These quantities are related by

$$f = \frac{1}{T}.$$

Glossary**period**

time it takes to complete one oscillation

periodic motion

motion that repeats itself at regular time intervals

frequency

number of events per unit of time

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6.1.2: Properties of Waves

Learning Objectives

- State the characteristics of a wave.
- Calculate the velocity of wave propagation, given frequency and wavelength



Figure 6.1.2.1: Waves in the ocean behave similarly to all other types of waves. (credit: Steve Jurveston, Flickr)

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 in 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.

A wave is a disturbance that propagates, or moves from the place it was created. 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 6.1.2.2](#). The wave is an up and down disturbance of the water surface. It causes a sea gull 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 $f = 1/T$, as usual. 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.

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.

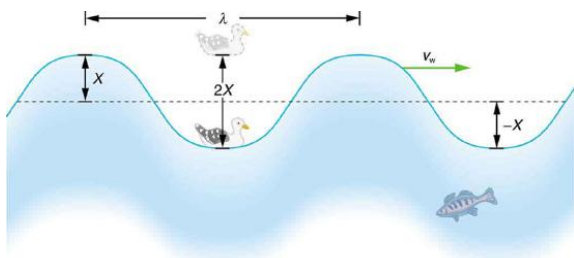


Figure 6.1.2.2: An idealized ocean wave passes under a sea gull that bobs up and down in simple harmonic motion. The wave has a wavelength λ , which is the distance between adjacent identical parts of the wave. The up and down disturbance of the surface propagates parallel to the surface at a speed v_w .

The water wave in the figure also has a length associated with it, called its **wavelength** λ , the distance between adjacent identical parts of a wave. (λ is the distance parallel to the direction of propagation.) The speed of propagation v_w is the distance the wave travels in a given time, which is one wavelength in the time of one period. In equation form, that is

$$v_w = \frac{\lambda}{T}$$

or

$$v_w = f\lambda. \quad (6.1.2.1)$$

This fundamental relationship holds for all types of waves. For water waves, v_w is the speed of a surface wave; for sound, v_w is the speed of sound; and for visible light, v_w is the speed of light, for example.

One note of caution with Equation (6.1.2.1): despite the appearance, wave speed v_w is a *not* a function of frequency f and wavelength λ . When f changes (or λ changes), v_w does not change (you will see this explicitly later with sound and light waves). Wave speed is a *property of the medium* the wave travels in. Unless the medium itself changes, the wave speed remains constant. This is true of sound and light waves, as well as almost all other types of waves you will see in this textbook (water waves, waves on a string, etc.). So, what happens when frequency or wavelength does change, then? The wavelength and frequency change to compensate. For example, if the frequency doubles for a given wave, instead of the wave speed doubling, the wavelength will decrease to half, so that the product $f\lambda$ remains constant.

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?

Calculate the Velocity of Wave Propagation: Gull in the Ocean

Calculate the wave velocity of the ocean wave in Figure 6.1.2.2 if the distance between wave crests is 10.0 m and the time for a sea gull 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 $T = 5.00 \text{ s}$. Therefore, we can use $v_w = \frac{\lambda}{T}$ to find the wave velocity.

Solution

1. Enter the known values into $v_w = \frac{\lambda}{T}$:

$$v_w = \frac{10.0 \text{ m}}{5.00 \text{ s}}.$$

2. Solve for v_w to find $v_w = 2.00 \text{ m/s}$.

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 sea gull moves up and down.

Transverse and Longitudinal Waves

A simple wave consists of a periodic disturbance that propagates from one place to another. The wave in Figure 6.1.2.3 propagates in the horizontal direction while the surface is disturbed in the vertical direction. Such a wave is called a **transverse wave** or shear wave; in such a wave, the disturbance is perpendicular to the direction of propagation. In contrast, in a **longitudinal wave** or compressional wave, the disturbance is parallel to the direction of propagation. Figure 6.1.2.4 shows an example of a longitudinal wave. The size of the disturbance is its amplitude X and is completely independent of the speed of propagation v_w .



Figure 6.1.2.3: In this example of a transverse wave, the wave propagates horizontally, and the

disturbance in the cord is in the vertical direction.

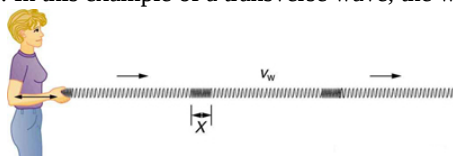


Figure 6.1.2.4: In this example of a

longitudinal wave, the wave propagates horizontally, and the disturbance in the cord is also in the horizontal direction.

Waves may be transverse, longitudinal, or *a combination of the two*. (Water waves are actually a combination of transverse and longitudinal. The simplified water wave illustrated in [Figure 6.1.2.2](#) shows no longitudinal motion of the bird.) The waves on the strings of musical instruments are transverse—so are electromagnetic waves, such as visible light.

Sound waves in air and water are longitudinal. Their disturbances are periodic variations in pressure that are transmitted in fluids. Fluids do not have appreciable shear strength, and thus the sound waves in them must be longitudinal or compressional. Sound in solids can be both longitudinal and transverse.

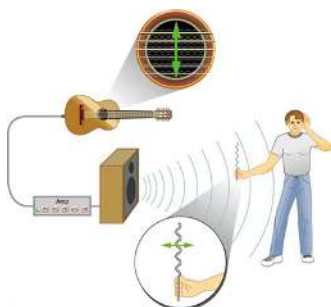


Figure 6.1.2.5: The wave on a guitar string is transverse. The sound wave rattles a sheet of paper in a direction that shows the sound wave is longitudinal.

Earthquake waves under Earth's surface also have both longitudinal and transverse components (called compressional or P-waves and shear or S-waves, respectively). These components have important individual characteristics—they propagate at different speeds, for example. Earthquakes also have surface waves that are similar to surface waves on water.

Exercise 6.1.2.1

Why is it important to differentiate between longitudinal and transverse waves?

Answer

In the different types of waves, energy can propagate in a different direction relative to the motion of the wave. This is important to understand how different types of waves affect the materials around them.

Section Summary

- A wave is a disturbance that moves from the point of creation with a wave velocity v_w .
- A wave has a wavelength λ , which is the distance between adjacent identical parts of the wave.
- Wave velocity and wavelength are related to the wave's frequency and period by $v_w = \frac{\lambda}{T}$ or $v_w = f\lambda$. However, wave velocity is a property of the medium and remains constant as wavelength or frequency changes.
- A transverse wave has a disturbance perpendicular to its direction of propagation, whereas a longitudinal wave has a disturbance parallel to its direction of propagation.

Glossary

longitudinal wave

a wave in which the disturbance is parallel to the direction of propagation

transverse wave

a wave in which the disturbance is perpendicular to the direction of propagation

wave velocity

the speed at which the disturbance moves; also called wave speed, propagation velocity, or propagation speed

wavelength

the distance between adjacent identical parts of a wave

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6.1.3: Interference

Learning Objectives

- Distinguish between constructive interference and destructive interference.
- Explain standing waves in terms of interference, nodes, and antinodes.
- Explain the beat phenomenon in terms of interference.



Figure 6.1.3.1: These waves result from the superposition of several waves from different sources, producing a complex pattern. (credit: waterborough, Wikimedia Commons)

Most waves do not look very simple. They look more like the waves in [Figure 6.1.3.1](#) than like the simple water wave considered in "Waves". (Simple waves may be created by a simple harmonic oscillation, and thus have a sinusoidal shape). Complex waves are more interesting, even beautiful, but they look formidable. Most waves appear complex because they result from several simple waves adding together. Luckily, the rules for adding waves are quite simple.

When two or more waves arrive at the same point, they superimpose themselves on one another. More specifically, the disturbances of waves are superimposed when they come together—a phenomenon called **superposition**. Each disturbance corresponds to a force, and forces add. If the disturbances are along the same line, then the resulting wave is a simple addition of the disturbances of the individual waves—that is, their amplitudes add. [Figure 6.1.3.2](#) and [Figure 6.1.3.3](#) illustrate superposition in two special cases, both of which produce simple results.

[Figure 6.1.3.2](#) shows two identical waves that arrive at the same point exactly in phase. The crests of the two waves are precisely aligned, as are the troughs. This superposition produces pure **constructive interference**. Because the disturbances add, pure constructive interference produces a wave that has twice the amplitude of the individual waves, but has the same wavelength.

[Figure 6.1.3.3](#) shows two identical waves that arrive exactly out of phase—that is, precisely aligned crest to trough—producing pure **destructive interference**. Because the disturbances are in the opposite direction for this superposition, the resulting amplitude is zero for pure destructive interference—the waves completely cancel.

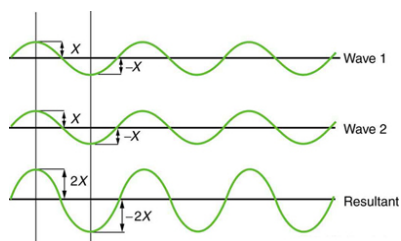
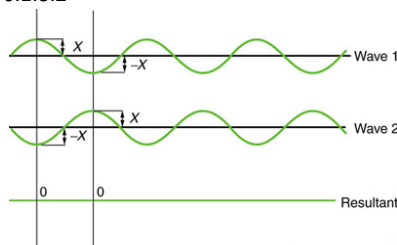


Figure 6.1.3.2: Pure constructive interference of two identical waves produces one with



twice the amplitude, but the same wavelength. of two identical waves produces zero amplitude, or complete cancellation.

Figure 6.1.3.3: Pure destructive interference

While pure constructive and pure destructive interference do occur, they require precisely aligned identical waves. The superposition of most waves produces a combination of constructive and destructive interference and can vary from place to place and time to time. Sound from a stereo, for example, can be loud in one spot and quiet in another. Varying loudness means the sound

waves add partially constructively and partially destructively at different locations. A stereo has at least two speakers creating sound waves, and waves can reflect from walls. All these waves superimpose. An example of sounds that vary over time from constructive to destructive is found in the combined whine of airplane jets heard by a stationary passenger. The combined sound can fluctuate up and down in volume as the sound from the two engines varies in time from constructive to destructive. These examples are of waves that are similar.

An example of the superposition of two dissimilar waves is shown in [Figure 6.1.3.4](#). Here again, the disturbances add and subtract, producing a more complicated looking wave.

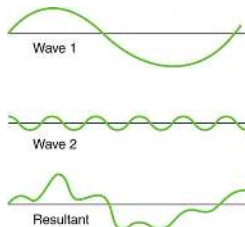


Figure 6.1.3.4: Superposition of non-identical waves exhibits both constructive and destructive interference.

Standing Waves

Sometimes waves do not seem to move; rather, they just vibrate in place. Unmoving waves can be seen on the surface of a glass of milk in a refrigerator, for example. Vibrations from the refrigerator motor create waves on the milk that oscillate up and down but do not seem to move across the surface. These waves are formed by the superposition of two or more moving waves, such as illustrated in [Figure 6.1.3.5](#) for two identical waves moving in opposite directions. The waves move through each other with their disturbances adding as they go by. If the two waves have the same amplitude and wavelength, then they alternate between constructive and destructive interference. The resultant looks like a wave standing in place and, thus, is called a **standing wave**. Waves on the glass of milk are one example of standing waves. There are other standing waves, such as on guitar strings and in organ pipes. With the glass of milk, the two waves that produce standing waves may come from reflections from the side of the glass.

A closer look at earthquakes provides evidence for conditions appropriate for resonance, standing waves, and constructive and destructive interference. A building may be vibrated for several seconds with a driving frequency matching that of the natural frequency of vibration of the building—producing a resonance resulting in one building collapsing while neighboring buildings do not. Often buildings of a certain height are devastated while other taller buildings remain intact. The building height matches the condition for setting up a standing wave for that particular height. As the earthquake waves travel along the surface of Earth and reflect off denser rocks, constructive interference occurs at certain points. Often areas closer to the epicenter are not damaged while areas farther away are damaged.

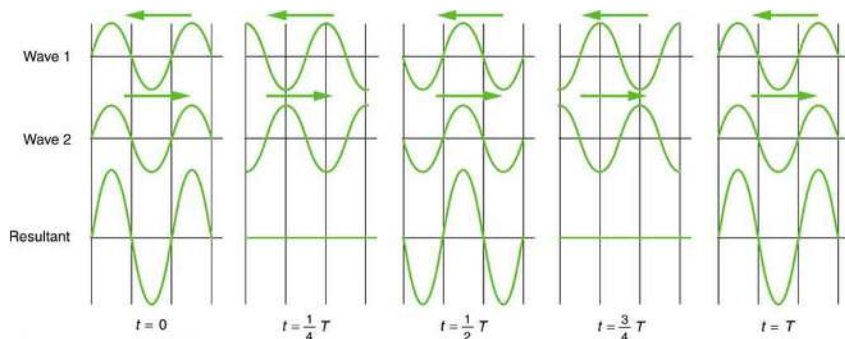


Figure 6.1.3.5: Standing wave created by the superposition of two identical waves moving in opposite directions. The oscillations are at fixed locations in space and result from alternately constructive and destructive interference.

Standing waves are also found on the strings of musical instruments and are due to reflections of waves from the ends of the string. [Figure 6.1.3.6](#) and [Figure 6.1.3.7](#) show three standing waves that can be created on a string that is fixed at both ends. **Nodes** are the points where the string does not move; more generally, nodes are where the wave disturbance is zero in a standing wave. The fixed ends of strings must be nodes, too, because the string cannot move there. The word **antinode** is used to

denote the location of maximum amplitude in standing waves. Standing waves on strings have a frequency that is related to the propagation speed v_w of the disturbance on the string. The wavelength λ is determined by the distance between the points where the string is fixed in place.

The lowest frequency, called the **fundamental frequency**, is thus for the longest wavelength, which is seen to be $\lambda_1 = 2L$. Therefore, the fundamental frequency is $f_1 = v_w/\lambda_1 = v_w/2L$. In this case, the **overtone** or harmonics are multiples of the fundamental frequency. As seen in Figure 6.1.3.7, the first harmonic can easily be calculated since $\lambda_2 = L$. Thus, $f_2 = v_w/\lambda_2 = v_w/L = 2f_1$. Similarly, $f_3 = 3f_1$, and so on. All of these frequencies can be changed by adjusting the tension in the string. The greater the tension, the greater v_w is and the higher the frequencies. This observation is familiar to anyone who has ever observed a string instrument being tuned. We will see in later chapters that standing waves are crucial to many resonance phenomena, such as in sounding boxes on string instruments.

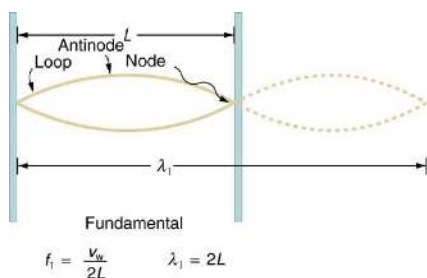


Figure 6.1.3.6: The figure shows a string oscillating at its fundamental frequency.

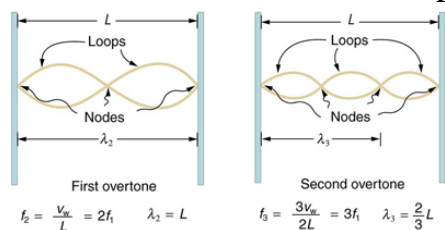


Figure 6.1.3.7: First and second harmonic frequencies are shown.

Beats

Striking two adjacent keys on a piano produces a warbling combination usually considered to be unpleasant. The superposition of two waves of similar but not identical frequencies is the culprit. Another example is often noticeable in jet aircraft, particularly the two-engine variety, while taxiing. The combined sound of the engines goes up and down in loudness. This varying loudness happens because the sound waves have similar but not identical frequencies. The discordant warbling of the piano and the fluctuating loudness of the jet engine noise are both due to alternately constructive and destructive interference as the two waves go in and out of phase. Figure 6.1.3.8 illustrates this graphically.

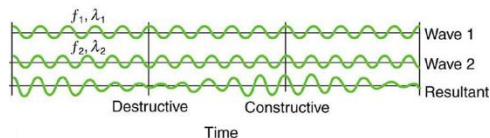


Figure 6.1.3.8: Beats are produced by the superposition of two waves of slightly different frequencies but identical amplitudes. The waves alternate in time between constructive interference and destructive interference, giving the resulting wave a time-varying amplitude.

The wave resulting from the superposition of two similar-frequency waves has a frequency that is the average of the two. This wave fluctuates in amplitude, or **beats**, with a frequency called the **beat frequency**. We can determine the beat frequency by adding two waves together mathematically. The result,

$$f_B = |f_1 - f_2|$$

is the beat frequency. The resultant wave has the average frequency of the two superimposed waves, but it also fluctuates in overall amplitude at the beat frequency f_B . This result is valid for all types of waves. However, if it is a sound wave, providing the two frequencies are similar, then what we hear is an average frequency that gets louder and softer (or warbles) at the beat frequency.

MAKING CAREER CONNECTIONS

Piano tuners use beats routinely in their work. When comparing a note with a tuning fork, they listen for beats and adjust the string until the beats go away (to zero frequency). For example, if the tuning fork has a **256 Hz** frequency and two beats per second are heard, then the other frequency is either **254** or **258 Hz**. Most keys hit multiple strings, and these strings are actually adjusted until they have nearly the same frequency and give a slow beat for richness. Twelve-string guitars and mandolins are also tuned using beats.

While beats may sometimes be annoying in audible sounds, we will find that beats have many applications. Observing beats is a very useful way to compare similar frequencies. There are applications of beats as apparently disparate as in ultrasonic imaging and radar speed traps.

Exercise 6.1.3.1

Imagine you are holding one end of a jump rope, and your friend holds the other. If your friend holds her end still, you can move your end up and down, creating a transverse wave. If your friend then begins to move her end up and down, generating a wave in the opposite direction, what resultant wave forms would you expect to see in the jump rope?

Answer

The rope would alternate between having waves with amplitudes two times the original amplitude and reaching equilibrium with no amplitude at all. The wavelengths will result in both constructive and destructive interference.

Exercise 6.1.3.2

Define nodes and antinodes.

Answer

Nodes are areas of wave interference where there is no motion. Antinodes are areas of wave interference where the motion is at its maximum point.

Exercise 6.1.3.3

You hook up a stereo system. When you test the system, you notice that in one corner of the room, the sounds seem dull. In another area, the sounds seem excessively loud. Describe how the sound moving about the room could result in these effects.

Answer

With multiple speakers putting out sounds into the room, and these sounds bouncing off walls, there is bound to be some wave interference. In the dull areas, the interference is probably mostly destructive. In the louder areas, the interference is probably mostly constructive.

Section Summary

- Superposition is the combination of two waves at the same location.
- Constructive interference occurs when two waves are superimposed in phase (both disturbances in the same direction).
- Destructive interference occurs when two waves are superimposed exactly out of phase (two disturbances in opposite directions).
- A standing wave is one in which two waves superimpose to produce a wave that forms a standing pattern of nodes and antinodes, locations of no displacement and of maximum displacement.
- Standing waves form at specific resonance frequencies, with the lowest frequency being the fundamental frequency, and higher frequencies being called overtones or harmonics.
- Beats occur when waves of similar frequencies f_1 and f_2 are superimposed. The resulting amplitude oscillates with a beat frequency given by

$$f_B = |f_1 - f_2|.$$

Glossary

antinode

the location of maximum amplitude in standing waves

beat frequency

the frequency of the amplitude fluctuations of a wave

constructive interference

when two waves arrive at the same point exactly in phase; that is, the crests of the two waves are precisely aligned, as are the troughs

destructive interference

when two identical waves arrive at the same point exactly out of phase; that is, precisely aligned crest to trough

fundamental frequency

the lowest frequency of a periodic waveform

nodes

the points where the string does not move; more generally, nodes are where the wave disturbance is zero in a standing wave

overtones

multiples of the fundamental frequency of a sound

superposition

the phenomenon that occurs when two or more waves arrive at the same point

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6.1.4: Properties of Sound Waves

Learning Objectives

- Define pitch.
- Describe the relationship between the speed of sound, its frequency, and its wavelength.
- Describe the effects on the speed of sound as it travels through various media.
- Describe the effects of temperature on the speed of sound.



Figure 6.1.4.1: When a firework explodes, the light energy is perceived before the sound energy. Sound travels more slowly than light does. (credit: Dominic Alves, Flickr)

Sound, like all waves, travels at a certain speed and has the properties of frequency and wavelength. You can observe direct evidence of the speed of sound while watching a fireworks display. The flash of an explosion is seen well before its sound is heard, implying both that sound travels at a finite speed and that it is much slower than light. You can also directly sense the frequency of a sound. Perception of frequency is called **pitch**. The wavelength of sound is not directly sensed, but indirect evidence is found in the correlation of the size of musical instruments with their pitch. Small instruments, such as a piccolo, typically make high-pitch sounds, while large instruments, such as a tuba, typically make low-pitch sounds. High pitch means small wavelength, and the size of a musical instrument is directly related to the wavelengths of sound it produces. So a small instrument creates short-wavelength sounds. Similar arguments hold that a large instrument creates long-wavelength sounds.

The relationship of the speed of sound, its frequency, and wavelength is the same as for all waves:

$$v_w = f\lambda,$$

where v_w is the speed of sound, f is its frequency, and λ is its wavelength. The wavelength of a sound is the distance between adjacent identical parts of a wave—for example, between adjacent compressions as illustrated in Figure 6.1.4.2. The frequency is the same as that of the source and is the number of waves that pass a point per unit time.

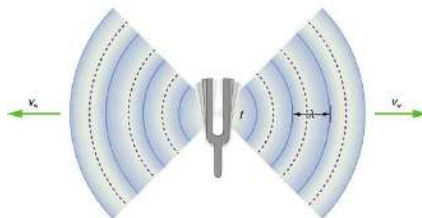


Figure 6.1.4.2: A sound wave emanates from a source vibrating at a frequency f , propagates at v_w , and has a wavelength λ .

Table 6.1.4.1 makes it apparent that the speed of sound varies greatly in different media. The speed of sound in a medium is determined by a combination of the medium's rigidity (or compressibility in gases) and its density. The more rigid (or less compressible) the medium, the faster the speed of sound. This observation is analogous to the fact that the frequency of a simple harmonic motion is directly proportional to the stiffness of the oscillating object. The greater the density of a medium, the slower the speed of sound. This observation is analogous to the fact that the frequency of a simple harmonic motion is inversely proportional to the mass of the oscillating object. The speed of sound in air is low, because air is compressible. Because liquids and solids are relatively rigid and very difficult to compress, the speed of sound in such media is generally greater than in gases.

Table 6.1.4.1: Speed of Sound in Various Media

Medium	$v_w(\text{m/s})$
Gases at 0°C	
Air	331
Carbon dioxide	259
Oxygen	316
Helium	965
Hydrogen	1290
Liquids at 20°C	
Ethanol	1160
Mercury	1450
Water, fresh	1480
Sea water	1540
Human tissue	1540
Solids (longitudinal or bulk)	
Vulcanized rubber	54
Polyethylene	920
Marble	3810
Glass, Pyrex	5640
Lead	1960
Aluminum	5120
Steel	5960

Earthquakes, essentially sound waves in Earth's crust, are an interesting example of how the speed of sound depends on the rigidity of the medium. Earthquakes have both longitudinal and transverse components, and these travel at different speeds. The bulk modulus of granite is greater than its shear modulus. For that reason, the speed of longitudinal or pressure waves (P-waves) in earthquakes in granite is significantly higher than the speed of transverse or shear waves (S-waves). Both components of earthquakes travel slower in less rigid material, such as sediments. P-waves have speeds of 4 to 7 km/s, and S-waves correspondingly range in speed from 2 to 5 km/s, both being faster in more rigid material. The P-wave gets progressively farther ahead of the S-wave as they travel through Earth's crust. The time between the P- and S-waves is routinely used to determine the distance to their source, the epicenter of the earthquake.

The speed of sound is affected by temperature in a given medium. While not negligible, this is not a strong dependence. At 0°C, the speed of sound is 331 m/s, whereas at 20.0°C it is 343 m/s, less than a 4% increase. Figure 6.1.4.3 shows a use of the speed of sound by a bat to sense distances. Echoes are also used in medical imaging.



Figure 6.1.4.3: A bat uses sound echoes to find its way about and to catch prey. The time for the echo to return is directly proportional to the distance.

One of the more important properties of sound is that its speed is nearly independent of frequency. This independence is certainly true in open air for sounds in the audible range of 20 to 20,000 Hz. If this independence were not true, you would certainly notice it for music played by a marching band in a football stadium, for example. Suppose that high-frequency sounds traveled faster—then the farther you were from the band, the more the sound from the low-pitch instruments would lag that from the high-pitch ones. But the music from all instruments arrives in cadence independent of distance, and so all frequencies must travel at nearly the same speed. Recall that

$$v_w = f\lambda.$$

In a given medium under fixed conditions, v_w is constant, so that there is a relationship between f and λ ; the higher the frequency, the smaller the wavelength. See [Figure 6.1.4.4](#) and consider the following example.

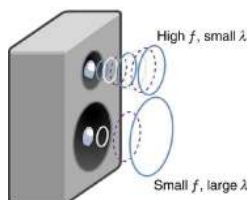


Figure 6.1.4.4: Because they travel at the same speed in a given medium, low-frequency sounds must have a greater wavelength than high-frequency sounds. Here, the lower-frequency sounds are emitted by the large speaker, called a woofer, while the higher-frequency sounds are emitted by the small speaker, called a tweeter.

The speed of sound can change when sound travels from one medium to another. However, the frequency usually remains the same because it is like a driven oscillation and has the frequency of the original source. If v_w changes and f remains the same, then the wavelength λ must change. That is, because $v_w = f\lambda$, the higher the speed of a sound, the greater its wavelength for a given frequency.

MAKING CONNECTIONS: TAKE-HOME INVESTIGATION—VOICE AS A SOUND WAVE

Suspend a sheet of paper so that the top edge of the paper is fixed and the bottom edge is free to move. You could tape the top edge of the paper to the edge of a table. Gently blow near the edge of the bottom of the sheet and note how the sheet moves. Speak softly and then louder such that the sounds hit the edge of the bottom of the paper, and note how the sheet moves. Explain the effects.

Exercise 6.1.4.1

Imagine you observe two fireworks explode. You hear the explosion of one as soon as you see it. However, you see the other firework for several milliseconds before you hear the explosion. Explain why this is so.

Answer

Sound and light both travel at definite speeds. The speed of sound is slower than the speed of light. The first firework is probably very close by, so the speed difference is not noticeable. The second firework is farther away, so the light arrives at your eyes noticeably sooner than the sound wave arrives at your ears.

Exercise 6.1.4.2

You observe two musical instruments that you cannot identify. One plays high-pitch sounds and the other plays low-pitch sounds. How could you determine which is which without hearing either of them play?

Answer

Compare their sizes. High-pitch instruments are generally smaller than low-pitch instruments because they generate a smaller wavelength.

Section Summary

The relationship of the speed of sound v_w , its frequency f , and its wavelength λ is given by

$$v_w = f\lambda,$$

which is the same relationship given for all waves.

In air, the speed of sound is related to air temperature.

v_w is the same for all frequencies and wavelengths.

Glossary

pitch

the perception of the frequency of a sound

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6.1.4.1: The Doppler Effect and Sonic Booms

Learning Objectives

- Define Doppler effect, Doppler shift, and sonic boom.
- Describe the sounds produced by objects moving faster than the speed of sound.

The characteristic sound of a motorcycle buzzing by is an example of the Doppler effect. The high-pitch scream shifts dramatically to a lower-pitch roar as the motorcycle passes by a stationary observer. The closer the motorcycle brushes by, the more abrupt the shift. The faster the motorcycle moves, the greater the shift. We also hear this characteristic shift in frequency for passing race cars, airplanes, and trains. It is so familiar that it is used to imply motion and children often mimic it in play.

The Doppler effect is an alteration in the observed frequency of a sound due to motion of either the source or the observer. Although less familiar, this effect is easily noticed for a stationary source and moving observer. For example, if you ride a train past a stationary warning bell, you will hear the bell's frequency shift from high to low as you pass by. This change in frequency due to *relative* motion of source and observer is called a **Doppler shift**. The Doppler effect is named for the Austrian physicist and mathematician Christian Johann Doppler (1803–1853), who did experiments with both moving sources and moving observers. Doppler, for example, had musicians play on a moving open train car and also play standing next to the train tracks as a train passed by. Their music was observed both on and off the train, and changes in frequency were measured.

What causes the Doppler shift? [Figure 6.1.4.1.1](#), [Figure 6.1.4.1.2](#), and [Figure 6.1.4.1.3](#) compare sound waves emitted by stationary and moving sources in a stationary air mass. Each disturbance spreads out spherically from the point where the sound was emitted. If the source is stationary, then all of the spheres representing the air compressions in the sound wave centered on the same point, and the stationary observers on either side see the same wavelength and frequency as emitted by the source, as in [Figure 6.1.4.1.1](#). If the source is moving, as in [Figure 6.1.4.1.2](#), then the situation is different. Each compression of the air moves out in a sphere from the point where it was emitted, but the point of emission moves. This moving emission point causes the air compressions to be closer together on one side and farther apart on the other. Thus, the wavelength is shorter in the direction the source is moving (on the right in [Figure 6.1.4.1.2](#)), and longer in the opposite direction (on the left in [Figure 6.1.4.1.2](#)). Finally, if the observers move, as in [Figure 6.1.4.1.3](#), the frequency at which they receive the compressions changes. The observer moving toward the source receives them at a higher frequency, and the person moving away from the source receives them at a lower frequency.

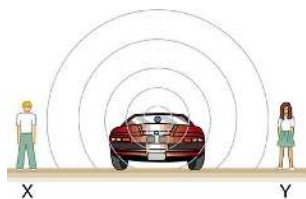


Figure 6.1.4.1.1: Sounds emitted by a source spread out in spherical waves. Because the source, observers, and air are stationary, the wavelength and frequency are the same in all directions and to all observers.

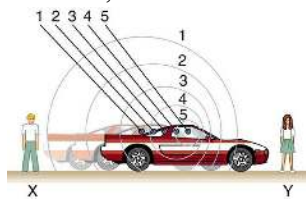


Figure 6.1.4.1.2: Sounds emitted by a source moving to the right spread out from the points at which they were emitted. The wavelength is reduced and, consequently, the frequency is increased in the direction of motion, so that the observer on the right hears a higher-pitch sound. The opposite is true for the observer on the left, where the wavelength is increased



and the frequency is reduced. **Figure 6.1.4.1.3:** The same effect is produced when the observers move relative to the source. Motion toward the source increases frequency as the observer on the right passes through more wave crests

than she would if stationary. Motion away from the source decreases frequency as the observer on the left passes through fewer wave crests than he would if stationary.

We know that wavelength and frequency are related by $v_w = f\lambda$, where v_w is the fixed speed of sound. The sound moves in a medium and has the same speed v_w in that medium whether the source is moving or not. Thus f multiplied by λ is a constant. Because the observer on the right in Figure 6.1.4.1.2 receives a shorter wavelength, the frequency she receives must be higher. Similarly, the observer on the left receives a longer wavelength, and hence he hears a lower frequency. The same thing happens in Figure 6.1.4.1.3. A higher frequency is received by the observer moving toward the source, and a lower frequency is received by an observer moving away from the source. In general, then, relative motion of source and observer toward one another increases the received frequency. Relative motion apart decreases frequency. The greater the relative speed is, the greater the effect.

THE DOPPLER EFFECT

The Doppler effect occurs not only for sound but for any wave when there is relative motion between the observer and the source. There are Doppler shifts in the frequency of sound, light, and water waves, for example. Doppler shifts can be used to determine velocity, such as when ultrasound is reflected from blood in a medical diagnostic. The recession of galaxies is determined by the observed frequency shift of light received from them measured against the frequency emitted by atoms in the laboratory. This implies much about the origins of the universe. Modern physics has been profoundly affected by observations of Doppler shifts.

Sonic Booms to Bow Wakes

What happens to the sound produced by a moving source, such as a jet airplane, that approaches or even exceeds the speed of sound? The answer to this question applies not only to sound but to all other waves as well.

Suppose a jet airplane is coming nearly straight at you, emitting a sound of frequency f_s . The greater the plane's speed v_s , the greater the Doppler shift and the greater the value observed for f_{obs} . Now, as v_s approaches the speed of sound, f_{obs} approaches infinity. At the speed of sound, this result means that in front of the source, each successive wave is superimposed on the previous one because the source moves forward at the speed of sound. The observer gets them all at the same instant, and so the frequency is infinite. (Before airplanes exceeded the speed of sound, some people argued it would be impossible because such constructive superposition would produce pressures great enough to destroy the airplane.) If the source exceeds the speed of sound, no sound is received by the observer until the source has passed, so that the sounds from the approaching source are mixed with those from it when receding. This mixing appears messy, but something interesting happens—a sonic boom is created. (See Figure 6.1.4.1.4.)

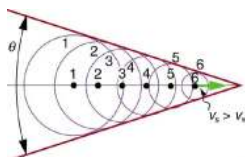


Figure 6.1.4.1.4: Sound waves from a source that moves faster than the speed of sound spread spherically from the point where they are emitted, but the source moves ahead of each. Constructive interference along the lines shown (actually a cone in three dimensions) creates a shock wave called a sonic boom. The faster the speed of the source, the smaller the angle θ .

There is constructive interference along the lines shown (a cone in three dimensions) from similar sound waves arriving there simultaneously. This superposition forms a disturbance called a **sonic boom**, a constructive interference of sound created by an object moving faster than sound. Inside the cone, the interference is mostly destructive, and so the sound intensity there is much less than on the shock wave. An aircraft creates two sonic booms, one from its nose and one from its tail. (See Figure 6.1.4.1.5.) During television coverage of space shuttle landings, two distinct booms could often be heard. These were separated by exactly the time it would take the shuttle to pass by a point. Observers on the ground often do not see the aircraft creating the sonic boom, because it has passed by before the shock wave reaches them, as seen in Figure 6.1.4.1.5. If the aircraft flies close by at low altitude, pressures in the sonic boom can be destructive and break windows as well as rattle nerves. Because of how destructive sonic booms can be, supersonic flights are banned over populated areas of the United States.

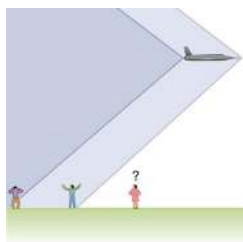


Figure 6.1.4.1.5: Two sonic booms, created by the nose and tail of an aircraft, are observed on the ground after the plane has passed by.

Sonic booms are one example of a broader phenomenon called bow wakes. A **bow wake**, such as the one in [Figure 6.1.4.1.6](#), is created when the wave source moves faster than the wave propagation speed. Water waves spread out in circles from the point where created, and the bow wake is the familiar V-shaped wake trailing the source. A more exotic bow wake is created when a subatomic particle travels through a medium faster than the speed of light travels in that medium (the speed of light in vacuum is $c = 2.998 \times 10^8 \text{ m/s}$, and this is the maximum speed anything can move, but in water, for example, speed of light slows to $0.75c$, and other particles can move faster). If the particle creates light in its passage, that light spreads on a cone with an angle indicative of the speed of the particle, as illustrated in [Figure 6.1.4.1.7](#). Such a bow wake is called Cherenkov radiation and is commonly observed in particle physics.



Figure 6.1.4.1.6: Bow wake created by a duck. Constructive interference produces the rather structured wake, while there is relatively little wave action inside the wake, where interference is mostly destructive. (credit: Horia Varlan,



Figure 6.1.4.1.7: The blue glow in this research reactor pool is Cherenkov radiation caused by subatomic particles traveling faster than the speed of light in water. (credit: U.S. Nuclear Regulatory Commission)

Doppler shifts and sonic booms are interesting sound phenomena that occur in all types of waves. They can be of considerable use. For example, the Doppler shift in ultrasound can be used to measure blood velocity, while police use the Doppler shift in radar (a microwave) to measure car velocities. In meteorology, the Doppler shift is used to track the motion of storm clouds; such “Doppler Radar” can give velocity and direction and rain or snow potential of imposing weather fronts. In astronomy, we can examine the light emitted from distant galaxies and determine their speed relative to ours. As galaxies move away from us, their light is shifted to a lower frequency, and so to a longer wavelength—the so-called red shift. Such information from galaxies far, far away has allowed us to estimate the age of the universe (from the Big Bang) as about 14 billion years.

Exercise 6.1.4.1.1

Why did scientist Christian Doppler observe musicians both on a moving train and also from a stationary point not on the train?

Answer

Doppler needed to compare the perception of sound when the observer is stationary and the sound source moves, as well as when the sound source and the observer are both in motion.

Exercise 6.1.4.1.2

Describe a situation in your life when you might rely on the Doppler shift to help you either while driving a car or walking near traffic.

Answer

If I am driving and I hear Doppler shift in an ambulance siren, I would be able to tell when it was getting closer and also if it has passed by. This would help me to know whether I needed to pull over and let the ambulance through

Section Summary

- The Doppler shift is change in observed frequency of a wave (for example, a sound wave) due to relative motion of the source and the observer, with approaching source and/or observer increasing the observed frequency and receding source and/or observer decreasing the observed frequency.
- A sonic boom is constructive interference of sound created by an object moving faster than sound.
- A shock wave is a type of bow wake created when any wave source moves faster than the wave propagation speed.

Glossary

Doppler shift

the change in wave frequency due to relative motion of source and observer

sonic boom

a constructive interference of sound created by an object moving faster than sound

bow wake

V-shaped disturbance created when the wave source moves faster than the wave propagation speed

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6.2: Light

Light from this page or screen is formed into an image by the lens of your eye, much as the lens of the camera that made this photograph. Mirrors, like lenses, can also form images that in turn are captured by your eye.



Figure 6.2.1: Image seen as a result of reflection of light on a plane smooth surface. (credit: NASA Goddard Photo and Video, via Flickr)

Our lives are filled with light. Through vision, the most valued of our senses, light can evoke spiritual emotions, such as when we view a magnificent sunset or glimpse a rainbow breaking through the clouds. Light can also simply amuse us in a theater, or warn us to stop at an intersection. It has innumerable uses beyond vision. Light can carry telephone signals through glass fibers or cook a meal in a solar oven. Life itself could not exist without light's energy. From photosynthesis in plants to the sun warming a cold-blooded animal, its supply of energy is vital.



Figure 6.2.2: Double Rainbow over the bay of Pocitos in Montevideo, Uruguay. (credit: Madrax, Wikimedia Commons)

We will start our discussion of visible light as a type of electromagnetic wave. This knowledge will help us answer questions regarding the nature of light and vision. What is color, and how do our eyes detect it? Why do diamonds sparkle? How does light travel? How do lenses and mirrors form images? These are but a few of the questions that are answered by the study of optics. Optics is the branch of physics that deals with the behavior of visible light and other electromagnetic waves.

It is convenient to divide optics into two major parts based on the features we are interested in. The wave characteristics of light, such as frequency and wavelength, relates to the colors we perceive and to how we characterize different types of electromagnetic wave along the electromagnetic spectrum. The wave nature of light is also responsible for phenomena such as diffraction and interference. We call this part of optics "wave optics" or "physical optics." But when light interacts with an object that is several times as large as the light's wavelength, its observable behavior is like that of a ray; it does not prominently display its wave characteristics, and we can look at primarily how it refracts and reflects. We call this part of optics "geometric optics" or "ray optics."

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6.2.1: Electromagnetic Radiation

Learning Objectives

- List three “rules of thumb” that apply to the different frequencies along the electromagnetic spectrum.
- Explain why the higher the frequency, the shorter the wavelength of an electromagnetic wave.
- Draw a simplified electromagnetic spectrum, indicating the relative positions, frequencies, and spacing of the different types of radiation bands.

In this module we examine how electromagnetic waves are classified into categories such as radio, infrared, ultraviolet, and so on, so that we can understand some of their similarities as well as some of their differences. We will also find that there are many connections with previously discussed topics, such as wavelength and resonance. A brief overview of the production and utilization of electromagnetic waves is found in [Table 6.2.1.1](#). Note that the vast majority of the different types of electromagnetic waves originate from atomic and/or molecular electron transitions—that is, from electrons changing their energy levels within atoms or molecules.

Table 6.2.1.1: Electromagnetic Waves

Type of EM wave	Production	Applications	Life sciences aspect	Issues
Radio and TV	Accelerating charges	Communications, Remote controls	MRI	Requires controls for band use
Microwaves	Accelerating charges and thermal agitation	Communications, Ovens, Radar	Deep heating	Cell phone use
Infrared	Thermal agitations and atomic/molecular electron transitions	Thermal imaging, Heating	Absorbed by atmosphere	Greenhouse effect
Visible light	Thermal agitations and atomic/molecular electron transitions	All pervasive	Photosynthesis, Human vision	
Ultraviolet	Thermal agitations and atomic/molecular electron transitions	Sterilization, Cancer control	Vitamin D production	Ozone depletion, Cancer causing
X-rays	Inner atomic electron transitions and fast collisions	Medical, Security	Medical diagnosis, Cancer therapy	Cancer causing
Gamma rays	Nuclear decay	Nuclear medicine, Security	Medical diagnosis, Cancer therapy	Cancer causing, Radiation damage

CONNECTIONS: WAVES

There are many types of waves, such as water waves and even earthquakes. Among the many shared attributes of waves are propagation speed, frequency, and wavelength. These are always related by the expression $v_w = f\lambda$. This module concentrates on EM waves, but other modules contain examples of all of these characteristics for sound waves and submicroscopic particles.

As noted before, an electromagnetic wave has a frequency and a wavelength associated with it and travels at the speed of light, or c . The relationship among these wave characteristics can be described by $v_w = f\lambda$, where v_w is the propagation speed of the wave, f is the frequency, and λ is the wavelength. Here $v_w = c$, so that for all electromagnetic waves,

$$c = f\lambda.$$

Thus, for all electromagnetic waves, the greater the frequency, the smaller the wavelength.

Figure 6.2.1.1 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.

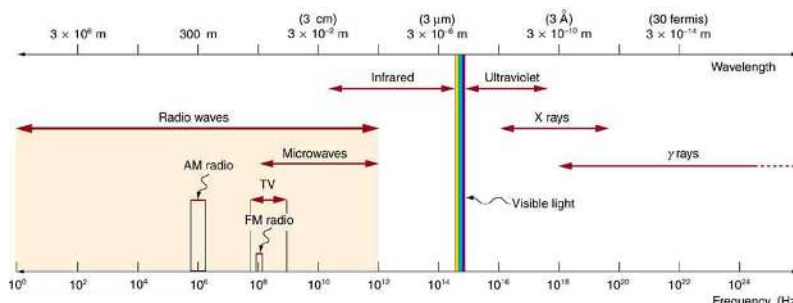


Figure 6.2.1.1: The electromagnetic spectrum, showing the major categories of electromagnetic waves. The range of frequencies and wavelengths is remarkable. The dividing line between some categories is distinct, whereas other categories overlap.

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.

Section Summary

- The relationship among the speed of propagation, wavelength, and frequency for any wave is given by $v_w = f\lambda$, so that for electromagnetic waves,

$$c = f\lambda,$$

where f is the frequency, λ is the wavelength, and c is the speed of light.

- The electromagnetic spectrum is separated into many categories and subcategories, based on the frequency and wavelength, source, and uses of the electromagnetic waves.

Glossary

electromagnetic spectrum

the full range of wavelengths or frequencies of electromagnetic radiation

radio waves

electromagnetic waves with wavelengths in the range from 1 mm to 100 km; they are produced by currents in wires and circuits and by astronomical phenomena

microwaves

electromagnetic waves with wavelengths in the range from 1 mm to 1 m; they can be produced by currents in macroscopic circuits and devices

infrared radiation (IR)

a region of the electromagnetic spectrum with a frequency range that extends from just below the red region of the visible light spectrum up to the microwave region, or from 0.74 μm to 300 μm

ultraviolet radiation (UV)

electromagnetic radiation in the range extending upward in frequency from violet light and overlapping with the lowest X-ray frequencies, with wavelengths from 400 nm down to about 10 nm

visible light

the narrow segment of the electromagnetic spectrum to which the normal human eye responds

X-ray

invisible, penetrating form of very high frequency electromagnetic radiation, overlapping both the ultraviolet range and the γ -ray range

gamma ray

(γ ray); extremely high frequency electromagnetic radiation emitted by the nucleus of an atom, either from natural nuclear decay or induced nuclear processes in nuclear reactors and weapons. The lower end of the γ -ray frequency range overlaps the upper end of the X-ray range, but γ rays can have the highest frequency of any electromagnetic radiation

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- [11.4: The Electromagnetic Spectrum- an Overview](#) by [OpenStax](#) is licensed [CC BY 4.0](#).

6.2.1.1: Sources of Light

Learning Objectives

- List and explain the characteristics and applications of different parts of the electromagnetic spectrum.

In this module, we look at the properties of different types of electromagnetic waves. Again, [Figure 6.2.1.1.1](#) shows the electromagnetic spectrum. The characteristics of the various types of electromagnetic waves you will read about below are related to their frequencies and wavelengths.

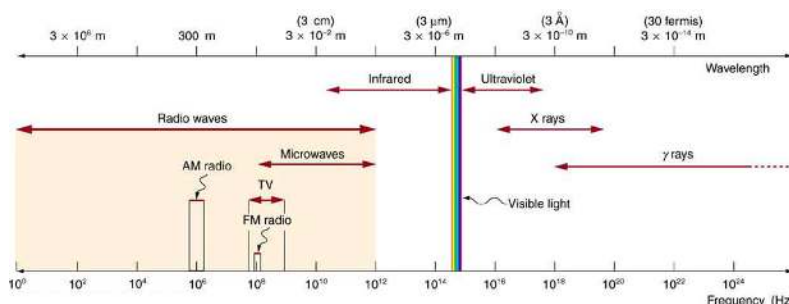


Figure 6.2.1.1.1: The electromagnetic spectrum, showing the major categories of electromagnetic waves. The range of frequencies and wavelengths is remarkable. The dividing line between some categories is distinct, whereas other categories overlap.

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 6.2.1.1.2](#).) These extremely long wavelength electromagnetic waves (about 6000 km!) are one means of energy loss in long-distance power transmission.



Figure 6.2.1.1.2: This high-voltage traction power line running to Eutingen Railway Substation in Germany radiates electromagnetic waves with very long wavelengths. (credit: Zonk43, Wikimedia Commons)

There is an ongoing controversy regarding potential health hazards associated with exposure to these electromagnetic fields (**E**-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 **E**-fields.

Extremely low frequency (ELF) radio waves of about 1 kHz are used to communicate with submerged submarines. The ability of radio waves to penetrate salt water is related to their wavelength (much like ultrasound penetrating tissue)—the longer the wavelength, the farther they penetrate. Since salt water is a good conductor, radio waves are strongly absorbed by it, and very long wavelengths are needed to reach a submarine under the surface. (See [Figure 6.2.1.1.3](#).)

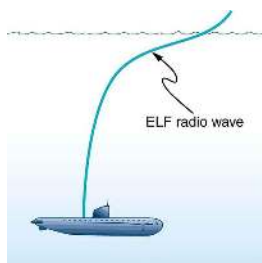


Figure 6.2.1.1.3: Very long wavelength radio waves are needed to reach this submarine, requiring extremely low frequency signals (ELF). Shorter wavelengths do not penetrate to any significant depth.

AM radio waves are used to carry commercial radio signals in the frequency range from 540 to 1600 kHz. The abbreviation AM stands for **amplitude modulation**, which is the method for placing information on these waves. (See [Figure 6.2.1.1.4](#).) A **carrier wave** having the basic frequency of the radio station, say 1530 kHz, is varied or modulated in amplitude by an audio signal. The resulting wave has a constant frequency, but a varying amplitude.

A radio receiver tuned to have the same resonant frequency as the carrier wave can pick up the signal, while rejecting the many other frequencies impinging on its antenna. The receiver's circuitry is designed to respond to variations in amplitude of the carrier wave to replicate the original audio signal. That audio signal is amplified to drive a speaker or perhaps to be recorded.

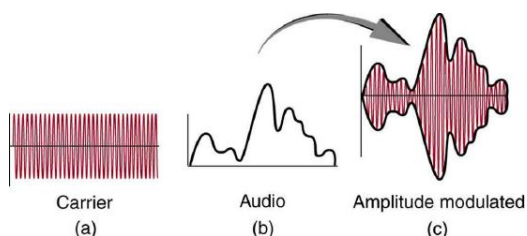


Figure 6.2.1.1.4: Amplitude modulation for AM radio. (a) A carrier wave at the station's basic frequency. (b) An audio signal at much lower audible frequencies. (c) The amplitude of the carrier is modulated by the audio signal without changing its basic frequency.

FM Radio Waves

FM radio waves are also used for commercial radio transmission, but in the frequency range of 88 to 108 MHz. FM stands for **frequency modulation**, another method of carrying information. (See [Figure 6.2.1.1.5](#).) Here a carrier wave having the basic frequency of the radio station, perhaps 105.1 MHz, is modulated in frequency by the audio signal, producing a wave of constant amplitude but varying frequency.

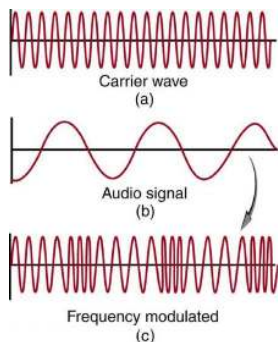


Figure 6.2.1.1.5: Frequency modulation for FM radio. (a) A carrier wave at the station's basic frequency. (b) An audio signal at much lower audible frequencies. (c) The frequency of the carrier is modulated by the audio signal without changing its amplitude.

Since audible frequencies range up to 20 kHz (or 0.020 MHz) at most, the frequency of the FM radio wave can vary from the carrier by as much as 0.020 MHz. Thus the carrier frequencies of two different radio stations cannot be closer than 0.020 MHz. An FM receiver is tuned to resonate at the carrier frequency and has circuitry that responds to variations in frequency, reproducing the audio information.

FM radio is inherently less subject to noise from stray radio sources than AM radio. The reason is that amplitudes of waves add. So an AM receiver would interpret noise added onto the amplitude of its carrier wave as part of the information. An FM receiver can

be made to reject amplitudes other than that of the basic carrier wave and only look for variations in frequency. It is thus easier to reject noise from FM, since noise produces a variation in amplitude.

Television is also broadcast on electromagnetic waves. Since the waves must carry a great deal of visual as well as audio information, each channel requires a larger range of frequencies than simple radio transmission. TV channels utilize frequencies in the range of 54 to 88 MHz and 174 to 222 MHz. (The entire FM radio band lies between channels 88 MHz and 174 MHz.) These TV channels are called VHF (for **very high frequency**). Other channels called UHF (for **ultra high frequency**) utilize an even higher frequency range of 470 to 1000 MHz.

The TV video signal is AM, while the TV audio is FM. Note that these frequencies are those of free transmission with the user utilizing an old-fashioned roof antenna. Satellite dishes and cable transmission of TV occurs at significantly higher frequencies and is rapidly evolving with the use of the high-definition or HD format.

Example 6.2.1.1.1: Calculating Wavelengths of Radio Waves

Calculate the wavelengths of a 1530-kHz AM radio signal, a 105.1-MHz FM radio signal, and a 1.90-GHz cell phone signal.

Strategy

The relationship between wavelength and frequency is $c = f\lambda$, where $c = 3.00 \times 10^8 \text{ m/s}$ is the speed of light (the speed of light is only very slightly smaller in air than it is in a vacuum). We can rearrange this equation to find the wavelength for all three frequencies.

Solution

Rearranging gives

$$\lambda = \frac{c}{f}.$$

(a) For the $f = 1530 \text{ kHz}$ AM radio signal, then,

$$\begin{aligned}\lambda &= \frac{3.00 \times 10^8 \text{ m/s}}{1530 \times 10^3 \text{ cycles/s}} \\ &= 196 \text{ m.}\end{aligned}$$

(b) For the $f = 105.1 \text{ MHz}$ FM radio signal,

$$\begin{aligned}\lambda &= \frac{3.00 \times 10^8 \text{ m/s}}{105.1 \times 10^6 \text{ cycles/s}} \\ &= 2.85 \text{ m.}\end{aligned}$$

(c) And for the $f = 1.90 \text{ GHz}$ cell phone,

$$\begin{aligned}\lambda &= \frac{3.00 \times 10^8 \text{ m/s}}{1.90 \times 10^9 \text{ cycles/s}} \\ &= 0.158 \text{ m.}\end{aligned}$$

Discussion

These wavelengths are consistent with the spectrum in [Figure 6.2.1.1.1](#). The wavelengths are also related to other properties of these electromagnetic waves, as we shall see.

The wavelengths found in the preceding example are representative of AM, FM, and cell phones, and account for some of the differences in how they are broadcast and how well they travel. The most efficient length for a linear antenna, such as discussed in "Production of Electromagnetic Waves", is $\lambda/2$, half the wavelength of the electromagnetic wave. Thus a very large antenna is needed to efficiently broadcast typical AM radio with its carrier wavelengths on the order of hundreds of meters.

One benefit to these long AM wavelengths is that they can go over and around rather large obstacles (like buildings and hills), just as ocean waves can go around large rocks. FM and TV are best received when there is a line of sight between the broadcast antenna

and receiver, and they are often sent from very tall structures. FM, TV, and mobile phone antennas themselves are much smaller than those used for AM, but they are elevated to achieve an unobstructed line of sight. (See [Figure 6.2.1.1.6.](#))

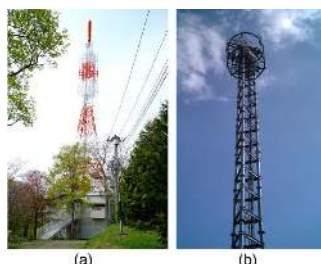


Figure 6.2.1.1.6: (a) A large tower is used to broadcast TV signals. The actual antennas are small structures on top of the tower—they are placed at great heights to have a clear line of sight over a large broadcast area. (credit: Ozizo, Wikimedia Commons) (b) The NTT Dokomo mobile phone tower at Tokorozawa City, Japan. (credit: tokoroten, Wikimedia Commons)

Radio Wave Interference

Astronomers and astrophysicists collect signals from outer space using electromagnetic waves. A common problem for astrophysicists is the “pollution” from electromagnetic radiation pervading our surroundings from communication systems in general. Even everyday gadgets like our car keys having the facility to lock car doors remotely and being able to turn TVs on and off using remotes involve radio-wave frequencies. In order to prevent interference between all these electromagnetic signals, strict regulations are drawn up for different organizations to utilize different radio frequency bands.

One reason why we are sometimes asked to switch off our mobile phones (operating in the range of 1.9 GHz) on airplanes and in hospitals is that important communications or medical equipment often uses similar radio frequencies and their operation can be affected by frequencies used in the communication devices.

For example, radio waves used in magnetic resonance imaging (MRI) have frequencies on the order of 100 MHz, although this varies significantly depending on the strength of the magnetic field used and the nuclear type being scanned. MRI is an important medical imaging and research tool, producing highly detailed two- and three-dimensional images. Radio waves are broadcast, absorbed, and reemitted in a resonance process that is sensitive to the density of nuclei (usually protons or hydrogen nuclei).

The wavelength of 100-MHz radio waves is 3 m, yet using the sensitivity of the resonant frequency to the magnetic field strength, details smaller than a millimeter can be imaged. This is a good example of an exception to a rule of thumb (in this case, the rubric that details much smaller than the probe’s wavelength cannot be detected). The intensity of the radio waves used in MRI presents little or no hazard to human health.

Microwaves

Microwaves are the highest-frequency electromagnetic waves that can be produced by currents in macroscopic circuits and devices. Microwave frequencies range from about 10^9 Hz to the highest practical LC resonance at nearly 10^{12} Hz. Since they have high frequencies, their wavelengths are short compared with those of other radio waves—hence the name “microwave.”

Microwaves can also be produced by atoms and molecules. They are, for example, a component of electromagnetic radiation generated by **thermal agitation**. The thermal motion of atoms and molecules in any object at a temperature above absolute zero causes them to emit and absorb radiation.

Since it is possible to carry more information per unit time on high frequencies, microwaves are quite suitable for communications. Most satellite-transmitted information is carried on microwaves, as are land-based long-distance transmissions. A clear line of sight between transmitter and receiver is needed because of the short wavelengths involved.

Radar is a common application of microwaves that was first developed in World War II. By detecting and timing microwave echoes, radar systems can determine the distance to objects as diverse as clouds and aircraft. A Doppler shift in the radar echo can be used to determine the speed of a car or the intensity of a rainstorm. Sophisticated radar systems are used to map the Earth and other planets, with a resolution limited by wavelength. (See [Figure 6.2.1.1.7.](#)) The shorter the wavelength of any probe, the smaller the detail it is possible to observe.

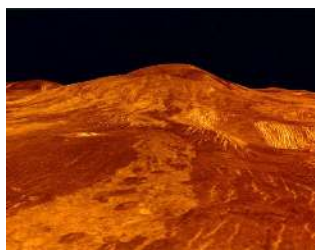


Figure 6.2.1.1.7: An image of Sif Mons with lava flows on Venus, based on Magellan synthetic aperture radar data combined with radar altimetry to produce a three-dimensional map of the surface. The Venusian atmosphere is opaque to visible light, but not to the microwaves that were used to create this image. (credit: NSSDC, NASA/JPL)

Heating with Microwaves

How does the ubiquitous microwave oven produce microwaves electronically, and why does food absorb them preferentially? Microwaves at a frequency of 2.45 GHz are produced by accelerating electrons. The microwaves are then used to induce an alternating electric field in the oven.

Water and some other constituents of food have a slightly negative charge at one end and a slightly positive charge at one end (called polar molecules). The range of microwave frequencies is specially selected so that the polar molecules, in trying to keep orienting themselves with the electric field, absorb these energies and increase their temperatures—called dielectric heating.

The energy thereby absorbed results in thermal agitation heating food and not the plate, which does not contain water. Hot spots in the food are related to constructive and destructive interference patterns. Rotating antennas and food turntables help spread out the hot spots.

Another use of microwaves for heating is within the human body. Microwaves will penetrate more than shorter wavelengths into tissue and so can accomplish “deep heating” (called microwave diathermy). This is used for treating muscular pains, spasms, tendonitis, and rheumatoid arthritis.

MAKING CONNECTIONS: TAKE-HOME EXPERIMENT—MICROWAVE OVENS

1. Look at the door of a microwave oven. Describe the structure of the door. Why is there a metal grid on the door? How does the size of the holes in the grid compare with the wavelengths of microwaves used in microwave ovens? What is this wavelength?
2. Place a glass of water (about 250 ml) in the microwave and heat it for 30 seconds. Measure the temperature gain (the ΔT). Assuming that the power output of the oven is 1000 W, calculate the efficiency of the heat-transfer process.
3. Remove the rotating turntable or moving plate and place a cup of water in several places along a line parallel with the opening. Heat for 30 seconds and measure the ΔT for each position. Do you see cases of destructive interference?

Microwaves generated by atoms and molecules far away in time and space can be received and detected by electronic circuits. Deep space acts like a blackbody with a 2.7 K temperature, radiating most of its energy in the microwave frequency range. In 1964, Penzias and Wilson detected this radiation and eventually recognized that it was the radiation of the Big Bang’s cooled remnants.

Infrared Radiation

The microwave and infrared regions of the electromagnetic spectrum overlap (see [Figure 6.2.1.1.1](#)). **Infrared radiation** is generally produced by thermal motion and the vibration and rotation of atoms and molecules. Electronic transitions in atoms and molecules can also produce infrared radiation.

The range of infrared frequencies extends up to the lower limit of visible light, just below red. In fact, infrared means “below red.” Frequencies at its upper limit are too high to be produced by accelerating electrons in circuits, but small systems, such as atoms and molecules, can vibrate fast enough to produce these waves.

Water molecules rotate and vibrate particularly well at infrared frequencies, emitting and absorbing them so efficiently that the emissivity for skin is $\epsilon = 0.97$ in the infrared. Night-vision scopes can detect the infrared emitted by various warm objects, including humans, and convert it to visible light.

We can examine radiant heat transfer from a house by using a camera capable of detecting infrared radiation. Reconnaissance satellites can detect buildings, vehicles, and even individual humans by their infrared emissions, whose power radiation is proportional to the fourth power of the absolute temperature. More mundanely, we use infrared lamps, some of which are called quartz heaters, to preferentially warm us because we absorb infrared better than our surroundings.

The Sun radiates like a nearly perfect blackbody (that is, it has $e = 1$), with a 6000 K surface temperature. About half of the solar energy arriving at the Earth is in the infrared region, with most of the rest in the visible part of the spectrum, and a relatively small amount in the ultraviolet. On average, 50 percent of the incident solar energy is absorbed by the Earth.

The relatively constant temperature of the Earth is a result of the energy balance between the incoming solar radiation and the energy radiated from the Earth. Most of the infrared radiation emitted from the Earth is absorbed by CO_2 and H_2O in the atmosphere and then radiated back to Earth or into outer space. This radiation back to Earth is known as the greenhouse effect, and it maintains the surface temperature of the Earth about 40°C higher than it would be if there is no absorption. Some scientists think that the increased concentration of CO_2 and other greenhouse gases in the atmosphere, resulting from increases in fossil fuel burning, has increased global average temperatures.

Visible Light

Visible light is the narrow segment of the electromagnetic spectrum to which the normal human eye responds. Visible light is produced by vibrations and rotations of atoms and molecules, as well as by electronic transitions within atoms and molecules. The receivers or detectors of light largely utilize electronic transitions. We say the atoms and molecules are excited when they absorb and relax when they emit through electronic transitions.

Figure 6.2.1.1.8 shows this part of the spectrum, together with the colors associated with particular pure wavelengths. We usually refer to visible light as having wavelengths of between 400 nm and 750 nm. (The retina of the eye actually responds to the lowest ultraviolet frequencies, but these do not normally reach the retina because they are absorbed by the cornea and lens of the eye.)

Red light has the lowest frequencies and longest wavelengths, while violet has the highest frequencies and shortest wavelengths. Blackbody radiation from the Sun peaks in the visible part of the spectrum but is more intense in the red than in the violet, making the Sun yellowish in appearance.

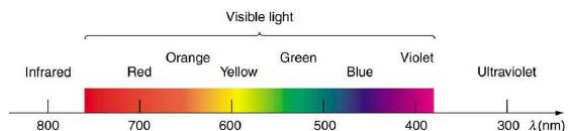


Figure 6.2.1.1.8: A small part of the electromagnetic spectrum that includes its visible components. The divisions between infrared, visible, and ultraviolet are not perfectly distinct, nor are those between the seven rainbow colors.

Living things—plants and animals—have evolved to utilize and respond to parts of the electromagnetic spectrum they are embedded in. Visible light is the most predominant and we enjoy the beauty of nature through visible light. Plants are more selective. Photosynthesis makes use of parts of the visible spectrum to make sugars.

Optics is the study of the behavior of visible light and other forms of electromagnetic waves. Optics falls into two distinct categories. When electromagnetic radiation, such as visible light, interacts with objects that are large compared with its wavelength, its motion can be represented by straight lines like rays. Ray optics is the study of such situations and includes lenses and mirrors.

When electromagnetic radiation interacts with objects about the same size as the wavelength or smaller, its wave nature becomes apparent. For example, observable detail is limited by the wavelength, and so visible light can never detect individual atoms, because they are so much smaller than its wavelength. Physical or wave optics is the study of such situations and includes all wave characteristics.

TAKE-HOME EXPERIMENT: COLORS THAT MATCH

When you light a match you see largely orange light; when you light a gas stove you see blue light. Why are the colors different? What other colors are present in these?

Ultraviolet Radiation

Ultraviolet means “above violet.” The electromagnetic frequencies of **ultraviolet radiation (UV)** extend upward from violet, the highest-frequency visible light. Ultraviolet is also produced by atomic and molecular motions and electronic transitions. The wavelengths of ultraviolet extend from 400 nm down to about 10 nm at its highest frequencies, which overlap with the lowest X-ray frequencies. It was recognized as early as 1801 by Johann Ritter that the solar spectrum had an invisible component beyond the violet range.

Solar UV radiation is broadly subdivided into three regions: UV-A (320–400 nm), UV-B (290–320 nm), and UV-C (220–290 nm), ranked from long to shorter wavelengths (from smaller to larger energies). Most UV-B and all UV-C is absorbed by ozone (O_3) molecules in the upper atmosphere. Consequently, 99% of the solar UV radiation reaching the Earth’s surface is UV-A.

Human Exposure to UV Radiation

It is largely exposure to UV-B that causes skin cancer. It is estimated that as many as 20% of adults will develop skin cancer over the course of their lifetime. Again, treatment is often successful if caught early. Despite very little UV-B reaching the Earth’s surface, there are substantial increases in skin-cancer rates in countries such as Australia, indicating how important it is that UV-B and UV-C continue to be absorbed by the upper atmosphere.

All UV radiation can damage collagen fibers, resulting in an acceleration of the aging process of skin and the formation of wrinkles. Because there is so little UV-B and UV-C reaching the Earth’s surface, sunburn is caused by large exposures, and skin cancer from repeated exposure. Some studies indicate a link between overexposure to the Sun when young and melanoma later in life.

The tanning response is a defense mechanism in which the body produces pigments to absorb future exposures in inert skin layers above living cells. Basically UV-B radiation excites DNA molecules, distorting the DNA helix, leading to mutations and the possible formation of cancerous cells.

Repeated exposure to UV-B may also lead to the formation of cataracts in the eyes—a cause of blindness among people living in the equatorial belt where medical treatment is limited. Cataracts, clouding in the eye’s lens and a loss of vision, are age related; 60% of those between the ages of 65 and 74 will develop cataracts. However, treatment is easy and successful, as one replaces the lens of the eye with a plastic lens. Prevention is important. Eye protection from UV is more effective with plastic sunglasses than those made of glass.

A major acute effect of extreme UV exposure is the suppression of the immune system, both locally and throughout the body.

Low-intensity ultraviolet is used to sterilize haircutting implements, implying that the energy associated with ultraviolet is deposited in a manner different from lower-frequency electromagnetic waves. (Actually this is true for all electromagnetic waves with frequencies greater than visible light.)

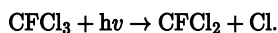
Flash photography is generally not allowed of precious artworks and colored prints because the UV radiation from the flash can cause photo-degradation in the artworks. Often artworks will have an extra-thick layer of glass in front of them, which is especially designed to absorb UV radiation.

UV Light and the Ozone Layer

If all of the Sun’s ultraviolet radiation reached the Earth’s surface, there would be extremely grave effects on the biosphere from the severe cell damage it causes. However, the layer of ozone (O_3) in our upper atmosphere (10 to 50 km above the Earth) protects life by absorbing most of the dangerous UV radiation.

Unfortunately, today we are observing a depletion in ozone concentrations in the upper atmosphere. This depletion has led to the formation of an “ozone hole” in the upper atmosphere. The hole is more centered over the southern hemisphere, and changes with the seasons, being largest in the spring. This depletion is attributed to the breakdown of ozone molecules by refrigerant gases called chlorofluorocarbons (CFCs).

The UV radiation helps dissociate the CFC’s, releasing highly reactive chlorine (Cl) atoms, which catalyze the destruction of the ozone layer. For example, the reaction of CFCl_3 with a photon of light ($h\nu$) can be written as:



The Cl atom then catalyzes the breakdown of ozone as follows:



A single chlorine atom could destroy ozone molecules for up to two years before being transported down to the surface. The CFCs are relatively stable and will contribute to ozone depletion for years to come. CFCs are found in refrigerants, air conditioning systems, foams, and aerosols.

International concern over this problem led to the establishment of the “Montreal Protocol” agreement (1987) to phase out CFC production in most countries. However, developing-country participation is needed if worldwide production and elimination of CFCs is to be achieved. Probably the largest contributor to CFC emissions today is India. But the protocol seems to be working, as there are signs of an ozone recovery. (See [Figure 6.2.1.1.9.](#))

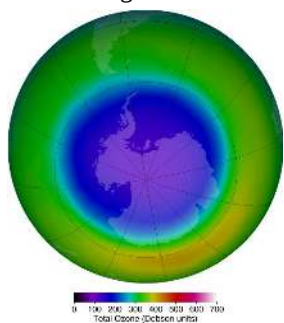


Figure 6.2.1.1.9: This map of ozone concentration over Antarctica in October 2011 shows severe depletion suspected to be caused by CFCs. Less dramatic but more general depletion has been observed over northern latitudes, suggesting the effect is global. With less ozone, more ultraviolet radiation from the Sun reaches the surface, causing more damage. (credit: NASA Ozone Watch)

Benefits of UV Light

Besides the adverse effects of ultraviolet radiation, there are also benefits of exposure in nature and uses in technology. Vitamin D production in the skin (epidermis) results from exposure to UVB radiation, generally from sunlight. A number of studies indicate lack of vitamin D can result in the development of a range of cancers (prostate, breast, colon), so a certain amount of UV exposure is helpful. Lack of vitamin D is also linked to osteoporosis. Exposures (with no sunscreen) of 10 minutes a day to arms, face, and legs might be sufficient to provide the accepted dietary level. However, in the winter time north of about **37°** latitude, most UVB gets blocked by the atmosphere.

UV radiation is used in the treatment of infantile jaundice and in some skin conditions. It is also used in sterilizing workspaces and tools, and killing germs in a wide range of applications. It is also used as an analytical tool to identify substances.

When exposed to ultraviolet, some substances, such as minerals, glow in characteristic visible wavelengths, a process called fluorescence. So-called black lights emit ultraviolet to cause posters and clothing to fluoresce in the visible. Ultraviolet is also used in special microscopes to detect details smaller than those observable with longer-wavelength visible-light microscopes.

X-Rays

In the 1850s, scientists (such as Faraday) began experimenting with high-voltage electrical discharges in tubes filled with rarefied gases. It was later found that these discharges created an invisible, penetrating form of very high frequency electromagnetic radiation. This radiation was called an **X-ray**, because its identity and nature were unknown.

THINGS GREAT AND SMALL: A SUBMICROSCOPIC VIEW OF X-RAY PRODUCTION

X-rays can be created in a high-voltage discharge. They are emitted in the material struck by electrons in the discharge current. There are two mechanisms by which the electrons create X-rays.

The first method is illustrated in [Figure 6.2.1.1.10](#). An electron is accelerated in an evacuated tube by a high positive voltage. The electron strikes a metal plate (e.g., copper) and produces X-rays. Since this is a high-voltage discharge, the electron gains sufficient energy to ionize the atom.

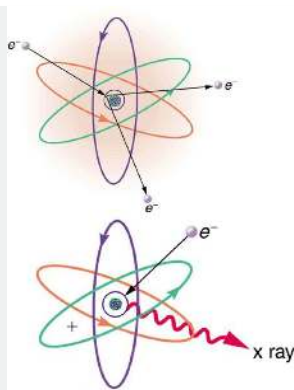


Figure 6.2.1.1.10: Artist's conception of an electron ionizing an atom followed by the recapture of an electron and emission of an X-ray. An energetic electron strikes an atom and knocks an electron out of one of the orbits closest to the nucleus. Later, the atom captures another electron, and the energy released by its fall into a low orbit generates a high-energy EM wave called an X-ray.

In the case shown, an inner-shell electron (one in an orbit relatively close to and tightly bound to the nucleus) is ejected. A short time later, another electron is captured and falls into the orbit in a single great plunge. The energy released by this fall is given to an EM wave known as an X-ray. Since the orbits of the atom are unique to the type of atom, the energy of the X-ray is characteristic of the atom, hence the name characteristic X-ray.

The second method by which an energetic electron creates an X-ray when it strikes a material is illustrated in [Figure 6.2.1.1.11](#). The electron interacts with charges in the material as it penetrates. These collisions transfer kinetic energy from the electron to the electrons and atoms in the material.

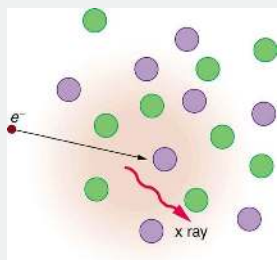


Figure 6.2.1.1.11: Artist's conception of an electron being slowed by collisions in a material and emitting X-ray radiation. This energetic electron makes numerous collisions with electrons and atoms in a material it penetrates. An accelerated charge radiates EM waves, a second method by which X-rays are created.

A loss of kinetic energy implies an acceleration, in this case decreasing the electron's velocity. Whenever a charge is accelerated, it radiates EM waves. Given the high energy of the electron, these EM waves can have high energy. We call them X-rays. Since the process is random, a broad spectrum of X-ray energy is emitted that is more characteristic of the electron energy than the type of material the electron encounters. Such EM radiation is called “bremsstrahlung” (German for “braking radiation”).

As described above, there are two methods by which X-rays are created—both are submicroscopic processes and can be caused by high-voltage discharges. While the low-frequency end of the X-ray range overlaps with the ultraviolet, X-rays extend to much higher frequencies (and energies).

X-rays have adverse effects on living cells similar to those of ultraviolet radiation, and they have the additional liability of being more penetrating, affecting more than the surface layers of cells. Cancer and genetic defects can be induced by exposure to X-rays. Because of their effect on rapidly dividing cells, X-rays can also be used to treat and even cure cancer.

The widest use of X-rays is for imaging objects that are opaque to visible light, such as the human body or aircraft parts. In humans, the risk of cell damage is weighed carefully against the benefit of the diagnostic information obtained. However, questions have risen in recent years as to accidental overexposure of some people during CT scans—a mistake at least in part due to poor monitoring of radiation dose.

The ability of X-rays to penetrate matter depends on density, and so an X-ray image can reveal very detailed density information. [Figure 6.2.1.1.12](#) shows an example of the simplest type of X-ray image, an X-ray shadow on film. The amount of information in a simple X-ray image is impressive, but more sophisticated techniques, such as CT scans, can reveal three-dimensional information with details smaller than a millimeter.



Figure 6.2.1.1.12: This shadow X-ray image shows many interesting features, such as artificial heart valves, a pacemaker, and the wires used to close the sternum. (credit: P. P. Urone)

The use of X-ray technology in medicine is called radiology—an established and relatively cheap tool in comparison to more sophisticated technologies. Consequently, X-rays are widely available and used extensively in medical diagnostics. During World War I, mobile X-ray units, advocated by Madame Marie Curie, were used to diagnose soldiers.

Because they can have wavelengths less than 0.01 nm, X-rays can be scattered (a process called X-ray diffraction) to detect the shape of molecules and the structure of crystals. X-ray diffraction was crucial to Crick, Watson, and Wilkins in the determination of the shape of the double-helix DNA molecule.

X-rays are also used as a precise tool for trace-metal analysis in X-ray induced fluorescence, in which the energy of the X-ray emissions are related to the specific types of elements and amounts of materials present.

Gamma Rays

Soon after nuclear radioactivity was first detected in 1896, it was found that at least three distinct types of radiation were being emitted. The most penetrating nuclear radiation was called a **gamma ray (γ ray)** (again a name given because its identity and character were unknown), and it was later found to be an extremely high frequency electromagnetic wave.

In fact, γ rays are any electromagnetic radiation emitted by a nucleus. This can be from natural nuclear decay or induced nuclear processes in nuclear reactors and weapons. The lower end of the γ -ray frequency range overlaps the upper end of the X-ray range, but γ rays can have the highest frequency of any electromagnetic radiation.

Gamma rays have characteristics identical to X-rays of the same frequency—they differ only in source. At higher frequencies, γ rays are more penetrating and more damaging to living tissue. They have many of the same uses as X-rays, including cancer therapy. Gamma radiation from radioactive materials is used in nuclear medicine.

[Figure 6.2.1.1.13](#) shows a medical image based on γ rays. Food spoilage can be greatly inhibited by exposing it to large doses of γ radiation, thereby obliterating responsible microorganisms. Damage to food cells through irradiation occurs as well, and the long-term hazards of consuming radiation-preserved food are unknown and controversial for some groups. Both X-ray and γ -ray technologies are also used in scanning luggage at airports.



Figure 6.2.1.1.13: This is an image of the γ rays emitted by nuclei in a compound that is concentrated in the bones and eliminated through the kidneys. Bone cancer is evidenced by nonuniform concentration in similar structures. For example, some ribs are darker than others. (credit: P. P. Urone)

Detecting Electromagnetic Waves from Space

A final note on star gazing. The entire electromagnetic spectrum is used by researchers for investigating stars, space, and time. As noted earlier, Penzias and Wilson detected microwaves to identify the background radiation originating from the Big Bang. Radio telescopes such as the Arecibo Radio Telescope in Puerto Rico and Parkes Observatory in Australia were designed to detect radio waves.

Infrared telescopes need to have their detectors cooled by liquid nitrogen to be able to gather useful signals. Since infrared radiation is predominantly from thermal agitation, if the detectors were not cooled, the vibrations of the molecules in the antenna would be stronger than the signal being collected.

The most famous of these infrared sensitive telescopes is the James Clerk Maxwell Telescope in Hawaii. The earliest telescopes, developed in the seventeenth century, were optical telescopes, collecting visible light. Telescopes in the ultraviolet, X-ray, and γ -ray regions are placed outside the atmosphere on satellites orbiting the Earth.

The Hubble Space Telescope (launched in 1990) gathers ultraviolet radiation as well as visible light. In the X-ray region, there is the Chandra X-ray Observatory (launched in 1999), and in the γ -ray region, there is the new Fermi Gamma-ray Space Telescope (launched in 2008—taking the place of the Compton Gamma Ray Observatory, 1991–2000.).

Section Summary

- The electromagnetic spectrum is separated into many categories and subcategories, based on the frequency and wavelength, source, and uses of the electromagnetic waves.
- Any electromagnetic wave produced by currents in wires is classified as a radio wave, the lowest frequency electromagnetic waves. Radio waves are divided into many types, depending on their applications, ranging up to microwaves at their highest frequencies.
- Infrared radiation lies below visible light in frequency and is produced by thermal motion and the vibration and rotation of atoms and molecules. Infrared's lower frequencies overlap with the highest-frequency microwaves.
- Visible light is largely produced by electronic transitions in atoms and molecules, and is defined as being detectable by the human eye. Its colors vary with frequency, from red at the lowest to violet at the highest.
- Ultraviolet radiation starts with frequencies just above violet in the visible range and is produced primarily by electronic transitions in atoms and molecules.

- X-rays are created in high-voltage discharges and by electron bombardment of metal targets. Their lowest frequencies overlap the ultraviolet range but extend to much higher values, overlapping at the high end with gamma rays.
- Gamma rays are nuclear in origin and are defined to include the highest-frequency electromagnetic radiation of any type.

Glossary

electromagnetic spectrum

the full range of wavelengths or frequencies of electromagnetic radiation

radio waves

electromagnetic waves with wavelengths in the range from 1 mm to 100 km; they are produced by currents in wires and circuits and by astronomical phenomena

microwaves

electromagnetic waves with wavelengths in the range from 1 mm to 1 m; they can be produced by currents in macroscopic circuits and devices

thermal agitation

the thermal motion of atoms and molecules in any object at a temperature above absolute zero, which causes them to emit and absorb radiation

radar

a common application of microwaves. Radar can determine the distance to objects as diverse as clouds and aircraft, as well as determine the speed of a car or the intensity of a rainstorm

infrared radiation (IR)

a region of the electromagnetic spectrum with a frequency range that extends from just below the red region of the visible light spectrum up to the microwave region, or from 0.74 μm to 300 μm

ultraviolet radiation (UV)

electromagnetic radiation in the range extending upward in frequency from violet light and overlapping with the lowest X-ray frequencies, with wavelengths from 400 nm down to about 10 nm

visible light

the narrow segment of the electromagnetic spectrum to which the normal human eye responds

amplitude modulation (AM)

a method for placing information on electromagnetic waves by modulating the amplitude of a carrier wave with an audio signal, resulting in a wave with constant frequency but varying amplitude

extremely low frequency (ELF)

electromagnetic radiation with wavelengths usually in the range of 0 to 300 Hz, but also about 1kHz

carrier wave

an electromagnetic wave that carries a signal by modulation of its amplitude or frequency

frequency modulation (FM)

a method of placing information on electromagnetic waves by modulating the frequency of a carrier wave with an audio signal, producing a wave of constant amplitude but varying frequency

TV

video and audio signals broadcast on electromagnetic waves

very high frequency (VHF)

TV channels utilizing frequencies in the two ranges of 54 to 88 MHz and 174 to 222 MHz

ultra-high frequency (UHF)

TV channels in an even higher frequency range than VHF, of 470 to 1000 MHz

X-ray

invisible, penetrating form of very high frequency electromagnetic radiation, overlapping both the ultraviolet range and the γ -ray range

gamma ray

(γ ray); extremely high frequency electromagnetic radiation emitted by the nucleus of an atom, either from natural nuclear decay or induced nuclear processes in nuclear reactors and weapons. The lower end of the γ -ray frequency range overlaps the upper end of the X-ray range, but γ rays can have the highest frequency of any electromagnetic radiation

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6.2.1.2: Reflection

Learning Objectives

- Explain reflection of light from polished and rough surfaces.

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 6.2.1.2.1](#), 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 6.2.1.2.2](#) 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 6.2.1.2.3](#). 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 6.2.1.2.4](#). When the moon reflects from a lake, as shown in [Figure 6.2.1.2.5](#), a combination of these effects takes place.

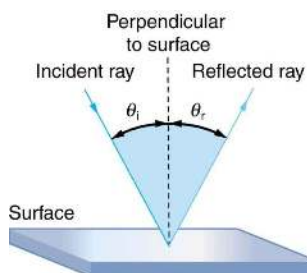


Figure 6.2.1.2.1: The law of reflection states that the angle of reflection equals the angle of incidence — $\theta_r = \theta_i$. The angles are measured relative to the perpendicular to the surface at the point where the ray strikes the surface.

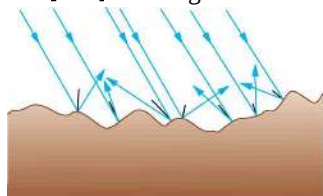


Figure 6.2.1.2.2: Light is diffused when it reflects from a rough surface. Here many parallel rays are

incident, but they are reflected at many different angles since the surface is rough.

When a sheet of paper is illuminated with many parallel incident rays, it can be seen at many different angles, because its surface is

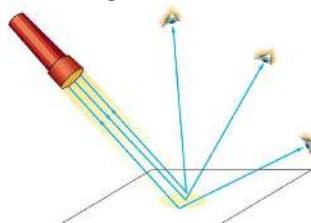


Figure 6.2.1.2.3:

rough and diffuses the light.

When a sheet of paper is illuminated with many parallel incident rays, it can be seen at many different angles, because its surface is

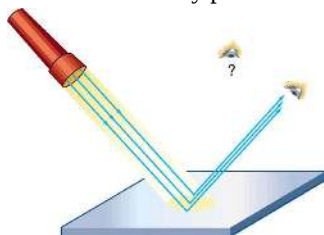


Figure 6.2.1.2.4: A mirror illuminated by many parallel rays reflects them in only one direction, since its surface is very smooth. Only the observer at a particular angle will see the reflected light.



Figure 6.2.1.2.5: Moonlight is spread out when it is reflected by the lake, since the surface is shiny but uneven. (credit: Diego Torres Silvestre, Flickr)

The law of reflection is very simple: The angle of reflection equals the angle of incidence.

THE LAW OF REFLECTION

The angle of reflection equals the angle of incidence.

When we see ourselves in a mirror, it appears that our image is actually behind the mirror. This is illustrated in [Figure 6.2.1.2.6](#). We see the light coming from a direction determined by the law of reflection. The angles are such that our image is exactly the same distance behind the mirror as we stand away from the mirror. If the mirror is on the wall of a room, the images in it are all behind the mirror, which can make the room seem bigger. Although these mirror images make objects appear to be where they cannot be (like behind a solid wall), the images are not figments of our imagination. Mirror images can be photographed and videotaped by instruments and look just as they do with our eyes (optical instruments themselves). The precise manner in which images are formed by mirrors and lenses will be treated in later sections of this chapter.

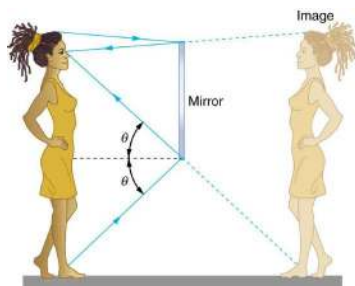


Figure 6.2.1.2.6: Our image in a mirror is behind the mirror. The two rays shown are those that strike the mirror at just the correct angles to be reflected into the eyes of the person. The image appears to be in the direction the rays are coming from when they enter the eyes.

TAKE-HOME EXPERIMENT: LAW OF REFLECTION

Take a piece of paper and shine a flashlight at an angle at the paper, as shown in [Figure 6.2.1.2.3](#). Now shine the flashlight at a mirror at an angle. Do your observations confirm the predictions in [Figure 6.2.1.2.3](#) and [Figure 6.2.1.2.4](#)? Shine the flashlight on various surfaces and determine whether the reflected light is diffuse or not. You can choose a shiny metallic lid of a pot or your skin. Using the mirror and flashlight, can you confirm the law of reflection? You will need to draw lines on a piece of paper showing the incident and reflected rays. (This part works even better if you use a laser pencil.)

Section Summary

- The angle of reflection equals the angle of incidence.
- A mirror has a smooth surface and reflects light at specific angles.
- Light is diffused when it reflects from a rough surface.
- Mirror images can be photographed and videotaped by instruments.

Glossary

mirror

smooth surface that reflects light at specific angles, forming an image of the person or object in front of it

law of reflection

angle of reflection equals the angle of incidence

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6.2.1.3: Refraction

Learning Objectives

- Determine the index of refraction, given the speed of light in a medium.

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 6.2.1.3.1.](#)) 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. 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.

Definition: REFRACTION

The changing of a light ray's direction (loosely called bending) when it passes through variations in matter is called refraction.

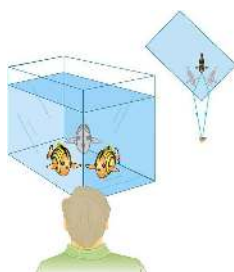


Figure 6.2.1.3.1: Looking at the fish tank as shown, we can see the same fish in two different locations, because light changes directions when it passes from water to air. In this case, the light can reach the observer by two different paths, and so the fish seems to be in two different places. This bending of light is called refraction and is responsible for many optical phenomena.

Why does light change direction when passing from one material (medium) to another? It is because light changes speed when going from one material to another. So before we study the law of refraction, it is useful to discuss the speed of light and how it varies in different media.

The Speed of Light

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. Roemer had noted that the average orbital period of one of Jupiter's moons, as measured from Earth, varied depending on whether Earth was moving toward or away from Jupiter. He correctly concluded that the apparent change in period was due to the change in distance between Earth and Jupiter and the time it took light to travel this distance. From his 1676 data, a value of the speed of light was calculated to be $2.26 \times 10^8 \text{ m/s}$ (only 25% different from today's accepted value). In more recent times, physicists have measured the speed of light in numerous ways and with increasing accuracy. One particularly direct method, used in 1887 by the American physicist Albert Michelson (1852–1931), is illustrated in [Figure 6.2.1.3.2.](#) Light reflected from a rotating set of mirrors was reflected from a stationary mirror 35 km away and returned to the rotating mirrors. The time for the light to travel can be determined by how fast the mirrors must rotate for the light to be returned to the observer's eye.

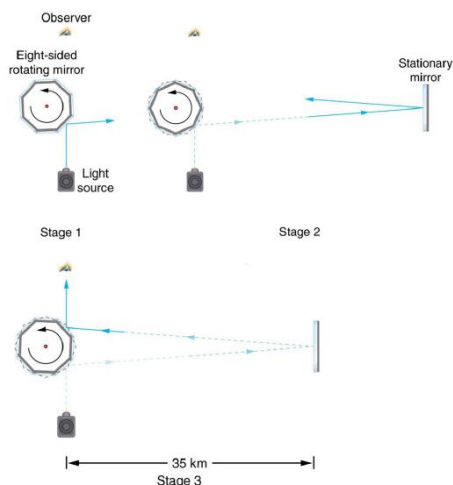


Figure 6.2.1.3.2: A schematic of early apparatus used by Michelson and others to determine the speed of light. As the mirrors rotate, the reflected ray is only briefly directed at the stationary mirror. The returning ray will be reflected into the observer's eye only if the next mirror has rotated into the correct position just as the ray returns. By measuring the correct rotation rate, the time for the round trip can be measured and the speed of light calculated. Michelson's calculated value of the speed of light was only 0.04% different from the value used today.

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

$$c = 2.9972458 \times 10^8 \text{ m/s} \approx 3.00 \times 10^8 \text{ m/s},$$

where the approximate value of $3.00 \times 10^8 \text{ m/s}$ is used whenever three-digit accuracy is sufficient. The speed of light through matter is less than it is in a vacuum, because light interacts with atoms in a material. The speed of light depends strongly on the type of material, since its interaction with different atoms, crystal lattices, and other substructures varies. We define the **index of refraction** n of a material to be

$$n = \frac{c}{v},$$

where v is the observed speed of light in the material. Since the speed of light is always less than c in matter and equals c only in a vacuum, the index of refraction is always greater than or equal to one.

VALUE OF THE SPEED OF LIGHT

$$c = 2.9972458 \times 10^8 \text{ m/s} \approx 3.00 \times 10^8 \text{ m/s}$$

Definition: INDEX OF REFRACTION

$$n = \frac{c}{v}$$

That is, $n \geq 1$. Table 6.2.1.3.1 gives the indices of refraction for some representative substances. The values are listed for a particular wavelength of light, because they vary slightly with wavelength. (This can have important effects, such as colors produced by a prism.) Note that for gases, n is close to 1.0. This seems reasonable, since atoms in gases are widely separated and light travels at c in the vacuum between atoms. It is common to take $n = 1$ for gases unless great precision is needed. Although the speed of light v in a medium varies considerably from its value c in a vacuum, it is still a large speed.

Table 6.2.1.3.1: Index of Refraction in Various Media

Medium	n
Gases at 0°C, 1 atm	
Air	1.000293
Carbon dioxide	1.00045
Hydrogen	1.000139
Oxygen	1.000271
Liquids at 20°C	
Benzene	1.501
Carbon disulfide	1.628
Carbon tetrachloride	1.461
Ethanol	1.361
Glycerine	1.473
Water, fresh	1.333
Solids at 20°C	
Diamond	2.419
Fluorite	1.434
Glass, crown	1.52
Glass, flint	1.66
Ice at 20°C	1.309
Polystyrene	1.49
Plexiglas	1.51
Quartz, crystalline	1.544
Quartz, fused	1.458
Sodium chloride	1.544
Zircon	1.923

Example 6.2.1.3.1: Speed of Light in Matter

Calculate the speed of light in zircon, a material used in jewelry to imitate diamond.

Strategy

The speed of light in a material, v , can be calculated from the index of refraction n of the material using the equation $n = c/v$.

Solution

The equation for index of refraction states that $n = c/v$. Rearranging this to determine v gives

$$v = \frac{c}{n}$$

The index of refraction for zircon is given as 1.923 in [Table 6.2.1.3.1](#), and c is given in the equation for speed of light. Entering these values in the last expression gives

$$v = \frac{3.00 \times 10^8 \text{ m/s}}{1.923} = 1.56 \times 10^8 \text{ m/s}.$$

Discussion

This speed is slightly larger than half the speed of light in a vacuum and is still high compared with speeds we normally experience. The only substance listed in Table 6.2.1.3.1 that has a greater index of refraction than zircon is diamond. We shall see later that the large index of refraction for zircon makes it sparkle more than glass, but less than diamond.

Law of Refraction

Figure 6.2.1.3.3 shows how a ray of light changes direction when it passes from one medium to another. As before, the angles are measured relative to a perpendicular to the surface at the point where the light ray crosses it. (Some of the incident light will be reflected from the surface, but for now we will concentrate on the light that is transmitted.) The change in direction of the light ray depends on how the speed of light changes. The change in the speed of light is related to the indices of refraction of the media involved. In the situations shown in Figure 6.2.1.3.3, medium 2 has a greater index of refraction than medium 1. This means that the speed of light is less in medium 2 than in medium 1. Note that as shown in Figure 6.2.1.3.3(a), the direction of the ray moves closer to the perpendicular when it slows down. Conversely, as shown in Figure 6.2.1.3.3(b), the direction of the ray moves away from the perpendicular when it speeds up. The path is exactly reversible. In both cases, you can imagine what happens by thinking about pushing a lawn mower from a footpath onto grass, and vice versa. Going from the footpath to grass, the front wheels are slowed and pulled to the side as shown. This is the same change in direction as for light when it goes from a fast medium to a slow one. When going from the grass to the footpath, the front wheels can move faster and the mower changes direction as shown. This, too, is the same change in direction as for light going from slow to fast.

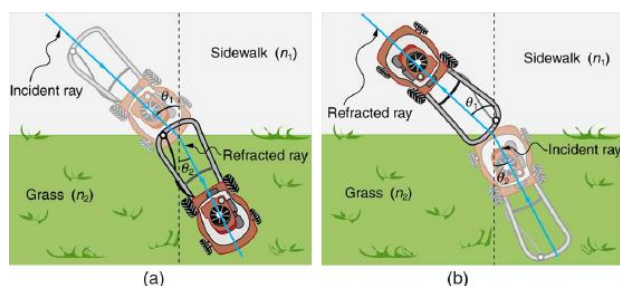


Figure 6.2.1.3.3: The change in direction of a light ray depends on how

the speed of light changes when it crosses from one medium to another. The speed of light is greater in medium 1 than in medium 2 in the situations shown here. (a) A ray of light moves closer to the perpendicular when it slows down. This is analogous to what happens when a lawn mower goes from a footpath to grass. (b) A ray of light moves away from the perpendicular when it speeds up. This is analogous to what happens when a lawn mower goes from grass to footpath. The paths are exactly reversible.

The amount that a light ray changes its direction depends both on the incident angle and the amount that the speed changes. For a ray at a given incident angle, a large change in speed causes a large change in direction, and thus a large change in angle.

TAKE-HOME EXPERIMENT: A BROKEN PENCIL

A classic observation of refraction occurs when a pencil is placed in a glass half filled with water. Do this and observe the shape of the pencil when you look at the pencil sideways, that is, through air, glass, water. Explain your observations. Draw ray diagrams for the situation.

Section Summary

- The changing of a light ray's direction when it passes through variations in matter is called refraction.
- The speed of light in vacuum $c = 2.9972458 \times 10^8 \text{ m/s} \approx 3.00 \times 10^8 \text{ m/s}$
- Index of refraction $n = \frac{c}{v}$, where v is the speed of light in the material, c is the speed of light in vacuum, and n is the index of refraction.

Glossary

refraction

changing of a light ray's direction when it passes through variations in matter

index of refraction

for a material, the ratio of the speed of light in vacuum to that in the material

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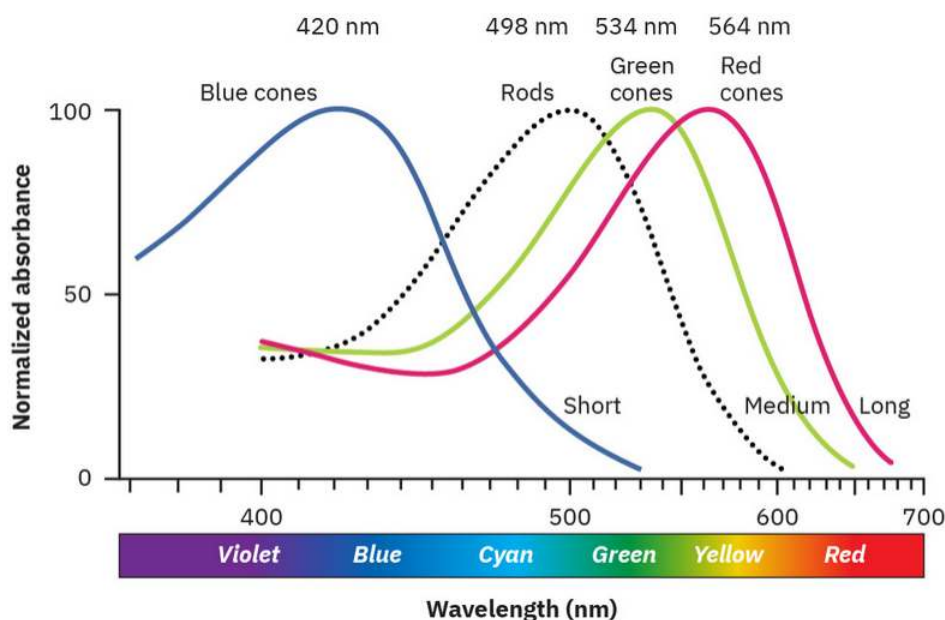
6.2.2: Color Perception

We've already discussed the wavelengths and frequencies of electromagnetic radiation and we've spoken of the narrow strip of frequencies that we can perceive in terms of colors. We can see the effect of a prism on beam of light as it splits into bands of color. But what does color mean? How is it that we can perceive wavelength differences as small as fifty-billionths of a meter? We've already explored sound and how it is perceived by us in different circumstances, now it is time to turn our attention to the visual.

For this portion of the textbook, we turn to human biology. When gathering data for experimental analysis it is important to understand the equipment being used, so you know how to interpret the data you've collected. Given that most of the information that we gather about the world comes to us through our eyes, understanding vision helps us better understand the data we collect about our world. As with most of this textbook, the treatment we give this topic is necessarily brief, as it encompasses biology, chemistry and psychology. What follows is a rudimentary discussion of the subject matter.

Light entering our eyes is focused on the retina. The retina is a collection of cells lining the eyeball that gather information from the light and send it to the occipital lobe of our brain via the optic nerve. There are two main types of *photoreceptors* (cells that sense light) in the retina:

- **Cones:** A cell concentrated near the center of the retina that work best in bright light and control high resolution vision. There are three types of cones and each type can detect a certain range of frequencies, peaking at a particular wavelength. The retina contains about six million cones.
- **Rods:** A cell that is far more sensitive to small energy differences than cones. They are essential for motion detection and for vision in dark environments. They do not record information about color, but provide vision in grey scale. They outnumber cones by a factor of twenty to one.



Plot of cone wavelengths taken from [OpenStax College Physics 2e](#) and is licensed under CC-BY

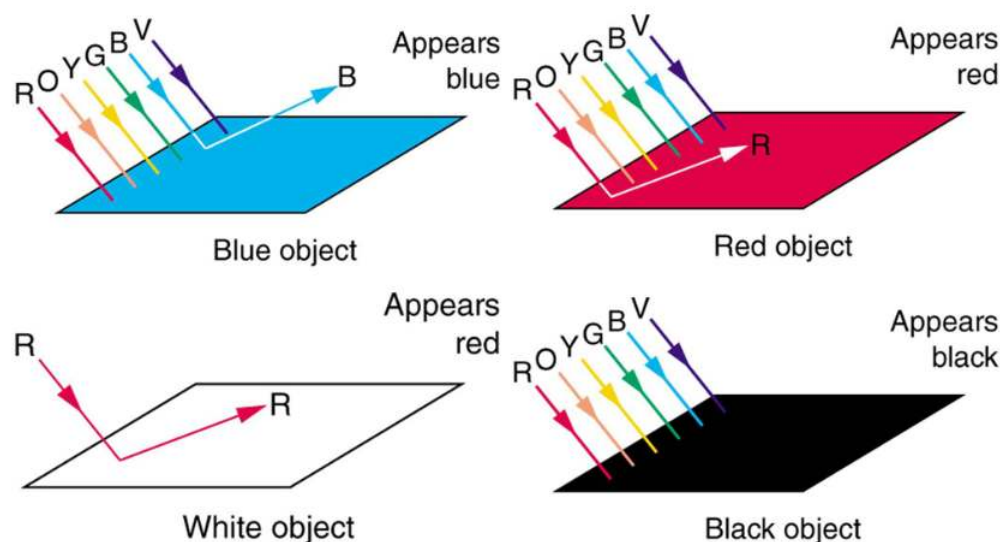
When light falls on the retina, certain cones respond to the wavelength of the incoming electromagnetic wave. The strength of the response has to do with the brightness (intensity) of the wave. In the simplest model of color vision, there are three primary colors that correspond to the red, blue and green cones. The colors we perceive depends on which cones are activated and the intensity of the cone activation.

Red and green cones can produce the perception of yellow if the intensities are properly balanced. If all three cone types are illuminated, we perceive white light. If blue and red cones are illuminated in one way, we perceive a bluish purple, another set of illuminations gives us a reddish purple. There is some evidence that our color vision is not quite as simple as this model would suggest, but the fundamental relationship between perceived color, cones and illumination levels is well established.

True Colors

In order for our eyes to see anything, light waves have to reach the object, be reflected from it and then reach our eyes where the rods and cones can then be activated. Opaque objects appear to be the colors we perceive them to be because of the light they reflect. White light is a mixture of all the visible colors of light, so if an object reflects all visible colors then it appears white to us. If an object is illuminated by white light and only reflects blue light (absorbing the rest of the colors) then the object appears blue. Similarly, if an object is illuminated by white light but only reflects red light, then it appears to be red. The *true color* of an object is defined to be the color perceived when illuminated by white light. A truly black object is one that absorbs all colors and reflects none.

However, if a white object is illuminated by red light, it will reflect that red light and appear red. If it is illuminated by red and green light, it could appear yellow. A red object illuminated by blue light will not reflect the blue light, and will appear black. Perceived colors can differ from true colors based on the color of the light that is illuminating the object.



Apparent colors taken from [OpenStax\(opens in new window\) College Physics 2e\(opens in new window\)](#) and is licensed under CC-BY

Light sources are assigned a color based on the frequencies (wavelengths) they emit. A neon sign appears red because it primarily emits red frequencies of light, while the Sun appears yellowish because the intensities of the yellow frequencies are greater than any of the other colors emitted.

Modified Model of Color Vision

Experiments have shown that the simple model presented above doesn't account for the full range of human color vision. It is known that there are nerve connections between rods and cones in the retina, and the optic nerve that carries information to the brain has fewer connections to the brain than there are rods and cones. This would imply that there is some sort of pre-processing that happens on the retina, before the information is sent to the occipital lobe. It is thought that the eye may make comparisons to information collected by adjacent receptors. If both receptors detect the same information, the redundant signals may be ignored. If the receptors report differences, additional processing power is allocated. This might help explain our sensitivity to detecting edges.

The founder of the Polaroid corporation, Edwin Land, proposed a model in which the three types of cones are organized into systems he called *retinexes*. Each retinex forms an image which is compared to the images formed by other retinexes. This may explain a visual ability called *color constancy*. Color constancy describes the ability of the eye to determine the color of an object under different ambient lighting conditions. It is possible to determine that a piece of paper is red whether it is illuminated by sunlight, candlelight or fluorescent light, even though the emitted color of each of these light sources is quite different. By comparing the paper to its surrounding in each case, the true color of the paper can be determined.

Our color vision is not completely understood, and it would appear that vision is far more complex than our initial model might suggest.

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