Introductory Chemistry

MONTGOMERY
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Content

This textbook is meant for students with very little or no background in chemistry. Concepts here are presented at a very basic level in order to provide the foundation for college-level general chemistry. This book is used in Introductory Chemistry, a non-credit, preparatory course at Montgomery College (Maryland).

If the reader is interested in an in-depth, college-level, general chemistry textbook, we highly recommend OpenStax Chemistry (Download for free at https://openstax.org/details/books/chemistry).

Contact Information

Questions and comments are welcome. Please contact us at chemistry@montgomerycollege.edu.

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Chapter 1
Introduction to Chemistry

1.1 The Science of Chemistry
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LEARNING OBJECTIVES

1. Describe the field of chemistry and its relationship to other scientific fields.
2. Describe the scientific method.
3. Differentiate among hypotheses, scientific laws, and scientific theories.
4. Create a strategy that will help you to succeed in your study of chemistry.

What is chemistry? Chemistry is the study of matter, its changes, and the energy involved in such changes. Chemistry is sometimes referred to as "the central science" due to its interconnectedness with a vast array of disciplines in science, technology and engineering. Chemistry and its subdisciplines play vital roles in biology, medicine, materials science, forensics, environmental science, and many other fields (Figure 1.1).

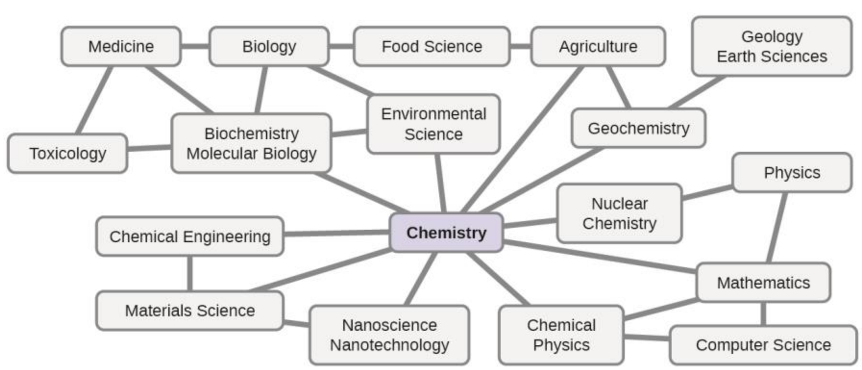

Figure 1.1 Knowledge of chemistry is central to understanding a wide range of scientific disciplines. This diagram shows just a few of the relationships between chemistry and other scientific fields. *Image Credit: OpenStax Chemistry*

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Scientific disciplines, or branches of science, involve studies of the natural universe. Mathematics, computer science, and information theory provide important tools that help us calculate, interpret, describe, and generally make sense of the sciences. Biology and chemistry converge in biochemistry, which is crucial to understanding the many complex factors and processes that keep living organisms alive. Chemical engineering, materials science, and nanotechnology combine chemical principles and empirical findings to produce useful substances, ranging from gasoline to fabrics to electronics. Agriculture, food science, veterinary science, and brewing and wine making help provide sustenance in the form of food and drink to the world's population. Medicine, pharmacology, biotechnology, and botany identify and produce substances that help keep us healthy. Environmental science, geology, oceanography, and atmospheric science incorporate many chemical ideas to help us better understand and protect our physical world. Chemical ideas are used to help understand the universe in astronomy and cosmology. A strong understanding of chemistry useful and essential for competence in every field of science

1.1 The Science of Chemistry

Chemistry is the branch of science dealing with the structure, composition, properties, and reactive characteristics of matter. In chemistry, **matter** is anything that has mass and occupies space. Thus, chemistry is the study of everything around us – the liquids that we drink, the gases we breathe, the composition of everything from the plastic case on your phone to the earth beneath your feet. More importantly, chemistry is the study of the **transformation** of matter. Crude oil is transformed into more useful petroleum products such as gasoline and kerosene by the process of refining. Crude metal ores are transformed into metals that can then be fashioned into everything from foil to automobiles. Potential drugs are identified from natural sources, isolated, and then modified to produce the pharmaceuticals that have led to advances in modern medicine. Chemistry is at the center of all of these processes and chemists are the people that study the nature of matter and learn to design, predict, and control these chemical transformations.

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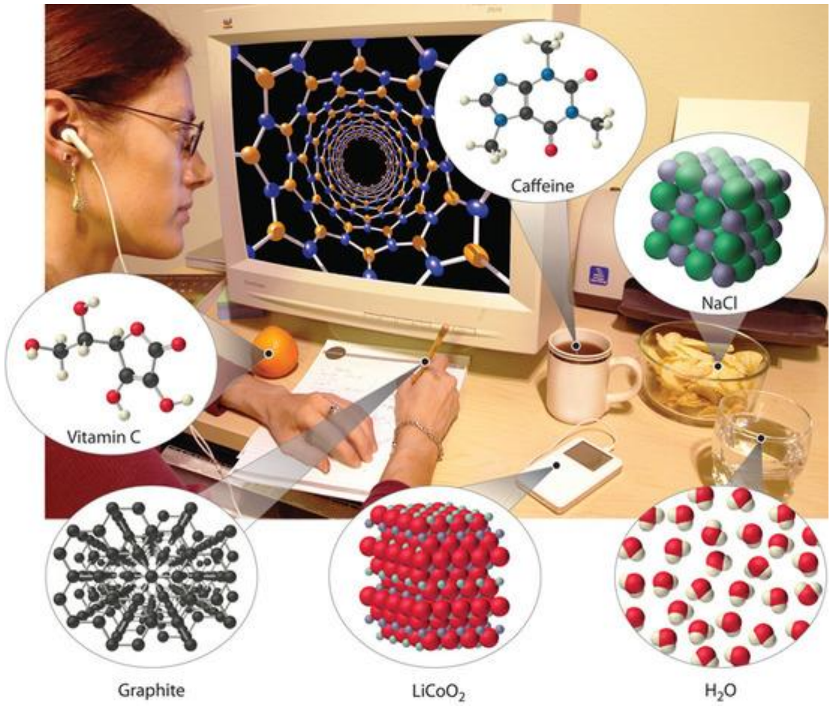

Figure 1.2 Chemistry is all around us. Even people not intending to become chemists can benefit from understanding its basic principles. *Image Credit: LibreTexts Chemistry*.

Examples of the practical applications of chemistry are everywhere (Figure 1.2). Engineers need to understand the chemical properties of the substances when designing biologically compatible implants for joint replacements or designing roads, bridges, buildings, and nuclear reactors that do not collapse because of weakened structural materials such as steel and cement. Archaeology and paleontology rely on chemical techniques to date bones and artifacts and identify their origins. Although law is not normally considered a field related to chemistry, forensic scientists use chemical methods to analyze blood, fibers, and other evidence as they investigate crimes. In particular, DNA matching—comparing biological samples of genetic material to see whether they could have come from the same person—has been used to solve many high-profile criminal cases as well as clear innocent people who have been wrongly accused or convicted. Forensics is a rapidly growing area of applied chemistry. In addition, the proliferation of chemical and biochemical innovations in industry is producing rapid growth in the area of patent law. Ultimately, the dispersal of information in all the fields in which chemistry plays a part requires experts who are able to explain complex chemical issues to the public through television, print journalism, the Internet, and popular books.

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1.2 The Scientific Method

Chemists, like all scientists, search for answers to questions and solutions to problems by using the **scientific method**. This procedure consists of making observations, formulating hypotheses, and conducting experiments in repeated cycles (Figure 1.3).

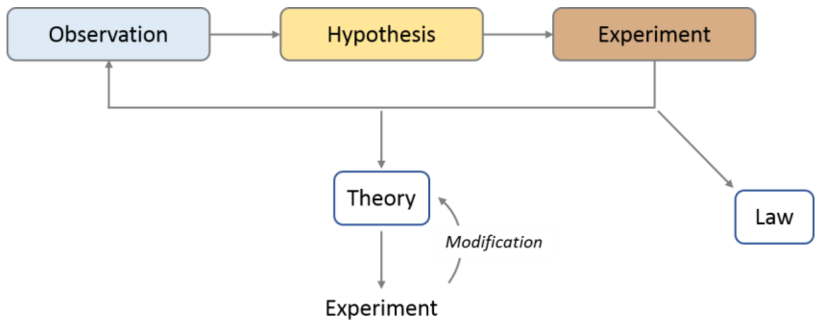

Figure 1.3 The Scientific Method. *Adapted from: Principles of General Chemistry*.

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, and sulfur crystals are yellow. **Quantitative** observations are measurements. Examples of quantitative observations include the following: the melting point of crystalline sulfur is 115.21 °C, and 35.9 grams of sodium chloride dissolve in 100 grams of water at 20 °C.

To begin an investigation of a set of observations or questions, a scientist formulates a **hypothesis**, an educated guess or tentative explanation of a natural phenomenon. Most good hypotheses are grounded in previously understood knowledge and must be testable. An untestable hypothesis is simply a guess and cannot be used as the basis for an experiment.

A scientist devises **experiments** to test if the hypothesis is or is not correct. Experiments must be carefully designed, properly executed, and reproducible. Results of an experiment might be used to amend a hypothesis, resulting in further rounds of experiments.

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Example 1.1

Consider the following statements and decide whether or not each is a hypothesis. If the statement is not a hypothesis, how would you rewrite it?

(a) Studying chemistry each night will make you a better student.

(b) The atmosphere, or coma, surrounding a comet contains water that has vaporized from the ice on the comet surface.

(c) Drinking coffee in the afternoon is bad for sleep.

Answers

(a) This is not a hypothesis and can be restated as: Studying chemistry for 1.5 hours every night will result in a score of at least 70% on the final exam.

(b) This is a statement that can be tested and is therefore a hypothesis.

(c) This is not a hypothesis and can be restated as: An 8 oz. cup of coffee in the late afternoon causes an increase in the amount of time it takes to fall asleep at night.

Example 1.2

Identify each of the following statements as either a qualitative description or a quantitative description.

(a) Gold metal is shiny and yellow.

(b) A ream of paper contains 500 sheets.

(c) The weather outside is hot and humid.

(d) The temperature outside is 85 °F.

Answers

(a) This statement describes the appearance of a piece of gold and is qualitative.

(b) This statement mentions a specific amount and is quantitative.

(c) Hot and humid are descriptions of the weather and the statement is qualitative.

(d) The temperature is given as a specific, measured quantity and is quantitative.

1.3 Scientific Laws and Theories

When experimental data are sufficiently reproducible, they are sometimes summarized in a **law**, a verbal or mathematical description of a phenomenon that allows for general predictions. One

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example, the **law of definite proportions**, was formulated by Joseph Proust (1754–1826) and 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. Scientific laws are unlikely to change greatly over time unless a major experimental error is discovered. A scientific theory is reliable and rigorous and has been tested and confirmed according to the scientific method. A scientific theory may change over time as new discoveries are made.

Example 1.3

Identify each of the following statements as either a law, a theory, a hypothesis, an experiment, or a set of observations.

(a) Ice always floats on liquid water.
(b) Birds evolved from dinosaurs.
(c) Hot air is less dense than cold air because hot gases expand when kept at constant pressure.
(d) When ten grams of ice were added to 100 mL of water at 25.0 °C, the temperature of the water decreased to 15.5 °C after the ice melted.
(e) The ingredients of Ivory soap were analyzed to see whether it is really 99.44% pure as advertised.

Answers

(a) This is a general statement of a relationship between the properties of solid and liquid water; it is a scientific law.
(b) This is a possible explanation for the origin of birds, so it is a hypothesis.
(c) This is a statement that explains the relationship between the temperature and the density of air based upon fundamental principles, so it is a scientific theory.
(d) The temperature is measured before and after a change is made to the system, so these are observations.
(e) This is an analysis designed to test a hypothesis (the manufacturer's claim of purity), so it is an experiment.

1.4 How to Succeed in Chemistry

Although the study of chemistry is challenging and complex, everyone has the potential to succeed with *hard work* and *perseverance.* There is no such thing as being bad at chemistry! If you have struggled to be successful in the past, now is your chance to excel and learn more than you ever thought you could. Keep in mind the following study tips:

1. Be a self-motivated learner! Why are you taking chemistry? How will it help you to achieve your educational goals? Are you willing to put in the time and effort to succeed in the course?

2. Find a quiet place to study. Turn off your phone, get away from distractions, and focus.

3. Read *before* attending class and read actively. Take notes as you read the material and outline each section of the text. *Simply highlighting text will not work*. Always stop at example problems and test yourself. Work out the problems with pencil and paper before looking at the solutions.

4. Always attend lecture. No excuses, no exceptions. It's a great idea to arrive early, get settled, and use the extra time to review notes and ask questions.

5. Engage during lecture. Sit at the front of the room with a pencil and paper ready for note taking, Have your calculator out (no phones). Use your calculator to work example problems along with your professor to make sure your keystrokes are correct. If you get an answer that is different from your professor's answer, raise your hand and ask about it.

6. Ask questions. If you have a question, ask it! Many other students will have exactly the same question and will be grateful to the person who speaks up!

7. Review and re-write your notes *right after class*. Go to a quiet study location, get a fresh sheet of paper, and re-work lecture example problems on your own. Only look at the answers in your class notes when you are finished. Note topics that are confusing and visit your professor or the learning center as soon as possible to get clarification.

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8. Set aside time *every single day* to practice what you've learned and work through problems. The number one reason why students fail chemistry is that they don't work enough problems. If you want extra practice, work through *all* the problems at the end of each chapter. If you want even more practice check out a different chemistry textbook from the learning center and work through its end of chapter problems.

9. A three-credit course requires about *nine hours of work each week* outside of class. Make sure that you schedule 1-2 hours of intensive, uninterrupted chemistry study *every day*. Chemistry is cumulative by nature so falling behind by even one lecture can be disastrous.

10. Form a study group. Working with other students makes chemistry more fun! Explain topics to each other. Ask each other questions. Try to make up your own quiz and exam questions and challenge each other to solve them.

11. A wrong answer is okay! Making mistakes is the best way to learn. Don't be afraid to make mistakes on an example problem or a homework problem. Don't let a challenging problem shut you down. *Attack the problem – don't let the problem attack you.*

12. Constantly ask yourself "Have I learned?" Your goal for every study session should be to learn a topic so deeply that you are able to explain it to another person. Going into a study session with a checklist of to-do items is not enough; you must reflect on the items that you completed and ask if you have really *learned* the material. Find a classmate and try to teach the material. You can even visit a tutor and request that they listen as you teach the topic.

This list is just as important as your periodic table for your success in this course. Keep coming back to these points during the semester to remind yourself how to study.

Video: How to Study Chemistry
https://youtu.be/v1uF3XB-9uc

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Chapter 1 Practice Problems

1.2 The Scientific Method

1. Which of the following fields are considered branches of science?
   (a) biophysics (b) art (c) business (d) astronomy

2. Which of the following fields are considered branches of science?
   (a) philosophy (b) accounting (c) anatomy (d) geochemistry

3. In your own words, describe the scientific method.

4. What is the scientific definition of a *hypothesis*? Why is a hypothesis more than just an "educated guess"?

5. Why do scientists need to perform experiments?

6. What are some characteristics of good scientific experiments?

7. Identify each of the following statements as either a qualitative description or a quantitative description.
   (a) The deepest part of the Pacific Ocean is the Mariana Trench.
   (b) The International Space Station completes 15.54 orbits around Earth each day.
   (c) Amethyst is a purple mineral.
   (d) A sample of amethyst contains 47% silicon by mass.

8. Identify each of the following statements as either a qualitative description or a quantitative description.
   (a) X-rays used in many chemistry experiments have a wavelength of 0.154 nanometers.
   (b) Gamma rays have very high energy.
   (c) Boiling water gives off a large amount of heat to the surrounding room.
   (d) Ethanol freezes at −114.1 °C.

1.3 Scientific Laws and Theories

9. What is the scientific definition of a *theory*?

10. What is the scientific definition of a *law*?

11. Identify each of the following statements as either a law, a theory, a hypothesis, an experiment, or a set of observations.
    (a) Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."

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(b) Heat always flows from hot objects to cooler ones, not in the opposite direction.
(c) The universe was formed by a massive explosion that propelled matter into a vacuum.
(d) Michael Jordan is the greatest pure shooter ever to play professional basketball.
(e) Limestone is insoluble in water but dissolves readily in dilute acid.

12. Identify each of the following statements as either a law, a theory, a hypothesis, or an experiment.
    (a) The pressure of a gas sample is inversely proportional to its volume at constant temperature.
    (b) The amount of carbon dioxide in the atmosphere is measured by voltage output of a gas analyzer situated near the peak of Mauna Loa at 3400 m above sea level.
    (c) An object in motion will remain in motion unless acted upon by a force.
    (d) Life forms can emerge and survive as a result of natural processes.
    (e) Indium reacts with hydrogen sulfide to form a bright orange powder.

1.4 How to Succeed in Chemistry

13. Create a study schedule for the next week. Be sure to account for each hour of your day, starting when you wake up and ending when you go to bed. Include hours spent at work, in other classes, caring for family members, or other obligations. How much time does this leave for studying chemistry?

14. Write five specific tasks that you are going to complete during your chemistry study time each week.

15. An excellent way to prepare for an exam is to organize your notes to connect all the ideas in a chapter. One way of doing so is by creating a word web. Using Figure 1.1 as an example, create a word web for each of the bold-face words in this chapter. Then expand outward by giving an example or definition for each word.

16. Reflect on your learning! Were you able to complete all of the practice problems for this chapter without looking at the solutions? Go back and rework difficult problems until you are confident with your answers.

17. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

18. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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Chapter 2
Measurements

LEARNING OBJECTIVES

1. Express numbers representing measurements.
2. Express measurements using scientific notation and units.
3. Represent the limits of a measurement by using significant figures.
4. Limit the results of calculations to the appropriate number of significant figures.
5. Use dimensional analysis to convert from one set of units to another.
6. Understand density and use it as a conversion factor.
7. Understand and convert between the various temperature scales used in chemistry.

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In 1999, the Mars Climate Orbiter spacecraft was lost attempting to orbit Mars because the thrusters were programmed in terms of English units, even though the engineers built the spacecraft using metric units. In 1993, a nurse mistakenly administered 23 units of morphine to a patient rather than the “2–3” units prescribed. (The patient ultimately survived.) In 1983, an Air Canada airplane had to make an emergency landing because it unexpectedly ran out of fuel; ground personnel had filled the fuel tanks with a certain number of pounds of fuel, not kilograms of fuel. These incidents occurred because people weren’t paying attention to the types of *measurement units* used.

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

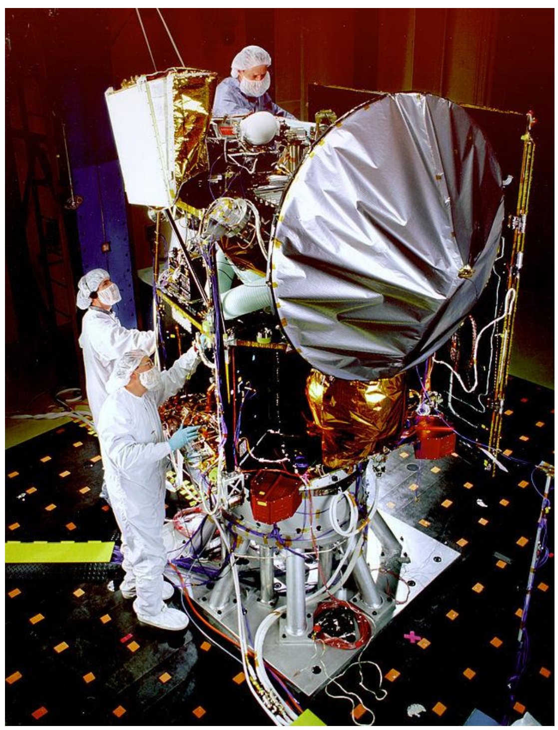

Figure 2.1 The Mars Climate Orbiter was launched in December 1998 and was supposed to determine the water distribution on Mars. The orbiter was lost due to a navigation error caused by a failure to translate the English units of pound-force-seconds to metric units of newton-seconds. Last contact was on September 23, 1999, and it is thought that the orbiter burned up in Mars’ atmosphere shortly thereafter. *Image Credit: NASA*

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2.1 Introduction to Measurements

Measurements provide the macroscopic information that is the basis for most of the hypotheses, theories, and laws that describe the behavior of matter. Every measurement provides three kinds of information:

1. the magnitude expressed as a number;
2. a standard of comparison included as units; and
3. a representation of the precision, or uncertainty.

The number included in a measurement indicates magnitude, but that number is meaningless without a standard of comparison, the units. For instance, the distance between Washington DC and New York City can be expressed as 230. The meaning of this number is not clear. Is it 230 kilometers? 230 blocks? 230 miles? In order to express the distance in a meaningful way, units must be included. The distance between Washington DC and New York City is *230 miles*.

2.2 Numbers and Notation

Numbers used in scientific measurements are expressed in either **standard notation** or **scientific notation**. Standard notation is the straightforward expression of a number, including any decimal places required. Numbers such as 17, 101.5, and 0.00446 are expressed in standard notation. For numbers fairly close to 1, standard notation works well. For very large numbers such as 306,000,000, or for very small numbers, such as 0.000000419, standard notation can be difficult to interpret because there are so many zeroes, or places to count.

For very large or very small numbers, scientific notation is used. Scientific notation is an expression of a number using powers of 10 and is used to abbreviate numbers having an extreme number of digits or decimal places. Suppose you need to communicate the number of drops of water there are in a river 12 kilometers long, 270 meters wide, and 38 meters deep. If a drop of water is a fraction of a milliliter, the answer is 123,120,000,000,000 drops of water; a number difficult to interpret quickly because the zeroes must be counted. In scientific notation, the number is 1.2312×10^{14}

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drops of water. When the number is expressed in scientific notation, the magnitude of that number is clear right away. Additionally, most of the digits in 123,120,000,000,000 are meaningless because the measurement is a very rough estimate. If we had given the answer the correct number of significant figures, the answer would have been 1.2 x 10^{14} drops of water. Table 2.1 gives several examples of measurements represented in scientific notation.

Table 2.1 Measurements in Scientific Notation

Object	Measurement	Standard Notation	Scientific Notation
human hair	width	0.000060 meters	6.0 x 10^{-5} m
hemoglobin	mass	0.00000000000000000000011 kg	1.1 x 10^{-22} kg
earth to sun	distance	150,000,000,000 meters	1.5 x 10^{11} m

In scientific notation, numbers are expressed as follows:

$$A \times 10^{\,n}$$

where A is the **coefficient,** and n is the **exponent,** an integer representing the number of factors of ten present in the number. The coefficient has only one digit before the decimal place, and that digit is not zero. In other words, 1.5 and 2.730 are possible coefficients, but 10.6 and 0.52 are not. The exponent is positive for very large numbers and it is negative for very small numbers, but it is always an integer. As was mentioned previously, scientific notation enables order of magnitude comparisons. A proton's mass is 0.0000000000000000000000016726 g. If written as 1.6726×10^{-24} g, the number is easier to read and compare with other masses such as the mass of an electron, 9.109×10^{-28} g. The order of magnitude of the ratio of the masses can be obtained by comparing the exponents instead of the more error-prone task of counting the leading zeros. In this case, −24 is larger than −27 and therefore the proton is roughly three orders of magnitude (10^3 = 1000 times) more massive than the electron.

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Example 2.1

A proton has a mass of 1.67262×10^{-24} grams and an electron has a mass of 0.00091×10^{-24} grams. Which is heavier? How many orders of magnitude heavier?

Solution
A proton has a mass that is larger than the mass of an electron. The calculation:

$$\frac{1.67262 \text{ x } 10^{-24} \text{ grams}}{9.1 \text{ x } 10^{-28} \text{ grams}} = 1.8 \times 10^{+3} = 1.8 \times 10^{3}$$

shows that a proton is about 1,800 times more massive. Note that the exponent is <u>*three*</u>. The mass of a proton is about <u>*three orders of magnitude*</u> larger than an electron.

2.3 Base Units

A number used to represent a measurement must be accompanied by **units**. In science, the International System of Units (SI) is used. This system specifies fundamental units for common measurements. Some SI **base units** are listed in Table 2.2.

For very large or very small measurements, both in base and derived units, the quantity may be modified by the addition of a numerical prefix that refers to a multiple of ten (Table 2.3). These prefixes make size or quantity comparisons easier. For instance, the diameter of the DNA double helix is approximately 2.3×10^{9} meters, or 23 nm. A computer may have 8×10^{9} bytes of memory, which can more easily be expressed as 8 GB of memory. Unit equivalencies may be expressed by using either positive or negative exponents, but always double-check to make sure that the sign of the exponent makes physical sense. A kilometer is a <u>*long*</u> distance and contains <u>*one thousand*</u> meters. A <u>*microgram*</u> is a tiny amount and contains <u>*one millionth*</u> (1×10^{-6}) of a gram.

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Table 2.2 Base Units in the International System of Units

Measurement	Base Unit	Abbreviation
mass	kilogram*	kg*
length	meter	m
time	second	s
temperature	Kelvin	K
current	ampere	amp
amount	mole	mol

* Note that the base unit for mass is the *kilogram* even though most masses in chemistry are expressed in grams.

Table 2.3 Unit Prefixes

Prefix	Abbreviation	Scientific Notation	Example
giga–	G	x 1,000,000,000 = x 10^{+9}	1 GB = 10^{+9} bytes
mega–	M	x 1,000,000 = x 10^{+6}	1 MB = 10^{+6} bytes
kilo–	k	x 1000 = x 10^{+3}	1 kg = 1000 g *or* 1 g = 10^{-3} kg
deci–	d	x 0.1 = x 10^{-1}	1 dm = 0.1 m *or* 1m = 10^{-1} dm
centi–	c	x 0.01 = x 10^{-2}	1 cm = 0.01 m *or* 1 cm = 10^{-2} m
milli–	m	x 0.001 = x 10^{-3}	1 mL = 10^{-3} L *or* 1000 mL = 1 L
micro–	µ	x 0.000001 = x 10^{-6}	1 µm = 10^{-6} m *or* 1 m = 10^{+6} µm
nano–	n	x 0.000000001 = x 10^{-9}	1 ng = 10^{-9} g *or* 1 g = 10^{9} ng
pico–	p	x 0.000000000001 = x 10^{-12}	1 ps = 10^{-12} s *or* 1 s = 10^{12} ps

* *The Greek letter µ (mu) is used for the prefix* micro.

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Example 2.2

Use the information in Table 2.3 to work the following problems.
(a) A yard is equivalent to 0.9144 meters. Express the length in centimeters.
(b) The width of a human hair is 60 micrometers. Express the width in meters.
(c) The amount of DNA needed for a PCR (polymerase chain reaction) experiment in biology is 250 nanograms. Express the amount in grams.

Solutions
(a) 0.9144 meters = 9.144×10^{-1} meters, which can be simplified to 9.144×10^{1} cm.
(b) 60 micrometers = 60×10^{-6} meters, which can be simplified to 6×10^{-5} m.
(c) 250 nanograms = 250×10^{-9} grams, which can be simplified to 2.5×10^{-7} g.

Test Yourself
(a) The volume of an Olympic swimming pool is 2.5 megaliters. Express the volume in liters.
(b) The width of a hemoglobin molecule is 550 nanometers. Express the width in meters.
(c) An IV bag of saline contains 1.051 L. Express this volume in milliliters.

Answers
(a) 2.5×10^{6} L (b) 5.5×10^{-7} m (c) 1051 mL

2.4 Derived Units

Measurements such as **volume** are **derived units**; these are constructed from combinations of base units. For instance, the volume of a rectangular box can be calculated from length x width x height. The length, the width, and the height each have a base unit of meters, and the volume can be expressed in terms of cubic meters, or m^3. Smaller object volumes can be expressed in units such as centimeters cubed, cm^3. The volumes for common solids can be calculated by using the geometric formulas presented in Table 2.4. Volumes for liquids are typically expressed in terms of liters. A liter is defined as the cube of one-tenth of a meter (Figure 2.2). Because one-tenth of a meter is 10 cm, a liter is 1000 cm^3. One liter is equal to 1000 mL, therefore 1 cm^3 = 1 mL. In medicine, the volume of one milliliter is referred to as one **cc**, or centimeter cubed.

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A rectangular box with a length of 3.00 cm, a width of 2.00 cm, and a height of 1.50 cm has a volume calculated as follows:

$$V = \ell \times \omega \times h = (3.00\ cm)(2.00\ cm)(1.50\ cm) = 9.00\ cm^3$$

The box has a volume of 9.00 centimeters cubed (cm^3). Because 1 cm^3 = 1 mL, the volume of the box can also be expressed as 9.00 mL.

Example 2.3

Find the volume in centimeters cubed for each of the following objects:
(a) a cube with a side measuring 2.0 cm
(b) a cylinder with a radius of 1.5 cm and a height of 8.0 cm
(c) a sphere with a diameter of 4.00 cm

Solutions
(a) A cube with a side measuring 2.0 cm has a volume of 8.0 cm^3.

$$V = x^3 = (2.0\ \text{cm})^3 = 8.0\ \text{cm}^3$$

(b) A cylinder with a radius of 1.5 cm and a height of 8.0 cm has a volume of 56 cm^3.

$$V = \pi r^2 h = \pi\ (1.5\ \text{cm})^2(8.0\ \text{cm}) = 56\ \text{cm}^3$$

(c) A sphere with a diameter of 4.00 cm has a volume of 33.5 cm^3.

$$V = \frac{4}{3}\pi r^3 = \frac{4}{3}\pi(2.00\ \text{cm})^3 = 33.5\ \text{cm}^3$$

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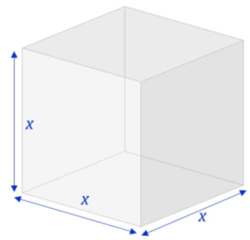
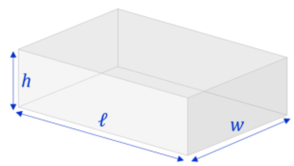
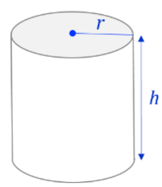
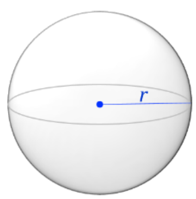
Table 2.4 Volumes of Common Solids

Solid	Shape	Volume (m^3)
cube		$V = x^3$
rectangular box		$V = \ell \times \omega \times h$
cylinder		$V = \pi r^2 h$
sphere		$V = \frac{4}{3}\pi r^3$

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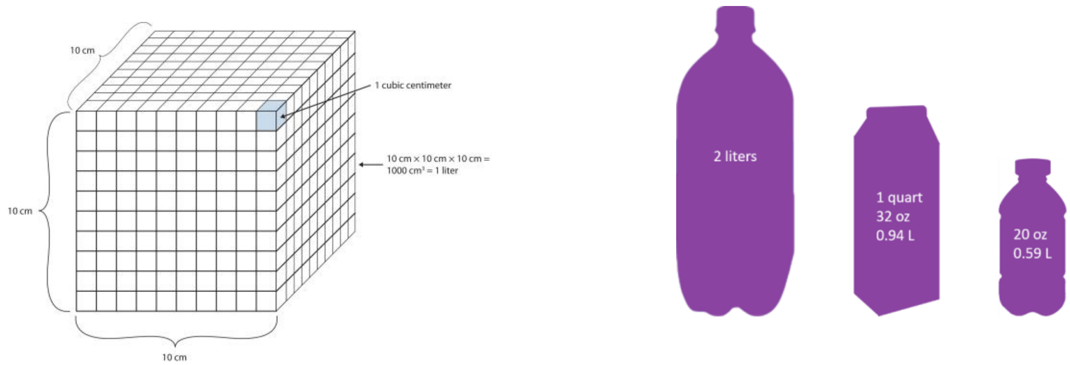

(a) **(b)**

Figure 2.2 (a) One liter is defined as one-tenth of a cubic meter, or 1000 cm^3. One cubic centimeter is equal to one milliliter. In medicine, a cubic centimeter is often represented as cc. (b) Common volumes for commercial products. *Image Credit (a): [OpenStax Chemistry]*

In a chemistry laboratory, it is important to recognize volumes and their relative sizes. In the United States, a common commercial volume for milk is one quart (32 fluid ounces). A sports drink such as Gatorade often comes in a 20-fluid ounce bottle. Compare these to the volume of large soda bottle, which is two liters (Figure 2.2). One liter is slightly larger than one quart (1 qt = 0.94 L).

Another derived unit used in chemistry is **density**. Density is a characteristic property of a substance and represents the amount of mass per unit volume. Consider a wedding ring that is shiny and silver-colored (Figure 2.3). How might the metal be identified? Is it silver, white gold, platinum, or some other metal? The density of the ring may offer a clue to its identity.

Figure 2.3 Density is a characteristic property of a material and can be used to help identify the composition of an object like this wedding ring. *Image Credit: Wikimedia Commons by Nemo*

Table 2.5 Examples of Derived Units

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Quantity	Units – Name	Units – Abbreviation	Derivation
area *(A)*	square meter	*m^2*	$A = \ell \times \omega$
volume *(V)*	cubic meters	*m^3*	$V = \ell \times \omega \times h$
density *(ρ)*	kilograms per cubic meter	*kg/m^3 or $kg{\cdot}m^{-3}$*	$\rho = \frac{mass}{volume}$
rate or velocity *(v)*	meters per second	*m/s or $m{\cdot}s^{-1}$*	$v = \frac{distance}{time}$
concentration *(M)*	molarity, moles per liter	*mol/L or $mol{\cdot}L^{-1}$*	$M = \frac{mol}{volume}$

Density can be calculated for an object if its mass and volume are known. The wedding ring in Figure 2.3 has a volume of 0.350 cm^3 and a mass of 7.40 g. Its density can be calculated as follows:

$$density = \frac{mass}{volume} \quad or \quad \rho = \frac{m}{V}$$

where ρ (the Greek letter *rho*) is the density, m is the mass of an object and V is the volume of the object. The ring has a density of:

$$\rho = \frac{m}{V} = \frac{7.40\ g}{0.350\ cm^3} = 21.1\frac{g}{cm^3} = 21.1g/cm^3 = 21.1\ g \cdot cm^{-3}$$

Note the equivalent ways in which the units for density can be written; they all convey the information that the metal has a mass of 21.1 grams for every cm^3 of volume. Compare the calculation result to the densities of some common substances listed in Table 2.6. What is one possible identity for the substance used to make the ring?

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Table 2.6 Densities of Common Substances at 25 °C

Substance	Density (g/mL)
ethanol	0.789
acetone	0.791
rubbing alcohol	0.786
motor oil	0.899
olive oil	0.920
water	1.00
seawater	1.03
honey	1.45
bromine	3.12
mercury	13.6

Substance	Density (g/cm^3)
balsa wood	0.160
cork	0.210
magnesium	1.74
aluminum	2.70
iron	7.86
copper	8.92
silver	10.5
lead	11.34
gold	19.32
platinum	21.45

Example 2.4

What is the density of a metal sample that has a mass of 44.6 g and a volume of 5.00 cm^3? Use the data in Table 2.6 to identify what the metal might be.

Solution

$$\rho = \frac{m}{V} = \frac{44.6\ g}{5.00\ cm^3} = 8.92 \frac{g}{cm^3}$$

The density of the sample matches the density of copper, so the metal might be copper. More experiments would be needed to identify the metal conclusively.

Example 2.5

A hiker found a shiny, gold-colored solid on a mountain trail. It looked like a piece of pure gold, so he took it to an inorganic chemistry professor at a nearby college to have it analyzed. The professor made some measurements and said that the solid had a mass of 15.7 grams and a volume of 3.08 cm^3. Was it gold?

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Solution

$$\rho = \frac{m}{V} = \frac{15.7\ g}{3.08\ cm^3} = 5.10\,\frac{g}{cm^3}$$

The solid picked up by the hiker had a density of 5.10 g/cm^3, but gold has a density of 19.32 g/cm^3. A possibility is that the solid was the mineral pyrite, FeS_2, which looks like gold but is much less dense. For this reason, pyrite is also known as fool's gold.

2.5 Significant Figures: Measurements

In the example just given (Example 2.5), the density of the hiker's solid was calculated by dividing the mass of the object by its volume.

$$\rho = \frac{m}{V} = \frac{15.7\ g}{3.08\ cm^3} = 5.097402597\,\frac{g}{cm^3} = 5.10\,\frac{g}{cm^3}$$

The original answer given by a calculator was 5.097402597 g/cm^3, but the final answer was given as 5.10 g/cm^3. If the density were stated as 5.097402597 g/cm^3, it would imply that measurements resulted in a very precise determination (to nine decimal places!) of the object's density. The object in question has a mass of 15.7 grams, with three digits used to express the mass to the *tenths place*. The volume of the object is 3.08 cm^3, with three digits used to express the volume to only the *hundredths place*. From these limitations, there is no way to know the density to nine decimal places. The calculated density must reflect the precision of the mass and volume measurements. In this case, both the mass and the volume are known to three digits and therefore the calculated density must be limited to three digits and rounded to 5.10 g/cm^3. The meaningful digits in a measurement or calculation are called **significant figures**. Significant figures in measurements will be covered in this section and applying the significant figures to calculations will be covered in the next section.

When conducting an experiment, a scientist must determine how precise all the measurements are by looking at what possible errors or sources of uncertainty exist. The amount of uncertainty depends both upon the skill of the scientist and the quality of the measuring tool or instrument.

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While some balances are capable of measuring masses only to the nearest 0.1 g, highly sensitive analytical balances (Figure 2.5) are capable of measuring to the nearest 0.001 g or better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be examined carefully when reading a measurement. A scientific measurement consists of **certain digits** plus one uncertain **or estimated digit**.

Consider the gray box in Figure 2.4. A ruler can be used to measure the length of the box. The box is a little longer than 2 cm. Each small division on the ruler represents one-tenth of a centimeter, and the edge of the box falls very close to the first line past 2 cm. This means that the length of the box is around 2.1 cm. The ruler has a mark or gradation at the 2.1 cm point, so both digits are **certain**. The actual measurement should also include one estimated digit. Because the edge of the box is exactly at the 2.1 cm mark, the length is 2.10 cm. Look again at the gray box. Does it make sense to report the length of box as simply 2 cm? What about 2.1000 cm? A measurement of 2 cm is approximate does not convey the precision to which the measurement was made. A measurement of 2.1000 cm is not possible with the detail on the given ruler.

The purple box in Figure 2.4 starts at the 3.0 cm mark and ends after the 3.8 cm mark. For the length of the box, the estimated digit falls between 3.8 cm and 3.9 cm. This means that the purple box is somewhere between 0.8 and 0.9 cm long. The edge of the purple box appears to fall half-way between 0.8 and 0.9, so the correct measurement is 0.85 cm.

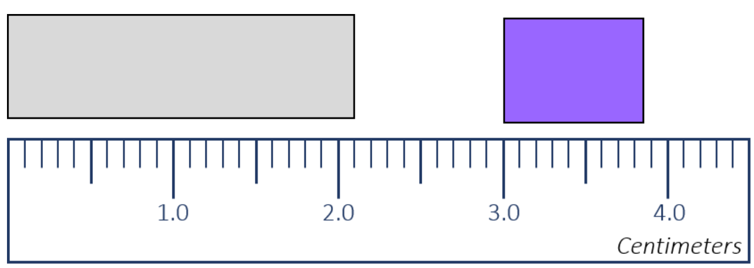

Figure 2.4 A ruler can be used to measure the length of the gray box and the purple box. The length of the gray box can be measured to three significant figures and is 2.10 cm. The length of the purple box can be measured to two significant figures and is 0.85 cm.

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Example 2.6

A machmeter is used by pilots to determine airspeed as a Mach number. The Mach number is the ratio of the true airspeed to the speed of sound. What is the Mach number displayed on the machmeter?

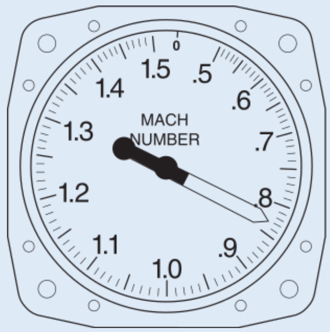

Image Credit: Wikimedia Commons by the Federal Aviation Administration

Solution

The numbers 0.83 are certain digits. The third decimal place can be estimated. A reasonable answer for the Mach number is 0.832 or a number close to 0.832 as long as it contains three significant figures. The Mach number is a ratio and therefore dimensionless.

The precision of a measurement is conveyed by the number of significant figures reported. For instance, the current population of the United States is 326,000,000. This number is an estimate because the population is not *exactly* 326 million people. The population can be expressed in scientific notation as 3.26×10^8 and the fact that there are only three significant figures is clear.

This concept of significant figures holds true for all measurements, even if an estimate is not made. If a sample is placed on an analytical balance and a reading of 0.1502 g is obtained (Figure 2.5), the digits 1, 5, and 0 are certain. The last digit, 2, indicates that the mass of the sample is likely between 0.1501 g and 0.1503 g. The sample therefore has a mass of 0.1502 g with a nominal uncertainty in the measurement of ±0.001 gram. All digits in a measurement, even the last or uncertain digit, are significant.

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Example 2.7

What is the volume of the liquid in the graduated cylinder before the piece of metal is added? What is the volume after the metal is added?

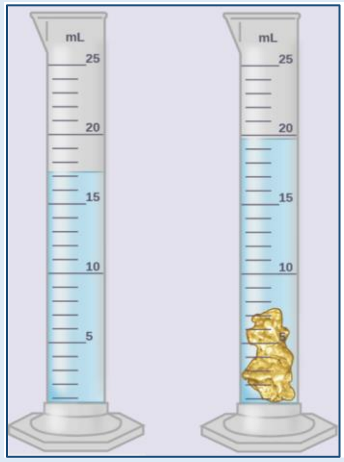

Image Credit: [OpenStax Chemistry]

Solution

The graduated cylinder to the left has a volume of 17.1 mL. The 17 is certain because there is a graduation mark as 17 on the cylinder. The third digit is estimated by the reader. The level of the liquid is slightly above the 17 mark, so 17.1 mL is a reasonable answer. The graduated cylinder containing liquid and metal has a volume of 19.8 mL.

Figure 2.5 An analytical balance can be used to take the mass of samples in a chemistry lab. In this figure, a sample with a mass of 0.1502 grams has been placed in the flask on the balance. The mass of the flask would have been subtracted, or tared, before the sample was added to it. *Image Credit: Mettler-Toledo.*

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The number of significant figures should be determined according to the following rules:

1. All non-zero digits are significant.

1385 four significant figures

1.575 four significant figures

2. Captive zeros or internal zeros (zeros between numbers) are always significant.

2025 four significant figures

3. Trailing zeros after a decimal point are significant.

10.50 four significant figures

1.9500 five significant figures

4. Leading zeros are not significant.

0.00001 one significant figure

0.0050 two significant figures

5. Trailing zeros before a decimal point may or may not be significant. To resolve ambiguity, express numbers in scientific notation. A trailing zero followed by a decimal point indicates that the zero is significant. For instance, 1400. means that both trailing zeros are significant. The number would be written as 1.400×10^3 in scientific notation.

1400
? ? ambiguous

1.4×10^3 two significant figures

Use a decimal to indicate that all digits are significant.

1400. four significant figures

6. Exact numbers such as counting numbers or defined quantities have an infinite number of significant figures.

24 students

185 cars

1 cm = 10 mm

1 in = 2.54 cm

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Example 2.8

Determine the number of significant figures in each of the following measurements.

(a) 36.7 m	(b) 0.006606 s	(c) 2,005 kg	(d) 36,490,000 people
(e) 27 cars	(f) 315.6 mi	(g) 0.0025 ns	(h) 22 min

Solutions

(a) Three significant figures. Rule 1

(b) Four significant figures. Rules 2 and 4

(c) Four significant figures. Rule 2

(d) Four significant figures. Rule 5

(e) Unlimited significant figures. Rule 6

(f) Four significant figures. Rule 1

(g) Two significant figures. Rule 4

(h) Two significant figures. Rule 1.

Test Yourself

How many significant figures are in each of the following measurements?

(a) 0.000601 m (b) 65.080 kg (c)100.00 g

Answers

(a) three significant figures (b) five significant figures (c) five significant figures

Results of scientific measurements or calculations must be analyzed for both **accuracy** and **precision**. Accuracy describes of how close a measurement is the accepted or actual value of the quantity being measured. Precision describes how close a series of measurements are to one another. Figure 2.6 shows how to properly classify shots at a target as either accurate, precise, neither, or both.

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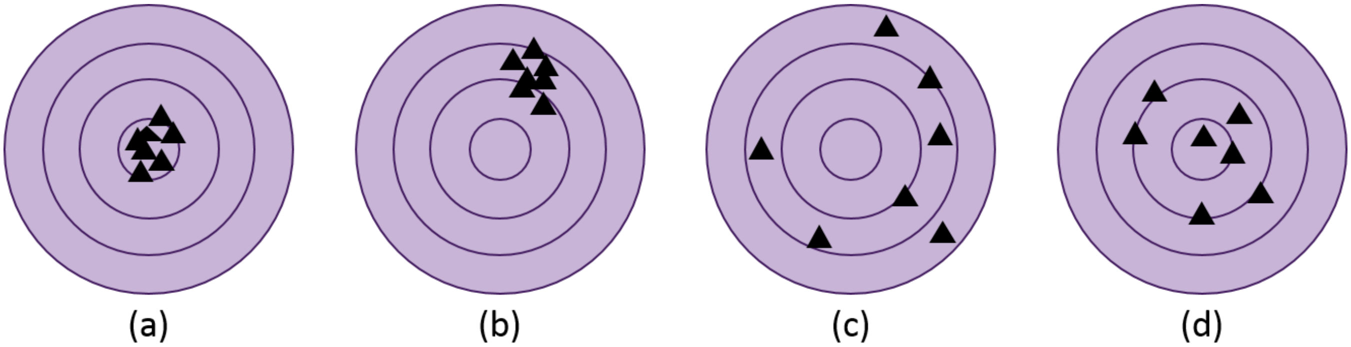
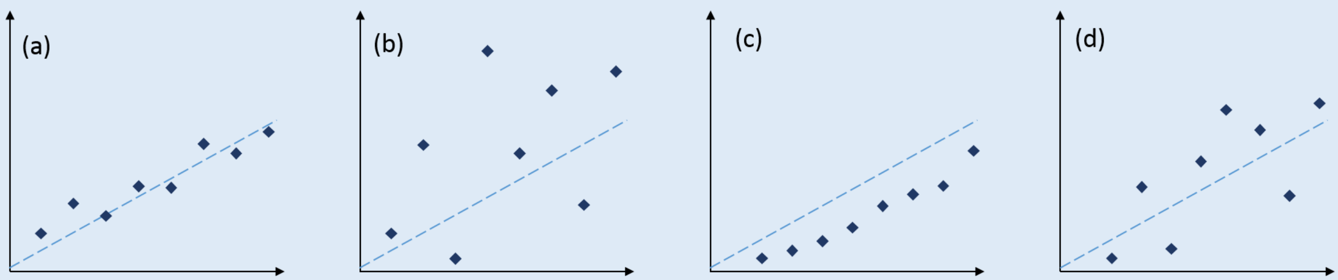


Figure 2.6 (a) The shots are close to the center of the target and also close together, so they are both accurate and precise. (b) The shots are close to one another but not to the center of the target, so they are precise but not accurate. (c) The shots are neither near the center of the target nor are they close together. These shots are neither accurate nor precise. (d) The shots are evenly distributed and tend to center around the center of the target. These shots are accurate but not precise.

Example 2.9

Scientists want results that are both accurate and precise. In terms of accuracy and precision, what is the best description for the data shown? Published data are depicted by the dashed line in each.



Solutions

(a) The data points are close to the line and clustered at the dashed line. These data are both accurate and precise.

(b) The data points are scattered and are nowhere near the dashed line. These data are neither accurate nor precise.

(c) The data points are clustered in the shape of the line but not near or around the dashed line. These data are precise but not accurate.

(d) Overall, the data points average out to the dashed line, but they are not clustered near the line. These data are accurate but not precise.

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Test Yourself

Students collected melting point data during an experiment. What is the best description for each student's data?

Melting Point Data (°C)

Accepted Value: 92.4 °C

	Student A	Student B	Student C	Student D
Trial 1	97.6	92.3	96.8	88.5
Trial 2	91.8	91.9	97.1	69.9
Trial 3	89.6	93.1	95.9	98.6
Trial 4	94.6	92.6	96.9	96.8
Trial 5	88.4	91.8	97.0	45.8
Mean	92.4	92.3	96.7	80.0

Answer

Student A has melting points that are accurate but not precise. The data that student B collected are both accurate and precise. Student C's melting points are precise but not accurate. Student D has data that are neither accurate nor precise.

What could have caused Student C to obtain data that are precise but not accurate?

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2.6 Significant Figures: Calculations

Results calculated from measurements must reflect the uncertainty of those measurements and be reported to the correct number of significant figures. In some cases, reporting to the correct number of significant figures requires **rounding** a number.

Rounding Rules

(a) If the digit to be dropped is *less than 5*, leave the retained digit *unchanged*.

(b) If the digit to be dropped is *greater than 5*, round up and *increase* the retained digit by one.

(c) If the digit to be dropped is equal to 5, the *retained digit should be even*. This may require leaving the retained digit alone in some cases and increasing the retained digit in others. For example:

6.875 rounded to three significant figures is 6.88.
The last retained digit (seven) is odd, so the number is rounded up to 6.88.

15.6325 rounded to five significant figures is 15.632.
The last retained digit (two) is even, so it does not change.

0.505 rounded to two significant figures is 0.50.
The retained digit (zero) is even, so it does not change.

Addition and Subtraction

When adding and subtracting, round the result to the same number of decimal places as the measurement with the *least number of decimal places*.

1.2 + 4.41 = 5.6

```
  1.2     one decimal place
+ 4.41    two decimal places
  5.61
```

one decimal place

77.210 − 30.46 = 46.8

```
  77.210   three decimal places
− 30.46    two decimal places
  46.750
```

two decimal places

Multiplication and Division

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When multiplying and dividing, round to the result to the same number of digits as the measurement with the *least number of significant figures*.

$2.5 \times 4.41 = 5.6$

two sig figs | three sig figs | two significant figures

$14.422 / 3.35 = 4.30$

five sig figs | three sig figs | three significant figures

*When dropped digit is five, the retained digit should be *even*.

Mixed Calculations

If a calculation requires a mixture of addition/subtraction and multiplication/division, apply the standard order of operations (PEMDAS: parentheses, exponents, multiplication, division, addition, then subtraction). Several extra digits should be carried through the calculation to the final result before rounding. Rounding intermediate results may introduce significant errors in calculations.

Example 2.10

Perform the following calculations and round the answer to the correct number of significant figures.

(a) 101.2 + 18.702 = (b) 202.88 − 1.013 =

Solutions

(a) Adding the numbers on a calculator gives 119.902. The final answer should be limited to the tenths place: 119.9.

(b) Subtracting the numbers on a calculator gives 201.867. The final answer should be limited to the hundredths place: 201.87.

Test Yourself

(a) 3.445 + 90.83 – 72.4 = (b) 15 + 2.75 – 9.253 =

Answers

(a) 21.9 (b) 8

Example 2.11

Perform the following calculations and round the answer to the correct number of significant figures.

(a) $76.4 \times 180.4 =$ (b) $0.152/0.05061 =$

Solutions

(a) The first number has three significant figures and the second number has four significant figures. The final answer (13782.56) should be rounded to three significant figures (13800) and expressed in scientific notation 1.38×10^4 to avoid ambiguity with the trailing zeros.

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(b) The numerator has three significant figures and the denominator has four significant figures. The final answer (3.003) should be rounded to three significant figures: 3.00.

Test Yourself

(a) 22.4 × 8.314 = (b) 1.381/6.02 =

Answers

(a) 186 (b) 0.229

Example 2.12

Perform the following calculations and record your answer to correct number of significant figures.

(a) 2.8 x 0.532 +12.680 = (b) (15.803 – 14.73)/9.3 =

Solutions

(a) Following the order of operations, the multiplication should be performed first. 2.8 x 0.532 = 1.<u>4</u>896 with the answer is limited to two significant figures (the last significant digit is underlined). Then the addition should be performed. 1.<u>4</u>896 + 12.680 = 14.2 with the final answer limited to the tenths place.

(b) Following the order of operations, the calculation in the parentheses should be performed first. 15.803 – 14.73 = 1.0<u>7</u>3 with the answer is limited to the hundredths place (underlined). The division follows. 1.0<u>7</u>3/9.3 = 0.12 with the answer limited to two significant figures.

Test Yourself

(a) 155 x 2.32 + 12 (b) 0.575 ÷ 1.2 + 3.275 =

Answers

(a) 372 (b) 3.75

Example 2.13

If the metal in Example 2.7 has a mass of 11.3 grams, what is its density?

Solution

$$\textit{Volume of metal} = 19.8\ cm^3 - 17.1\ cm^3 = 2.7\ cm^3$$

$$\rho = \frac{m}{V} = \frac{11.3\ g}{2.7\ cm^3} = 4.2\,\frac{g}{cm^3}$$

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2.7 Converting Units: Dimensional Analysis

When calculating the density of the wedding ring shown in Figure 2.3, the mass was given as 7.40 grams. If the mass were given instead as 7.40 x 10^{-3} kg, the density could still be calculated and expressed in g/cm^3 if the units of kilogram were **converted** to grams by using **dimensional analysis**.

A mass of 7.40 x 10^{-3} kg can be converted to grams by using the fact that one thousand grams is equal to one kilogram: 1000 g = 1 kg. The mass of 7.40 x 10^{-3} kg can be converted as follows:

$$7.40 \times 10^{-3}\ kg \times \frac{1000\ g}{1\ kg} = 7.40 \times 10^{-3}\ kg \times (1)$$

Because 1000 g = 1 kg and the mass 7.40 x 10^{-3} kg is being multiplied by <u>*one*</u>, the measured quantity does not change when the multiplication is performed. The units, however, do change. The units for kilogram appear in both the numerator and denominator of the calculation and therefore cancel. With the *kg* unit gone, the only unit left is *g*. The calculation results in an answer that is now in grams.

$$7.40 \times 10^{-3}\ \cancel{kg} \times \frac{1000\ \boldsymbol{g}}{1\ \cancel{kg}} = 7.40\ \boldsymbol{g}$$

Keeping track of units and using conversion factors is a fundamental mathematical technique in science. This type of calculation is used at all levels of chemistry and must be mastered properly and used consistently.

Table 2.7 Useful Conversion Factors

Mass	Length	Volume	Pressure
1 kg = 2.205 lb	1 mi = 1.60934 km	1 qt = 0.9643 L	1 atm = 760 mmHg*
1 lb = 0.4536 kg	1 m = 39.37 in	1 gal = 4 qt*	1 torr = 1 mmHg*
1 oz = 28.35 g	1 m = 1.094 yd	1 qt = 32 liquid oz*	1 atm = 101,325 Pa*
1 ton (metric) = 1000 kg*	1 mi = 5280 ft*	1 qt = 4 cups*	1 Pa = 1 kg m^{-1} s^{-2}*
1 ton (imperial) = 2000 lb*	12 in = 1 ft*	1 cup = 236.588 mL	1 bar = 10^5 Pa*
1 mg = 1000 μg*	1 in = 2.54 cm*	1 mL = 1 cm^3 = 1 cc*	1 atm = 1.01325 bar*

**The marked conversions are defined and have an infinite number of significant figures.*

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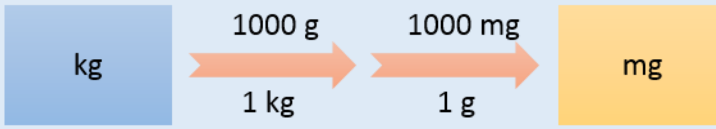
Example 2.14

How many milligrams are in 2.77 kilograms?

Solution

First convert kilograms to grams (1 kg = 1000 g) and then convert grams to milligrams (1 g = 1000 mg). The concept map for this calculation is:



$$2.77\ \cancel{kg} \times \frac{1000\ \cancel{g}}{1\ \cancel{kg}} \times \frac{1000\ \boldsymbol{mg}}{1\ \cancel{g}} = 2.77 \times 10^6\ \boldsymbol{mg}$$

Test Yourself

How many pounds are in 2.77 kg? *Answer* 6.11 lb

How many nanoseconds are in 2.75 minutes? *Answer* 1.65 x 10^{11} ns

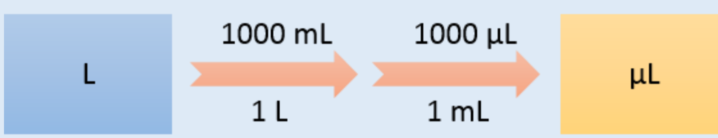
Example 2.15

(a) Convert 35.9 liters to microliters.

(b) Convert 555 nanometers to meters.

Solutions

(a) The conversion to be used are 1 L = 1000 mL and 1 mL = 1000 μL Starting with the measurement given, 35.9 liters, a conversion to milliliters and then microliters can be performed. The concept map for this conversion is:



$$35.9\ \cancel{L} \times \frac{1000\ \cancel{mL}}{1\ \cancel{L}} \times \frac{1000\ \boldsymbol{\mu L}}{1\ \cancel{mL}} = 3.59 \times 10^7\ \boldsymbol{\mu L}$$

Recall that exact numbers or definitions such as 1 L = 1000 mL do not count when determining significant figures, so the answer should contain three significant figures.

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(b) One meter is 10^9 nanometers (1 m = 10^{+9} nm). Starting with the measurement given, 555 nm, a conversion to meters can be performed:

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

1 m = 10^{+9} nm is exact, so the answer should be expressed in scientific notation to three significant figures.

Test Yourself

(a) Convert 67.08 μL to liters.

(b) Convert 56.8 m to kilometers.

Answers

(a) 6.708×10^{-5} L

(b) 5.68×10^{-2} km

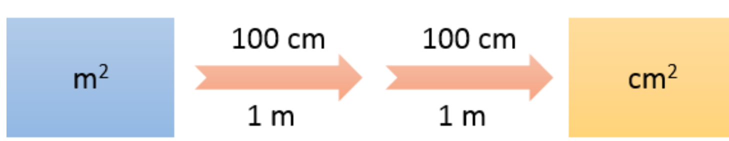
For derived units such as area and volume, more than one conversion factor may have to be applied. For instance, what if a measurement is 20.3 meters squared (m^2) and needs to be converted to centimeters squared (cm^2)? First, the measurement 20.3 m^2 can be written as 20.3 m·m. Next, each meter unit must be converted to centimeters using the fact that 1 m = 100 cm.

$$20.3\ m^2 = 20.3\ \cancel{m} \cdot \cancel{m} \times \frac{100\ \boldsymbol{cm}}{1\ \cancel{m}} \times \frac{100\ \boldsymbol{cm}}{1 \cancel{m}} = 2.03 \times 10^5\ \boldsymbol{cm^2}$$

When using an equality, the measurement is multiplied by one, so this conversion can also be completed by using 1 cm = 10^{-2} m.

$$20.3\ m^2 = 20.3\ \cancel{m} \cdot \cancel{m} \times \frac{1\ \boldsymbol{cm}}{10^{-2}\ \cancel{m}} \times \frac{1\ \boldsymbol{cm}}{10^{-2} \cancel{m}} = 2.03 \times 10^5\ \boldsymbol{cm^2}$$

When planning the calculations in a dimensional analysis problem, a concept map that lays out the specific steps or conversions is useful. For the problem just shown, a concept map might look like:



Density, a derived unit introduced in Section 2.4, has SI units of kg/m^3, but chemists often use density in more convenient units such as g/cm^3. Tungsten has a density of 19300 kg/m^3. What is

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the density of tungsten in g/cm³? There are two major steps for this conversion: (1) kg must be converted to grams in the numerator and (2) m³ must be converted to cm³ in the denominator. The concept map for this calculation is:

$$\frac{kg}{m^3} \xrightarrow[1\ kg]{1000\ g} \frac{g}{m^3} \xrightarrow[100\ cm]{1\ m} \xrightarrow[100\ cm]{1\ m} \xrightarrow[100\ cm]{1\ m} \frac{g}{cm^3}$$

(1) *kilograms to grams, m³ remains in the denominator*

$$19300\ \frac{\cancel{kg}}{m^3} \times \frac{1000\ g}{1\cancel{kg}} = 1.93 \times 10^7\ \frac{g}{m^3}$$

(2) *m³ to cm³, the already-converted grams remain in the numerator*

$$1.93 \times 10^7\ \frac{g}{m^3} = 1.93 \times 10^7\ \frac{g}{m \cdot m \cdot m}$$

$$1.93 \times 10^7\ \frac{g}{\cancel{m} \cdot \cancel{m} \cdot \cancel{m}} \times \frac{1\ \cancel{m}}{100\ cm} \times \frac{1\ \cancel{m}}{100\ cm} \times \frac{1\ \cancel{m}}{100\ cm} = 19.3\ \frac{g}{cm^3}$$

Using density written as g/cm³ or g/mL (recall that 1 cm³ = 1mL) as a conversion factor is very common in chemistry. Suppose a sample of copper has a mass of 5.75 grams. What is the volume of the sample of copper? As with any dimensional analysis problem, the calculation begins with the given measurement, 5.75 grams. The fact that the sample is copper means that it has a density of 8.92 g/cm³ (Table 2.6), which means that 8.92 g = 1 cm³.

$$5.75\ \cancel{g} \times \frac{1\ cm^3}{8.92\ \cancel{g}} = 0.645\ cm^3$$

Example 2.16

What is 883 m³ expressed in cubic centimeters?

Solution

Each of the meter units must be converted to centimeters using the equality 1 m = 100 cm. The concept map for this calculation is:

$$m^3 \xrightarrow[1\ m]{100\ cm} \xrightarrow[1\ m]{100\ cm} \xrightarrow[1\ m]{100\ cm} cm^3$$

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$$0.883\ m^3 = 0.883\ \cancel{m} \cdot \cancel{m} \cdot \cancel{m} \times \frac{100\ \mathbf{cm}}{1\ \cancel{m}} \times \frac{100\ \mathbf{cm}}{1\ \cancel{m}} \times \frac{100\ \mathbf{cm}}{1\ \cancel{m}} = 8.83 \times 10^5\ \mathbf{cm^3}$$

$$or \qquad 0.883\ \cancel{m}^3 \times \left(\frac{100\ \mathbf{cm}}{1\ \cancel{m}}\right)^{\mathbf{3}} = 8.83 \times 10^5\ \mathbf{cm^3}$$

The answer contains three significant figures. Recall that defined values such a 1 m = 100 cm are exact and are not included in significant figure determination.

Test Yourself

How many cubic millimeters are present in 0.0923 m^3? Draw a concept map to get started.

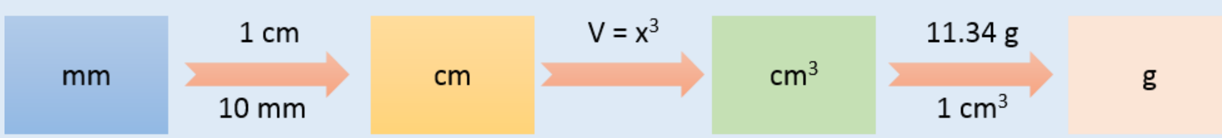
Answer 9.23 x 10^7 mm^3

Example 2.17

A cube of pure lead (density = 11.34 g/cm^3) has a side measuring 54.5 mm. What is the mass of the cube? Express your answer in grams and use the proper number of significant figures.

Solution

Density can be used to convert volume to mass. The volume of the lead cube must be determined in cm^3 in order to use the density (g/ cm^3) to calculate mass. The concept map for this calculation is:



$$x = 54.5\ mm \times \frac{1\ \mathbf{cm}}{10\ \cancel{mm}} = 5.45\ \mathbf{cm} \quad and \quad V = x^3 = (5.45\ cm)^3 = 161.9\ cm^3$$

$$161.9\ \cancel{cm}^3 \times \frac{11.34\ \mathbf{g}}{1\ \cancel{cm}^3} = 1.84 \times 10^3\ \mathbf{g}$$

The side measurement (54.5mm) has three significant figures so the answer must have three significant figures

Test Yourself

A cylinder of pure iron (density =7.86 g/cm^3) has a diameter of 0.105 in and a height of 2.6 in. What is the mass of the cylinder expressed in kilograms? *Answer* 2.9 x 10^{-3} kg

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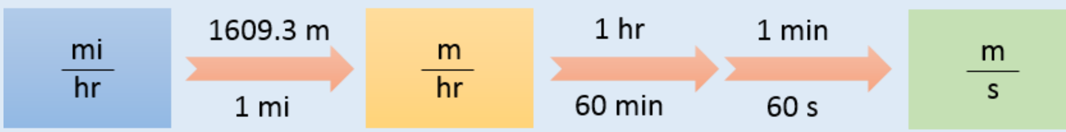
Example 2.18

At his fastest, Usain Bolt ran at a speed of 27.7 mi/hr (miles per hour). What is this speed in units of meters per second?

Solutions

First, the distance in miles in the numerator must be converted to meters (1 mi = 1609.3 m). Then the time in hours in the denominator must be converted to seconds (1 hr = 60 min and 1 min = 60 s). The concept map for this calculation is:



(1) *miles to meters, hours remain in the denominator*

$$27.7\frac{\cancel{mi}}{hr}\times\frac{1609.3\ \boldsymbol{m}}{1\ \cancel{mi}}=4.4578\times10^4\,\frac{\boldsymbol{m}}{hr}$$

(2) *hours to seconds, already-converted meters remain in the numerator*

$$4.4578\times10^4\,\frac{m}{\cancel{hr}}\times\frac{1\ \cancel{hr}}{60\ \cancel{min}}\times\frac{1\ \cancel{min}}{60\ \boldsymbol{s}}=12.4\frac{m}{\boldsymbol{s}}$$

Test Yourself

What is 0.203 m/min in units of cm/s? Draw a concept map to get started.

Answer 0.00338 m/s or 3.38×10^{-3} m/s = 3.38×10^{-1} cm/s

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2.8 Temperature

One of the fundamental quantities in science is **temperature** – the measure of the average amount of **kinetic energy**, the energy due to motion, in a system. In the United States, temperatures are usually expressed in degrees Fahrenheit (°F). On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F. In science and engineering, temperatures are expressed in degrees Celsius (°C). On the Celsius scale, water freezes at 0 °C and boils at 100 °C. To convert between the Celsius and Fahrenheit temperature scales, use the following equations:

$$°C = (°F - 32) \times \frac{5}{9} \qquad °F = \left(°C \times \frac{9}{5}\right) + 32$$

The SI unit for temperature is the kelvin (K). By convention, the word kelvin (all lowercase) and unit K (uppercase) are used for the unit. Neither the word "degree" nor the degree symbol (°) are used with kelvin temperatures. The phrase Kelvin temperature scale is itself capitalized just like the Fahrenheit and Celsius scales. The divisions on the Celsius and Kelvin temperature scales are the same, so a change of 1.0 °C is equal to a change of 1.0 K. The difference is that a temperature on the Celsius scale is 273.15 degrees above the temperature on the Kelvin scale. Water freezes at 273.15 K and boils at 373.15 K (Figure 2.7). To convert between temperatures on the Celsius and Kelvin scales, use the equations shown in Figure 2.7.

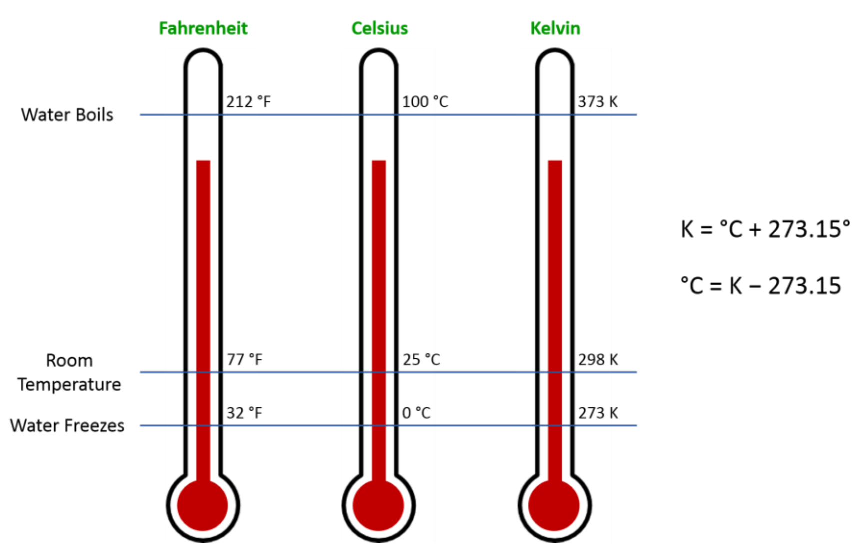

Figure 2.6 7 The relationships between the Fahrenheit, Celsius, and Kelvin temperature scales are shown. Note that a change of 1.0 °C is equal to a change of 1.0 K.

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Example 2.19

(a) Normal temperature for the human body is 98.6 °F. What is this temperature in °C?

(b) Standard laboratory temperature is 25.0 °C. What is this temperature in °F.

Solutions

(a)

$$°C = (°F - 32) \times \frac{5}{9} = (98.6 - 32) \times \frac{5}{9} = 37.0\ °C$$

(b)

$$°F = \left(°C\ \times \frac{9}{5}\right) + 32 = \left(25.0 \times \frac{9}{5}\right) + 32 = 77.0\ °F$$

Test Yourself	(a) Convert 0.00 °F to °C.	(b) Convert 212 °C to °F.
Answers	(a) −17.8 °C	(b) 414 °F

Example 2.20

Classrooms are usually kept at 72.0 °F. What is this temperature expressed in °C and K?

Solution

$$°C = (72.0 - 32)\ \times \frac{5}{9}\ =\ 40.0\ \times \frac{5}{9}\ =\ 22.2\ °C$$

$$K = 22.2°C + 273.15 = 295.4\ K$$

Test Yourself What is -12.8 °F on the Kelvin scale?

Answer 248.3 K *Remember that a Kelvin temperature cannot be negative!*

Units are Everywhere: The Gimli Glider

On July 23, 1983, an Air Canada Boeing 767 jet had to glide to an emergency landing at Gimli Industrial Park Airport in Gimli, Manitoba, because it unexpectedly ran out of fuel during flight. There was no loss of life in the course of the emergency landing, only some minor injuries associated in part with the evacuation of the craft after landing. For the remainder of its operational life (the plane was retired in 2008), the aircraft was nicknamed "the Gimli Glider."

Figure 2.8 The Gimli Glider is the Boeing 767 that ran out of fuel and glided to safety at Gimli Airport. The aircraft ran out of fuel because of confusion over the units used to express the amount of fuel. *Image Credit: Wikimedia Commons by WillF.*

The 767 took off from Montreal on its way to Ottawa, heading for Edmonton, Canada. About halfway through the flight, all the engines on the plane began to shut down because of a lack of fuel. When the final engine cut off, all electricity (which was generated by the engines) was lost; the plane became a powerless glider. Captain Robert Pearson was an experienced glider pilot, although he had never flown a glider the size of a 767. First Officer Maurice Quintal quickly determined that the aircraft would not be able make it to Winnipeg, the next large airport. He suggested his old Royal Air Force base at Gimli Station, one of whose runways was still being used as a community airport. Between the efforts of the pilots and the flight crew, they managed to get the airplane safely on the ground (although with buckled landing gear) and all passengers off safely.

What happened? At the time, Canada was transitioning from the older English system to the metric system. The Boeing 767s were the first aircraft whose gauges were calibrated in the metric system of units (liters and kilograms) rather than the English system of units (gallons and pounds). Thus, when the fuel gauge read 22,300, the gauge meant kilograms, but the ground crew mistakenly fueled the plane with 22,300 *pounds* of fuel. This ended up being just less than half of the fuel needed to make the trip, causing the engines to quit about halfway to Ottawa. Quick thinking and extraordinary skill saved the lives of 61 passengers and 8 crew members.

This incident would not have occurred if people were watching their units.

Test Yourself If the ground crew filled the fuel tank with 22,300 pounds of fuel, how many kilograms of fuel were loaded? If the tank was meant to hold 22,300 kilograms of fuel, what percentage of the tank was filled?

Answer 10,100 kilograms and 45.4%

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Chapter 2 Practice Problems

2.2 Numbers

1. Express each of the following numbers in scientific notation. Use all the digits given.
   (a) 98653 (b) 145.3 (c) 0.00175 (d) 0.189

2. Express each of the following numbers in scientific notation. Use all the digits given.
   (a) 0.0598 (b) 76.89 (c) 237 (d) 12

3. Express each of the following numbers in standard notation Use all the digits given.
   (a) 1.5×10^{-2} (b) 7.78×10^{2} (c) 0.645×10^{-4} (d) 645×10^{6}

4. Express each of the following numbers in standard notation. Use all the digits given.
   (a) 4.55×10^{-1} (b) 8.15×10^{3} (c) 0.157×10^{-3} (d) 581×10^{4}

5. Express each of the following numbers in scientific notation. Use all the digits given.
   (a) 34.05 (b) 0.0607 (c) 2,567 (d) 0.0010200

6. Express each of the following numbers in scientific notation. Use all the digits given.
   (a) 200 (b) 0.000980 (c) 708 (d) 0.02030

7. Write each of the following measurements in standard notation. Use all the digits given.
   (a) 9.80×10^{-1} (b) 5.75×10^{3} (c) 965×10^{-3} (d) 75.031×10^{4}

8. Write each of the following measurements in standard notation. Use all the digits given.
   (a) 758×10^{3} (b) 0.0025×10^{-4} (c) 1.0156×10^{3} (d) 258×10^{-3}

9. The diameter of Jupiter is 142,984 kilometers. The diameter of Mercury is 4,878 kilometers.
   (a) Express each of these diameters in scientific notation. Use all the digits given.
   (b) Which planet has the larger diameter? How many orders of magnitude larger?

10. The Empire State Building is 443.2 meters. The height of Kilimanjaro is 19,341 meters.
   (a) Express each of these heights in scientific notation and use all the digits given.
   (b) Which is taller? How many orders of magnitude taller?

2.3 Base Units

11. Name the SI base unit for each of the following:
   (a) amount (b) mass (c) length

12. Name the SI base unit for each of the following:
   (b) distance (b) current (c) time

13. Name the SI base unit that would be used for each of the measurements described.

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(a) the distance between Washington DC and Los Angeles
(b) the mass of a penny
(c) the number of eggs in a dozen

14. Name the SI base unit that would be used for each of the measurements described.
(a) the mass of the sun
(b) the temperature of boiling water
(c) the time it takes for a chemical reaction to occur

15. Use the unit prefix information in Table 2.3 to answer the following questions. Express each answer in scientific notation. Use three digits.
(a) A mile is equivalent to 1609.34 meters. What is this distance in millimeters?
(b) A red blood cell has a diameter of 8 micrometers. What is the diameter is centimeters?
(c) If a computer disk can store 3.5 gigabytes of data, how many bytes can it store?
(d) Airborne asbestos fibers have an average diameter of about 2.25 micrometers. What is the average diameter in millimeters?

16. Use the unit prefix information in Table 2.3 to work the following problems. Express each answer in scientific notation. Use two digits.
(a) The distance between Los Angeles and Washington D.C. is 4,297 kilometers. Express this distance in meters, centimeters, and millimeters.
(b) The James Webb Space Telescope (JWST) has mirrors coated with gold. The thickness of the gold coating is 100 nanometers. What is the thickness of the gold coating in meters?
(c) A gold atom has a diameter of 0.166 nanometers. What is the diameter in centimeters?
(d) The distance from Earth to Mars is 54.6 million kilometers. What is the distance in meters?

2.4 Derived Units

17. Start with SI base units and write the derived units used for each of the following. You may need to consult a reference book or the internet.
(a) volume (b) density (c) energy (d) area

18. Start with SI base units and write the derived units used for each of the following. You may need to consult a reference book or the internet.
(a) pressure (b) frequency (c) velocity (d) force

19. Find the volume in SI units for each of the following objects. Express each answer in scientific notation. Use three digits.
(a) a sphere with a radius of 4.55 mm
(b) a cylinder with a diameter of 34.5 cm and a height of 225 cm
(c) a box with a length of 2.25 m, a width of 3.62 mm, and a height of 1.57 m

20. Find the volume in SI units for each of the following objects. Express each answer in scientific notation. Use two digits.
(a) a sphere with a radius of 4.5 mm
(b) a cylinder with a diameter of 34 cm and a height of 225 cm

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(c) a box with a length of 2.2 m, a width of 365 cm, and a height of 1560 mm

21. Acceleration is defined as a change in velocity per unit time, a derived unit. What combination of base units might be used to express acceleration?

22. Density is defined as the mass of an object per unit volume. What combination of base units might be used to express density?

23. A metal sphere has a mass of 378 g and a volume of 75.5 cm^3. What is the density of the metal? Use three digits in your answer.

24. 5.00 mL of a liquid has a mass of 3.95 g. What is the density of the liquid? Use three digits in your answer.

2.5 Significant Figures: Measurements

25. Identify each of the following numbers as exact or inexact. If the number is inexact, state the number of significant figures it contains.
   (a) It takes 45 minutes to drive from Rockville, MD to Arlington, VA.
   (b) A metal cylinder has a diameter of 14 cm.
   (c) The Kennedy Center Opera House has 2362 seats.
   (d) There are 13 cookies in a baker's dozen.
   (e) A paper clip weights 0.675 grams
   (f) A metal cube has a volume of 8 cm^3.

26. Identify each of the following numbers as exact or inexact. If the number is inexact, state the number of significant figures it contains.
   (a) A car is traveling at 45 miles per hour.
   (b) A baby weighs 7 and ½ pounds.
   (c) A newspaper article contains 189 words.
   (d) A flight from New York to Los Angeles takes 6 hours.
   (e) A plastic bottle contains 20 ounces of liquid.
   (f) A kindergarten student is 5 years old.

27. Write the following measurements in scientific notation and to four significant figures.
   (a) 0.0550102 cm (b) 300,410 g (c) 10.008 in (d) 507.05 kg

28. Write the following measurements in scientific notation and to three significant figures.
   (a) 342.79513 m (b) 9,750.87 g (c) 0.000045389 mL (d) 65.85 km

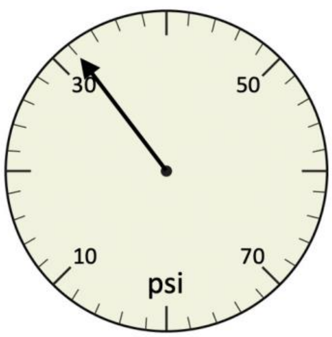
29. Write each of the following measurements in scientific notation and to the appropriate number of significant figures. Use the correct abbreviation for each unit.
   (a) The distance between carbon atoms in an ethane molecule is 154 picometers.
   (b) The atomic radius for helium is 0.000031 micrometers.
   (c) A hemoglobin molecule has a diameter of 51,000 picometers.

(d) The wavelength of an X-ray beam is 0.00015406 micrometers.
(e) The visible light spectrum starts at 390 nanometers
(f) The visible light spectrum ends at a wavelength of 0.0000070 decimeters.

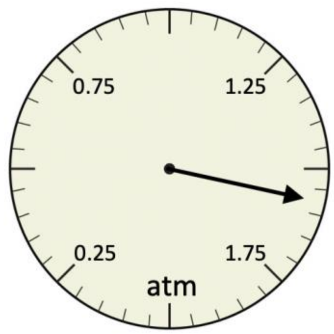
30. Write each of the following measurements in scientific notation. Use the appropriate number of significant figures and the correct abbreviation for each unit.
    (a) The distance from Earth to the sun is 149.6 million km.
    (b) The diameter of the moon is 347.4 million cm.
    (c) Voyager 1, which launched in September 1977, was the first Voyager spacecraft to reach Jupiter. Voyager 1 is currently 21.1 billion km from Earth currently in interstellar space.
    (d) Voyager 2, which launched in August 1977 and is the only spacecraft to have visited the ice giants of Uranus and Neptune, is currently 17,700 Gm (gigameters) from Earth.
    (e) Radio waves can have wavelengths of up to 1,000,000 dm
    (f) The primary mirror on the Hubble Space Telescope is 2,400 millimeters in diameter.

31. Express the pressure depicted on each gauge using the correct number of significant figures and the correct units.

(a)

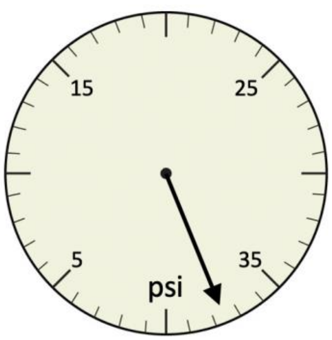


(b)

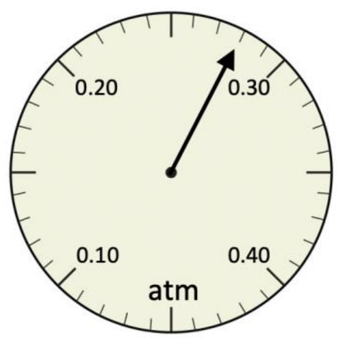


32. Express the pressure depicted on each gauge using the correct number of significant figures and the correct units.

(a)



(b)



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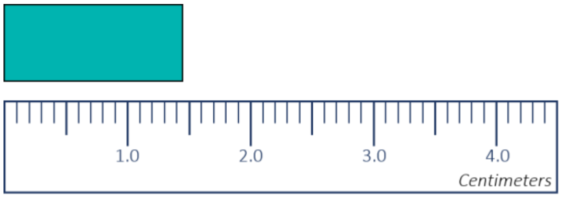
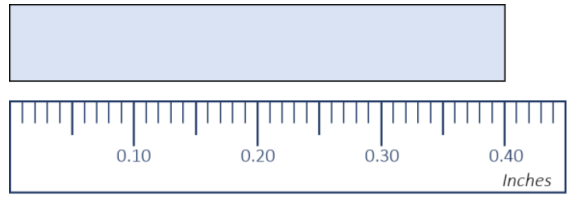
33. Express the length of each box using the correct number of significant figures and the correct units.

(a)

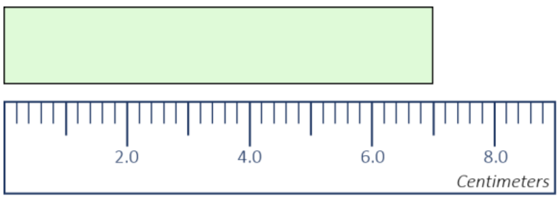


(b)

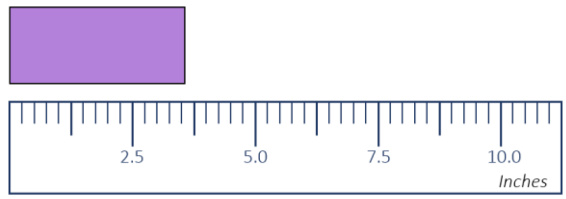


34. Express the length of each box using the correct number of significant figures and the correct units.

(a)

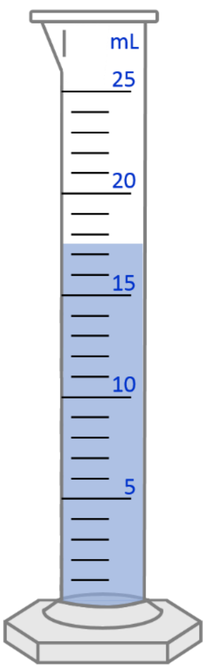


(b)

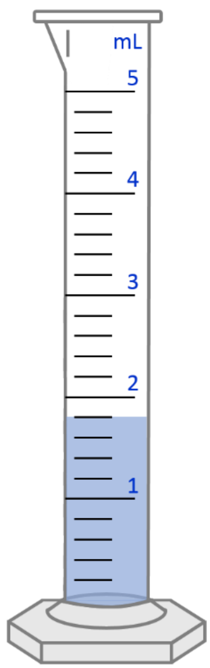


35. Express the volume of the liquid in each graduated cylinder using the correct number of significant figures and the correct units.

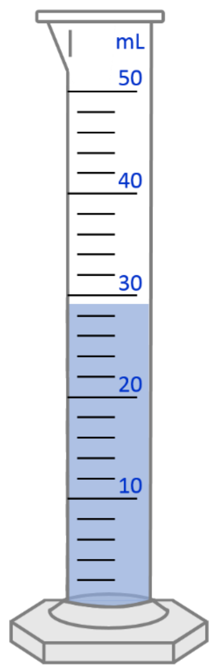
(a)



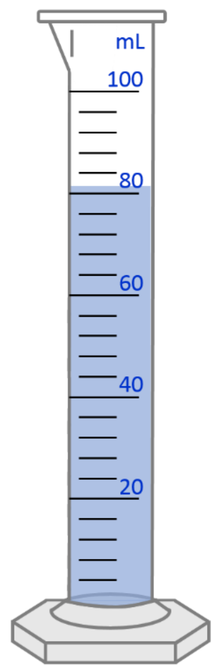
(b)



(c)



(d)



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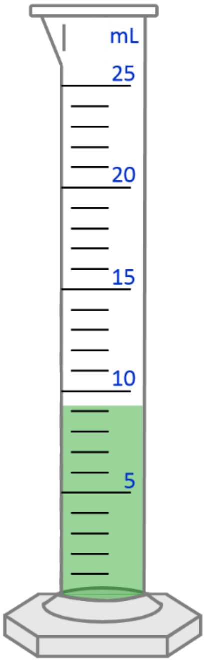
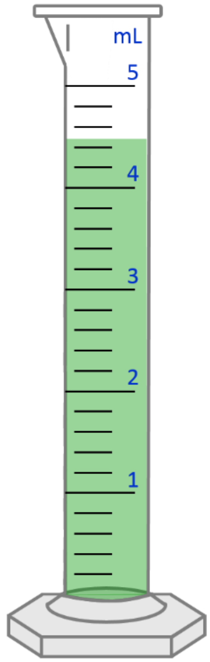
36. Express the volume of the liquid in each graduated cylinder using the correct number of significant figures and the correct units.

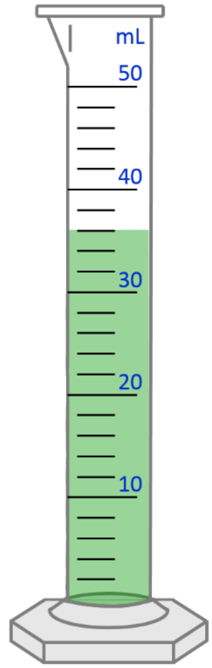
(a)



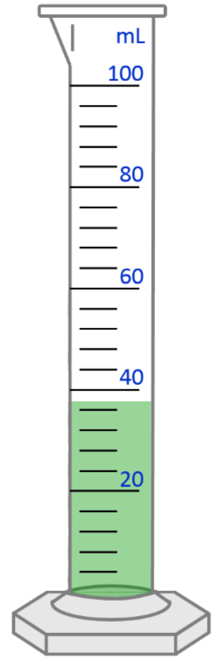
(b)



(c)

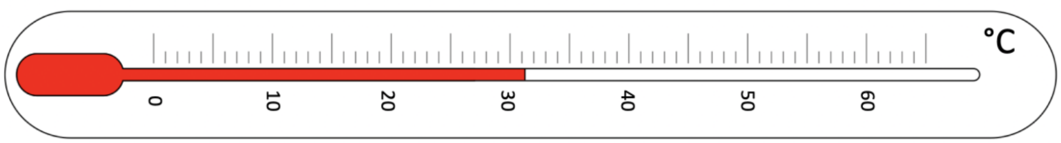


(d)

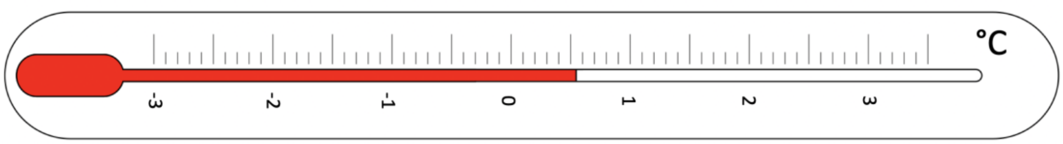


37. Express the temperature of each thermometer to the correct number of significant figures.

a.

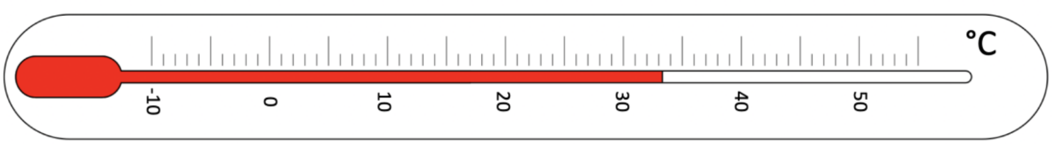


b.

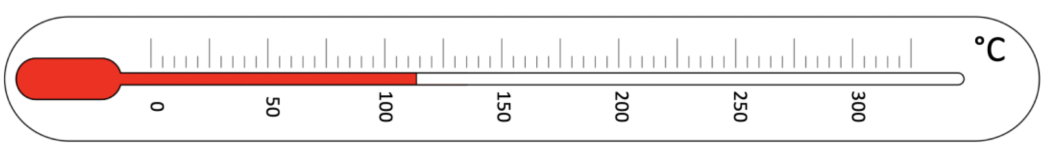


38. Express the temperature of each thermometer to the correct number of significant figures.

a.



b.



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2.6 Significant Figures: Calculations

39. Perform the following mathematical operations and write each answer with the correct number of significant figures. Use scientific notation where appropriate.
    (a) 56.0 + 3.44 = (b) 0.00665 + 1.004 = (c) 45.99 − 32.8 =
    (d) 32.59 − 21.8 + 75.02 = (e) 0.358 + 1.259 + 13.5 = (f) 21.56 – 0.00758 =

40. Perform the following mathematical operations and write each answer with the correct number of significant figures. Use scientific notation where appropriate.
    (a) 0.865 + 2.19 = (b) 14.56 + 0.596 = (c) 25.858 – 1.23 =
    (d) 17.36 – 1.26 + 0.0780 = (e) 0.00685 – 0.001255 = (f) 7.56 + 0.0456 + 1.20 =

41. Perform the following mathematical operations and write each answer with the correct number of significant figures. Use scientific notation where appropriate.
    (a) 121.5 x 0.12 = (b) 0.698 ÷ 0.001356 = (c) 18.2 x 0.02658 =
    (d) 21.585 ÷ 56.2 = (e) 0.0100 x 0.0010 = (f) 0.0368 ÷ 0.0021 =

42. Perform the following mathematical operations and write each answer with the correct number of significant figures. Use scientific notation where appropriate.
    (a) 21.25 x 0.39 = (b) 14.00 ÷ 2.000 = (c) 0.0025 x 2.65 =
    (d) 250. ÷ 50. = (e) 6.0 x 7.25 = (f) 70.0 ÷ 35.0 =

43. Perform the following mathematical operations and write each answer with the correct number of significant figures. Use scientific notation where appropriate.
    (a) (35.0 x 0.789) + 49.38 = (b) (24.0 x 0.125) + 5.38 = (c) 0.12 x 6.22/12.6 =
    (d) (0.0025 + 1.23)/2.75 = (e) (3.75 x 10^3)/0.25 = (f) 1.75/5.0 + 0.875 =

44. Perform the following mathematical operations and write each answer with the correct number of significant figures. Use scientific notation where appropriate.
    (a) (18,900 x 76.33) ÷ 0.036 = (b) 2.25 x 10^{-1} + 1.38 = (c) 22.0 – 7.5 x 0.75 =
    (d) (5.50 x 0.25)/(0.201 ÷ 8.2) = (e) 7.58 x 10^4 + 0.75 = (f) 0.215 x 13 + 1.7 =

45. Find the volume for each of the following objects. Use the units given and express each answer to the correct number of significant figures. Use scientific notation where appropriate.
    (a) a sphere with a radius of 0.0254 m
    (b) a cylinder with a diameter of 1.25 mm and a height of 2.75 mm
    (c) a box with a length of 0.220 cm, a width of 3.6 cm, and a height of 2.250 cm

46. Find the volume for each of the following objects. Use the units given and express each answer to the correct number of significant figures. Use scientific notation where appropriate.
    (a) a sphere with a radius of 21.56 cm
    (b) a cylinder with a diameter of 0.025 m and a height of 17.56 m
    (c) a box with a length of 0.050 mm, a width of 0.0036 mm, and a height of 0.100 cm. Find volume in units of cm^3.

2.7 Converting Units: Dimensional Analysis

Report all answers to the correct number of significant figures. Refer to Table 2.7 for useful conversion factors. You may also need to look up conversion factors on your own.

47. Perform the following conversions.
    (a) 5.4 km to meters
    (b) 0.665 m to millimeters
    (c) 7.60 cm to km
    (d) 53.7 mL to liters
    (e) 0.0965 μL to milliliters
    (f) 987.5 dL to liters

48. Perform the following conversions.
    (a) 85.5 kg to grams
    (b) 0.986 kg to μg
    (c) 765.0 mg to kg
    (d) 86.7 s to milliseconds
    (e) 2.375 ms to s
    (f) 77.56 s to μs

49. Perform the following conversions.
    (a) 9.44 m^2 to square centimeters
    (b) 3.44×10^8 mm^3 to cubic meters
    (c) 0.00444 cm^2 to square meters
    (d) 0.445 nm^3 to cubic centimeters

50. Perform the following conversions.
    (a) 8.11 x 10^2 mm^2 to square centimeters
    (b) 425.8 μm^3 to cubic meters
    (c) 5.6 x 10^{-15} nm^2 to square meters
    (d) 728 m^3 to cubic millimeters

51. Perform the following conversions.
    (a) 45.0 m/min to meters per second
    (b) 25.0 cm/hr to inches per second
    (c) 0.92 mi/hr to millimeters per minute
    (d) 7.5 μm/s to feet per hour

52. Perform the following conversions.
    (a) 55.5 cm/s to meters per hour
    (b) 45.89 μm/min to inches per hour
    (c) 65. mi/hr to meters per minute
    (d) 685 m/s to km/min

53. Perform the following conversions.
    (a) 8.96 g/cm^3 to kg per cubic meter
    (b) 21.89 $kg \cdot m \cdot s^{-2}$ to $g \cdot cm \cdot min^{-2}$
    (c) 0.575 m^3/kg to cubic feet per pound
    (d) 4.5 x 10^{-5} $m^{-1} \cdot g \cdot s^{-2}$ to $mm^{-1} \cdot \mu g \cdot hr^{-2}$

54. Perform the following conversions.
    (a) 0.701 kg/cm^3 to grams per cubic meter
    (b) 5.65 x 105 s^{-1} to min^{-1}
    (c) 987 m/s^2 to inches per hour squared
    (d) 0.096 $m^2 \cdot s^{-2}$ to $mm^2 \cdot \mu s^{-2}$

55. How many grams are in a 0.124 kg sample of table salt?

56. How many meters are in a distance of 1.25 x 10^3 km?

57. If a farm occupies 435 acres, how many square kilometers does it occupy?

58. The Great Salt Lake has a surface area of 2.15 x10^5 mi^2. How many square kilometers is this?

59. The Statue of Liberty is 46. meters tall. How tall is she in inches? How tall is she in feet?

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60. The highest point on the Golden Gate Bridge is 227 meters above the water. What is this distance in inches? What is the distance in feet?

61. How many seconds are there in 7.25 days? Assume that the conversion factors given are exact. 1 year = 365.24 days; 1 day = 24 hr

62. How many seconds are there in 25.5 days? Assume that the conversion factors given are exact. 1 year = 365.24 days; 1 day = 24 hr

63. A stick of butter weighs 0.25 lbs. What is the weight of the stick of butter in milligrams?

64. A garnet weighs 0.85 mg. What is the weight of the garnet in pounds?

65. An Italian recipe for making creamy pasta sauce calls for 0.75 liters of cream. How many cups of cream are needed?

66. A recipe for chicken soup call for 4.5 cups of chicken broth. How many liters of chicken broth are needed?

67. A running track measures 125 m per lap. How many laps must an athlete run in order to run 2.0 miles?

68. An Olympic pool is 164.0 feet in length. How many laps must an athlete swim in order to swim 800.0 meters?

69. The price of gas in Paris is 1.29 euro per liter. What is the price of gasoline in dollars per gallon? Assume that one euro is equal to $1.25.

70. The price of one gallon of milk in Munich is 0.95 euro per liter. What is the price of milk in dollars per gallon? Assume that one euro is equal to $1.25.

71. A cyclist rides at an average speed of 24 miles per hour. If she wants to bike 195 km, how many minutes must she ride?

72. Along a highway in France, the speed limit is 115 km per hour. How many minutes will it take to travel 5.75 miles?

Density as a Conversion Factor

73. A piece of metal ore weighs 8.25 g. When placed into a graduated cylinder containing water, the liquid level rises from 21.25 mL to 26.47 mL. What is the density of the ore?

74. A cube of nickel metal has a side of length 2.25 cm. How many grams of nickel metal are present in this cube? The density of nickel metal is 8.90 g/cm^3.

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75. An insoluble solid has a mass of 121.7576 g. A graduated cylinder contained 8.2 mL of water. When the solid was dropped into the graduated cylinder, the total volume rose to 25.8 mL. What is the density of the solid? Assume that water has a density of 1.00 g/cm^3.

76. A piece of pure copper is placed into a graduated cylinder filled with water. The copper displaces 4.70 mL of water. What is the mass of the piece of copper? The density of copper is 8.96 g/cm^3.

77. Ancient Roman coins were made of pure silver (density = 10.5 g/cm^3). What was the mass of a pure silver coin that had a diameter of 2.55 cm and thickness of 1.75 mm?

78. The United States Mint Coin Specifications for coins currently in production state that a penny has a diameter of 0.750 inches and a thickness of 1.52 mm. What is the volume of a penny in mm^3 and in cm^3? If the penny is 100% copper, what is its mass?

79. A 35.0 mL sample of wine (density = 0.789 g/mL) is added to a glass that has a mass of 49.38 g. What will be the mass of the glass plus the wine?

80. A glass bottle filled with 500. mL of milk (density = 1.034 g/mL) has a mass of 1.75 kg. What is the mass of just the glass bottle?

2.7 Converting Units: Temperature

81. Convert each of the following temperatures to degrees Celsius.
    (a) 32.0 °F (b) 212.0 °F (c) –100.0 °F
    (d) 0.0 °F (e) 100.0 °F (f) –45.0 °F

82. Convert each of the following temperatures to degrees Celsius.
    (a) 75.0 °F (b) 95.0 °F (c) –10.0 °F
    (d) 98.6 °F (e) 200. °F (f) –25.0 °F

83. Convert each of the following temperatures to kelvin.
    (a) 0.0 °C (b) 100. °C (c) –173. °C
    (d) 75.5 °C (e) 250. °C (f) –55. °C

84. Convert each of the following temperatures to kelvin.
    (a) 25.0 °C (b) 99.0 °C (c) –75.0 °C
    (d) 400. °C (e) 125. °C (f) –150. °C

85. Convert each of the following temperatures degrees Celsius.
    (a) 175. K (b) 298. K (c) 350. K
    (d) 25.0 K (e) 400. K (f) 50.0 K

86. Convert each of the following temperatures degrees Celsius.
    (a) 0.0 K (b) 75.0 K (c) 750. K
    (d) 173. K (e) 10.0 K (f) 1075. K

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87. The freezing point of water is 0.00 °C. Express this temperature in degrees Fahrenheit and kelvin.

88. A temperature of 98.6 °F is considered normal for the human body. Express this temperature in degrees Celsius and kelvin.

89. If a patient has a temperature of 38.1 °C, would it be considered a fever?

90. The highest recorded air temperature on earth was 56.7 °C on July 10, 1913 in Death Valley, CA. What is this temperature in degrees Fahrenheit? In kelvin?

91. The lowest naturally-occurring temperature on earth was −128.6 °F on July 21, 1983 at Vostok Station in Antarctica. What is this temperature in degrees Celsius? In Kelvin?

Extra Practice Problems

92. The dosage of a certain painkiller is 0.25 mg per kg of body weight. If a patient requires 9.70 mg of the painkiller, what is the patient's body weight in pounds?

93. Acetaminophen for children comes dissolved in a liquid containing 160 mg of medicine per 5.00 mL of liquid. One dose of acetaminophen is 15.0 mg of medicine/kg of body weight. If a child weighs 37.5 pounds and requires one dose, how many milliliters of liquid must the child be given?

94. An infant ibuprofen suspension contains 100 mg of medicine/5.0 mL of liquid. The recommended dosage of medicine is 10 mg of medicine/kg of patient body weight. How many mL of liquid should be given to an infant weighing 18 pounds?

95. A nurse gave a patient some liquid antihistamine. The dose was 0.25 mg of medicine/kg of body weight to be given twice a day. Over a 72-hour period, the nurse administered a total of 141.1 mg. What is the weight of the patient in pounds?

96. The speed of light is 1.86×10^5 mi $\cdot$ s^{-1}. How many millimeters does light travel in one year? Assume that the conversion factors given are exact. 1 year = 365.24 days; 1 day = 24 hr

97. A nerve impulse travels at a speed of 400 ft $\cdot$ s^{-1}. What is the speed of the impulse in mi $\cdot$ hr^{-1}?

98. You're planning to buy a new car. The car that you like gets 32 miles to a gallon of gasoline (32 miles/gallon). The car that your spouse likes gets 15 kilometers to the liter (15 km/liter). Which car has the better gas mileage?

99. One type of plane can travel from Washington DC to Tokyo, a distance of 6733 miles, in 13 hours and 55 minutes. A different type of plane can travel from San Francisco to Paris, a distance of 8965 km, in 9 hours and 55 minutes. Which plane travels as the highest speed?

100. A hiker finds a golden nugget along a trail but is unsure if it is gold or pyrite (FeS_2), often referred to as fool's gold. The nugget has a total volume of 67.5 mm^3. Gold has a density of 19.3 g/cm^3 and pyrite has a density of 5.01 g/cm^3. What should the mass of the nugget be if it is really gold? If it is pyrite?

101. A jeweler offers to sell a ring to a woman and tells her that it is made of platinum (density = 21.4 g/cm^3). The woman decides to perform a test to determine the ring's density. She places the ring on a balance and finds that it has a mass of 5.84 g. She then finds that the ring displaces 0.556 mL of water. Is the ring platinum?

102. The Honda Insight, a hybrid electric vehicle, has an EPA gas mileage rating of 57 miles/gallon in the city. How many kilometers can the Insight travel on the amount of gasoline that would fit in a soda can? The volume of a soda can is 355 mL. Gasoline has a density of 719.7 kg/m^3.

103. A moving truck has an EPA gas mileage rating of 6.5 miles/gallon of diesel fuel. How many kilometers can the moving truck drive without refueling if its fuel tank holds 675 liters? Diesel has a density of 0.832 kg/L.

Review: How to Succeed in Chemistry

104. Review the study plan that you created at the end of Chapter 1. Adjust as necessary and keep track of your study time over the next several weeks.

105. In each study session, did you complete the five specific tasks you set forth in your original study plan?

106. Create a word web for each of the bold-face words in this chapter. Be sure to include an example or definition for each word.

107. Reflect on your learning! Were you able to complete all of the practice problems for this chapter without looking at the solutions? Go back and rework difficult problems until you are confident with your answers.

108. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

109. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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Chapter 3
Matter

Chemistry involves the study of matter. Chemists study the composition of matter and the processes used to change one type of matter into another type of matter. The earliest human use of chemistry involved controlling fire to cook food, convert clay into pottery, and smelt metals. Today, chemists study the behavior and composition of matter by using sophisticated equipment and highly sensitive instruments. Research into specific types of matter and how matter behaves under various conditions plays a part in biology, medicine, materials science, geology, forensics, environmental science, and many other fields.

Matter is defined as anything that occupies space and has mass. Matter exists in many different forms including solids, liquids, and gases. In this chapter, you will learn how to describe and classify matter in several different ways. You will also learn to identify the processes used to change matter.

LEARNING OBJECTIVES

1. Describe the basic properties of solids, liquids, and gases.
2. Classify examples of matter as elements, compounds, and mixtures.
3. Classify changes in matter as either physical or chemical changes.
4. Apply the law of conservation of matter.
5. Explain the energy loss or gain accompanying a change in matter.

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3.1 States of Matter

The three common physical states of matter are solid, liquid, and gas (Figure 3.1). In the solid phase, particles of matter are packed tightly together. Solid matter contains particles that vibrate but remain in fixed positions. This means that solid matter has a definite volume and definite, rigid shape. Solids are incompressible and have a high density compared to gases. When a chemical is in the solid phase, its formula can be followed by the symbol (s).

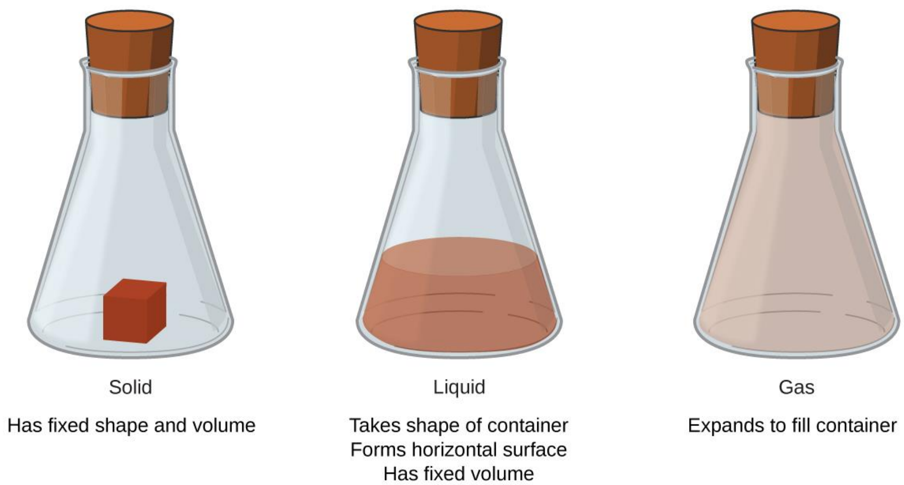

Figure 3.1 The three common states or phases of matter are solid, liquid, and gas. *Image credit: OpenStax Chemistry*.

In the liquid phase, matter also has a definite volume, but it takes the shape of the bottom of its container. Liquid particles are close together and non-compressible, just like in solids, but particles in a liquid can flow and move past one another with ease. For the same substance, the density of the liquid state is generally lower than the density of the solid state. A notable exception to this trend is ice and water where ice can be less dense than the water. Substances in the liquid phase can be denoted by the symbol (ℓ) following the chemical formula.

A gas will expand to fill its container. Gas particles are far apart and the distance between particles will increase until the container fills completely. In a gas sample, particles can be forced closer together by compression, but the density of gas remains very low compared to solids and liquids. The chemical formula for a gas can be followed by the symbol (g).

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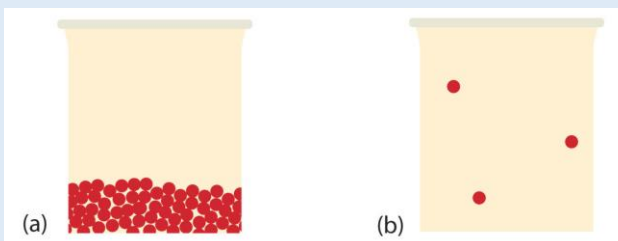
Table 3.1 States of Matter

Phase	Shape	Volume	Density	Compressibility
Solid (s)	definite rigid	definite	high	incompressible
Liquid (ℓ)	indefinite takes shape of container	definite	medium-high	incompressible
Gas (g)	indefinite fills container completely	indefinite	low	compressible

Example 3.1

Which state of matter does each of the following figures depict?



Answers

In (a), particles are in contact but are also able to move around each other. The sample shown is a liquid. In (b), most of the volume is empty space, and the particles are separated and freely moving about. The sample shown is a gas.

3.2 Descriptions of Matter

Matter is defined as anything that occupies space and has mass. The terms solid, liquid, and gas can be used as descriptors of the space occupied by matter. The **mass** of an object is used to describe the quantity of matter it contains. Mass is a fundamental property of an object unrelated to its location. Mass differs from an object's **weight**, which is a force that gravity exerts on an object. For instance, if the mass of Milli the Mole is 45 kg, she will have a mass of 45 kg on Earth and a mass of

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45 kg on the moon. Milli's weight on Earth, however, will be six times her weight on the moon (Figure 3.2).

Figure 3.2 Mass *vs.* Weight. Milli the Mole has a mass of 45 kg on Earth and mass of 45 kg on the moon. Milli's *weight* on Earth will be *six times* her weight on the moon. *Image credit. Milli is a trademark of the American Chemical Society. Earth and moon: nasa.gov.*

Substances *vs*. Mixtures

Chemists use many additional terms to describe matter and its composition. A sample of matter that has fixed chemical composition and characteristic properties is called a **substance**. Oxygen, for example, is a substance that is a colorless, odorless gas at 25 °C. Sodium chloride and glucose are also substances. By definition, a substance is pure because it contains only one type of formula unit, for example, atoms or molecules (Figure 3.3).

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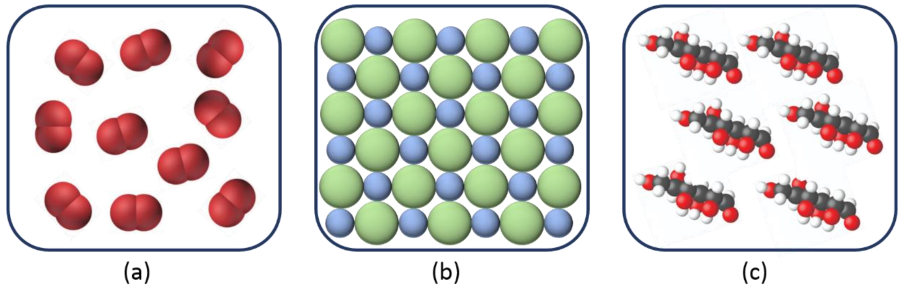

Figure 3.3 A substance is pure because it contains only one type of formula unit, for example, atoms or molecules. Substances depicted are (a) oxygen gas, made of molecules (b) sodium chloride, made of formula units and (c) glucose, made of molecules

A sample of air, however, is *not* a pure substance. Air contains a **mixture** of gases such as hydrogen, nitrogen, oxygen, argon, carbon dioxide, and several others. Any sample of matter that contains two or more substances is called a **mixture**. Individual substances within a mixture retain their properties and the substances can be present in variable proportions. A mixture can further be described as **homogeneous** when all components are in the same physical state, or phase, and have no visible boundaries. A sample of air contains many different gases, but its components cannot be visualized separately. Another example of a homogenous mixture is table salt dissolved in water. If no solid salt particles are visible, the mixture looks uniform throughout (Figure 3.4).

Homogeneous mixtures such as air and salt water are also referred to as **solutions**. Note that solutions do not have to be liquid! As was mentioned, air is a solution of gases. Solder is a solid solution made from lead and tin.

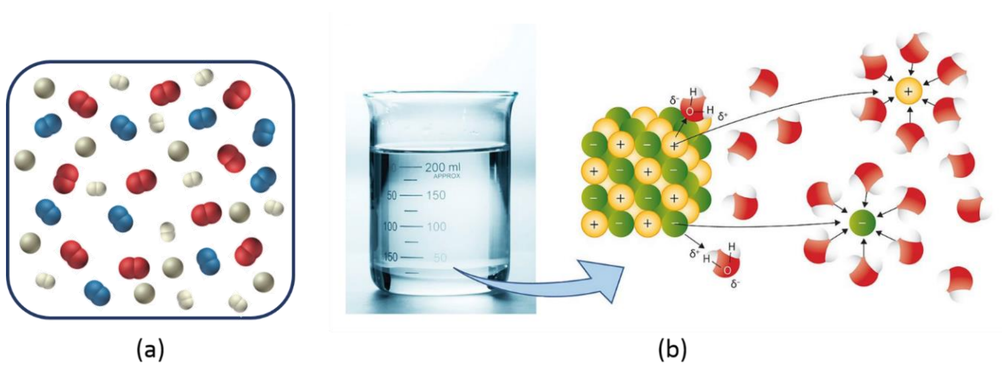

Figure 3.4 A mixture contains two or more substances. In the figure, (a) depicts a homogeneous mixture containing the gases hydrogen, nitrogen, oxygen, and argon. A mixture of table salt (sodium chloride) dissolved in water (b) is homogeneous because it is uniform throughout and there are no visible boundaries between the substances. Each of these homogeneous mixtures is also a *solution*. *Image credit: [Principles of General Chemistry](#).*

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Mixtures that are not completely uniform are termed **heterogeneous**. Mixtures such as ramen or soda appear heterogeneous to the eye (Figure 3.5). Both ramen and soda are mixtures because they contain different phases of matter. Some mixtures must be examined under a microscope to determine whether they are homogeneous or heterogeneous. Milk, for example, appears homogeneous at first glance, but under a microscope, tiny globules of fat and protein are visible. Because there are definite boundaries for the particles, milk is a heterogeneous mixture.

Figure 3.5 Noodle soup and carbonated water are heterogeneous mixtures. Image credit: *Wikimedia Commons.*

Elements *vs*. Compounds

Most mixtures can be separated into pure substances, which may be either elements or compounds. If a substance contains only one type of atom, it is an **element**. Examples of elemental substances are helium gas, oxygen gas, liquid mercury, solid copper wire, and a block of iron. An element cannot be broken down easily. Substances containing formula units made from two or more different kinds of atoms are called **compounds**. Carbon dioxide, for example, is a substance that is made up of identical molecules, each of which contains carbon and oxygen in a fixed proportion: one carbon atom to two oxygen atoms. Water is compound that contains one oxygen atom to two hydrogen atoms. Chemical processes can be used to break down and separate different types of atoms in a compound.

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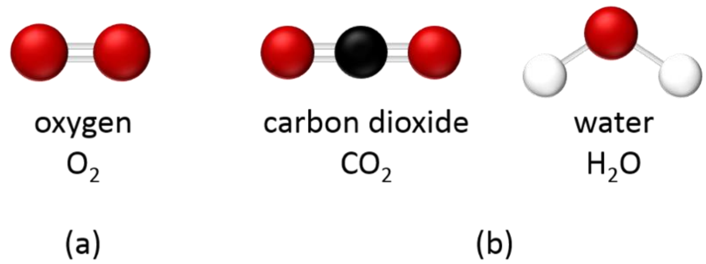

Figure 3.6 A sample of oxygen gas contains only oxygen atoms (a) and therefore is an elemental substance. Carbon dioxide and water (b) are compounds and contain elements in fixed proportions.

If a molecule consists of only one type of atom, it is an element. Elements such as hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine usually exist as pairs of atoms bound together. These elements are **diatomic** molecules and are represented as H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2, respectively. The element phosphorus usually exists as the molecule P_4, and elemental sulfur usually exists as the molecule S_8 (Figure 3.7).

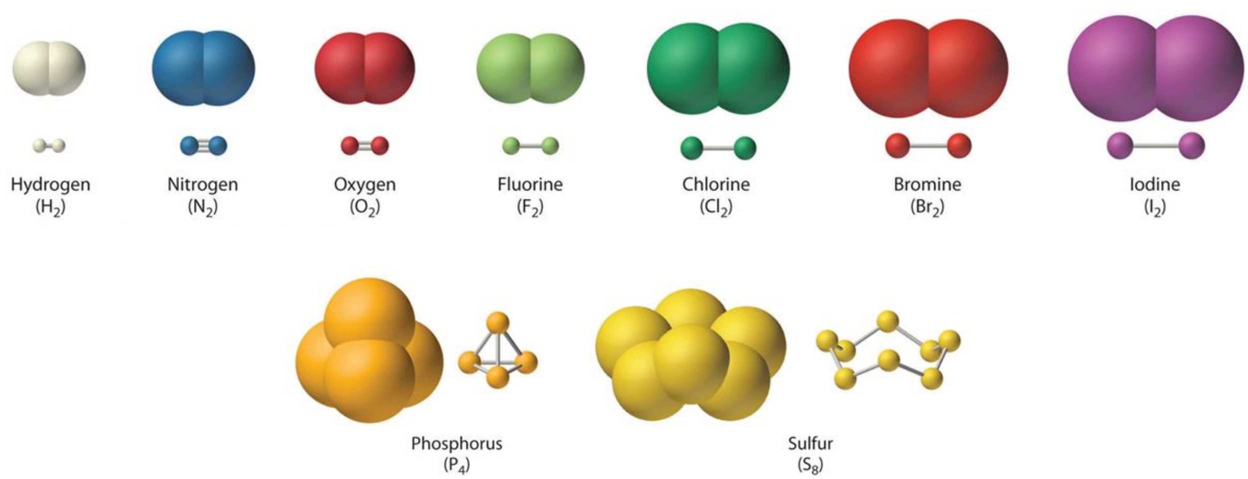

Figure 3.7 The elements hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine exist as pairs of atoms called diatomic molecules. The element phosphorus usually exists as a molecule containing four phosphorus atoms, and the element sulfur usually exists as a ring-shaped molecule containing eight sulfur atoms. *Image credit: OpenStax Chemistry.*

A flowchart showing a method for the classification of matter is shown in Figure 3.8.

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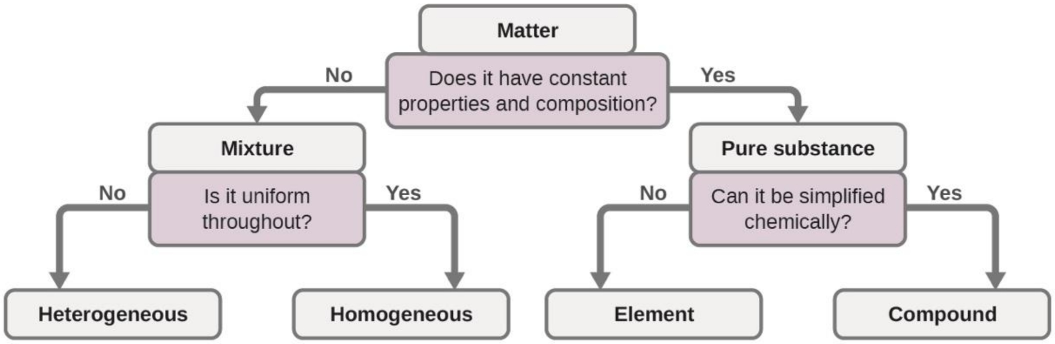

Figure 3.8 Matter can be classified according to the method shown. *Image credit: OpenStax Chemistry.*

Example 3.2

A picture of chocolate candies is shown below. What terms would you use to describe this sample of matter?

Image Credit Wikimedia Commons by Evan Amos.

Answer

The sample of matter has varied composition, so it should first be described as a *mixture*. The sample is not uniform and is therefore a *heterogeneous mixture*.

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Example 3.3

A sample of solid iodine contains dark purple, shiny crystals. What terms would you use to describe a sample of solid iodine?

Image Credit Wikimedia Commons by Benjah.

Answer

The sample has constant composition and is made of only one type of molecule, so it is a *pure substance*. It contains only one type of atom, iodine, so the sample is an *element*.

3.3 Descriptions of Matter: Properties and Changes

Matter can also be described and classified by referring to its **properties**. A **physical property** is a characteristic of matter that is not associated with a change in its chemical composition. Familiar examples of physical properties include density, color, hardness, melting and boiling points, and electrical conductivity.

Some physical properties, such as density and color, can be observed without changing the physical state of the matter being examined. Other physical properties, such as the melting temperature of iron or the freezing temperature of water, can only be observed as matter undergoes a physical change. A **physical change** involves a single substance and is a change in the state, form, or properties of matter without any accompanying changes in the chemical composition of the substances contained in the matter. A physical change occurs when ice melts and when dry ice sublimates into carbon dioxide gas (Figure 3.9). Common physical changes are summarized in Table 3.2.

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Table 3.2 Examples of Common Physical Changes

Process	Phase Change	Example
Melting	solid (s) → liquid (ℓ)	ice on a hot day
Freezing	liquid (ℓ) → solid (s)	a filled ice cube tray in a freezer
Boiling	liquid (ℓ) → gas (g)	water turning to steam when heated
Condensation	gas (g) → liquid (ℓ)	dew forming on grass
Sublimation	solid (s) → gas (g)	dry ice → carbon dioxide gas
Deposition	gas (g) → solid (s)	water vapor forming ice on windows

(a) (b)

Figure 3.9 (a) Water undergoes a physical change when ice is heated and forms liquid water. (b) At room temperature, dry ice undergoes sublimation and produced carbon dioxide gas. *Image credit: Principles of General Chemistry and Wikimedia Commons.*

The ability to change from one type of substance into another is a chemical property. Examples of chemical properties include flammability, digestibility, acidity, and reactivity. Iron, for example, combines with oxygen in the presence of water to form rust, but chromium does not react with oxygen readily. Nitroglycerin is very dangerous because it explodes easily, but neon is an unreactive gas.

A chemical property is identified by the **chemical change** that occurs when one type of matter is transformed into another type of matter. A chemical change always results in one or more substances that have different identities from the original sample. The process that takes place during a chemical change involves the making or breaking of chemical bonds within a substance. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The combustion of sugar is a chemical change because it produces carbon dioxide and water. Examples of chemical changes include redox reactions, all forms of combustion, and food being cooked.

During any change, whether physical or chemical, the **law of conservation of matter** applies. The law states:

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

Brewing beer and the operation of batteries provide examples of the conservation of matter. During the brewing of beer, the ingredients (water, yeast, grains, malt, hops, and sugar) are converted into beer (water, alcohol, carbonation, and flavoring substances) with no actual loss of matter. This is most clearly seen during the bottling process, when glucose turns into ethanol and carbon dioxide, and the total mass of the substances does not change. The conservation of matter can also be seen in a lead-acid car battery. The original substances (lead, lead oxide, and sulfuric acid), which are capable of producing electricity, are changed into other substances (lead sulfate and water) that do not produce electricity, with no change in the actual amount of matter (Figure 3.10).

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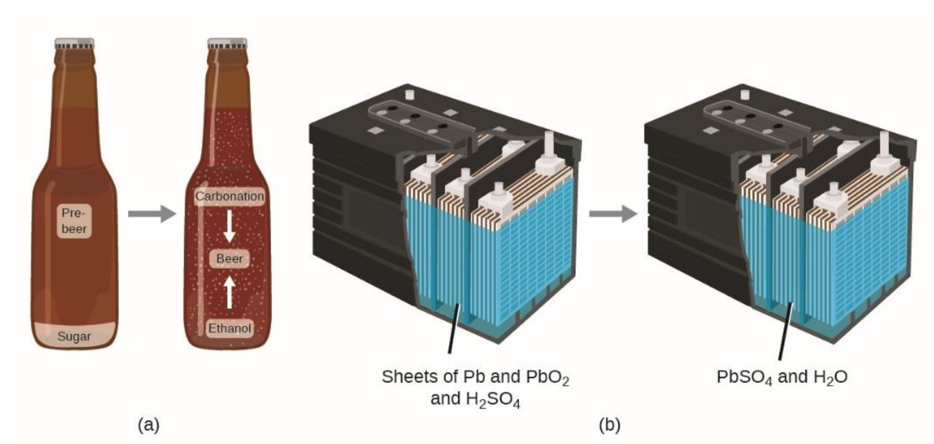

Figure 3.10 (a) The mass of materials used to make beer is the same as the mass of beer produced. Sugar becomes alcohol and carbonation during the brewing process. (b) The mass of lead, lead oxide plates, and sulfuric acid that goes into producing electricity from a car battery is exactly equal to the mass of lead(II) sulfate and water formed. *Image credit: OpenStax Chemistry.*

Extensive *vs*. Intensive Properties

Physical properties of matter fall into one of two categories. If the property depends on the amount of matter present, it is an **extensive property**. The mass and volume of a substance are examples of extensive properties; for instance, a gallon of milk has a larger mass and volume than a cup of milk. The value of an extensive property is directly proportional to the amount of matter in question. If the property of a sample of matter does not depend on the amount of matter present, it is an **intensive property**. Temperature is an example of an intensive property. If the gallon and cup of milk are each at 25 °C, when they are combined, the temperature remains at 25 °C. As another example, consider the distinct but related properties of heat and temperature. A drop of hot cooking oil spattered on your arm causes brief, minor discomfort, whereas a pot of hot oil yields severe burns. Both the drop and the pot of oil are at the same temperature (an intensive property), but the pot clearly contains much more heat (an extensive property).

Another example of extensive and intensive properties involves sulfur. Figure 3.11 shows extensive properties of samples of sulfur crystals and powder. Mass and volume are extensive properties and are related to the amount of material in the sample. The intensive properties of color and melting point are independent of the amount of the sample.

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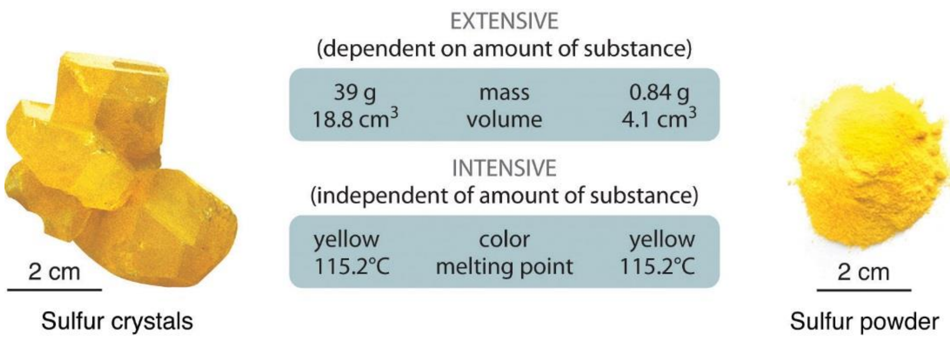

Figure 3.11 Because they differ in size, the two samples of sulfur have different extensive properties, such as mass and volume. In contrast, their intensive properties, including color and melting point are identical. *Image credit: Principles of General Chemistry.*

Example 3.4

Classify each of the following as a *physical* property or a *chemical* property. For physical properties, identify each as either *extensive* or an *intensive*.

(a) density (d) volume

(b) flammability (e) viscosity

(c) toxicity (f) melting point

Answers

(a) physical, intensive (d) physical, extensive

(b) chemical (e) physical, intensive

(c) chemical (f) physical, intensive

Example 3.5

Classify each of the following as a physical change or a chemical change.

(a) an ice cube melting (b) dry ice turning into carbon dioxide gas

(c) chopping wood for a campfire (d) a campfire burning (e) a steak cooking

Answers

(a) physical (b) physical (c) physical (d) chemical (e) chemical

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3.4 Matter and the Periodic Table

While many elements differ dramatically in their chemical and physical properties, some sets of elements have similar properties. We can identify sets of elements that exhibit common behaviors. For example, many elements conduct heat and electricity well, whereas others are poor conductors. These properties can be used to sort the elements into three classes: metals (elements that conduct well), nonmetals (elements that conduct poorly), and metalloids (elements that have properties of both metals and nonmetals).

The periodic table is a table of elements that places elements with similar properties close together. You will learn more about the periodic table as you continue your study of chemistry.

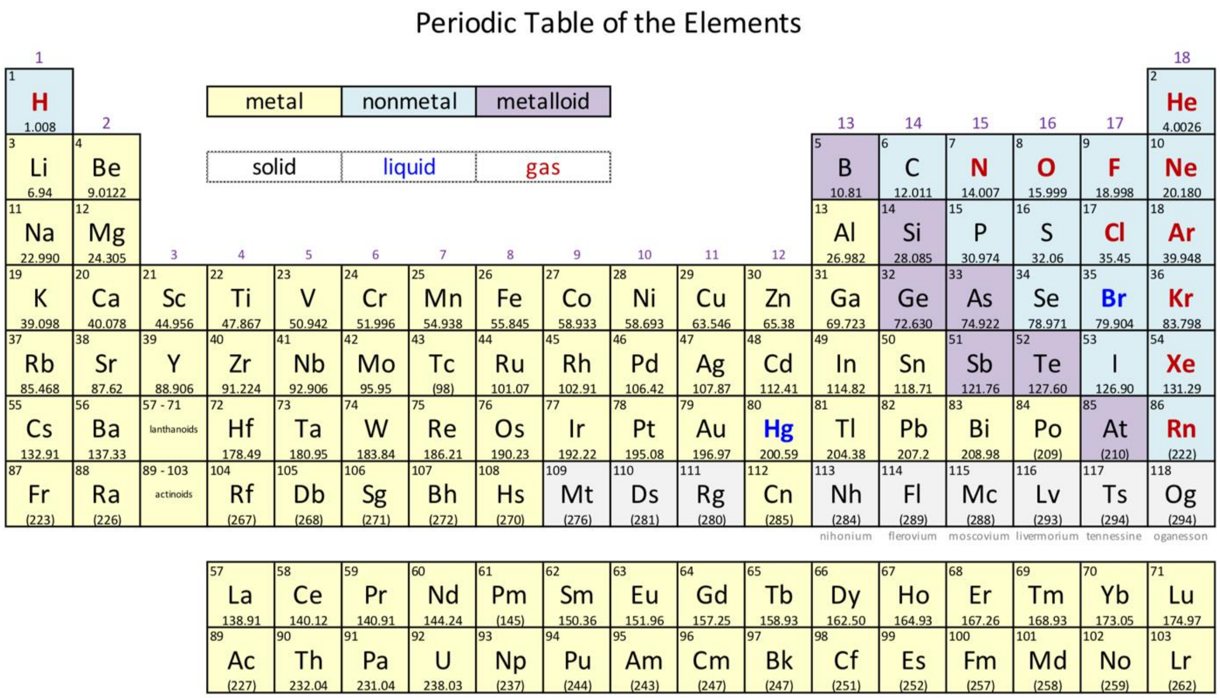

Figure 3.12 The periodic table shows how elements may be grouped according to certain similar properties. Note that the background color denotes whether an element is a metal, nonmetal, or metalloid, whereas the symbol color indicates whether it is a solid (black), liquid (blue), or gas (red) at standard laboratory conditions. *Based upon data from the International Union of Pure and Applied Chemistry (IUPAC).*

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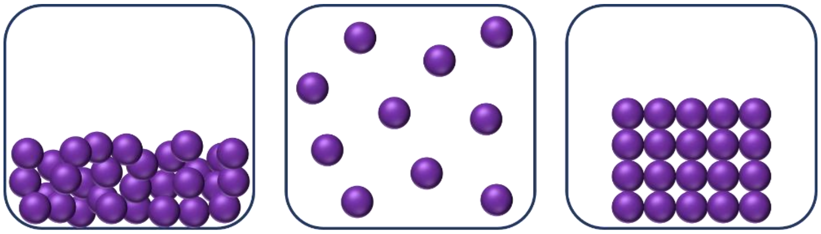
Chapter 3 Practice Problems

3.1 States of Matter

1. Classify each of the following as a solid, a liquid, or a gas.



(a) (b) (c)

2. Classify each of the following as a solid, a liquid, or a gas.

(a) (b) (c)

3. Which physical state has variable volume and variable shape?

4. Which physical state is compressible and has variable volume?

5. For each of the following descriptions, classify the phase of matter.
   (a) definite volume, indefinite shape
   (b) indefinite volume, indefinite shape
   (c) definite volume, definite shape

6. For each of the following descriptions, classify the phase of matter.
   (a) incompressible, definite volume, definite shape
   (b) compressible, indefinite volume, indefinite shape
   (c) incompressible, indefinite shape, definite volume

7. Think about the solids, liquids, and gases encountered in everyday life. What properties distinguish solids from liquids?

8. What properties distinguish liquids from gases? Solids from gases?

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3.2 Descriptions of Matter

9. Why do chemists use an object's mass, rather than its weight, to indicate the amount of matter it contains?

10. If a rock from the earth is taken to the moon, will its weight be smaller or larger on the moon? What about its mass? If the same rock is taken to Jupiter, will its weight be smaller or larger than on earth? What about its mass?

11. How does a heterogeneous mixture differ from a homogeneous mixture?

12. How does a homogeneous mixture differ from a pure substance?

13. Classify each of the following mixtures as heterogeneous or homogeneous.
(a) mineral water (b) freshly squeezed orange juice (c) table salt

14. Classify each of the following mixtures as heterogeneous or homogeneous.
(a) potting soil (b) chicken soup (c) distilled water

15. How does an element differ from a compound?

16. How does an atom differ from a molecule?

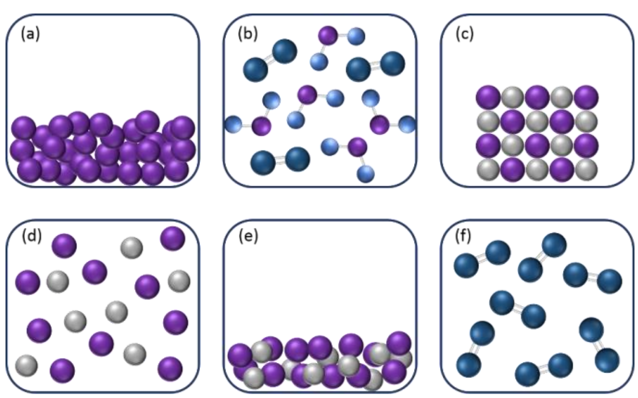
17. Does a molecule always have to be a compound? Explain.

18. Name three elements that exist as atoms rather than as molecules.

19. Classify each of the following as an element, a compound, or a mixture.
(a) copper (b) baking soda (c) sucrose

20. Classify each of the following as an element, a compound, or a mixture.
(a) air (b) distilled water (c) dry ice

21. Classify each of the following as an element, a compound, or a mixture.
Decide whether each figure represents a solid, a liquid, or a gas?



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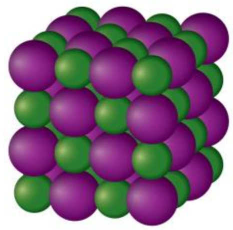
22. Does the picture below represent a solid, liquid, or gas? Explain.
    Does the picture represent an element or a compound? Explain.



3.3 Descriptions of Matter: Properties and Changes

23. Classify each of the following as a chemical or a physical change.
    (a) water freezing (b) TNT exploding (c) firewood burning (d) lead melting

24. Classify each of the following as a chemical or a physical change.
    (a) ice melting (b) meat cooking (c) wax melting (d) gasoline burning

25. Do the following descriptions concern a chemical property or a physical property?
    (a) shiny (b) soluble (c) corrosive (d) powdery

26. Do the following descriptions concern a chemical property or a physical property?
    (a)perishable (b) flammable (c) malleable (d) solid

27. Classify each of the following as a physical change or a chemical change.
    (a) $H_2O(\ell) \rightarrow H_2O(g)$ (b) $2\ NH_3(g) \rightarrow N_2(g) + 3\ H_2(g)$
    (c) $C(s) + 2\ H_2(g) \rightarrow CH_4(g)$ (d) $C_6H_6(\ell) \rightarrow C_6H_6(s)$

28. Classify each of the following as a physical change or a chemical change.
    (a) $C(s) + O_2(g) \rightarrow CO_2(g)$ (b) $CO_2(s) \rightarrow CO_2(g)$
    (c) $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$ (d) $CH_4(g) \rightarrow CH_4(\ell)$

3.4 Matter and the Periodic Table

29. List five metals from the periodic table. Write the symbol as well as the name of each element.

30. List five nonmetals from the periodic table. Write the symbol as well the name of each element.

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31. Which elements exist as gases at standard laboratory conditions?

32. Which elements exist as liquids at standard laboratory conditions?

33. Classify each of the following elements as metals/nonmetals/metalloids and solid/liquid/gas at standard laboratory conditions.
    (a) aluminum (b) sulfur (c) iron (d) antimony
    (e) iodine (f) nickel (g) magnesium (h) radon

34. Classify each of the following elements as metals/nonmetals/metalloids and solid/liquid/gas at standard laboratory conditions.
    (a) platinum (b) selenium (c) potassium (d) krypton
    (e) fluorine (f) titanium (g) calcium (h) tin

35. What is the physical state at room temperature and pressure for chlorine (Cl_2), bromine (Br_2), iodine (I_2), and carbon (C)?

36. Arsenic and germanium are used as semiconductors. What terms can be used to describe arsenic and germanium? Use the periodic table in Figure 3.8 to answer this question.

37. Name a diatomic element that is a solid at standard laboratory conditions?

38. Name a diatomic element that is a liquid at standard laboratory conditions?

Review: How to Succeed in Chemistry

39. Identify the chapters and sections that are difficult for you. Make sure to spend some extra review time on these topics and see a chemistry tutor if necessary.

40. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

41. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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Chapter 4
Atoms and Elements

The earliest recorded discussion of the basic structure of matter comes from ancient Greek philosophers, the scientists of their day. In the fifth century BC, Leucippus and Democritus argued that all matter was composed of small, finite particles that they called *atomos*, a term derived from the Greek word for "indivisible." They thought of **atoms** as moving particles that differed in shape and size, and which could join together. Later, Aristotle and others concluded that matter consisted of various combinations of the four "elements"—fire, earth, air, and water—and could be infinitely divided. Interestingly, these philosophers thought about atoms and "elements" as philosophical concepts, but it wasn't until the early 1800s that scientists performed experiments to develop an understanding of atoms, atomic structure, and how atoms interact to form clusters of atoms such as **molecules**.

LEARNING OBJECTIVES

1. Describe modern atomic theory.
2. Summarize the discoveries made by Thomson and Rutherford.
3. Become familiar with the components and the structure of the atom.
4. Use isotope notation.
5. Use isotope abundance to calculate average atomic mass.

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4.1 Atomic Theory

In 1807, John Dalton, and English schoolteacher, helped to revolutionize chemistry with his ideas that are now known as **Dalton's atomic theory**:

1. Matter is composed of exceedingly small, indivisible particles called **atoms.**
2. An **element** consists of only one type of atom, and all the atoms have identical properties.
3. Atoms of one element differ in properties from atoms of all other elements.
4. A compound consists of atoms of two or more elements combined in a small, whole-number ratio (Figure 4.1).

Dalton's atomic theory provides an explanation for many of the properties of matter that were presented in Chapter 3. For example, if elemental copper consists of only one kind of atom, it cannot be broken down into simpler substances. During a chemical reaction matter can be neither created nor destroyed; the total mass of matter present at the beginning of the chemical process will be the same as at the end of the process according to the **law of conservation of matter**.

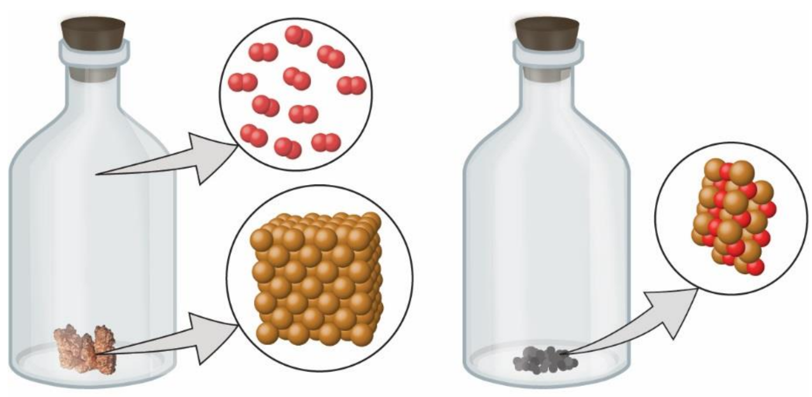

Figure 4.1 When the elements copper and oxygen react, atoms combine in a one-to-one ratio to produce copper(II) oxide, CuO. *Image credit: OpenStax Chemistry.*

In the late 1800s, a number of scientists probing the structure of the atom investigated the electrical discharges that could be produced in low-pressure gases, with the most significant discovery made by English physicist **J. J. Thomson** in 1897. Thomson's apparatus consisted of a sealed glass tube containing two metal electrodes. He noted that an unusual form of energy was emitted from the

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cathode, the negatively charged electrode. Thomson proved that some subatomic particles could be deflected, or bent, by magnetic or electric fields, suggesting that the particles were charged (Figure 4.2).

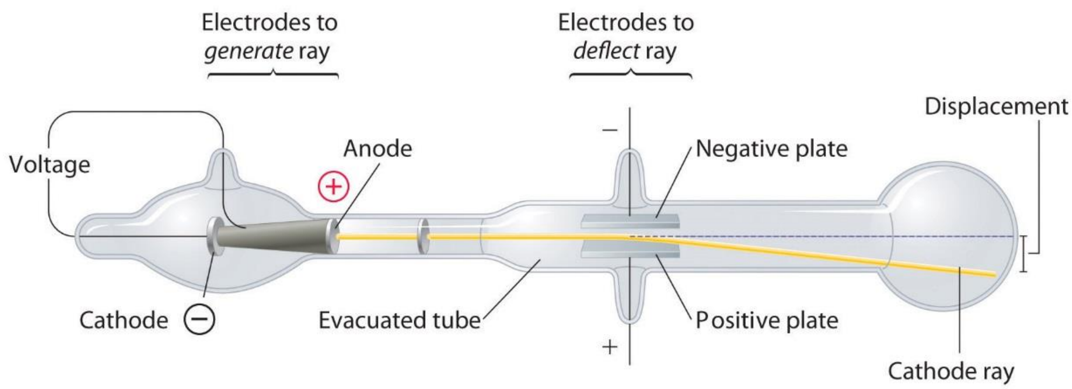

Figure 4.2 As the beam emitted from the cathode travels toward the right, particles, now known as electrons, are deflected toward the positive electrode, demonstrating that they are negatively charged. *Image credit: Principles of General Chemistry.*

Based on his observations, Thomson concluded that the particles were attracted by positive (+) charges and repelled by negative (−) charges, so they must be negatively charged. He also noted that the particles were the same no matter what metal he used for the electrodes, so they must be fundamental, subatomic constituents of all atoms. Although controversial at the time, Thomson's idea was gradually accepted, and his cathode ray particle is what we now call an **electron**, a negatively charged, subatomic particle with a mass more than one thousand-times less that of an atom.

In 1909, more information about the electron was uncovered by American physicist **Robert A. Millikan** via his oil drop experiments. Millikan created microscopic oil droplets, which could be electrically charged as they formed. These droplets initially fell due to gravity, but their downward progress could be slowed or even reversed by an electric field present in the apparatus. By adjusting the electric field strength and making careful measurements and appropriate calculations, Millikan was able to determine the charge on individual drops.

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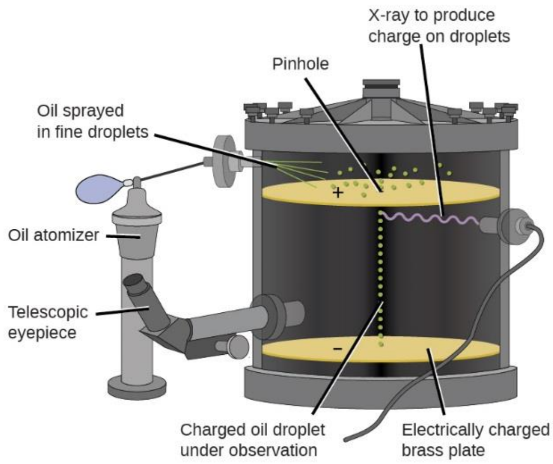

Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

Figure 4.3 Millikan's experiment measured the charge of individual oil drops. The charge of a particular oil drop is always a multiple of 1.6 x 10^{-19} coulombs. Millikan concluded that this value must be the fundamental charge of an electron.
Image credit: OpenStax Chemistry.

Looking at the charge data that Millikan gathered, you may have recognized that the charge of an oil droplet is always a multiple of a specific charge, 1.6 x 10^{-19} coulombs. Millikan concluded that this value must therefore be a fundamental charge—the charge of a single electron. When combined with the results of Thomson's research the mass of a single electron was determined to be 9.109 x 10^{-31} kg.

Thomson and Millikan established that the atom was not indivisible as Dalton had believed, but contained at least one subatomic particle, the negatively charged electron. Because atoms are neutral, some part of the atom had to be positively charged, but the positive charge had yet to be explained. In 1904, Thomson proposed the **plum pudding** model of an atom, which described a positively charged mass with an equal amount of negative charge in the form of electrons embedded in it. At about the same time, a Japanese scientist named **H. Nagaoka** proposed a **Saturnian** model of the atom consisting of a positively charged sphere surrounded by a halo of electrons (Figure 4.4).

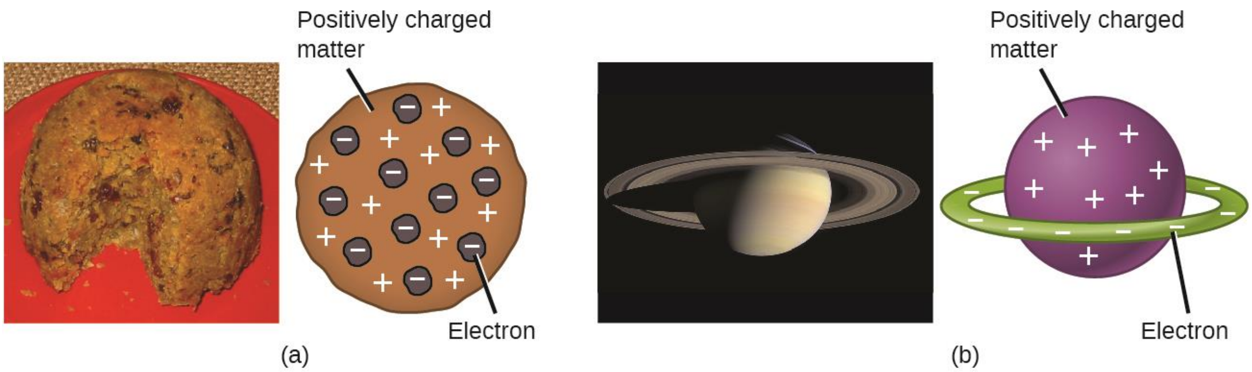

Figure 4.4 (a) Thomson proposed an atomic model that resembled plum pudding, with positive and negative charges distributed evenly. (b) Nagaoka proposed a model that resembled the planet Saturn with a positively charged center with a ring of electrons. *Image credit: OpenStax Chemistry.*

The next major development in understanding the atom came from **Ernest Rutherford**. He performed a series of experiments using a beam of high-speed, positively charged particles, called alpha particles, that were produced by the radioactive decay of radium. Rutherford aimed a beam of the positively charged particles at a very thin piece of gold foil. The resultant scattering was observed by using a luminescent screen that glowed briefly where the particles hit (Figure 4.5).

Most particles passed right through the foil without being deflected at all, meaning that atoms were largely empty space. Some particles, however, were deflected slightly, and a very small number had large deflections. Rutherford's observations were at odds with Thomson's plum pudding model of the atom where mass is distributed evenly throughout an atom. Instead, his results indicated that the mass of a gold atom is highly concentrated in a very small area of space, which he termed the **nucleus**.

In 1920, Rutherford established that the nucleus of the hydrogen atom was a positively charged particle, which he called a **proton**. He also deduced that, in order to obtain the correct mass for an atom, the nucleus had to contain electrically neutral particles with the same approximate mass as the proton. The **neutron**, however, was not discovered until James Chadwick, a student of Rutherford, discovered it in 1932.

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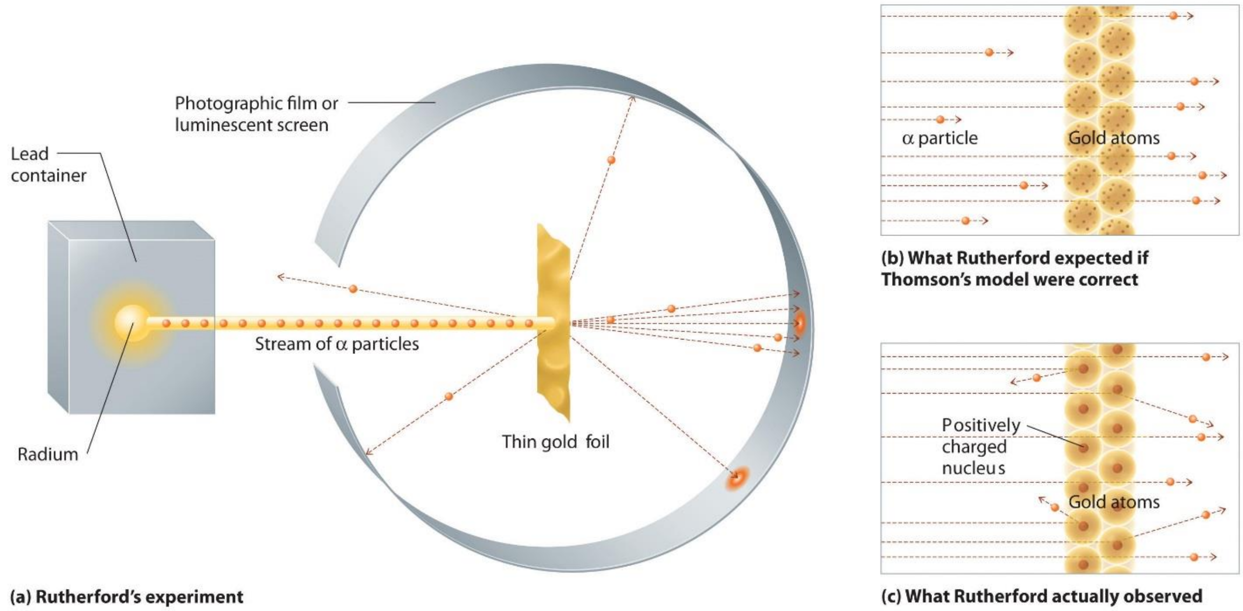

Figure 4.5 (a) A representation of the apparatus Rutherford used to detect deflections in a stream of α particles aimed at a thin gold foil target. The particles were produced by a sample of radium. (b) If Thomson's model of the atom were correct, the α particles should have passed straight through the gold foil. (c) But a small number of α particles were deflected in various directions, including right back at the source. This could be true only if the positive charge were much more massive than the α particle. It suggested that the mass of the gold atom is concentrated in a very small region of space, which he called the nucleus. *Image credit: Principles of General Chemistry.*

4.2 Atomic Structure and Symbolism

The development of modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume (Figure 4.6).

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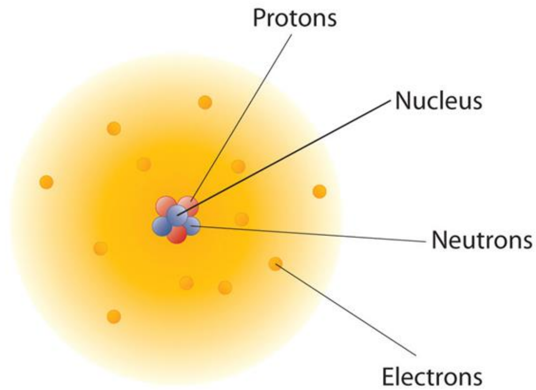

Figure 4.6 An atom has a nucleus containing protons and neutrons, and therefore the majority of its mass, in a very small area of space. The volume of an atom is largely occupied by electrons. *Image credit: Principles of General Chemistry.*

Atoms—and the protons, neutrons, and electrons that they contain—are extremely small in size and mass. For example, one carbon atom is about 1.4 x 10^{-12} meters across and weighs less than 2 x 10^{-23} g. When describing the properties of tiny objects such as atoms, we use an appropriately small unit of measure known as the atomic mass unit (**amu**). The Dalton (**Da**) is a measure often used in biology and biochemistry and is equivalent to one amu.

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087 amu and is neutral with a charge of zero. The electron has a charge of 1− and is a much lighter particle with a mass of about 0.00055 amu. The properties of these fundamental particles are summarized in Table 4.1.

Table 4.1 Properties of Subatomic Particles

Particle	Location	Symbol	Unit Charge	Mass (amu)	Mass (g)
Proton	nucleus	p or p^+	1+	1.00727	1.67262 x 10^{-24}
Neutron	nucleus	n or n^o	0	1.00866	1.67493 x 10^{-24}
Electron	outside nucleus	e^-	1−	0.00055	0.00091 x 10^{-24}

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Example 4.1

A proton has a mass of 1.00727 amu. How many electrons (0.00055 amu) would it take to equal the mass of one proton?

Solution

$$1.00727\ \cancel{\text{amu}} \times \frac{1\ \text{electron}}{0.00055\ \cancel{\text{amu}}} = 1800\ \text{amu} = 1.8 \times 10^3\ \text{amu}$$

Example 4.2

The enzyme methane monooxygenase (MMO) is an iron-containing protein that bacteria use to convert methane into methanol. The mass of MMO is 245 kDa (kilodaltons). What is the mass of MMO in amu?

Solution

$$245\ \cancel{\text{kDa}} \times \frac{1000\ \cancel{\text{Da}}}{1\ \cancel{\text{kDa}}} \times \frac{1\text{amu}}{1\ \cancel{\text{Da}}} = 245{,}000\ \text{amu} = 2.45 \times 10^5\ \text{amu}$$

Chemical Symbols

A **chemical symbol** is the abbreviation used indicate an element or an atom of an element. For example, the symbol for mercury is Hg. The symbols for several common elements and their atoms are listed in Table 4.2. Some symbols are derived from the common name of the element; others are abbreviations of the name in Latin, Greek, or another language. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the notation for the compound carbon monoxide, which contains atoms of the elements carbon (C) and oxygen (O). All elements that have been discovered thus far are shown in the periodic table in Figure 4.7.

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Table 4.2 Selected Elements and Symbols

Element	Symbol
Aluminum	Al
Calcium	Ca
Cobalt	Co
Copper	Cu (from *cuprum*)
Gold	Au (from *aurum*)

Element	Symbol
Iron	Fe (from *ferrum*)
Lead	Pb (from *plumbum*)
Magnesium	Mg
Silver	Ag (from *argentum*)
Tin	Sn (from *stannum*)

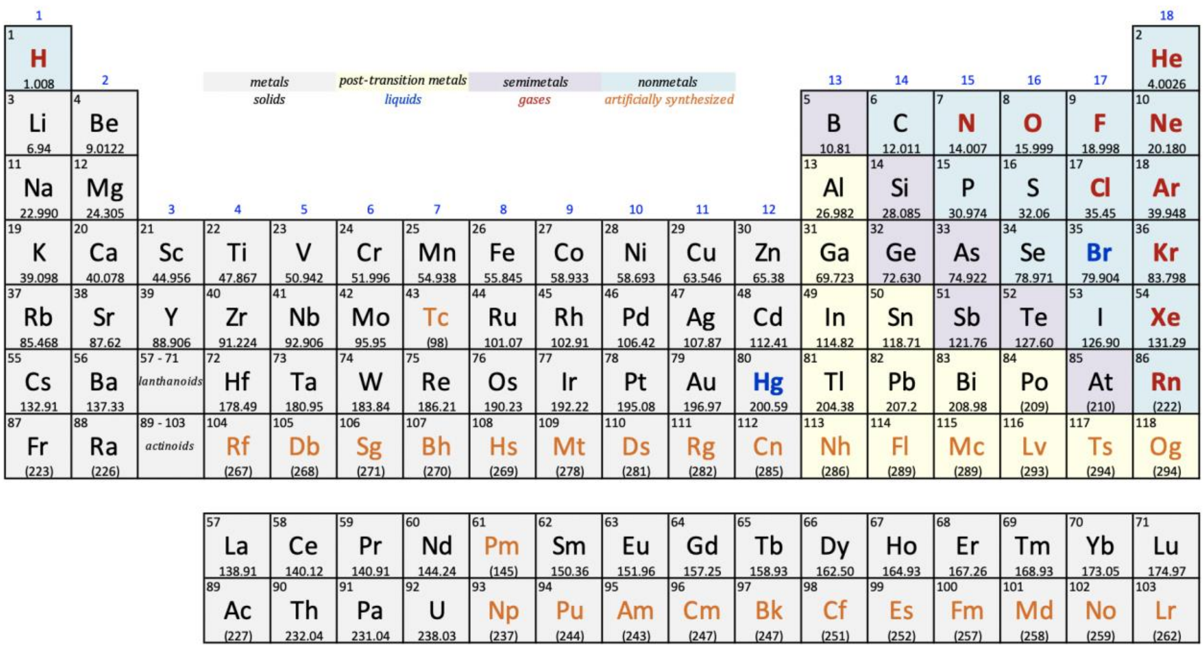

Figure 4.7 The Periodic Table of the Elements. The physical states of the elements are depicted for standard laboratory temperature and pressure. The search for elements beyond 118 has begun and researchers predict that 119 and 120 will be found in the next five years. *Based upon data from the International Union of Pure and Applied Chemistry (IUPAC).*

Protons, Neutrons, and Electrons

The number of protons in the nucleus of an atom is its **atomic number (Z),** which is the number of the element as listed on the periodic table. *The atomic number is different for each element, and the elements are arranged in order of increasing Z.* For example, any atom that contains 25 protons has an atomic number of 25 and must be the element manganese. The number of neutrons may vary for different isotopes of manganese and the number of electrons may vary for different charges of manganese, but any atom or ion containing 25 protons will be some version of manganese.

A neutral atom must have the same number of protons (positive charges) and electrons (negative charges). The atomic number for a neutral atom therefore indicates the number of electrons in that atom.

The total number of protons and neutrons in an atom is called the **mass number (A)**. A proton has a mass of ~1 amu and a neutron has a mass of ~1 amu; the mass number is an integer indicating the total count of protons plus neutrons. Recall that electrons have a very small mass compared to protons and neutrons and do not contribute to the mass number. As a result the mass of an atom and an ion of the same element are approximately the same.

Isotopes

The symbol for an isotope of any element, sometimes called the full atomic symbol, is written by using the atomic symbol for the element as listed on the periodic table (X), the mass number (A) and the number of protons (Z). An isotope of carbon (Z = 6) that has a mass number of 12 (A = 12) would be represented by the symbol ${}^{12}_{6}C$ and is shown in Figure 4.8.

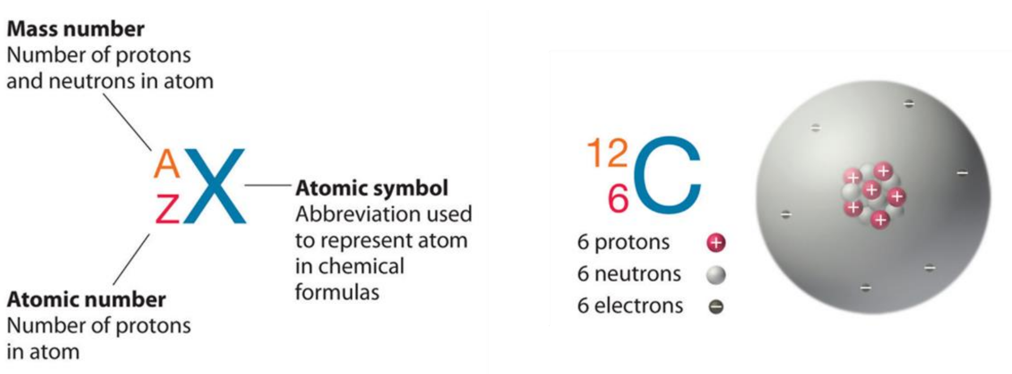

Figure 4.8 The symbol used to represent an isotope involves the atomic symbol from the periodic table (X), the mass number (A), and the number of protons (Z). Z is also the atomic number for the element as listed on the periodic table.
Image credit: Principles of General Chemistry.

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Example 4.3

Fill in the table below:

	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
A			
Z			
X			
# protons			
# neutrons			
# electrons			

Answers

These are isotopes of carbon. Each has Z = 6, but the number of neutrons, and therefore the mass number varies.

	$^{12}_{6}C$	$^{13}_{6}C$	$^{14}_{6}C$
A	12	13	14
Z	6	6	6
X	C	C	C
# protons	6	6	6
# neutrons	6	7	8
# electrons	6	6	6

Example 4.4

Write the symbol for each of the following isotopes.

(a) carbon with six protons and seven neutrons

(b) the neutral isotope with five protons and six neutrons

(c) iron with a mass number of 57

(d) the neutral isotope with a mass number of 238 and 92 protons

(e) chlorine with 18 neutrons

(f) the neutral isotope with seven electrons and eight neutrons

Solutions

(a) $^{13}_{6}C$ (b) $^{11}_{5}B$ (c) $^{57}_{26}Fe$ (d) $^{238}_{92}U$ (e) $^{35}_{17}Cl$ (f) $^{15}_{7}N$

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If the atomic symbol is given, the number of protons (Z) provides redundant information and is sometimes omitted. For example, magnesium has an atomic number (Z) of 12, and an isotope of magnesium with a mass number (A) of 24 can be written as $^{24}_{12}Mg$ or simply as ^{24}Mg because Z = 12 is implied from the atomic symbol. In print, the isotope might also be listed as 24-Mg. When read aloud, ^{24}Mg or 24-Mg would be called "magnesium 24."

PhET Interactive Simulation for Section 4.2

Practice isotope symbols and concepts with *Build an Atom.*

https://phet.colorado.edu/en/simulation/build-an-atom

4.3 Average Atomic Mass

Because each proton and each neutron contribute approximately one amu to the mass of an atom, and each electron contributes far less, the **atomic mass** of a single atom is *approximately* equal to its mass number (a whole number). Inspection of the atomic masses listed on the periodic table, reveals that the average masses of atoms of most elements are *not* whole numbers. The average atomic mass of an element shown in a periodic table is a **weighted average** of the masses of all the isotopes present in a naturally occurring sample of that element. The weighted average can be calculated from the mass of each isotope and the fractional abundance of that isotope (Equation 4.1) where m_1 is the mass of the first isotope and a_1 is the fractional abundance of the first isotope. The fractional bundance of each isotope is 1/100 of the percentage abundance. As an example, if an isotope has an abundance of 30.0%, it has a fractional abundance of 0.300. The terms are repeated for each of the naturally occurring isotopes of the element.

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$$average\ atomic\ mass = m_1a_1 + m_2a_2 + m_3a_3 + \cdots \qquad \text{(Equation 4.1)}$$

For example, the element boron is composed of two isotopes as shown in Table 4.3. The average atomic mass for boron can be calculated by using Equation 4.1 with only two terms because there are only two naturally occurring isotopes.

$$average\ atomic\ mass = m_1a_1 + m_2a_2$$

Table 4.3 Isotope Masses and Abundances for Boron

Isotope	Mass (amu)	Abundance (%)	Fractional Abundance
^{10}B	10.0129	19.9%	0.199
^{11}B	11.0093	80.1%	0.801

$$average\ atomc\ mass = m_1a_1 + m_2a_2$$

$$average\ atomic\ mass = (10.0129\ \text{amu})(0.199) + (11.0093\ \text{amu})(0.801)$$

$$average\ atomic\ mass = (1.99\ \text{amu}) + (8.82\ \text{amu}) = \mathbf{10.81\ amu}$$

Example 4.5

Magnesium has three naturally occurring isotopes: ^{24}Mg, ^{25}Mg, and ^{26}Mg. The isotope masses and abundances are given in the data table below.

Isotope Masses and Abundances for Magnesium

Isotope	Mass (amu)	Abundance (%)	Fractional Abundance
^{24}Mg	23.9850	78.99%	0.7899
^{25}Mg	24.9858	10.00%	0.1000
^{26}Mg	25.9826	11.01%	0.1101

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Solution

Magnesium has three stable isotopes, so the average atomic mass equation will have three terms:

$average\ atomic\ mass = m_1a_1 + m_2a_2 + m_3a_3$

$$average\ atomic\ mass = (23.9850\ \text{amu})(0.7899) + (24.9858\ \text{amu})(0.1000) + (25.9826\ \text{amu})(0.1101)$$

$average\ atomic\ mass = (18.95\ \text{amu}) + (2.499\ \text{amu}) + (2.860\ \text{amu}) = \mathbf{24.31\ amu}$

Check the periodic table. Does this average atomic mass make sense?

There are some elements whose average atomic masses are very close to whole numbers. For example, carbon has an atomic mass of 12.011 amu and helium has an atomic mass of 4.0026 amu. This usually occurs when there is one isotope that is so much more abundant than any other isotope that the others are essentially negligible in calculating the average atomic mass. Naturally occurring carbon is overwhelmingly carbon 12 and helium is overwhelmingly helium 4.

PhET Interactive Simulation for Section 4.3

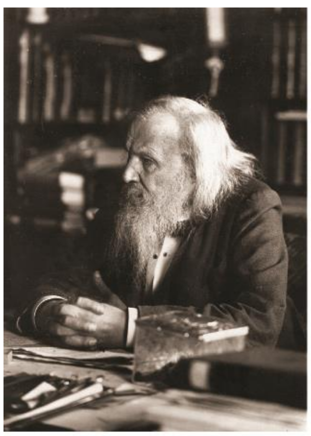
Practice average atomic mass calculations with *Isotopes and Atomic Mass.*

https://phet.colorado.edu/en/simulation/isotopes-and-atomic-mass

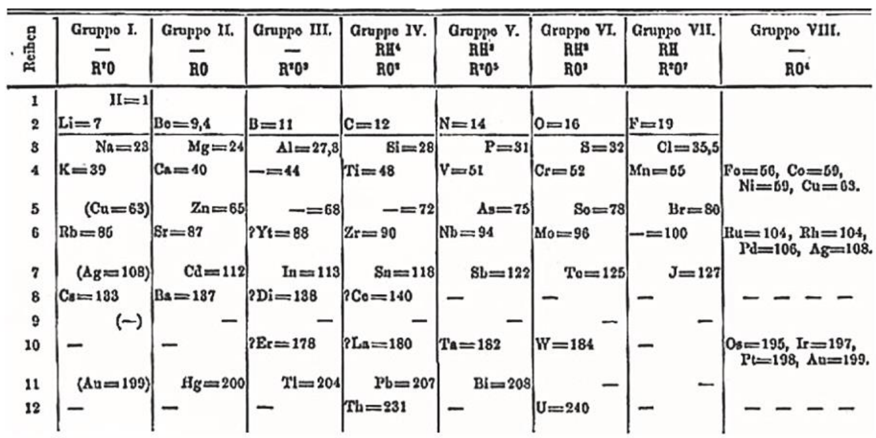
4.4 Atoms and the Periodic Table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped by similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K). These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive than are Ca, Sr, and Ba. Li, Na, and K form compounds with oxygen in a ratio of two metal atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one metal atom to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of the groups mentioned above.

In 1869, Dimitri **Mendeleev** recognized that there was a periodic relationship among the properties of the elements known at that time. He published tables with the elements arranged according to increasing atomic mass. Then he used his table to predict the existence of elements that would have the properties similar to aluminum and silicon but were yet unknown. The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work, and eventually lead to the widespread adoption of his periodic arrangement of the elements.



(a)



Reihen	Gruppo I. — R^2O	Gruppo II. — RO	Gruppo III. — R^2O^3	Gruppo IV. RH^4 RO^2	Gruppo V. RH^3 R^2O^5	Gruppo VI. RH^2 RO^3	Gruppo VII. RH R^2O^7	Gruppo VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

(b)

Figure 4.9 Mendeleev's early periodic table had gaps where he predicted new elements would fit in when they were discovered. *Image credit: Principles of General Chemistry.*

In 1913 Henry G. J. Moseley presented a new and more accurate road map of the elements. Versions of Moseley's periodic table now hang in almost every chemistry classroom in the world. His new table was not based on atomic weights of the elements as was Mendeleev's, but his table was organized on the atomic numbers of the elements, or the number of protons in the atoms of each element. The discovery of this rhythm, or periodicity, in the physical and chemical properties of the elements has proven invaluable to the progress of chemistry.

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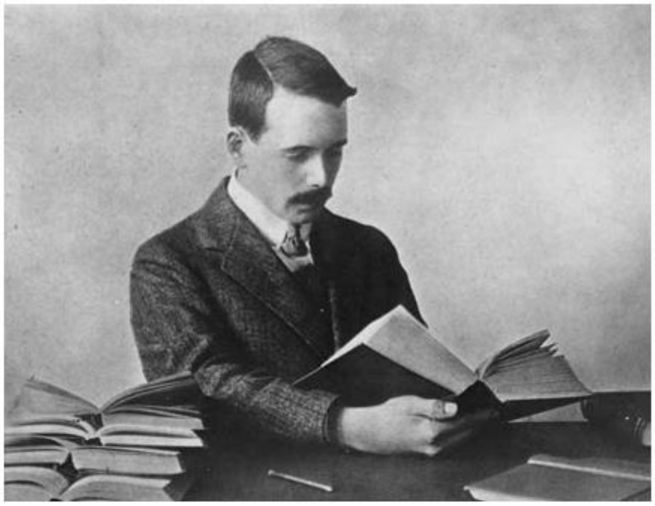

Figure 4.10 The periodic table used in modern chemistry was developed by Henry G. J. Mosely. His insight into the periodicity of the elements was groundbreaking. Sadly, Moseley was killed at the age of 28 during the Battle of Gallipoli during World War I.
Image credit: Royal Society of Chemistry.

As was mentioned in the previous chapter (Figure 3.8), elements having common properties can be sorted in three general groups on the periodic table: **metals** (elements that are shiny, malleable, good conductors of heat and electricity); **nonmetals** (elements that appear dull, poor conductors of heat and electricity); and **metalloids** (elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals).

The elements can further be classified into groups defined by similar reactivity and the propensity to form specific types of compounds. The columns of the periodic table are known as **groups** while the rows are called **periods**. Group 1, the **alkali metals**, contains elements such as lithium, sodium, and potassium. Group 2 elements are known as the **alkaline earths**. The terms *alkali* and *alkaline* refer to the basic solutions that form when the metals or their oxides react with water. Other groups with specific names include Group 15 (**pnictogens**), Group 16 (**chalcogens**), and Group 17 (**halogens**). Group 18 consists of inert gases known as the **noble gases**.

Groups 3 through 12 are the **transition metals**. These metals have similar properties within each group. For instance, vanadium (V), niobium (Nb), and tantalum (Ta) are often studied together due to similar chemical reactivity. Near this area of the periodic table are **rare earths** or **inner**

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transition metals, also known as the **lanthanides** and the **actinides**. Two common terms for groups 1-2 and groups 13-18 are the **representative elements** or **main group elements**. All of the known elements are either representative elements, transition metals, or rare earths. It is generally best to use the most specific term you can for a given set of elements. For example, “fluorine, chlorine, and bromine” are halogens while fluorine, oxygen, and silicon” are main group elements.

The location of these groups on the periodic table will prove very useful when predicting how neutral atoms of each element react to form either positively charged **cations** or negatively charged **anions**.

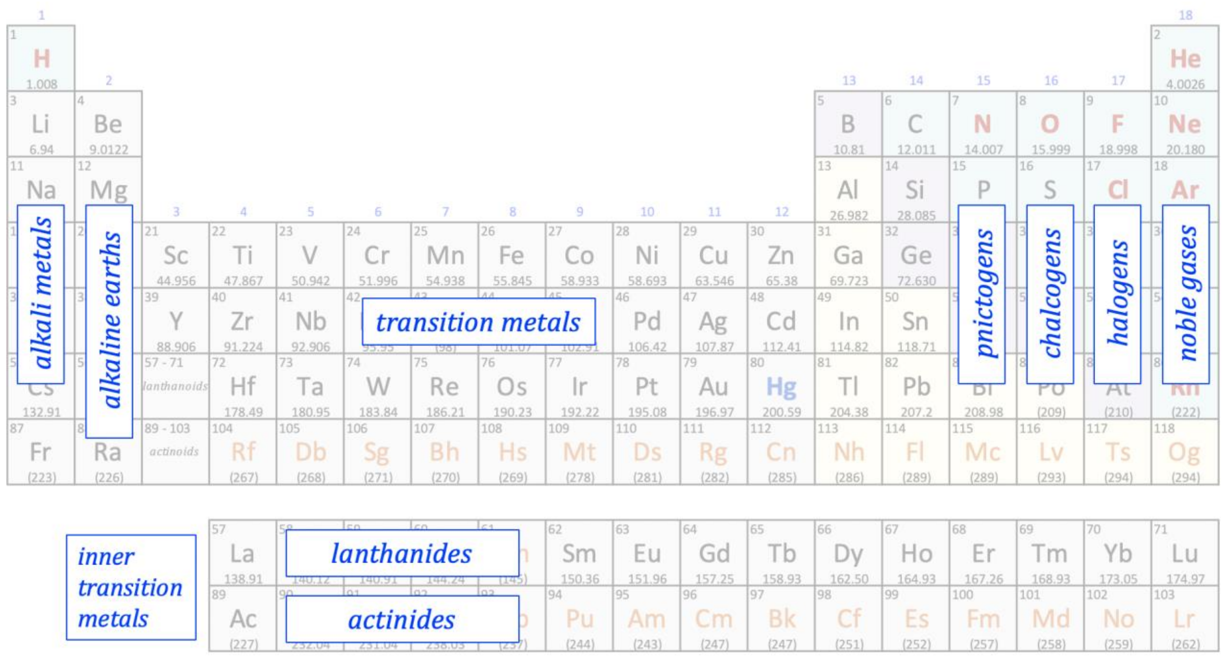

Figure 4.11 Elements on the periodic table can be classified as metals, nonmetals, and metalloids. The elements can further be classified into groups defined by similar reactivity and the propensity to form specific types of compounds.

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4.5 Ion Formation

The **charge** of an atom is defined as the number of protons minus the number of electrons. Atoms acquire charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an **anion**. An atom that loses one or more electrons exhibits a positive charge and is called a **cation**.

For example, a neutral sodium atom has Z = 11 and therefore 11 protons and 11 electrons. If a neutral sodium atom loses one electron, it becomes a cation with a 1+ charge (11 protons – 10 electrons = 1+ overall ion charge). A neutral oxygen atom has Z=8 and therefore has eight protons and eight electrons. If the oxygen atom gains two electrons it will become an anion with a 2− charge (8 protons – 10 electrons = 2− overall charge).

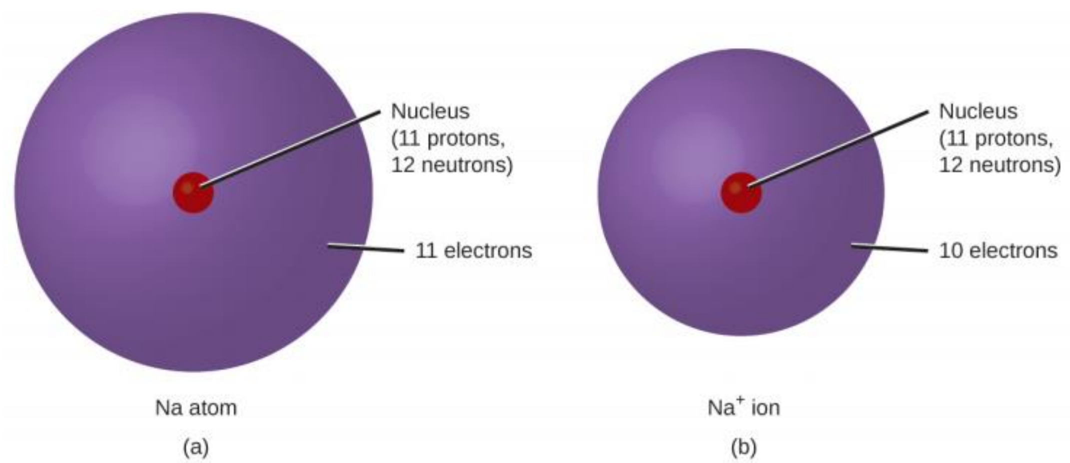

Figure 4.12 A neutral atom of sodium, containing 11 protons and 11 electrons can lose one electron to become the cation Na^+. *Image credit: Principles of General Chemistry*.

The periodic table can be used to predict the formation and resulting charges of many anions and cations. Elements in the alkali metals, the alkaline earths, and the nonmetals other than the noble gases gain or lose a specific number of electrons to leave them with the same number of electrons, or the same *electron configuration*, as an atom of the nearest noble gas. For instance, sodium, an alkali metal, contains 11 electrons and loses one of these electrons to form Na^+. The cation Na^+ has 10 electrons, the same number of electrons as neon, a noble gas. Noble gases have a particularly

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stable arrangement of electrons and this is the reason that atoms in the alkali metal group tend to be most stable as cations with a 1+ charge.

A neutral calcium atom, with 20 protons and 20 electrons, readily loses _two_ electrons. This results in a cation with 20 protons, 18 electrons, and a 2+ charge. The calcium cation, Ca^{2+} has the same number of electrons as the noble gas argon and is particularly stable. Similarly, aluminum typically loses three electrons to become Al^{3+}.

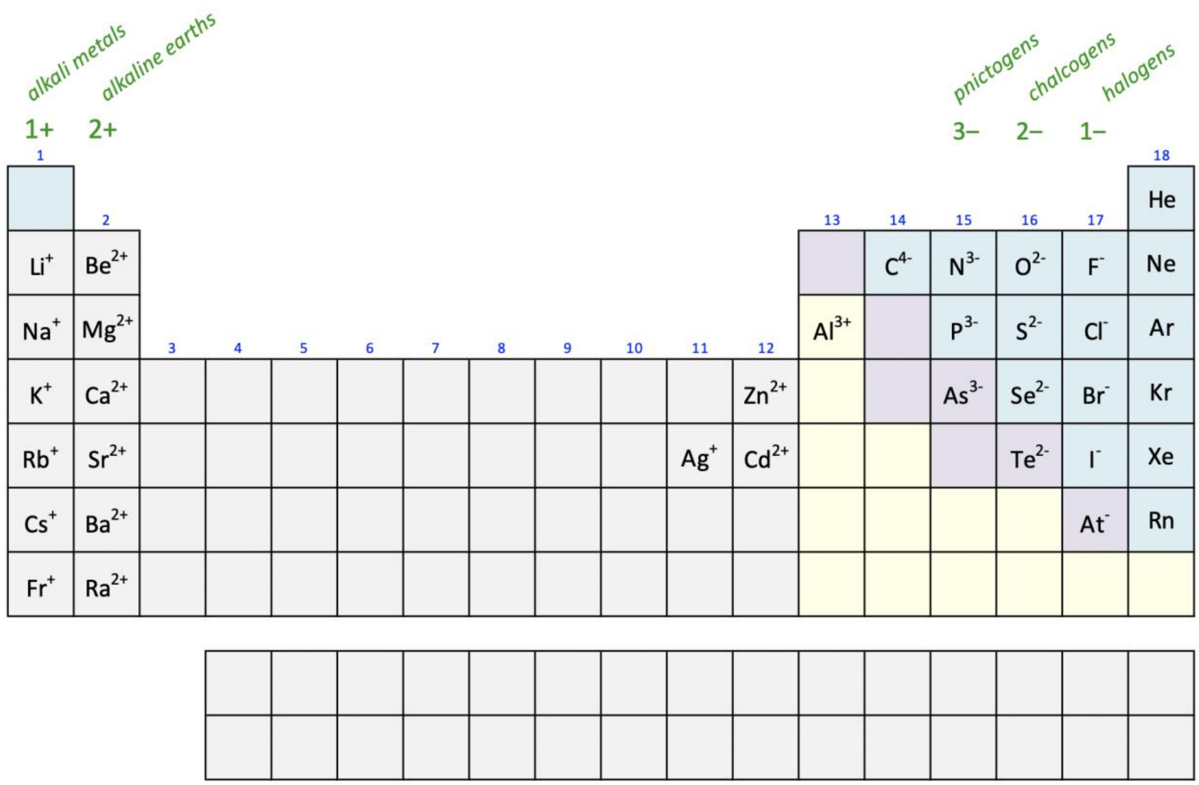

Figure 4.13 The periodic table is useful for predicting ion formation and charge. The ions shown in this periodic table are those with fixed charges. *Adapted from* *OpenStax Chemistry*.

Atoms of nonmetal elements (halogens and most of the chalcogens, and pnictogens) tend to gain electrons to give an anion with the same number of electrons as the nearest noble gas. Halogens tend to gain one electron to gain the number of electrons of the next noble gas. Fluorine, for instance, gains one electron to become the anion F^- with an electron configuration that is similar to neon. Sulfur gains two electrons to form the anion S^{2-}, which has the same electron configuration as argon. Arsenic gains three electrons to become the anion As^{3-}, which has the same electron

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configuration as krypton. In general, an atom in group 17 gains one electron to become an anion with a 1− charge; an atom in group 16 gains two electrons to be a 2− anion; and an atom in group 15 will tend to gain three electrons to form a 3− anion. This general rule only applies to the nonmetals in each group, however, and some of the lower elements are metalloids or metals. The metals and some of the metalloids tend to act more like elements mentioned in the next paragraph.

Elements in the central area of the periodic table containing the transition metals and the metallic elements in groups 13-16 as well as the lanthanides and actinides exhibit variable charges that are not easily predicted. For example, copper can form ions with a 1+ or 2+ charge, iron can form ions with a 2+ or 3+ charge. There are, however, certain transition metals that typically have a "fixed" charge. Silver ions are almost always Ag^{+}, zinc ions are almost always Zn^{2+}, and cadmium ions are almost always Cd^{2+}. Among the main group metals, aluminum ions are almost always Al^{3+}.

Ions can be represented by the same ${}^{A}_{Z}X$ isotope notation (full atomic symbol) presented earlier in the chapter. For an ion, however, the overall charge for the ion must be included. A sodium ion with a mass number of 23 would be written as ${}^{23}_{11}Na^{+}$. This isotope would have 11 protons, 12 neutrons, 10 electrons, and overall charge of 1+.

Example 4.6

Write the isotope symbol for each of the following ions.

(a) oxygen with 8 neutrons and 10 electrons

(b) titanium with 24 neutrons and 18 electrons

(c) bromine with a mass number of 80 and 36 electrons

(d) the ion with 13 protons, 14 neutrons, and 10 electrons

(e) the ion with Z = 15, A = 32, and an overall charge of 3−.

(f) the ion with 76 protons, 116 neutrons, and 68 electrons.

Solutions

(a) ${}^{16}_{8}O^{2-}$ (b) ${}^{46}_{22}Ti^{4+}$ (c) ${}^{80}_{35}Br^{-}$ (d) ${}^{27}_{13}Al^{3+}$ (e) ${}^{32}_{15}P^{3-}$ (f) ${}^{192}_{76}Os^{8+}$

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Chapter 4 Practice Problems

4.1 Atomic Theory

1. In your own words, state the tenets of Dalton's atomic theory.

2. Describe the experiments that provided evidence that electrons carry a negative charge.

3. Explain the "plum pudding" model of the atom and explain why this model is no longer used.
4. Describe the gold foil experiment done by Rutherford and how the results contributed to modern atomic theory.

5. What is the meaning of the term "nucleus" in chemistry?

6. What observation led Rutherford to propose the existence of the neutron?

4.2 The Structure of the Atom

7. An isotope of uranium has an atomic number of 92 and a mass number of 235. How many protons, neutrons, and electrons are present in this atom? Write the symbol for this isotope.

8. An isotope of boron has an atomic number of 5 and a mass number of 11. How many protons, neutrons, and electrons are present in this atom? Write the symbol for this isotope.

9. The number of protons in the nucleus of a tin atom is 50, and the number of neutrons in the nucleus is 68. What is the mass number of this isotope? Write the symbol for this isotope.

10. The number of protons in the nucleus of a calcium atom is 20, and the number of neutrons in the nucleus is 26. What is the mass number of this isotope? Write the symbol for this isotope.

11. Write the symbol for the isotope of sodium containing 12 neutrons.

12. Write the symbol for the isotope of sulfur containing 18 neutrons.

13. Write the full atomic symbols for "iodine 129" and "iodine 131." How many protons, neutrons, and electrons are present in each?

14. Write the full atomic symbols for "chlorine 35" and "chlorine 37." How many protons, neutrons, and electrons are present in each?

15. Which of the following pairs represent isotopes?
    $^{40}_{20}\text{Ca}\ and\ ^{40}_{19}\text{K}$ $\quad$ $^{56}_{26}\text{Fe}\ and\ ^{58}_{26}\text{Fe}$ $\quad$ $^{238}_{92}\text{U}\ and\ ^{235}_{92}\text{U}$ $\quad$ $^{15}_{7}\text{N}\ and\ ^{15}_{8}\text{O}$

16. Which of the following pairs represent isotopes?

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$^{50}_{22}Ti$ *and* $^{50}_{24}Cr$ $^{36}_{18}Ar$ *and* $^{40}_{18}Ar$ $^{162}_{66}Dy$ *and* $^{162}_{68}Er$ $^{10}_{5}B$ *and* $^{11}_{5}B$

17. Determine the number of protons, neutrons, and electrons in a neutral atom of each of the following isotopes:
    (a) $^{10}_{5}B$ (b) $^{199}_{80}Hg$ (c) $^{63}_{29}Cu$ (d) $^{14}_{7}N$

18. Determine the number of protons, neutrons, and electrons in a neutral atom of each of the following isotopes:
    (a) $^{7}_{3}Li$ (b) $^{125}_{52}Te$ (c) $^{77}_{34}Se$ (d) $^{31}_{15}P$

19. Identify the element X for each of the following isotopes:
    (a) $^{131}_{53}X$ (b) $^{238}_{92}X$ (c) $^{187}_{73}X$ (d) $^{50}_{20}X$

20. Identify the element X for each of the following isotopes:
    (a) $^{179}_{74}X$ (b) $^{22}_{10}X$ (c) $^{10}_{5}X$ (d) $^{86}_{36}X$

4.3 Isotopes and Average Atomic Mass

21. An unknown element X has two stable isotopes. The first isotope has a mass of 14.00307 amu and the second isotope has a mass number of 15.00011 amu. If the average atomic mass of X is 14.01 amu, which isotope has a larger percent abundance? Explain.

22. An unknown element Z has two stable isotopes. The first isotope has a mass of 6.0151 amu and the second isotope has a mass number of 7.0160 amu. If the average atomic mass of X is 6.97 amu, which isotope has a larger percent abundance? Explain.

23. Use the average atomic mass from the periodic table to choose the most abundant isotope in the each of the following pairs.
    (a) $^{6}_{3}Li$ *or* $^{7}_{3}Li$ (b) $^{35}_{17}Cl$ *or* $^{37}_{17}Cl$ (c) $^{10}_{5}B$ *or* $^{11}_{5}B$

24. Use the average atomic mass from the periodic table to choose the most abundant isotope in the each of the following pairs.
    (a) $^{14}_{7}N$ *or* $^{15}_{7}N$ (b) $^{63}_{29}Cu$ *or* $^{65}_{29}Cu$ (c) $^{121}_{51}Sb$ *or* $^{123}_{51}Sb$

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25. Bromine has two stable, naturally occurring isotopes. ^{79}Br has a mass of 78.9183 amu and an abundance of 50.69%. ^{81}Br has a mass of 80.9163 amu and an abundance of 49.31%. Calculate the average atomic mass of bromine.

Isotope	Mass (amu)	Abundance
^{79}Br	78.9183	50.69%
^{81}Br	80.9163	49.31%

26. The element iridium (Ir) has two stable isotopes. ^{191}Ir has a mass of 190.9606 amu and a percent abundance of 37.3%. ^{193}Ir has a mass of 192.9629 amu and an abundance of 62.7%. Calculate the average atomic mass of iridium.

27. Calculate the average atomic mass of silicon from its three naturally occurring isotopes. One isotope, ^{28}Si, has a mass of 27.976927 amu and an abundance of 92.23%. Another isotope, ^{29}Si, has a mass of 28.976495 amu and an abundance of 3.10%. The third isotope, ^{30}Si, has a mass of 29.973770 amu.

Isotope	Mass (amu)	Abundance
^{28}Si	27.976927	92.23%
^{29}Si	28.976495	4.63%
^{30}Si	29.973770	?

28. Calculate the average atomic mass of neon from its three naturally occurring isotopes. One isotope, ^{20}Ne, has a mass of 19.9924 amu and an abundance of 90.48%. Another isotope, ^{21}Ne, has a mass of 20.9938 amu and an abundance of 0.27%. The third isotope, ^{22}Ne, has a mass of 21.9913 amu.

29. Naturally occurring chlorine consists of ^{35}Cl (mass 34.96885 amu) and ^{37}Cl (mass 36.96590 amu), with an average mass of 35.453 amu. What is the percent abundance of each of these isotopes in a sample of chlorine?

Isotope	Mass (amu)	Abundance
^{35}Cl	34.96885	?
^{37}Cl	36.96590	?

30. Naturally occurring copper consists of ^{63}Cu (mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu), with an average mass of 63.546 amu. What is the percent abundance of each of these isotopes in a sample of copper?

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31. The ^{18}O:^{16}O abundance ratio on some meteorites is greater than that used to calculate the average atomic mass of oxygen on earth. Is the average mass of an oxygen atom in these meteorites greater than, less than, or equal to that of a terrestrial oxygen atom?

32. On a planet in a galaxy far, far away, a certain element X has the stable isotopes given in the table below. What is the average atomic mass of the element X on this planet?

Isotope	Mass (amu)	Abundance
^{127}X	126.95	25.26%
^{129}X	128.89	19.25%
^{132}X	131.85	7.86%
^{133}X	132.96	?

4.4 The Periodic Law

33. Based on its position on the periodic table, classify selenium as metal, nonmetal, or metalloid. What is the expected charge on a selenium ion? Explain.

34. Based on its position on the periodic table, classify indium as metal, nonmetal, or metalloid. What is the expected charge on an indium ion? Explain.

35. Write down the chemical symbols and names for the halogens.

36. Write down the chemical symbols and names for the alkaline earths.

37. Which of these sets of elements are all in the same period?
    (a) potassium, vanadium, ruthenium
    (b) lithium, carbon, chlorine
    (c) sodium, magnesium, sulfur
    (d) chromium, nickel, krypton
    (e) potassium, calcium, zinc

38. Which of these sets of elements are all in the same group?
    (a) barium, tungsten, argon
    (b) sodium, rubidium, barium
    (c) nitrogen, phosphorus, bismuth
    (d) copper, silver, gold
    (e) magnesium, strontium, samarium

39. Classify each of the following as a transition metal, a halogen, or a noble gas.
    (a) manganese (b) fluorine (c) xenon (d) zinc

40. Classify each of the following as a transition metal, a halogen, or a noble gas.
    (a) mercury (b) chlorine (c) iron (d) neon

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4.5 Ion Formation

41. Write the isotope symbol for each of the following ions:
    (a) the ion with a 1+ charge, atomic number 55, and mass number 133
    (b) the ion with 54 electrons, 53 protons, and 74 neutrons
    (c) the ion with atomic number 15, mass number 31, and a 3− charge
    (d) the ion with 24 electrons, 30 neutrons, and a 3+ charge

42. Write the isotope symbol for each of the following ions:
    (a) the ion with a 3+ charge, 28 electrons, and a mass number of 71
    (b) the ion with 36 electrons, 35 protons, and 45 neutrons
    (c) the ion with 86 electrons, 142 neutrons, and a 4+ charge
    (d) the ion with a 2+ charge, atomic number 38, and mass number 87

43. The following isotopes are used in medicine. For each, determine the number of protons, neutrons, and electrons and write the isotope symbol.
    (a) atomic number 9, mass number 18, charge of 1−
    (b) atomic number 43, mass number 99, charge of 7+
    (c) atomic number 78, mass number 195, charge of 2+
    (d) atomic number 81, mass number 201, charge of 1+

44. The following ions are essential in for human health. For each, determine the number of protons, neutrons, and electrons and write the isotope symbol.
    (a) atomic number 26, mass number 58, charge of 2+
    (b) atomic number 53, mass number 127, charge of 1−
    (c) atomic number 19, mass number 41, charge 1+
    (d) atomic number 30, mass number 68, charge 2+

Review: How to Succeed in Chemistry

45. Identify the chapters and sections that are difficult for you. Make sure to spend some extra review time on these topics and see a chemistry tutor if necessary.

46. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

47. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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Chapter 5
Molecules and Compounds

LEARNING OBJECTIVES

1. Define the terms *molecule* and *compound*.
2. Represent compounds as chemical formulas.
3. Distinguish between ionic compounds and molecular compounds.
4. Name simple substances based on their formulas.
5. Determine the formula for a molecule or ionic compound based on its name.
6. Calculate the formula masses of substances.
7. Calculate the molecular mass for a molecular compound.

5.1 Compounds

Recall from Chapter 3 that pure substances can be classified as either elements or compounds. A substance containing just one type of atom is an element. The noble gases are monatomic elements and can exist stably without bonding to each other or to other elements. For example, a neon atom (Ne) exists as a single neon atom. Several elements exist naturally as **molecules**, or identical clusters of atoms by strong interactions called bonding. Molecules are usually made up exclusively of non-metal atoms, and every molecule of a given type has an identical composition to every other

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molecule of that given type. For example, every water molecule (H_2O) has exactly two hydrogen atoms and one oxygen atom in it, no more and no less. Elements such as hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine usually exist as pairs of atoms bound together. These elements are **diatomic molecules** and are represented as H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The element phosphorus usually exists as the molecule P_4, and elemental sulfur usually exists as the molecule S_8 (Figure 5.1).

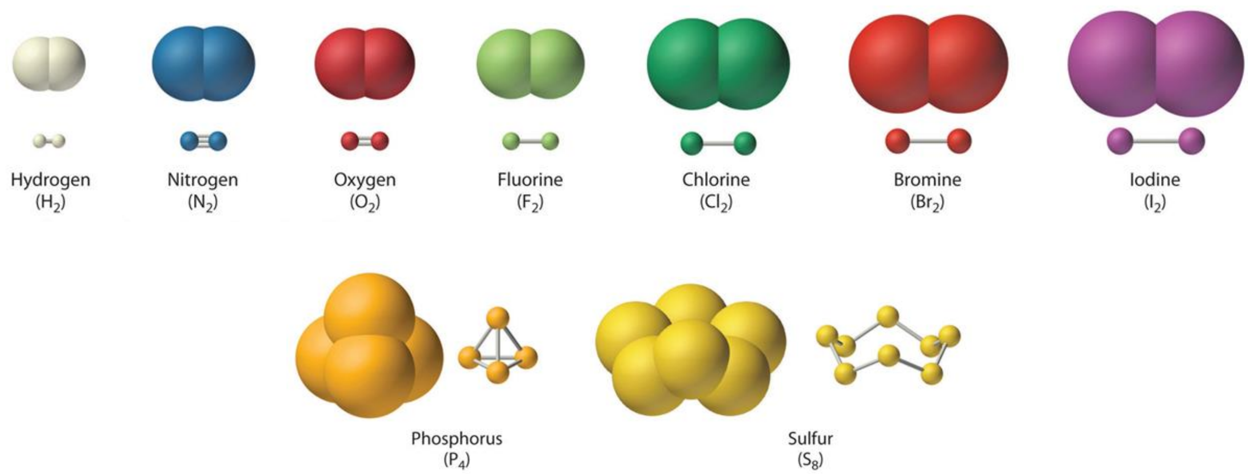

Figure 5.1 The elements hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine exist as pairs of atoms called diatomic molecules. The element phosphorus exists as a molecule containing four phosphorus atoms, and the element sulfur exists as a ring-shaped molecule containing eight sulfur atoms. *Image credit: OpenStax Chemistry.*

A **compound**, on the other hand, is a substance that contains two or more elements combined in fixed proportions and can only be broken down into its elements by using a **chemical process**. Physical processes such as filtration or distillation cannot break a compound down into its elements.

Compounds have characteristic physical and chemical properties that are distinct from those of its elements. Elemental sodium is a soft metal that bursts into flame when it comes into contact with water. Chlorine is a toxic gas that causes acute damage to the respiratory tract and can be fatal after just a few minutes of exposure. Sodium chloride, however, is a white crystalline material that dissolves readily in water and is a normal part of the human diet (Figure 5.2).

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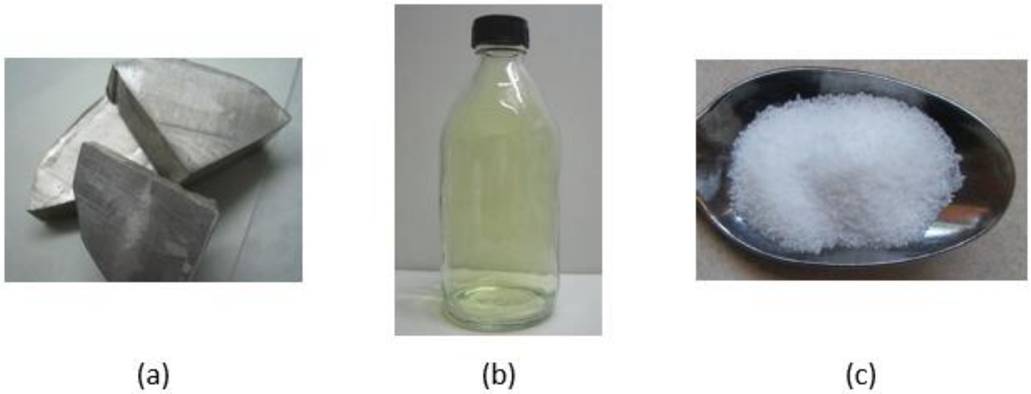

Figure 5.2 Table salt is the compound containing sodium chloride. (a) Pure sodium is a very reactive soft metal. (b) Chlorine is a yellow-green gas and is poisonous. (c) Table salt contains atoms of sodium and atoms of chlorine but has very different physical and chemical properties. *Image Credit: Wikimedia Commons.*

One of the simplest compounds containing carbon and hydrogen is called methane, in which there are always four hydrogen atoms for each carbon atom. The atoms that make up methane are bound together tightly and in a very specific arrangement to make up a molecule (Figure 5.3). Each molecule moves around as a unit containing exactly one carbon atom and four hydrogen atoms. The fact that the fundamental unit of methane is constant means that a flask full of methane has the same physical and chemical properties no matter where it is located. This is the **law of constant composition**.

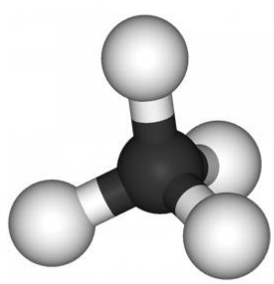

Figure 5.3 Methane is a compound made from carbon (black sphere) and hydrogen (white spheres). A sample of methane always has four hydrogen atoms for each carbon atom. *Image Credit: Wikimedia Commons.*

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5.2 Representing Compounds

Methane, with its one carbon atom and four hydrogen atoms, can be represented by using a model or a drawing (Figure 5.3). A simpler representation of the methane molecule makes use of the fixed proportion between carbon and hydrogen. A **chemical formula** is a convenient way to summarize the fixed ratios of elements needed for a compound; the *atomic symbol* is used for each *element* and a *subscript* indicates the ratio. An atomic symbol written without a subscript indicates a ratio of one for that component. Because methane always contains one carbon atom and four hydrogen atoms, its chemical formula is CH_4.

Carbon dioxide is a greenhouse gas that humans generate in large amounts every day. Carbon dioxide has two oxygen atoms for every carbon atom and is represented by the chemical formula CO_2. Ammonia is a substance that contains three hydrogen atoms for every nitrogen atom. Ammonia has the chemical formula NH_3 (Figure 5.4).

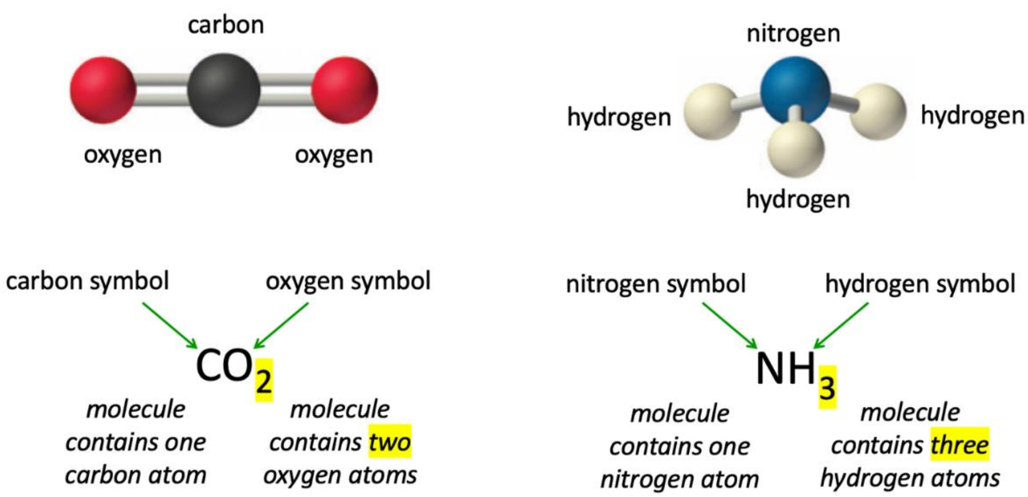

Figure 5.4 Carbon dioxide contains two oxygen atoms for each carbon atom and therefore has the chemical formula CO_2. Ammonia contains three hydrogen atoms for each nitrogen atom and has the chemical formula NH_3.

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5.3 Ionic Compounds

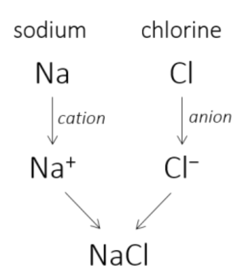
Compounds fall into two main categories:

(1) **Ionic compounds** contain positive ions and negative ions.

(2) **Molecular compounds** contain two or more only nonmetals (Section 5.4).

An **ionic compound** contains positively charged cations, usually made from metals which tend to lose electrons, and negatively charged anions, usually made from nonmetals which end to gain electrons. In some cases, ion charges can be determined by locating the element on a periodic table (Figure 4.13). The metal cation and non-metal anion must combine in a whole number ratio that results in a neutral compound.

Sodium and chlorine are the components of table salt. Since sodium is a metal and chlorine is a nonmetal, they are probably going to form an ionic compound. From the location of sodium in Group 1, the alkali metals, the charge of the ion can be predicted. In an ionic compound, sodium will be the cation and have a 1+ charge. Likewise, the charge of the anion of chlorine, a halogen in Group 17, will be 1−; this anion is called "chloride". When combined in a one-to-one ratio, a sodium cation and a chloride anion will form a neutral compound that is represented by the formula NaCl. Ionic compounds are expressed by using the lowest whole number ratio of its components, which is called an empirical formula. For example, a substance with a chemical formula of $H_2C_2O_4$ would have an empirical formula of HCO_2.



Magnesium and oxygen combine to form a stable ionic compound. Magnesium is an alkaline earth metal in Group 2 and therefore occurs in an ionic compound as a 2+ cation, Mg^{2+}. Oxygen is a

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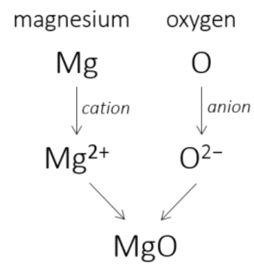
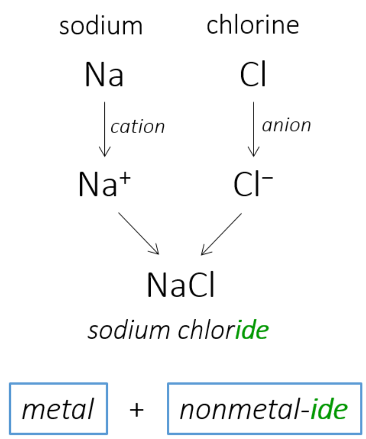
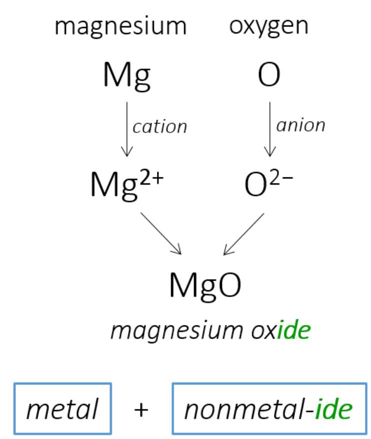
chalcogen and therefore occurs in an ionic compound with a 2− charge and can be represented as O^{2-}. When combined, one magnesium cation and one oxygen anion form a neutral compound, MgO.



Naming compounds in chemistry, or nomenclature, is easy because there are standard rules that can be applied once the composition of a compound is known. Nomenclature rules allow chemists all over the world to use the same terminology when referring to compounds. For simple ionic compounds like NaCl and MgO, the name of the cation is followed by the name of the anion. The metal cation is simply the name of the element, but the non-metal ion takes the ending *−ide*. The compound NaCl is *sodium chloride* and the compound MgO is *magnesium oxide*.





In the compounds discussed, the metal cation and the non-metal anion have charges of the same magnitude but opposite sign (1+ and 1− or 2+ and 2−), so finding the ratio of ions necessary for a neutral compound is simple. When the cation and anion have charges of different magnitudes, the total charge for all the cations present must equal the total charge for all the anions present. Whole number multiples of one or both components must be considered so that the charges balance out for a neutral compound. For example, consider the ionic compound formed from lithium and oxygen. The lithium cation has a 1+ charge but the oxide anion has a 2− charge. In order to make a

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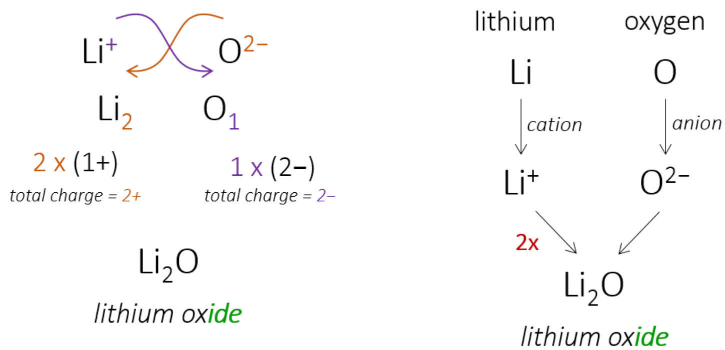
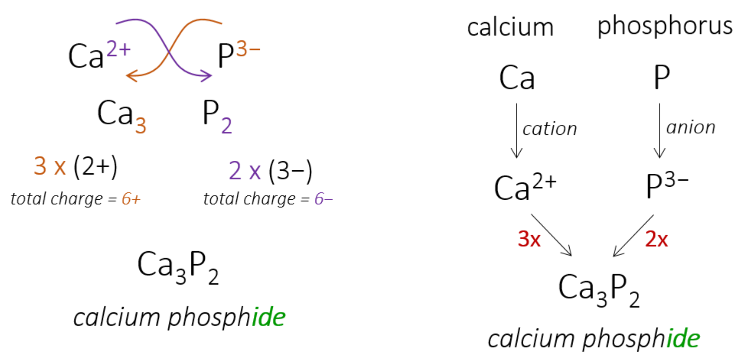
neutral compound, two lithium cations, each with a 1+ charge, are needed to balance out the 2− charge of the anion. One way to determine the formula if the cation and anion have different charges is to “criss-cross” and use the cation charge as the subscript for the anion and use the anion charge for the subscript of the cation. Remember that the number of atoms in a formula must be positive, so use the magnitude or absolute value of the anion charge. Also remember that ionic compounds are written as their empirical formula, so if you use the criss-cross method you need to check and reduce the components to the smallest whole number ratio. The name of the compound is still a combination of the name of the metal ion followed by the name of the non-metal ion with the ending *–ide*. The final compound has the formula Li_2O and the name of the compound is *lithium oxide*.



$\boxed{\textit{metal}}$ + $\boxed{\textit{nonmetal-ide}}$

Now consider the compound formed from calcium and phosphorus. Calcium forms a 2+ cation and phosphorus forms a 3− anion. The compound is *calcium phosphide*, Ca_3P_2.



$\boxed{\textit{metal}}$ + $\boxed{\textit{nonmetal-ide}}$

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Note that even if the atom has a subscript greater than one, the name of the compound remains simply the name of the metal cation followed by the name of the non-metal anion with the ending changed to *–ide*.

Example 5.1

Predict and name the formula for the ionic compound formed by each of the following combinations of metals and nonmetals.

(a) potassium and bromine
(b) calcium and chlorine
(c) sodium and sulfur
(d) magnesium and nitrogen
(e) lithium and phosphorus

Solutions

(a) Potassium is from Group 1 and the ion is therefore K^+. Bromine is from Group 17 and the ion is therefore Br^-. K^+ and Br^- combine to form KBr, *potassium bromide*.
(b) Calcium is from Group 2 and the ion is therefore Ca^{2+}. Chlorine is from Group 17 and the ion is therefore Cl^-. Ca^{2+} and Cl^- combine to form $CaCl_2$, *calcium chloride*.
(c) Sodium is from Group 1 and the ion is therefore Na^+. Sulfur is from Group 16 and the ion is therefore S^{2-}. Na^+ and S^{2-} combine to form Na_2S, *sodium sulfide*.
(d) Magnesium is from Group 2 and the ion is therefore Mg^{2+}. Nitrogen is from Group 15 and the ion is therefore N^{3-}. Mg^{2+} and N^{3-} combine to form Mg_3N_2, *magnesium nitride.*
(e) Lithium is from Group 1 and the ion is therefore Li^+. Phosphorus is from Group 15 and the ion is therefore P^{3-}. Li^+ and P^{3-} combine to form Li_3N, *lithium phosphide*.

Some metals have variable charges and require a bit more information in order for a formula to be written. The metals that almost always have a fixed charge in an ionic compound are group 1 (1+), group 2 (2+), silver (1+), cadmium (2+), zinc (2+), and aluminum (3+). The other metals can form cations with different charges, so the charge must be specified in the name by using a roman

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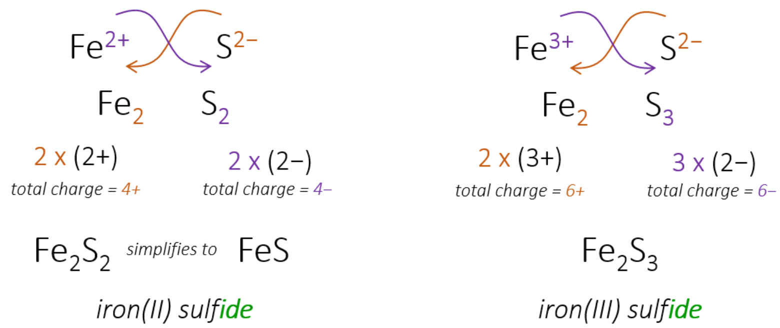
numeral. For example, if Fe^{2+} combines with S^{2-}, the formula is FeS. The charge on the iron is 2+, so the roman numeral *II* must be included in the name. The compound FeS is *iron(II) sulfide*.

If iron with a 3+ charge combines with S^{2-}, the chemical formula is Fe_2S_3. The 3+ charge is represented by the roman numeral *III* and the compound is named *iron(III) sulfide*.



$\boxed{\text{transition metal}}$ $\boxed{\text{(roman numeral)}}$ + $\boxed{\text{nonmetal-ide}}$

Among the metals with variable charges, there are a few that typically exist as a cation of a specific charge. Aluminum, for instance, is almost always Al^{3+}. Cadmium is usually Cd^{2+}, zinc is typically Zn^{2+}, and silver is generally Ag^{+}. If an ionic compound contains a fixed charge metal, a roman numeral designating charge is not required as part of the name. $AlCl_3$ is therefore written as simply *aluminum chloride* rather than *aluminum(III) chloride*. See Table 5.1 for a list of some of ionic compounds containing these cations. If you use a roman numeral for a fixed-charge metal it is technically incorrect, since it implies that the metal is variable-charge and thus was reasonably likely to have some alternative charge.

Table 5.1 Examples of Compounds Containing Al^{3+}, Zn^{2+}, Cd^{2+}, and Ag^{+}

Metal	Cation	Non-metal	Anion	Formula	Compound Name
Aluminum	Al^{3+}	Fluorine	F^{-}	AlF_3	aluminum fluoride *not aluminum(III) fluoride*
Cadmium	Cd^{2+}	Oxygen	O^{2-}	CdO	cadmium oxide *not cadmium(II) oxide*
Silver	Ag^{+}	Nitrogen	N^{3-}	Ag_3N	silver nitride *not silver(I) nitride*
Zinc	Zn^{2+}	Chlorine	Cl^{-}	$ZnCl_2$	zinc chloride *not zinc(II) chloride*

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Example 5.2

Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.

(a) nickel ion with a 2+ charge and bromide

(b) iron ion with a 3+ charge and sulfide

(c) zinc and oxide

(d) aluminum and nitride

(e) platinum ion with a 4+ charge and chloride

Answers

(a) The compound is $NiBr_2$, *nickel(II) bromide*.

(b) The compound is Fe_2S_3, *iron(III) sulfide*.

(c) The compound is ZnO, *zinc oxide*. Zinc is one of the metals with a cation that typically has fixed charge (2+) so a roman numeral is not necessary.

(d) The compound is AlN, *aluminum nitride*. Aluminum is one of the metals with a cation that typically has fixed charge (3+) so a roman numeral is not necessary.

(e) The compound is $PtCl_4$, *platinum(IV) chloride.*

Test Yourself

Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.

(a) titanium ion with a 4+ charge and oxide

(b) chromium ion with a 3+ charge and chloride

(c) copper ion with a 1+ charge and sulfide i

(d) aluminum ion with a 3+ charge and oxide

Answers

(a) TiO_2, *titanium(IV) oxide*

(b) $CrCl_3$, *chromium(III) chloride*

(c) Cu_2S, *copper(I) sulfide*

(d) Al_2O_3, *aluminum oxide*

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Some ions exist as clusters of atoms bound together in a tight bundle and carrying a single overall charge. These **polyatomic ions** travel as a single unit, act as a single ion, and combine with other ions to make ionic compounds (Figure 5.5). Polyatomic ions are important in myriad chemical and biological processes. Chemistry and biology students should be able to recognize polyatomic ions, their charges, and predict how they combine with other ions to form compounds. Some of the most important polyatomic ions are listed in Table 5.2.

Ionic compounds containing polyatomic ions follow the same rules regarding charge balance and nomenclature. The ionic compound made from sodium and hydroxide requires one sodium ion (Na^+) for each hydroxide ion (OH^-). The compound, NaOH, is named sodium hydroxide. When naming ionic compound containing polyatomic ions, use the same method as for naming ionic compound but simply substitute the name of the polyatomic ion for the anion or cation as appropriate.

Ammonium nitrate is an ionic compound containing one ammonium ion (NH_4^+) for each nitrate ion (NO_3^-). The chemical formula for the compound is NH_4NO_3.

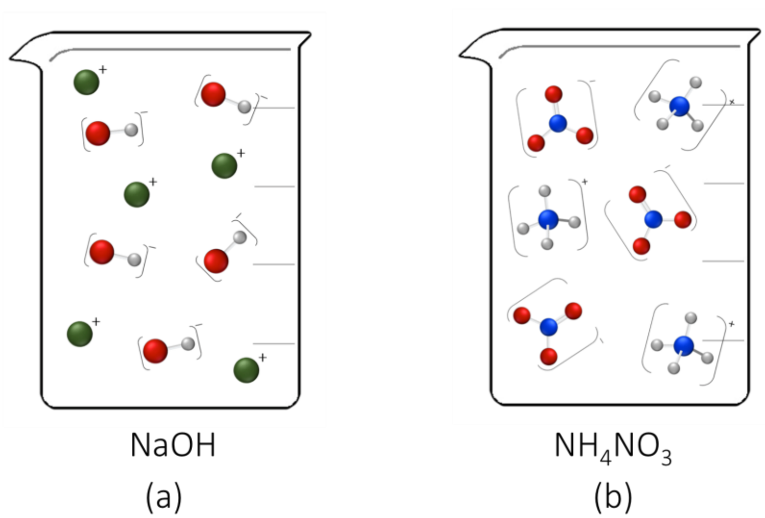

Figure 5.5 The atoms within a polyatomic ion are bound together tightly, and the ion moves as a single unit. (a) In sodium hydroxide, the sodium cation (green) and the hydroxide anion (red/gray) move around in aqueous solution. The hydroxide hydrogen and oxygen atoms stay together as a unit and carry an overall negative charge. (b) Ammonium moves as the group NH_4^+ (blue/gray) and nitrate moves as the group NO_3^- (blue/red).

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Table 5.2 Selected Polyatomic Ions

Name	Formula
ammonium	NH_4^+
hydronium	H_3O^+
hydroxide	OH^-
acetate	CH_3COO^-
cyanide	CN^-
chromate	CrO_4^{2-}
dichromate	$Cr_2O_7^{2-}$

Name	Formula
nitrate	NO_3^-
nitrite	NO_2^-
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}
phosphate	PO_4^{3-}
carbonate	CO_3^{2-}
bicarbonate	HCO_3^-

When multiple polyatomic ions are needed in a compound, the ion is enclosed in parentheses with the whole number ratio of the ion denoted with a subscript. Calcium (Ca^{2+}) combines with nitrate (NO_3^-) to form the ionic compound calcium nitrate. Calcium is located in Group 2 and carries a 2+ charge which means that two nitrate ions, each with a 1– charge, are needed to form a neutral compound. The compound is simply named calcium nitrate, but the formula must contain a subscript to indicate that two nitrate ions are present for each calcium ion. Recall that the nitrate ion travels as a bundle but is otherwise treated like a single-atom anion when constructing a chemical formula. The chemical formula for calcium nitrate is therefore $Ca(NO_3)_2$. When a compound containing a polyatomic anion is named, the name of the metal or cation is followed by the name of the anion as given in Table 5.2. The name of $Ca(NO_3)_2$ is *calcium nitrate*.

If the compound contains a single-atom non-metal anion, also called a **monatomic** anion, the ending changed to *–ide* as in previous examples. The compound made from ammonium (NH_4^+) and sulfur (S^{2-}) is $(NH_4)_2S$, *ammonium sulfide*. A compound containing a polyatomic ion and a variable-charge metal must include the roman numeral charge in the name unless the metal ion is Al^{3+}, Zn^{2+}, Cd^{2+}, or Ag^+. The ionic compound containing Cu^{2+} and phosphate (PO_4^{3-}) is $Cu_3(PO_4)_2$, *copper(II) phosphate*.

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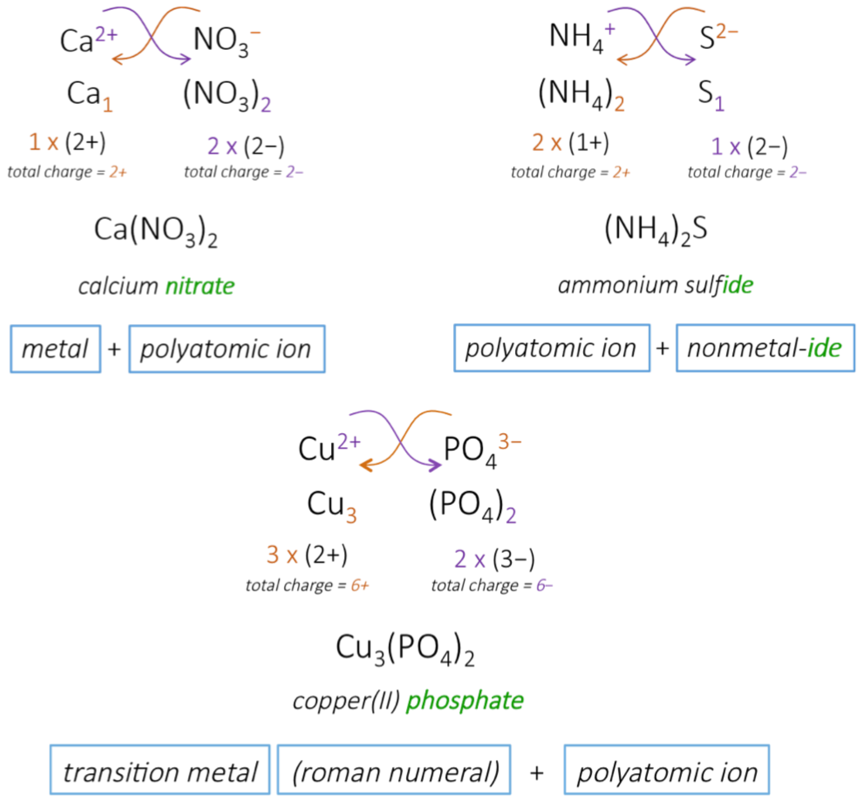

Example 5.3

Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.

(a) calcium and hydroxide

(b) ammonium and sulfate

(c) chromium with a 3+ charge and carbonate

(d) magnesium and phosphate

(e) zinc and cyanide

Answers

(a) The compound is $Ca(OH)_2$, *calcium hydroxide*.

(b) The compound is $(NH_4)_2SO_4$, *ammonium sulfate*.

(c) The compound is $Cr_2(CO_3)_3$, *chromium(III) carbonate.* Chromium has a variable charge, so a roman numeral must be used.

(d) The compound is $Mg_3(PO_4)_2$, *magnesium phosphate*.

(e) The compound is $Zn(CN)_2$, *zinc cyanide.* Zinc is one of the fixed-charge exceptions and has a charge that typically fixed as 2+, so a roman numeral is not necessary.

Test Yourself

Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.

(a) platinum ion with a 4+ charge and sulfate ion

(b) manganese ion with a 2+ charge and phosphate ion

(c) ammonium ion and oxygen ion

(d) silver ion and nitrate ion

Answers

(a) $Pt(SO_4)_2$, *platinum(IV) sulfate*

(b) $Mn_3(PO_4)_2$, *manganese(II) phosphate*

(c) $(NH_4)_2O$, *ammonium oxide*

(d) $AgNO_3$, *silver nitrate*

Ionic compounds are held together via electrostatic interactions, that is the attraction between two particles of opposite charge. In the solid state, ionic compounds also tend to exist in extended networks or lattices as opposed to single molecules. Recall that all molecules of a substance are exactly like all other molecules of that substance. Ionic crystals, on the other hand, can be bigger or smaller without being a different substance. A salt crystal containing 1000 Na^+ and 1000 Cl^- is chemically the same as a salt crystal containing 10 Na^+ and 10 Cl^-; one cluster is just larger than the other cluster. This is why ionic compounds are represented by their empirical formula. Sodium chloride contains one sodium cation for every chloride anion and is represented as NaCl. A sample of sodium chloride contains sodium and chloride in a 1:1 ratio but is an extended network of sodium cations and chloride anions. Similarly, zinc sulfide contains one zinc cation for every sulfide anion and has the formula ZnS (Figure 5.6).

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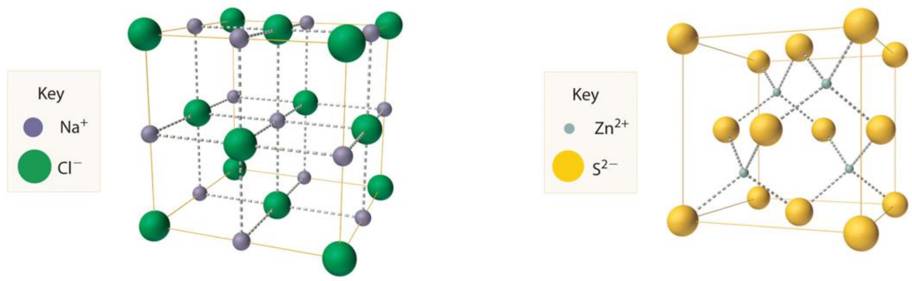

Figure 5.6 Most solid ionic compounds exist as extended lattices. (a) In sodium chloride, each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion is surrounded by six Na^+ ions. The network is regular and continuous but is broken up when NaCl is dissolved in water. (b) In zinc sulfide, a continuous lattice forms in which each Zn^{2+} ion is in contact with four S^{2-} ions, and each S^{2-} ion is in contact with four Zn^{2+} ions. *Adapted from [Principles of General Chemistry](#).*

5.4 Molecular Compounds

Compounds that exist as discrete molecules containing only non-metals are termed **molecular compounds.** Water (H_2O), carbon dioxide (CO_2), and ammonia (NH_3) are molecular compounds that are common in everyday life. These are very different from **ionic compounds** like sodium chloride (NaCl) and silver nitrate ($AgNO_3$). Ionic compounds typically form when a metal atom **donates** an electron to become a cation and a non-metal atom **accepts** that electron to form an anion. The cation and the anion are held together *via* electrostatic attraction. In a molecular compound, two atoms within the compound **share** electrons in the form of a **covalent bond**.

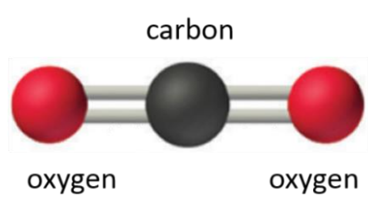

Figure 5.7 Carbon dioxide (CO_2) is a molecular compound containing carbon and oxygen. The central carbon atom forms two bonds to the oxygen atom on one side and two bonds to the oxygen atom on the opposite side of the molecule. The four bonds hold the atoms together in a single unit.

When a molecular compound contains just two non-metal elements, it is called a **binary molecular compound**. Naming binary molecular compounds is very similar to naming simple ionic compounds. The first element in the formula is simply written as the name of the element. The

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second element is simply the stem of the element name followed by *−ide*. The major difference between naming ionic compounds and naming binary molecular compounds is that when naming a binary molecular compound, a **prefix** is used to specify the number of atoms present in the molecule (Table 5.3).

Table 5.3 Prefixes for Naming Binary Molecular Compounds

Number	Prefix
1	*mono−*
2	*di−*
3	*tri−*
4	*tetra−*
5	*penta−*

Number	Prefix
6	*hexa−*
7	*hepta−*
8	*octa−*
9	*nona−*
10	*deca−*

There are a few rules to keep in mind when naming a binary molecular compound:

(1) The elements are typically written in the order C, P, N, H, S, I, Br, Cl, O, F.

(2) For the prefix for the second element, the letter "o" or the letter "a" is usually dropped when the second element is oxygen. The "*o*" of *mono−* and the "*a*" of *tetra−* are dropped when used to name an oxide. For example, four oxygen atoms would be named *tetr**o**xide* instead of *tetr**ao**xide*.

(3) If there is only one atom of the first element, the prefix (*mono−*) is omitted. CO_2 is therefore named *carbon dioxide* instead of *monocarbon dioxide*.

Several binary molecular compounds are referred to by their common names. For example, H_2O is simply *water*. NH_3 is *ammonia*. CH_4 is referred to as *methane*. N_2O is *nitrous oxide*. H_2O_2 is called hydrogen peroxide.

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Table 5.4 Examples of Binary Molecular Compounds

Formula	Name
CO	*carbon **mono**oxide*
SO_2	*sulfur **di**oxide*
PF_3	*phosphorus **tri**fluoride*
N_2O_4	***di**nitrogen **tetr**oxide*
As_2O_5	***di**arsenic **pent**oxide*
IF_7	*iodine **hepta**fluoride*
S_2F_{10}	***di**sulfur **deca**fluoride*

Example 5.4

Name each of the following binary molecular compounds.

(a) NO_2 (b) $BrCl_3$ (c) NO (d) NH_3 (e) SF_6

Solutions

(a) NO_2 has one nitrogen atom listed first, so it is simply *nitrogen* (not mononitrogen). There are two oxygen atoms, so the compound name will end with a *di*− (two) *oxide* (oxygen). NO_2 is *nitrogen dioxide*.

(b) The single bromine atom means that the name starts with simply bromine. The three chlorine atoms mean that the prefix *tri*− (three) will be needed. $BrCl_3$ is *bromine trichloride*.

(c) A single nitrogen at the beginning of the formula means that the name will start with *nitrogen*. Because the nitrogen is followed by a single oxygen atom, the prefix *mono*− (one) is needed. NO is nitrogen monoxide.

(d) NH_3 is a molecule that is always called by its common name, *ammonia*.

(e) The formula begins with one sulfur atom, so the name will start with *sulfur*. Six fluorine atoms mean that the prefix *hexa*− (six) will be used for the second part of the name. SF_6 is *sulfur hexafluoride*.

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Test Yourself

Name each of the following binary molecular compounds.

(a) CCl_4 (b) PCl_3 (c) P_4S_3 (d) $XeCl_4$ (e) S_2F_{10}

Answers

(a) *carbon tetrachloride* (b) *phosphorus trichloride* (c) *tetraphosphorus trisulfide*

(d) *xenon tetrachloride* (e) *disulfur decafluoride*

5.5 Acids

The simplest definition an **acid** is a compound containing at least one hydrogen atom that can dissociate into an H^+ cation and an anion. In their pure states, acids are **molecular compounds** but they break down into cations (H^+) and anions when dissolved in aqueous solution. The chemical formula for an acid is written with the hydrogen first, followed by an anion. The name of an acid depends upon the anion formed when H^+ dissociates in aqueous solution.

In this chapter, acids will be divided into two main categories:

(1) A **binary acid** has H^+ attached to a single element.

(2) An **oxoacid** has H^+ attached to an oxygen atom in a polyatomic ion.

The name of a binary acid begins with the prefix *hydro–* followed by the name of the anion adjusted with the suffix *–ic* then the word *acid*. The most common binary acids are listed in Table 5.5.

Table 5.5 Common Binary Acids

Formula	Acid Name
HF	***hydro**fluori**c** acid*
HCl	***hydro**chlori**c** acid*
HBr	***hydro**bromi**c** acid*
HI	***hydro**iodi**c** acid*

hydro- + *element-ic* + *acid*

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Another common acid named by using the prefix *hydro*− and suffix −*ic*. is HCN. HCN is known as ***hydro****cyan***ic** *acid*.

Oxyacids are compounds containing one or more H^+ ions attached to an oxygen-containing polyatomic anion. One of the most common oxyacids is HNO_3. The H^+ ion is attached to the nitrate anion, NO_3^-. Oxyacids are named based upon the polyatomic anion, in this case *nitrate*. If the polyatomic ion name ends with −*ate*, the ending becomes −*ic*. HNO_3 is therefore *nitric acid*.

Table 5.6 Oxyacids from –*ate* Anions

Formula	Polyatomic Ion	Acid Name
HNO_3	*nitrate*	*nitr**ic** acid*
CH_3COOH	*acetate*	*acet**ic** acid*
H_2SO_4	*sulfate*	*sulfur**ic** acid*
H_2CO_3	*carbonate*	*carbon**ic** acid*
H_3PO_4	*phosphate*	*phosphor**ic** acid*

oxyanion-ate → *oxyanion-ic* + *acid*

If the polyatomic ion name ends in −*ite*, the ending becomes −*ous*. For example, the acid formed from the sulfite ion, SO_3^{2-}, is H_2SO_3. The name of H_2SO_3 is *sulfurous acid*.

Table 5.6 Oxyacids from –*ite* Anions

Formula	Polyatomic Ion	Acid Name
HNO_2	*nitrite*	*nitr**ous** acid*
H_2SO_3	*sulfite*	*sulfur**ous** acid*
$HClO_2$	*chlorite*	*chlor**ous** acid*

oxyanion-ite → *oxyanion-ous* + *acid*

Example 5.4

Name each of the following acids.

(a) HBr (b) H_2CO_3 (c) HNO_2 (d) H_3PO_4 (e) HCl

Solutions

(a) HBr is a binary acid containing hydrogen and bromine. The name of the acid is therefore *hydrobromic acid*.

(b) H_2CO_3 is acid made from the carbonate ion, CO_3^{2-}. H_2CO_3 is *carbonic acid*.

(c) HNO_2 is acid made from the nitrite ion, NO_2^-. HNO_2 is *nitrous acid*.

(d) H_3PO_4 is acid made from the phosphate ion, PO_4^{3-}. H_3PO_4 is *phosphoric acid*.

(e) HCl is a binary acid containing hydrogen and chlorine. The name of the acid is therefore *hydrochloric acid*.

Test Yourself

Name each of the following acids.

(a) HF (b) H_2SO_4 (c) HNO_3 (d) HI (e) H_3PO_4

Answers

(a) *hydrofluoric acid* (b) *sulfuric acid* (c) *nitric acid*

(d) *hydroiodic acid* (e) *phosphoric acid*

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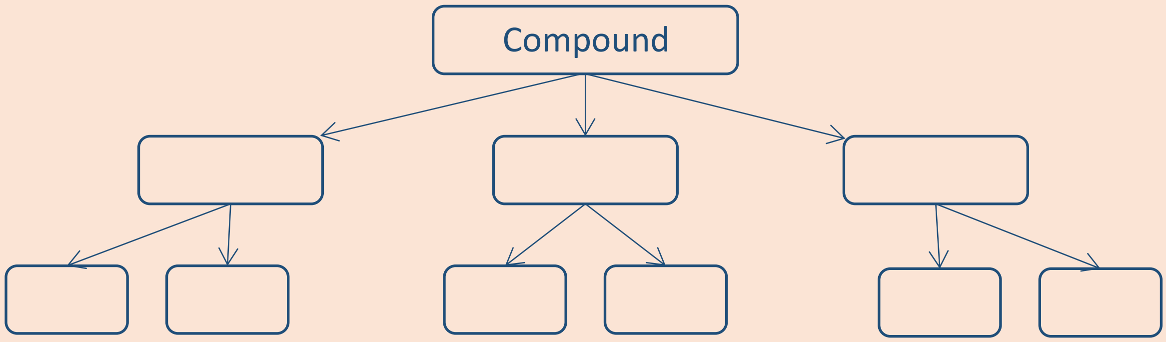
Nomenclature Review – Create a Decision Tree

Draw a decision tree for naming a substance given its chemical formula. To get started:

1. Classify the compound as ionic, molecular, or acid.
2. What are the components of the compound and how are they named?
3. Are there prefixes and/or suffixes involved?



Keep adding layers, arrows, and boxes!

Use the decision tree you've created to work through the example nomenclature examples, the end-of-chapter nomenclature problems, and the examples provided in lecture. If the names you are getting do not match the correct answers, revise your decision tree and try again.

5.6 Formula Mass

In chemistry, the mass of a formula unit of an ionic compound is known as the **formula mass** (FM) or the **molar mass** (M). Recall from Section 5.3 that the formula for an ionic compound is the smallest whole number ratio of its components. The formula mass is the sum of the atomic masses of each atom in a formula unit. For example, sodium chloride is NaCl. The formula mass for NaCl is determined by adding the atomic mass for sodium (Na) and the atomic mass for chlorine (Cl). The atomic masses, in atomic mass units or amu, for the elements are listed on the periodic table. Using the periodic table given earlier in this textbook (reproduced below), the atomic mass for sodium is 22.990 amu and the atomic mass for chlorine is 35.45 amu. Adding these together, and remembering the significant figures rule for addition, the total is 58.44 amu. The formula unit for sodium chloride is NaCl and the formula mass for NaCl is 58.44 amu.

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$$\begin{array}{rll} \text{Na:} & 22.990 & \text{amu} \\ \text{Cl:} & 35.45 & \text{amu} \\ \hline \text{NaCl:} & 58.44 & \text{amu} \end{array}$$

The ionic compound calcium bromide has the formula unit $CaBr_2$. For each calcium ion, there are two bromide ions. To obtain the formula mass for $CaBr_2$, the atomic mass for calcium (40.078 amu) must be added to 2x the atomic mass for bromine. Each Br atom is 79.907 amu, so two bromine atoms (2 x 79.904 amu) together are 159.808 amu. The total, or the formula mass, for $CaBr_2$ is 199.866 amu. In this case, both atomic masses are given to three decimal places, so the overall formula mass can be calculated to three decimal places.

$$\begin{array}{rll} \text{Ca:} & 40.078 & \text{amu} \\ 2 \times \text{Br:} & 159.808 & \text{amu} \\ \hline CaBr_2\text{:} & 199.866 & \text{amu} \end{array}$$

The formula mass for a compound containing one or more polyatomic ions are calculated in the same way. The subscripts in the formula must be read carefully to make sure that each atom in the formula unit is included the appropriate number of times. For example, the formula mass for iron(II) nitrate, $Fe(NO_3)_2$, must include the mass of *one* iron atom, *two* nitrogen atoms (14.007 amu each), and *six* oxygen atoms (15.999 amu each). Because the nitrate ion is followed by a subscript of two, everything in the parentheses must be multiplied by two. The formula mass for $Fe(NO_3)_2$ is 179.853 amu.

$$\begin{array}{rll} \text{Fe:} & 55.845 & \text{amu} \\ 2 \times \text{N:} & 28.014 & \text{amu} \\ 6 \times \text{O:} & 95.994 & \text{amu} \\ \hline Fe(NO_3)_2\text{:} & 179.853 & \text{amu} \end{array}$$

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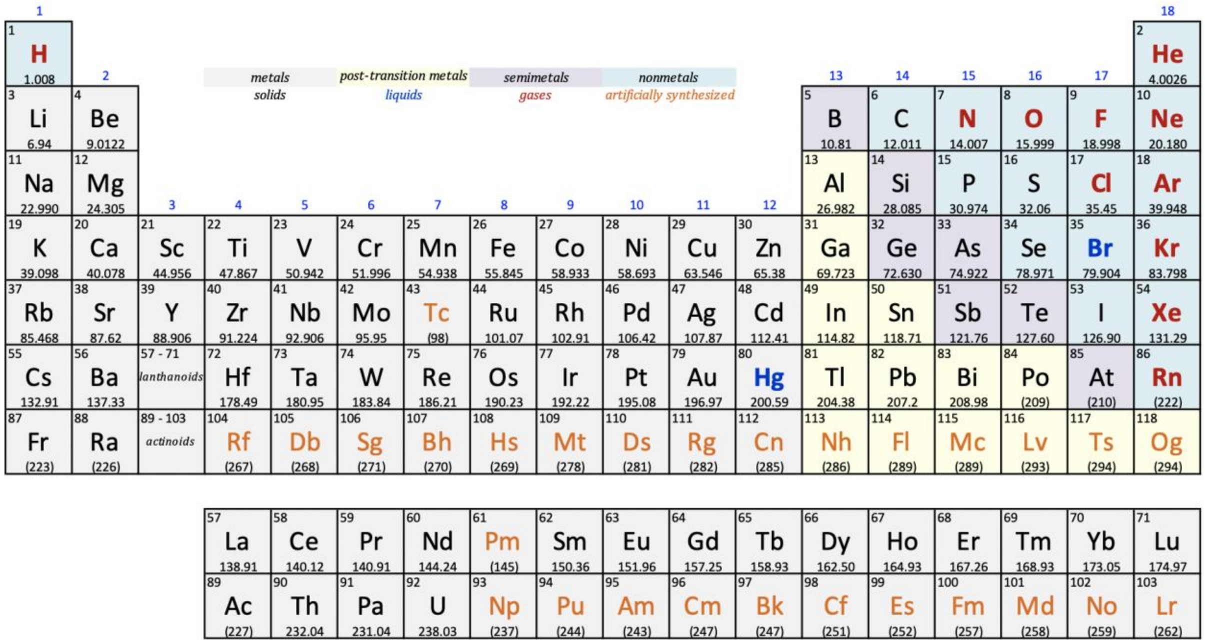

Figure 5.8 The Periodic Table of the Elements. The physical states of the elements are depicted for standard laboratory temperature and pressure. The search for elements beyond 118 has begun and researchers predict that 119 and 120 will be found in the next five years. *Based upon data from the International Union of Pure and Applied Chemistry (IUPAC).*

Example 5.6

Calculate the formula mass for each of the following compounds.

(a) KCl (b) H_2SO_4 (c) $NaNO_3$

(d) magnesium fluoride (e) lithium phosphate (f) ammonium nitrate

Solutions

Note: These answers may vary slightly if they are from a periodic table other than the one in Figure 5.8 is used.

(a) The atomic mass of potassium is 39.098 amu and the atomic mass for chlorine is 35.45 amu. The formula mass for KCl is therefore 74.55 amu.

(b) The mass for two hydrogen atoms is 2 x 1.008 amu = 2.016 amu. The mass for one sulfur atom is 32.06 amu, and the mass for four oxygen atoms is 4 x 15.999 amu = 63.996 amu. The formula mass for H_2SO_4 is therefore 98.07 amu. Using the significant figures rule for addition, there are two decimal places in the final answer.

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(c) The mass for one sodium atom is 22.990 amu. The mass for one nitrogen atom is 14.007 amu, and the mass for three oxygen atoms is 3 x 15.999 amu = 47.997 amu. The formula mass for $NaNO_3$ is therefore 84.994 amu. Using the significant figures rule for addition, there are three decimal places in the final answer.

(d) The formula unit for magnesium fluoride is MgF_2. Obtaining the correct formula unit from a compound name is important or it is impossible to determine the correct formula mass. The formula mass for MgF_2 is 62.301 amu.

(e) Lithium phosphate has the formula Li_3PO_4. The formula mass for Li_3PO_4 is 115.79 amu.

(f) Ammonium nitrate has the formula NH_4NO_3. The formula mass for NH_4NO_3 is 80.043 amu.

The chemical formula for an ionic compound is called a formula unit and the mass is the **formula mass**. The chemical formula for a molecular compound is calculated in much the same way. The main distinction is that the formula unit for a molecule represents the molecule itself and its mass can be referred to as the **molecular mass**. It is also technically correct to refer to the molecular mass as a formula mass as well, with the formula unit representing one molecule. The compound carbon tetrachloride is represented by the binary molecular formula CCl_4. Both carbon and chlorine are nonmetals, so this compound *cannot* be ionic in nature. The molecular mass for CCl_4 is calculated from the atomic mass of one carbon atom (12.011 amu) and four chlorine atoms (4 x 35.45 amu = 141.81 amu). The molecular mass for CCl_4 is therefore 153.81 amu. Again, when adding the atomic masses, the significant figures rule for addition must be applied.

$$\begin{array}{rll} \text{C:} & 12.011 & \text{amu} \\ 4 \times \text{Cl:} & 141.80 & \text{amu} \\ \hline \text{CCl}_4\text{:} & 153.81 & \text{amu} \end{array}$$

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Example 5.7

Calculate the molecular mass for each of the following compounds.

(a) CH_4 (b) SO_2 (c) PF_3

(d) dinitrogen tetroxide (e) ammonia (f) sulfur hexafluoride

Solutions

Note: These answers may vary slightly if they are from a periodic table other than the one in Figure 5.8 is used

(a) The atomic mass of carbon is 12.011 amu and the mass of four hydrogen atoms is 4 x 1.008 amu = 4.032 amu. The formula mass of the molecule CH_4 is therefore 16.043 amu.

(b) The atomic mass of sulfur is 32.06 amu and the mass of two oxygen atoms is 2 x 15.999 amu = 31.998 amu. The formula mass of the molecule SO_2 is therefore 64.06 amu.

(c) The atomic mass of phosphorus is 30.974 amu and the mass of three fluorine atoms is 3 x 18.998 amu = 56.994 amu. The formula mass of the molecule PF_3 is therefore 87.968 amu.

(d) The formula for dinitrogen tetroxide is N_2O_4. The formula mass is 92.010 amu.

(e) The formula for ammonia is NH_3. The formula mass is 17.031 amu.

(f) The formula for sulfur hexafluoride is SF_6. The formula mass is 146.048 amu.

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Chapter 5 Practice Problems

5.1 Molecules and Compounds

1. Define the following terms:
   (a) molecule (b) compound (c) element

2. Describe the difference between a compound and an element.

3. Does a molecule always have to be a compound? Explain and give examples.

4. Does an element always have to be a molecule? Explain and give examples.

5. Write a list of elements that exist at molecules.

6. Write a list of elements that exist as single atoms.

5.2 Representing Compounds

7. Which of these formulas represent molecules? Describe the components of each molecule in terms of the type and number of atoms present.
   (a) Fe (b) PCl_3 (c) P_4 (d) Ar

8. Which of these formulas represent molecules? Describe the components of each molecule in terms of the type and number of atoms present.
   (a) I_2 (b) He (c) H_2O (d) Al

9. What is the difference between CO and Co?

10. What is the difference between H_2O and H_2O_2 (hydrogen peroxide)?

11. Write a list of the names and formulas for the diatomic elements.

12. Sulfur and phosphorus are *not* diatomic elements. Write the formula for elemental sulfur. Write the formula for elemental phosphorus.

5.3 Ionic Compounds

13. Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.
    (a) potassium and fluoride (b) magnesium and sulfide
    (c) calcium and iodide (d) lithium and oxide

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14. Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.
    (a) strontium and oxygen
    (b) rubidium and sulfur
    (c) calcium and phosphorus
    (d) lithium and nitrogen

15. Predict the compound formed by each of the following combinations. Write the formula and the name of the compound. Remember to include a roman numeral if necessary.
    (a) Mg^{2+} and F^{-}
    (b) Ti^{4+} and O^{2-}
    (c) V^{3+} and O^{2-}
    (d) Mn^{4+} and P^{3-}
    (e) Fe^{2+} and Cl^{-}
    (f) Sr^{2+} and S^{2-}

16. Predict the compound formed by each of the following combinations. Write the formula and the name of the compound. Remember to include a roman numeral if necessary.
    (a) Zn^{2+} and Cl^{-}
    (b) Ca^{2+} and P^{3-}
    (c) Mg^{2+} and O^{2-}
    (d) Ni^{2+} and O^{2-}
    (e) Fe^{3+} and S^{2-}
    (f) Co^{2+} and Se^{2-}

17. Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.
    (a) lithium and nitrate
    (b) rubidium and sulfate
    (c) calcium and phosphate
    (d) magnesium and acetate
    (e) ammonium and sulfide
    (f) ammonium and sulfate

18. Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.
    (a) potassium and carbonate
    (b) lithium and cyanide
    (c) barium and hydroxide
    (d) sodium and phosphate
    (e) ammonium and hydroxide
    (f) strontium and nitrate

19. Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.
    (a) manganese ion with a 4+ charge and oxide
    (b) titanium ion with a 2+ charge and hydroxide
    (c) molybdenum ion with a 4+ charge and sulfate
    (d) zinc ion and phosphate
    (e) ammonium ion and sulfate
    (f) osmium ion with an 8+ charge and oxide

20. Predict the ionic compound formed by each of the following combinations. Write the formula and the name of the compound.
    (a) ammonium ion and hydroxide
    (b) tungsten ion with an 6+ charge and oxide
    (c) titanium ion with a 4+ charge and sulfate
    (d) cadmium and phosphate
    (e) rubidium and sulfate
    (f) cesium and oxide

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5.4 Molecular Compounds

21. Write the name of each of the following binary molecular compounds.
    (a) ClF_3 (b) CO (c) SF_4 (d) PF_3

22. Write the name of each of the following binary molecular compounds.
    (a) $TeCl_2$ (b) N_2O_3 (c) CS_2 (d) CF_4

23. Write the name of each of the following binary molecular compounds.
    (a) B_2O_3 (b) SeF_4 (c) XeF_6 (d) Cl_2O

24. Write the name of each of the following binary molecular compounds.
    (a) $XeCl_2$ (b) BF_3 (c) P_2S_3 (d) O_2F_2

25. Write the formula for each of the following binary molecular compounds.
    (a) silicon tetrachloride (b) phosphorus trichloride
    (c) dinitrogen disulfide (d) xenon hexafluoride

26. Write the formula for each of the following binary molecular compounds.
    (a) carbon disulfide (b) xenon trioxide
    (c) diselenium dibromide (d) disulfur decafluoride

27. Write the formula for each of the following binary molecular compounds.
    (a) iodine trichloride (b) diarsenic trisulfide
    (c) nitrogen triiodide (d) chlorine dioxide

28. Write the formula for each of the following binary molecular compounds.
    (a) dinitrogen trioxide (b) silicon disulfide
    (c) selenium dioxide (d) nitrogen dioxide

5.5 Acids

29. Name some of the properties that all acids have in common.

30. When is the suffix –*ate* used to name an oxoacid? When is the suffix –*ite* used?

31. Write the name of each of the following acids.
    (a) HF (b) HNO_3 (c) H_2SO_4 (d) HI

32. Write the name of each of the following acids.
    (a) CH_3COOH (b) HCN (c) H_2CO_3 (d) H_3PO_4

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33. Write the formula for each of the following acids.
   (a) nitric acid (b) acetic acid
   (c) hydrobromic acid (d) sulfuric acid

34. Write the formula for each of the following binary molecular compounds.
   (a) hydrochloric acid (b) carbonic acid
   (c) hydrocyanic acid (d) phosphoric acid

5.6 Formula Mass and Nomenclature Review

35. Fill in the table with the missing information. Write each mass to two decimal places and include units in your answer.

Compound	Name	Ionic or Molecular	Formula Mass
CO_2	*carbon dioxide*	*molecular*	*44.01 amu*
CaF_2	*calcium fluoride*	*ionic*	*78.07 amu*
TiO_2	*titanium(IV) oxide*	*ionic*	*79.86 amu*
	copper(I) sulfide		
$Zn(NO_3)_2$			
	manganese(II) phosphate		
Ag_2CO_3			
	phosphorus trichloride		
H_2O			
$FeSO_4$			
IF_7			
	ammonia		

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36. Fill in the table with the names and masses for the compound listed.

Compound	Name	Ionic or Molecular	Formula Mass
H_2CO_3			
	magnesium phosphate		
	sodium phosphide		
Cr_2S_3			
	calcium phosphate		
XeF_6			
SeO_2			
	carbon tetrachloride		
NH_3			
	molybdenum(IV) sulfide		
H_3PO_4			
	technetium(II) carbonate		
	acetic acid		

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Chapter 6
Counting Atoms

6.1 The Mole

6.2 Molar Mass

6.3 Empirical and Molecular Formulas

Figure 6.1 The sculpture *Atomium* in Brussels shows the arrangement of iron atoms in a crystal magnified 165 billion times. *Image credit: Wikimedia Commons by Mike Cattell.*

LEARNING OBJECTIVES

1. Describe the unit *mole*.
2. Relate the mole quantity of a substance to its mass.
3. Describe the composition of a compound in terms of its percent element composition.
4. Use percent composition to determine the empirical formula and the molecular formula of a compound.

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6.1 The Mole

Up until this point, chemical substances have been discussed in terms of individual formula units such as atoms and molecules. These formula units, however, are incredibly small; a single cup of water contains about a million million million million water molecules. As further concepts in chemistry are explored, it would be difficult and inconvenient to continue keeping track of substances molecule by molecule. What is needed is a way to deal with macroscopic, rather than microscopic, amounts of matter.

In chemistry, molecules and atoms are counted by groups, much like many other common items. For example, donuts are frequently purchased in groups called a *dozen*. A purchase of a dozen donuts equates to 12 donuts. Similarly, baking a dozen cookies, yields 12 cookies. A dozen of anything is twelve of those items, so a dozen is a group of 12. A similar concept applies in chemistry. When atoms and molecules are counted, they are counted by groups known as **moles** (abbreviated **mol**). One mole of something is a rather large number of that thing:

$$1 \text{ mol of atoms} = 6.022 \times 10^{23} \text{ atoms} = 602{,}200{,}000{,}000{,}000{,}000{,}000{,}000 \text{ atoms}$$

$$1 \text{ mol of molecules} = 6.022 \times 10^{23} \text{ molecules}$$

One mole is 602 sextillion of something, so it is clearly a very large number. If a mole of dollar bills needed to be counted, and everyone on earth (about 7.5 billion people) counted one bill per second, it would take about 2.5 million years to count all the bills. A mole of grains of sand would fill a cube about 32 km on a side. The height of a mole of pennies stacked on top of each other would be about the same diameter as the Milky Way galaxy. A mole is a lot of things—but atoms and molecules are very tiny. One mole of carbon atoms would make a cube that is 1.74 cm on a side, small enough to carry in a pocket.

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Figure 6.2 One mole of grains of sand would fill a cube about 32 km on a side and would contain about 80,000 times the number of grains of sand on all of the beaches and deserts on Earth. *Image credit: https://www.pexels.com.*

Chemists have a good reason to have chosen this particular number to count atoms and molecules. Known as **Avogadro's number** (N_A), it is an experimentally determined number relating the formula mass (or molar mass) of a substance in *amu* to the mass of the substance in *grams*. Another way of stating this is that amu is the same thing as g/mole. For example, the atomic mass of carbon is 12.011 amu, and the mass of one mole of carbon atoms is 12.011 grams. This relationship can be written as follows:

$$12.011 \text{ g C} = 1 \text{ mol C} = 6.022 \times 10^{23} \text{ atoms C}$$

This equality can be used as a conversion factor between the number of formula units of a substance and the number of moles of that substance (Example 6.1).

Example 6.1

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

Solution

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

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$$2.76\,\cancel{mol\ H_2O} \times \frac{6.022 \times 10^{23}\ molecules\ H_2O}{1\,\cancel{mol\ H_2O}} = 1.66 \times 10^{24}\ molecules\ H_2O$$

To determine the total number of atoms, we have the following:

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

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

Answer 2.78×10^{22} molecules

Equality statements such as these are possible for every pure substance. In section 5.6, formula masses were defined as the sum of the atomic masses for that substance. Now, thanks to Avogadro's number, masses of substances such as atoms and molecules can be expressed in the more familiar unit of grams. Remember that the formula mass or molar mass of an atom is often called its **atomic mass**, while the formula mass of a molecular substance is called its **molecular mass**. The atomic mass of an atom is the number of grams in one mole of atoms of that element while the molecular mass of a molecule is the number of grams in one mole of molecules. The formula mass of an ionic compound is the number of grams in one mole of the formula units of that ionic compound.

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Example 6.2

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

Solutions

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula but express the total in grams per mole (g/mol), not atomic mass units as in chapter 5. The masses of the atoms can be taken from the periodic table:

Element	Quantity		Atomic mass		Subtotal (g/mol)
C	6	x	12.01	=	72.06
H	12	x	1.008	=	12.096
O	6	x	16.00	=	96.00
			Total Molar Mass		180.16 g/mol

Per convention, the unit *grams per mole* is written as a fraction as g/mol or g mol^{-1}.

Test Yourself

Complete the following table:

Compound	Molar Mass
$AgNO_3$	
H_2SO_4	
$(NH_4)_3PO_4$	
C_2H_5OH	
$H_2C_2O_4$	

Answers

Compound	Molar Mass
$AgNO_3$	169.87 g/mol
H_2SO_4	98.08 g/mol
$(NH_4)_3PO_4$	149.09 g/mol
C_2H_5OH	46.07 g/mol
$H_2C_2O_4$	90.04 g/mol

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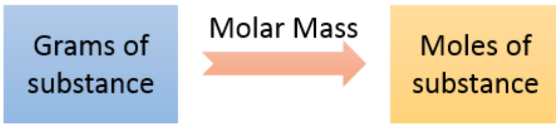
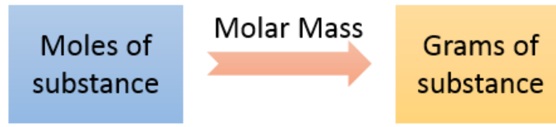
Using the molar mass of a substance, the number of moles of a certain substance can be calculated from its mass, and vice versa, as these examples illustrate. The molar mass is used as the conversion factor, as indicated below:



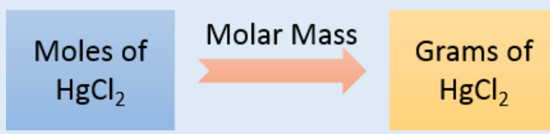


Example 6.3

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

Solution

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



$$3.56\ \cancel{mol\ HgCl_2} \times \frac{271.4\ g\ HgCl_2}{1\ \cancel{mol\ HgCl_2}} = 967\ g\ HgCl_2$$

Test Yourself What is the mass of 33.7 mol of H_2O?

Answer 607 g

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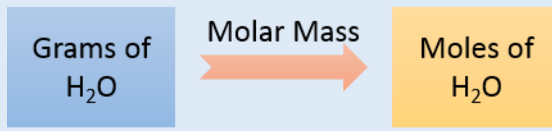
Example 6.4

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

Solution

Use the molar mass of H_2O as a conversion factor from mass to moles:



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

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

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

Answer 0.363 mol

Other conversion factors such as Avogadro's number or density can be combined with the definition of a mole.

Example 6.5

How many atoms are present in a copper penny with a mass of 2.501 g?

Figure 6.3 Pennies minted before 1982 were primarily composed of copper. Since that time, they have been made of zinc with a copper plating. *Image credit: Wikimedia Commons.*

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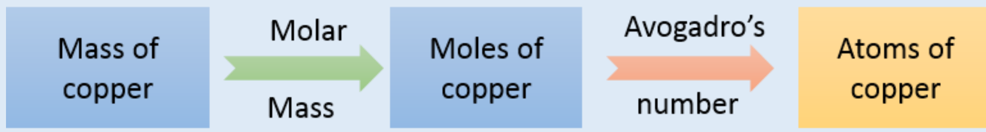
Solution

As noted earlier in the section, Avogadro's number is the number of atoms present in a mole. This can be used together with the molar mass of copper to find the number of atoms in the penny:



$$2.501\ \cancel{g\ Cu} \times \frac{1\ \cancel{mol\ Cu}}{63.55\ \cancel{g\ Cu}} \times \frac{6.022\ \times\ 10^{23}\ atoms\ Cu}{1\ \cancel{mol\ Cu}} = 2.370\ \times\ 10^{22}\ atoms\ Cu$$

Test Yourself What is the mass of 5.75×10^{24} atoms of aluminum?

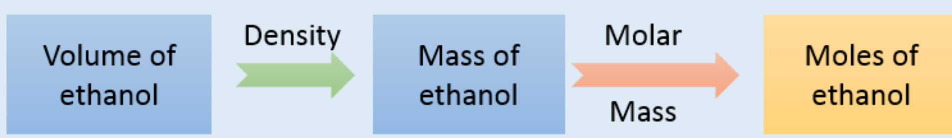
Answer 258 g Al

Example 6.6

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

Solution

Here, density is used to convert from volume to mass, and then the molar mass is used to determine the number of moles:



$$100.0\ \cancel{mL\ ethanol} \times \frac{0.789\ \cancel{g\ ethanol}}{\cancel{mL\ ethanol}} \times \frac{1\ mol\ ethanol}{46.08\ \cancel{g\ ethanol}} = 1.71\ mol\ ethanol$$

Test Yourself

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

Answer 0.201 mol

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6.2 The Composition of Compounds

Chapter 5 described how to write a chemical formula for a compound. For example, CCl_4, or carbon tetrachloride, is a molecule that contains one carbon atom and four chlorine atoms. As stated in the last section, however, it doesn't make sense to think in terms of individual molecules, since they are so small and trillions upon trillions of them are used at a time in the laboratory. Instead, the mole is used. This means that the chemical formula acts as a conversion factor – if the molecules are counted in terms of moles, the atoms making up those molecules can be counted in terms of moles as well. For example, one molecule of carbon tetrachloride contains one atom of carbon and four atoms of chlorine, since the molecular formula is CCl_4. This also means that one *mole* of carbon tetrachloride contains one *mole* of carbon atoms and four *moles* of chlorine atoms. For any chemical compound, the relationships in the chemical formula can be used to count the atoms present in this way.

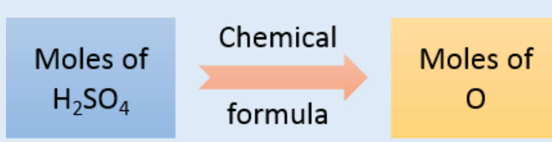
Recall that when polyatomic ions are present in an ionic compound, the entire polyatomic ion is placed in parentheses when there are more than one of them per formula unit. So the compound $(NH_4)_2SO_4$ contains one mole of sulfur atoms, four moles of oxygen atoms, two moles of nitrogen atoms, and eight moles of hydrogen atoms in one mole of the compound. There are two moles of nitrogen atoms because there are two moles of ammonium, each of which contains one nitrogen, and there are eight moles of hydrogen atoms because there are two moles of ammonium, each of which contains four hydrogen (2 x 4 = 8). This process of determining how many moles of one thing there are per mole of another thing is called **stoichiometry**.

Example 6.7

How many moles of oxygen atoms are present in 3.20 mol of sulfuric acid, H_2SO_4?

Solution

The chemical formula indicates that there are four moles of oxygen atoms present for every one mole of sulfuric acid; therefore:



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$$3.20\ \cancel{mol\ H_2SO_4} \times \frac{4\ mol\ O}{1\ \cancel{mol\ H_2SO_4}} = 12.8\ mol\ O$$

Test Yourself

How many moles of oxygen atoms are in 7.72 mol of barium nitrate? Start by writing the formula for barium nitrate.

Answer 46.3 $mol\ O$

In the laboratory, atoms cannot be counted; a chemist would measure the **mass** of a substance. Therefore, the composition of a compound is often given in terms of **mass percentage**. The mass percentage of a compound is a breakdown of the elements contained within that compound, and what percentage of the total mass is due to that element.

As an example, consider water, with the chemical formula H_2O. This formula indicates that there are two hydrogen atoms in every molecule of water, or two moles of hydrogen per mole of water. One might say that there is twice as much hydrogen in a water molecule as there is oxygen, and that would be correct, going by numbers of atoms. Keep in mind, though, that oxygen is much more massive than hydrogen, so the oxygen contributes a greater fraction of the mass of a water molecule than hydrogen does. Water, in fact, is only 11.19% hydrogen by mass, and 88.81% oxygen by mass.

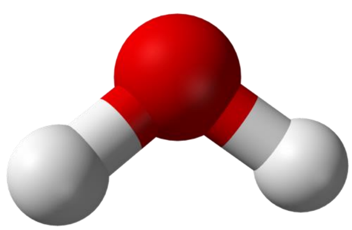

Figure 6.4 Molecular structure of water. There are twice as many hydrogen atoms as oxygen, but water is only 11.19% hydrogen by mass.

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To see how those numbers are calculated, consider first the definition of mass percentage. It can be determined by the following equation:

$$\frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

Mass percentage is an intensive property, much like density or molar mass. An intrinsic property does not vary with the amount of the substance that is present. Whether one gram, one hundred grams, or one million grams of water are present, it will always have the same mass percentage. To calculate the mass percentage, then, one can assume any amount of the compound. The easiest assumption to make is that *one mole* of the compound is present. If this is the case, then the equation above can be rewritten as:

$$\frac{\text{mass of element in one mole of the compound}}{\text{mass of one mole of the compound}} \times 100\%$$

Note the denominator of the new equation. There is an easy definition of the mass of one mole of a compound: the molar mass. The molar mass of water is 18.01 g/mol, which is the sum of the molar masses of two hydrogen and one oxygen.

What about the numerator? How can the mass of one element of a compound in a mole of that compound be determined? This is where the chemical formula is used. For water, the chemical formula H_2O indicates that there are two moles of hydrogen and one mole of oxygen in one mole of water. The mass, then, of oxygen in one mole of water is simply the molar mass of oxygen, or 15.999 g/mol. For hydrogen, since there are two moles of hydrogen atoms per mole of water, the mass in one mole of water is two times the molar mass of hydrogen, or 2 x 1.008 g/mol = 2.016 g/mol.

A general formula to determine the mass percentage of an element is a compound is:

$$\frac{\text{mole ratio} \times \text{molar mass of element}}{\text{molar mass of compound}} \times 100\%$$

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For hydrogen and oxygen in water, respectively:

$$\frac{2 \times 1.008\ \text{g/mol}}{18.01\ \text{g/mol}} \times 100\% = 11.19\%\ \text{H}$$

$$\frac{1 \times 15.999\ \text{g/mol}}{18.01\ \text{g/mol}} \times 100\% = 88.81\%\ \text{O}$$

Note that the mass percentages of all of the elements should add up to 100%, or very close to it (there may be a slight deviation due to rounding).

Example 6.8

What is the mass percentage of iron in rust, Fe_2O_3?

Figure 6.5 Iron(III) oxide is also known as rust. *Image credit: Wikimedia Commons.*

Solution

Multiply the mole ratio of iron in the compound by its molar mass, then divide by the formula mass of the compound:

$$\frac{2\ \times 55.845\ g/mol}{159.69\ g/mol} \times 100\% = 69.94\%\ \text{iron}$$

Test Yourself What is the mass percentage of fluorine in SF_6?

Answer 78.05% fluorine

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6.3 Empirical and Molecular Formulas

If it is possible to use the mole ratios within the molecular formula to determine the percent composition of a compound, as was covered in the last section, it is also possible to do the reverse: use percent composition to calculate the formula of a compound. However, absent any additional information, it is possible to calculate only the **empirical formula** from the mass percentage. The empirical formula is the simplest mole ratio of the elemental components of a compound, as opposed to the molecular formula, which is the actual elemental makeup of a compound. For example, hydrogen peroxide has the molecular formula H_2O_2, because the molecule consists of two atoms of oxygen and two atoms of hydrogen in a single chemically bound unit. The empirical formula, however, is the simplest ratio. Since hydrogen peroxide has two atoms of each, there is a 1:1 ratio of the atoms, making the empirical formula of hydrogen peroxide HO.

Example 6.9

What is the empirical formula of glucose, $C_6H_{12}O_6$?

Figure 6.6 A simple blood test is commonly used by diabetics to monitor glucose levels in the blood. *Image credit: Wikimedia Commons.*\

Solution

The mole ratios of each of the elements, 6 (for carbon), 12 (hydrogen), and 6 (oxygen) have a common denominator – each can be divided by 6. The empirical formula is the simplest ratio, which can be obtained by dividing each number by that common denominator:

$$C_{\frac{6}{6}}H_{\frac{12}{6}}O_{\frac{6}{6}} = CH_2O$$

Test Yourself What is the empirical formula of tetraphosphorus decoxide, P_4O_{10}? *Answer* P_2O_5

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To determine the mass percentage of an element in a compound, one begins by making the assumption of one mole of the compound. To find the empirical formula, it is necessary to again begin with an assumption; this time, the most convenient assumption is that there are 100 grams of the compound. With this assumption, one can infer the masses of each of the elements in the compound based on their mass percentages. The term *percent* means per one hundred, so the mass of each element in 100 grams of the compound is numerically equal to its mass percent. The mass percentages of all the elements in a substance is often called the **percent composition**. Water's percent composition is 11.19% hydrogen by mass and 88.81% oxygen by mass. With 100 grams of water, the hydrogen making up the water has a total mass of 11.19 grams, and the oxygen making up the water has a total mass of 88.81 grams.

The empirical formula of a compound is the *mole ratio* of the elements in the compound. The assumption of 100 g of the compound allows one to determine the relative masses, but it is the relative numbers of *moles* that will ultimately be used in the empirical formula (recall that water is composed of two moles of hydrogen atoms for each mole of oxygen atoms, not two grams of hydrogen for each gram of oxygen). Fortunately, we have already discussed a way to find the number of moles of an element by using the molar mass as a conversion factor.

The steps for determining the empirical formula of a compound are best illustrated via example:

Naphthalene, the compound causing the characteristic smell of mothballs, has a percent composition of 93.71% carbon and 6.29% hydrogen. Find the empirical formula of naphthalene.

Step 1. Assume 100 grams of naphthalene. Since naphthalene is 93.71 % carbon and 6.28% hydrogen, there are 93.71 g C and 6.29 g H, as discussed above.

Step 2. Find the number of moles of each element in the compound using the assumed masses and the molar masses of the elements:

$$93.71\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.01\ \cancel{\text{g C}}} = 7.813\ \text{mol C}$$

$$6.29\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 6.24\ \text{mol H}$$

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Step 3. Use the calculated number of moles to write an *interim formula* for the compound, using those moles as a subscript. This interim formula has the correct mole ratio of the elements in the compound, but it is *not* a valid empirical formula, because it does not contain whole numbers. The interim formula in this example is:

$$\mathrm{C_{7.813}H_{6.24}}$$

where the subscripts in the formula are identical to the numbers calculated in Step 2.

Step 4. Simplify the interim formula by dividing *all* numbers in the formula by the smallest calculated number of moles. In the case of naphthalene, the smallest calculated number of moles is for hydrogen, with 6.24 mol. The new formula will be:

$$\mathrm{C_{\frac{7.813}{6.24}}H_{\frac{6.24}{6.24}}} = \mathrm{C_{1.25}H}$$

Step 5. Very often, you will have a valid empirical formula after Step 4, at which point you are done with the problem. In this example case, however, the formula still does not contain entirely whole numbers, although the numbers are far more manageable. The final step, then, is to multiply all of the numbers in the formula by a whole number that will eliminate the fraction. The multiple that should be used is as follows:

Decimal in interim formula	Multiple to apply
#.5	2
#.33, #.67	3
#.25, #.75	4

In the simplified formula, the carbon has a subscript of 1.25. According to the table, a multiple of 4 should be applied to both of the elements in the compound:

$$\mathrm{C_{1.25\times4}H_{1\times4}} = \mathrm{C_5H_4}$$

Therefore, the empirical formula of naphthalene is C_5H_4,.

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Example 6.10

Iron(III) oxide, the chemical name for the substance more commonly known as rust, is 69.94% iron and 30.06% oxygen by mass. What is its empirical formula?

Solution

Following the steps outlined in the text:

Step 1. Convert percentages into masses, assuming 100 g of the compound:

69.94% Fe = 69.94 g Fe

30.06% O = 30.06 g O

Step 2. Convert masses of each element into moles.

$$69.94\ \cancel{\text{g Fe}} \times \frac{1\ \text{mol Fe}}{55.485\ \cancel{\text{g Fe}}} = 1.252\ \text{mol Fe}$$

$$30.06\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 1.879\ \text{mol O}$$

Step 3. Write the interim formula using the calculated numbers of moles as the subscripts.

$$\text{Fe}_{1.252}\text{O}_{1.879}$$

Step 4. Divide the subscripts in the interim formula by the smallest number of moles.

$$\text{Fe}_{\frac{1.252}{1.252}}\text{O}_{\frac{1.879}{1.252}} = \text{FeO}_{1.5}$$

Step 5. Clear fractions by multiplying by a common factor, if necessary.

$$\text{Fe}_{1\times2}\text{O}_{1.5\times2} = \text{Fe}_2\text{O}_3$$

Therefore, the empirical formula of iron(III) oxide is Fe_2O_3.

Test Yourself

CFC-12, one of the compounds responsible for the destruction of the ozone layer, has a percent composition of 9.93% carbon, 58.64% chlorine, and 31.43% fluorine by mass. Find its empirical formula.

Answer CCl_2F_2

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From the empirical formula, it is possible to determine the molecular formula but more information is needed: the formula mass of the compound. The molecular formula is a whole number multiple of the empirical formula. This whole number factor can be determined by finding the ratio of the formula mass of the compound to the formula mass using the empirical formula.

Consider the previous example using naphthalene. The empirical formula is C_5H_4. If the molar mass of the compound naphthalene is 128.17 g/mol, what is its molecular formula? First, find the formula mass using the empirical formula: (5 x 12.01 g/mol) + (4 x 1.008 g/mol) = 64.08 g/mol. Next, divide the molar mass of the compound by the mass calculated from the empirical formula.

$$\frac{128.17\ \mathrm{g/mol}}{64.08\ \mathrm{g/mol}} = 2$$

This is the factor that needed to multiply the empirical formula by in order to get the molecular formula:

$$(\mathrm{C_5H_4}) \times 2 = \mathrm{C_{10}H_8}$$

Therefore, the molecular formula of naphthalene is $C_{10}H_8$, as shown in Figure 6.7.

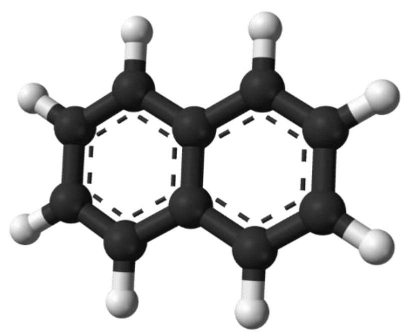

Figure 6.7 Naphthalene has the empirical formula C_5H_4 and molecular formula $C_{10}H_8$. *Image credit: Wikimedia Commons.*

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Example 6.11

Ethylene glycol, often used as a component of automobile antifreeze, is composed of 38.70% carbon, 9.74% hydrogen, and 51.55% oxygen by mass. Its molar mass is 62.07 g/mol. What is its molecular formula?

Solution

Begin by finding the empirical formula and assume 100 g total of the compound.

Step 1:

$$38.70\%\ \text{C} = 38.70\ \text{g C}$$
$$9.74\%\ \text{H} = 9.74\ \text{g H}$$
$$51.55\%\ \text{O} = 51.55\ \text{g O}$$

Step 2:

$$38.70\ \cancel{\text{g C}} \times \frac{1\ \text{mol C}}{12.011\ \cancel{\text{g C}}} = 3.22\ \text{mol C}$$

$$9.74\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.008\ \cancel{\text{g H}}} = 9.66\ \text{mol H}$$

$$51.55\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 3.22\ \text{mol O}$$

Step 3:

$$C_{3.22}H_{9.66}O_{3.22}$$

Step 4. Divide the subscripts in the interim formula by the smallest number of moles.

$$C_{\frac{3.22}{3.22}}H_{\frac{9.66}{3.22}}O_{\frac{3.22}{3.22}} = CH_3O$$

Step 5 is unnecessary since the empirical formula is composed entirely of whole numbers.

To find the molecular formula, divide the molar mass of the compound by the molar mass of the empirical formula. The molar mass of the empirical formula is:

$$(12.011\ \text{g/mol X 1}) + (1.008\ \text{g/mol X 3}) + (16.00\ \text{g/mol X 1}) = 31.035\ \text{g/mol}$$

Dividing the molar mass of the compound by this factor gives us:

$$\frac{62.07\ \text{g/mol}}{31.035\ \text{g/mol}} = 2$$

The molecular formula of ethylene glycol is: (CH_3O) X 2 = $C_2H_6O_2$

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Test Yourself

The refrigerant Freon 112 is composed of 11.78% carbon, 69.57% chlorine, and 18.64% fluorine. Its molar mass is 203.82 g/mol. What is its molecular formula?

Answer $C_2Cl_4F_2$

Chapter 6 Practice Problems

6.1 The Mole

1. How many atoms of Fe are present in 4.55 mol of Fe?
2. How many atoms of K are present in 0.0665 mol of K?
3. How many molecules of H_2S are present in 2.509 mol of H_2S?
4. How many molecules of acetylene (C_2H_2)are present in 0.336 mol of acetylene (C_2H_2)?
5. How many moles are present in 3.55×10^{24} Pb atoms?
6. How many moles of Pb are present in 2.09×10^{22} Ti atoms?
7. How many moles PF_3 are present in 1.00×10^{23} PF_3 molecules?
8. How many moles of NH_4OH are present in 5.52×10^{25} formula units of NH_4OH?

6.2 Molar Mass

9. Determine the molar mass of each substance.
   (a) Si (b) SiH_4 (c) K_2O
10. Determine the molar mass of each substance.
   (a) Cl_2 (b) $SeCl_2$ (c) $Ca(C_2H_3O_2)_2$
11. Determine the molar mass of each substance.
   (a) Al (b) Al_2O_3 (c) $CoCl_3$
12. Determine the molar mass of each substance.

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(a) O_3 (b) NaI (c) $C_{12}H_{22}O_{11}$

13. What is the mass of 4.44 mol of Rb?

14. What is the mass of 0.311 mol of Xe?

15. What is the mass of 12.34 mol of $Al_2(SO_4)_3$?

16. What is the mass of 0.0656 mol of $PbCl_2$?

17. How many moles of CO are present in 45.6 g of CO?

18. How many moles of LiF are present in 0.00339 g of LiF?

19. How many moles of SF_6 are present in 1.223 g of SF_6?

20. How many moles of $BaCO_3$are present in 48.8 g of $BaCO_3$?

21. How many atoms of C are present in 22.3 g of C?

22. How many molecules of H_2O are present in 55.3 g of H_2O?

23. How many moles of mercury are present in 54.8 mL of mercury if the density of mercury is 13.6 g/mL?

24. How many moles of O_2 are present in 56.83 mL of O_2 if the density of O_2 is 0.00133 g/mL?

25. How many moles of hydrogen atoms are present in 6.87 mol $(NH_4)_3PO_4$?

26. How many moles of carbon are present in 0.87 mol CH_3COOH?

27. What is the mass percent of carbon in C_2H_6?

28. What is the mass percent of oxygen in $Al(OH)_3$?

6.3 Empirical and Molecular Formula

29. A compound is composed of 27.29% C and 72.71% O by mass. What is its empirical formula?

30. A compound is composed of 40.0% C, 6.7% H, and 53.3% O by mass. What is its empirical formula?

31. A compound contains 24.3% C, 4.1% H, and 71.6% Cl by mass, and has a molar mass of 99 g/mol. What is its molecular formula?

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32. A compound contains 49.5% C, 5.2% H, 16.5% O, and 28.9% N by mass, and has a molar mass of 194.2 g/mol. What is its molecular formula?

Review: How to Succeed in Chemistry

33. How is your study plan working? Are you on track to earn the grade you want for the course? Schedule an appointment with your instructor if you need help getting back on track or have other concerns.

34. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

35. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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Chapter 7
Chemical Reactions

Figure 7.1 A space shuttle relies on burning a solid propellant in a chemical reaction to create enough energy to ascend into space. *Image credit: Wikimedia Commons by NASA.*

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LEARNING OBJECTIVES

1. Understand how chemical equations are used to represent chemical reactions.
2. Identify reactants, products, and write balanced chemical equations.
3. Write and balance complete ionic and net ionic reactions.
4. Given a set of reactants, predict the products of a chemical reaction.
5. Apply solubility rules in order to predict the products of a chemical reaction.
6. Identify acid-base reactions and predict the products.
7. Determine oxidation numbers and explain an oxidation-reduction reaction.
8. Classify chemical reactions.

7.1 Representing Chemical Reactions

A chemical change results from one type of matter being converted into a different type of matter during a **chemical reaction**. A chemical reaction generally occurs when bonds in the reactants are broken and different bonds in the products are formed. The substances designated as reactants have distinct composition, physical, and chemical properties compared to those of the products. Bonds between single atoms are difficult to see without specialized equipment, but changes in physical and chemical properties often indicate that a chemical reaction has taken place. Such changes include: heat production, precipitate formation, gas evolution, color changes, light emission, or fluorescence (Figure 7.2).

Figure 7.2 Indicators that may signal that chemical reaction has occurred: heat production, precipitate formation, gas evolution, color changes, or fluorescence. *Image credit: Royal Society of Chemistry and University of California (green fluorescent jellyfish).*

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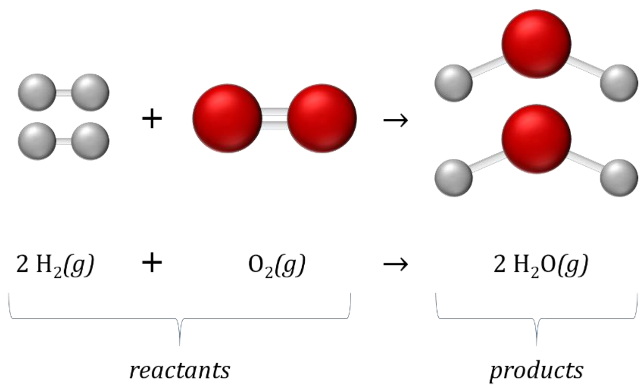
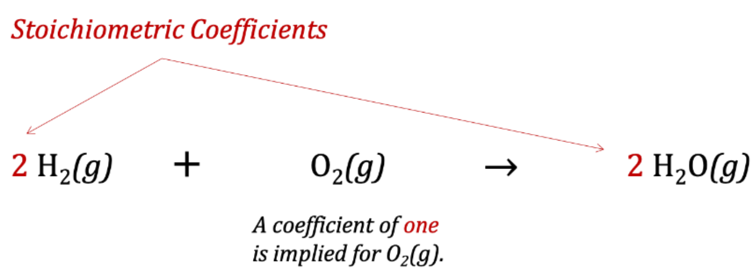
A chemical reaction is expressed by using a **chemical equation** representing the process taking place. For instance, hydrogen gas and oxygen gas react to form water. Hydrogen gas and oxygen gas are the **reactants** and are written on the **left** side of the equation. The reaction process is represented by an arrow, and the **product**, water, is listed to the **righ**t of the arrow. The physical state of each substance is often denoted by (s), (ℓ), (g), or (aq).



The numbers of molecules of each substance required for the reaction are represented by **stoichiometric coefficients** placed before each substance in the equation. If no number is written before a particular substance, the stoichiometric coefficient is one.



All chemical equations must be written with appropriate stoichiometric coefficients. In the end the number and type of each atom in the reactants must be exactly the same as the number and type of each atom in the products. The atoms may be connected differently and therefore the molecules may be different, but the atom count on each side of the equation must be the same. This is a result of the **law of conservation of matter** (Chapter 4.1). The process by which stoichiometric coefficients are determined for a chemical reaction is called **balancing**.

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Balancing Chemical Equations

Balancing a chemical equation begins with the identification of reactants and products. Consider the reaction between solid copper and aqueous silver nitrate to produce aqueous copper(II) nitrate and solid silver (Figure 7.3).

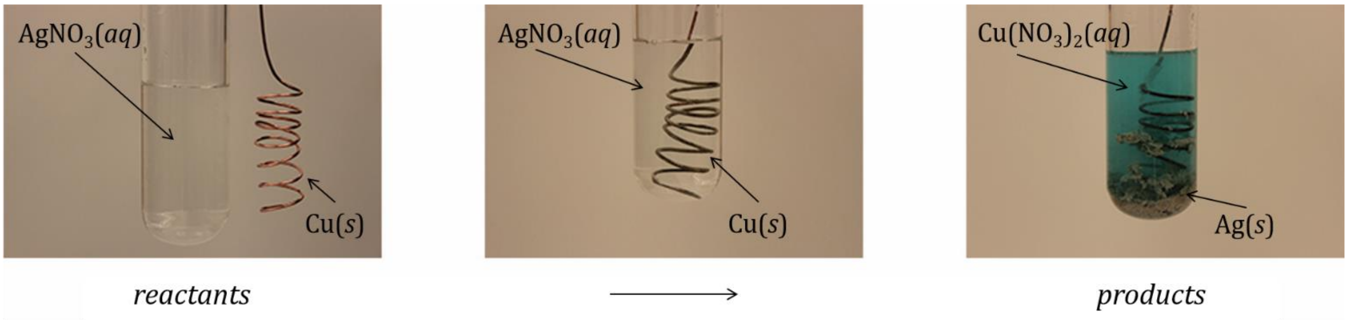

Figure 7.3 Solid copper and aqueous silver nitrate react to form solid silver and aqueous copper (II) nitrate. What changes indicate that a chemical reaction has taken place? *Image credit: OpenStax Chemistry.*

Some general steps to start balancing a chemical equation are:

1. Write the formula for each reactant and for each product. The correct formula, including the correct **subscripts**, must be used for each substance involved in the reaction. The physical states of substances are often noted when they are important for understanding the reaction.

Reactant	Formula		Product	Formula
copper *solid*	$Cu(s)$	→	copper(II) nitrate *aqueous*	$Cu(NO_3)_2(aq)$
silver nitrate *aqueous*	$AgNO_3(aq)$		silver *solid*	$Ag(s)$

2. Use the formula and physical state of each substance to write the **unbalanced** chemical equation. The reactants are written first, followed by an arrow and then the products. Within the reactants and within the products, chemical species may be listed in any order.

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

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3. Count the atoms in the **reactants**. Remember to take into account the number of atoms represented by each subscript.

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

Cu: 1

Ag: 1

N: 1

O: 3

4. Count the atoms in the **products**. Remember to take into account the number of atoms represented by each subscript.

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

Cu: 1

Ag: 1

N: 2

O: 6

It is easiest to see the fact that equation is unbalanced if the reactant atom count and the product atom count are written side by side in an atom count table.

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

reactants		*products*
1	Cu	1
1	Ag	1
1	N	2
3	O	6

5. The goal of balancing the equation is to adjust the stoichiometric coefficients such that the atom counts for the reactants match the atom counts for the products. The balancing process requires a bit of trial and error, but there are a few simple rules for getting started:

   (a) Begin with the most complicated substance. In this reaction, the product $Cu(NO_3)_2$ contains the largest number of atoms. Within this substance, start with a metal, if there is one. There is one copper atom in $Cu(NO_3)_2$. The reactant side of the equation also contains one copper atom, so no changes are needed at this point.

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(b) There are two nitrogen atoms in the product $Cu(NO_3)_2$. The reactants contain only one nitrogen atom in $AgNO_3$. To increase the nitrogen atom count on the reactant side of the equation, there must be two $AgNO_3$. A coefficient of "2" is placed in front of $AgNO_3$. When balancing, *only* change the stoichiometric coefficients, the *subscripts must not change* because the identity of the substances do not change. When updated, the atom count table becomes:

$$Cu(s) \quad + \quad \mathbf{2}\ AgNO_3(aq) \quad \rightarrow \quad Cu(NO_3)_2(aq) \quad + \quad Ag(s)$$

reactants		*products*
1	Cu	1
2 ~~1~~	Ag	1
2 ~~1~~	N	2
6 ~~3~~	O	6

(c) With this change, the silver atom count in the reactants increases and the coefficient for silver on the product side must be increased to match.

$$Cu(s) \quad + \quad \mathbf{2}\ AgNO_3(aq) \quad \rightarrow \quad Cu(NO_3)_2(aq) \quad + \quad \mathbf{2}\ Ag(s)$$

reactants		*products*
1	Cu	1
2 ~~1~~	Ag	~~1~~ 2
2 ~~1~~	N	2
6 ~~3~~	O	6

(d) When all the atom counts match, the equation is balanced.

6. Before continuing, *re-write the balanced equation*. Use this freshly re-written equation to check all atom counts. The balanced reaction will usually be the first step in a complex calculation, so taking a moment to check the atom counts is *critical and well worth the time.*

Example 7.1

When methane gas, $CH_4(g)$, is burned in oxygen, a process known as combustion occurs. The products of methane combustion are carbon dioxide gas and water vapor. Write and balance the chemical equation for this process.

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Solution

1. Write the formula for each reactant and each product.

Reactant	Formula		Product	Formula
methane *gas*	$CH_4(g)$	→	carbon dioxide *gas*	$CO_2(g)$
oxygen *gas*	$O_2(g)$		water *gas*	$H_2O(g)$

2. Write the *unbalanced* chemical equation. Include the physical state of each species.

$$CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

3. Count the atoms in the products.
4. Count the atoms in the reactants.

$CH_4(g)$	+	$O_2(g)$	→	$CO_2(g)$	+	$H_2O(g)$
		reactants		*products*		
		1	C	1		
		4	H	2		
		2	O	3		

5. Begin with CH_4. Carbon is present in only one reactant and in only one product, and CH_4 is most complex molecule containing carbon. The unbalanced equation has one carbon on the reactant side and one carbon on the product side. The hydrogen count however, indicates that the number of water molecules in the product must be multiplied by 2 so that there will be 4 hydrogen atoms on each side of the equation.

$CH_4(g)$	+	$O_2(g)$	→	$CO_2(g)$	+	**2** $H_2O(g)$
		reactants		*products*		
		1	C	1		
		4	H	~~2~~ 4		
		2	O	~~3~~ 4		

The only atom left to balance is oxygen. Adding a coefficient of 2 in front of the $O_2(g)$ completes the balancing process.

$CH_4(g)$	+	**2** $O_2(g)$	→	$CO_2(g)$	+	**2** $H_2O(g)$
		reactants		*products*		
		1	C	1		
		4	H	~~2~~ 4		
		4 ~~2~~	O	~~3~~ 4		

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Note that oxygen is present in both CO_2 and H_2O; starting with either of these molecules results in a constantly changing oxygen count and makes balancing difficult. Also note that elemental O_2 is present on the reactant side. Elements and single atoms are easy to adjust and should be left to the end.

6. Re-write the final balanced reaction and check all atom counts.

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

C: one on each side *H: four on each side* *O: four on each side*

Test Yourself

Balance the chemical equation for each of the following reactions:

(a) $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

(b) Aqueous sodium chloride reacts with aqueous lead(II) nitrate to form aqueous sodium nitrate and solid lead(II) chloride.

(c) $H_2SO_4(aq) + KOH(aq) \rightarrow K_2SO_4(aq) + H_2O(\ell)$

(d) Solid cobalt reacts with aqueous platinum(II) nitrate to form solid platinum and aqueous cobalt(II) nitrate.

Answers

(a) $\mathbf{2}\ C_2H_6(g) + \mathbf{7}\ O_2(g) \rightarrow \mathbf{4}\ CO_2(g) + \mathbf{6}\ H_2O(g)$

(b) $\mathbf{2}\ NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow \mathbf{2}\ NaNO_3(aq) + PbCl_2(s)$

(c) $H_2SO_4(aq) + \mathbf{2}\ KOH(aq) \rightarrow K_2SO_4(aq) + \mathbf{2}\ H_2O(\ell)$

(d) $Co(s) + Pt(NO_3)_2(aq) \rightarrow Co(NO_3)_2(aq) + Pt(s)$

Some chemical equations can be balanced by using fractional coefficients. For example:

$$2\ NH_3(g) + 5/2\ O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)$$

In introductory chemistry, it is best to balance all chemical equations by using whole number coefficients; calculations in Chapters 8 and 9 will be simpler with whole numbers. To balance the equation above by using whole numbers, simply multiply the entire equation by 2.

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$$2 \times [2\ NH_3(g) + 5/2\ O_2(g) \rightarrow 2\ NO(g) + 3\ H_2O(g)]$$

The balanced equation becomes:

$$\mathbf{4}\ NH_3(g) + \mathbf{5}\ O_2(g) \rightarrow \mathbf{4}\ NO(g) + \mathbf{6}\ H_2O(g)$$

The question as to why a reaction takes place is at the core of chemistry. Over time, chemists have studied many chemical processes in the laboratory and tried to develop overarching concepts to explain chemical reactivity and help predict the products of reactions that are yet to be tried. As more and more data are gathered, chemists are able to improve predictions, but there are a set of common factors known to cause a reaction to take place; these can be used to predict the products of some simple chemical processes.

7.2 Reactions in Aqueous Solution

Chemistry in aqueous solution is the basis for life on Earth. The human body is essentially a large bag containing chemicals dissolved in water. Various organs and tissues contain different solutions and mixtures of substances, but the common solvent is water. For instance, blood is an aqueous solution containing many large molecules responsible for transporting oxygen and carbon dioxide during metabolic processes. Gastric juice produced in the stomach is a colorless solution containing acids and special molecules needed for digestion. Cerebrospinal fluid is a clear, colorless solution containing ions and proteins that must be circulated through the brain and spinal column. Aside from human physiology, it is important to have an in-depth understanding of how substances act when dissolved in water in order to understand the chemistry important to everyday life. Topics such as climate change, ocean acidification, mineral formation, corrosion, and water pollution require a strong grasp of aqueous chemistry.

When certain substances dissolves in water, a solution containing ions is produced. Compounds that dissolve to produce ions are called **electrolytes**. If nearly all of the dissolved compound becomes ions, then the substance is known as a **strong electrolyte**. If only a relatively small fraction of the dissolved substance undergoes the ion-producing process, it is called a **weak electrolyte**. Substances that dissolve in water and produce no ions are termed **nonelectrolytes**.

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Substances may be identified as strong, weak, or nonelectrolytes by measuring the electrical conductance of an aqueous solution containing the substance. To conduct electricity, a substance must contain mobile, charged species. Most familiar is the conduction of electricity through metallic wires, in which case the mobile, charged entities are electrons. Solutions may also conduct electricity if they contain dissolved ions, with conductivity increasing as ion concentration increases. Applying a voltage to electrodes immersed in a solution permits assessment of the relative concentration of dissolved ions, either quantitatively, by measuring the electrical current flow, or qualitatively, by observing the brightness of a light bulb included in the circuit (Figure 7.4).

Figure 7.4 When a nonelectrolyte such as ethanol is dissolved in water, no ions form and the solution cannot conduct electricity. When a strong electrolyte such as potassium chloride is dissolved in water, all the atoms are present in ionic form, and there is strong conductivity. When a weak electrolyte such as acetic acid is dissolved in water, few ions form and the resultant conductivity is weak. *Image credit: OpenStax Chemistry.*

When a strong electrolyte such as sodium chloride is dissolved in water, the process can be represented by the chemical equation:

$$NaCl(s) + H_2O(\ell) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(\ell)$$

If an item appears on both sides of a chemical equation, it can be canceled just like in mathematical equations. In this case, liquid water appears on both sides and the equation can be simplified:

$$NaCl(s) + \cancel{H_2O(\ell)} \rightarrow Na^+(aq) + Cl^-(aq) + \cancel{H_2O(\ell)}$$

$$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$$

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If a substance has canceled but is important for understanding the reaction, then the substance can be written over the arrow. In this case, however, it is not absolutely necessary because the aqueous state of Na^+ and Cl^- implies that they are dissolved in water.

$$NaCl(s) \xrightarrow{H_2O(\ell)} Na^+(aq) + Cl^-(aq)$$

If two electrolytes are mixed in aqueous solution, a reaction between ions may take place. The chemical equation for the process can be constructed in a similar manner. For instance, when lead(II) nitrate and potassium chloride are mixed in aqueous solution, solid lead(II) chloride and aqueous potassium nitrate form. The reaction can be represented in three ways:

1. Molecular Equation: All compounds are listed.

   Aqueous lead(II) nitrate and potassium chloride react to form solid lead(II) chloride and aqueous potassium nitrate.

   Unbalanced: $Pb(NO_3)_2(aq) + KI(aq) \rightarrow PbI_2(s) + KNO_3(aq)$

   Balanced: $Pb(NO_3)_2(aq) + \mathbf{2}\ KI(aq) \rightarrow PbI_2(s) + \mathbf{2}\ KNO_3(aq)$

   **Whenever writing an equation to represent a reaction, make sure that it is balanced!*

2. Complete Ionic Equation: All ions are listed for electrolytes, and a phase designation is used for non-electrolytes.

   $Pb^{2+}(aq) + 2\ NO_3^-(aq) + 2\ K^+(aq) + 2\ Cl^-(aq) \rightarrow PbI_2(s) + 2\ K^+(aq) + 2\ NO_3^-(aq)$

3. Net Ionic Equation: Ions that are common to both sides cancel; only ions that participate in the reaction are listed. The ions that cancel are termed **spectator ions**.

   $Pb^{2+}(aq) + \cancel{2\ NO_3^-(aq)} + \cancel{2\ K^+(aq)} + 2\ Cl^-(aq) \rightarrow PbI_2(s) + \cancel{2\ K^+(aq)} + \cancel{2\ NO_3^-(aq)}$

   K^+ and NO_3^- are spectator ions and the net ionic equation is:

   $Pb^{2+}(aq) + 2\ Cl^-(aq) \rightarrow PbI_2(s)$

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Now that the conventions for representing chemical reactions have been introduced, the rules and concepts governing chemical reactions will be presented. When two or more substances are mixed, it is often possible to predict whether or not a reaction will take place, and if so, the products that will result. A chemical reaction will occur if the conditions are favorable and the resultant products are lower in energy or otherwise more stable. A common driving force for a chemical reaction is the production of a substance such as a solid, liquid, or gas. The rules for predicting the products of a chemical reaction are based upon the body of knowledge developed by chemists over time. These predictions are, however, just that. The actual products that form under experimental conditions may differ, and that is why doing chemistry in the lab is such fun!

7.3 Precipitation Reactions

Figure 7.5 When $Pb(NO_3)_2$(*aq*) and KI(*aq*) are mixed, PbI_2(*s*) forms.
Image credit: Wikimedia Commons.

A precipitation reaction occurs when two dissolved substances react to form a solid. The solid product is called a **precipitate**, and its formation often drives a reaction to completion. Because a precipitate is an **insoluble** solid, it falls to the bottom of the reaction flask (Figure 7.5). The products of many precipitation reactions may be predicted by using the **solubility rules** listed in Table 7.1.

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Table 7.1 General Solubility Rules for Aqueous Solutions

Ion	Rule	Exceptions
Li^+ Na^+ K^+ NH_4^+	soluble	---
NO_3^- CH_3COO^-	soluble	---
Cl^- Br^- I^-	soluble except with …	Ag^+ Hg_2^{2+} Pb^{2+}
SO_4^{2-}	soluble except with …	Ag^+ Hg_2^{2+} Pb^{2+} Ca^{2+} Sr^{2+} Ba^{2+}
S^{2-} O^{2-} OH^-	insoluble except with …	Li^+ Na^+ K^+ NH_4^+ Ca^{2+} Sr^{2+} Ba^{2+}
CO_3^{2-} PO_4^{3-} CrO_4^{2-}	insoluble except with …	Li^+ Na^+ K^+ NH_4^+

* *Solubilities cover a wide range from completely soluble to somewhat soluble to completely insoluble. These rules are simply a guide to be used for predicting the products of a reaction.*

* *Ra ions have the same solubility rules as calcium, strontium, and barium. Rb, Cs, and Fr ions have the same solubility rules as lithium through potassium.*

Example 7.2

Use the solubility rules in Table 7.1 to determine whether or not each of the following compounds is soluble or insoluble in aqueous solution.

(a) barium sulfate (b) calcium nitrate (c) magnesium hydroxide

Solutions

(a) Sulfates are soluble except when paired with ions of silver, mercury, lead, calcium, strontium, or barium. $BaSO_4$ is therefore insoluble.

(b) Nitrates are always soluble. $Ca(NO_3)_2$ is therefore soluble.

(c) Hydroxides are insoluble except when paired with ammonium, the group 1 cations, calcium, strontium, or barium. $Mg(OH)_2$ is therefore soluble.

Test Yourself

(a) silver bromide (b) lead(II) sulfate (c) zinc sulfide

(d) ammonium sulfate (e) lithium carbonate (f) calcium hydroxide

Answers

(a) insoluble (b) insoluble (c) insoluble (d) soluble (e) soluble (f) soluble

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Example 7.3

Write the chemical reaction that occurs for the following aqueous mixtures. For each, write the balanced molecular equation, the complete ionic equation, and the net ionic equation.

(a) silver nitrate and magnesium bromide

(b) sodium phosphate and calcium nitrate

(c) ammonium carbonate and zinc chloride

Solutions

(a) An aqueous solution containing $AgNO_3$ and $MgBr_2$ will result in formation of AgBr precipitate.

Molecular Equation: $2\ AgNO_3(aq) + MgBr_2(aq) \rightarrow 2\ AgBr(s) + Mg(NO_3)_2(aq)$

Complete Ionic Equation:

$2\ Ag^+(aq) + 2\ NO_3^-(aq) + Mg^{2+}(aq) + 2\ Br^-(aq) \rightarrow AgBr(s) + Mg^{2+}(aq) + 2\ NO_3^-(aq)$

Net Ionic Equation:

$2\ Ag^+(aq) + 2\ Br^-(aq) \rightarrow 2\ AgBr(s)$ simplifies to $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$

(b) An aqueous solution containing Na_3PO_4 and $Ca(NO_3)_2$ will result in formation of $Ca_3(PO_4)_2(s)$.

Molecular Equation: $2\ Na_3PO_4(aq) + 3\ Ca(NO_3)_2(aq) \rightarrow Ca_3(PO_4)_2(s) + 6\ NaNO_3(aq)$

Complete Ionic Equation:

$6\ Na^+(aq) + 2\ PO_4^{3-}(aq) + 3\ Ca^{2+}(aq) + 6\ NO_3^-(aq)$

$\rightarrow Ca_3(PO_4)_2(s) + 6\ Na^+(aq) + 6\ NO_3(aq)$

Net Ionic Equation: $3\ Ca^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$

(c) An aqueous solution containing $(NH_4)_2CO_3$ and $ZnCl_2$ will result in formation of $ZnCO_3(s)$.

Molecular Equation: $(NH_4)_2CO_3(aq) + ZnCl_2(aq) \rightarrow ZnCO_3(s) + 2\ NH_4Cl(aq)$

Complete Ionic Equation:

$2\ NH_4^+(aq) + CO_3^{2-}(aq) + Zn^{2+}(aq) + 2\ Cl^-(aq) \rightarrow ZnCO_3(s) + 2\ NH_4^+(aq) + 2\ Cl^-(aq)$

Net Ionic Equation: $Zn^{2+}(aq) + CO_3^{2-}(aq) \rightarrow ZnCO_3(s)$

7.4 Acid-Base Reactions

The word acid comes from the Latin word *acidus*, which means "sour." Acids familiar to most people are vinegar (acetic acid) and lemon juice (citric acid). When carbon dioxide gas dissolves in water, carbonic acid (H_2CO_3); soda water and even regular water that has been exposed to air over time taste sour. Acids were first described by Arrhenius as he experimented with the conductivity of ions in aqueous solution. He found that compounds such as HCl, HNO_3, and H_2SO_4 are strong electrolytes and completely dissolve in solution to form one or more H^+ cations and an associated anion when in aqueous solution.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

$$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$$

$$H_2SO_4(aq) \rightarrow 2\ H^+(aq) + SO_4^{2-}(aq)$$

The simplest definition of a base is a compound that produces OH^- ion in aqueous solution. Basic solutions are often referred to as **alkaline** solutions. Alkali metals and alkaline earths are named as such because they form strongly basic oxides. Common bases are the glass cleaner ammonium hydroxide, NH_4OH, and the antacid magnesium hydroxide, $Mg(OH)_2$. Basic compounds often feel slippery and basic foods often have a bitter taste. Some bases that are strong electrolytes include NaOH, KOH, and $Ba(OH)_2$.

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

$$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$$

$$Ba(OH)_2(aq) \rightarrow Ba^{2+}(aq) + 2\ OH^-(aq)$$

When an acid and a base are mixed in aqueous solution, a **neutralization** reaction occurs in which a salt and liquid water form. The formation of water, a very stable molecule, is the driving force behind acid-base reactions. For example, when aqueous sodium hydroxide and aqueous hydrochloric acid are mixed, the neutralization reaction is:

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(\ell)$$

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Sodium chloride is a soluble compound and is therefore designated as aqueous. Water is always the product of an aqueous neutralization reaction, but salt produced during a neutralization reaction is not always soluble. The solubility rules from Table 7.1 must be always be considered when writing the equation for an aqueous reaction. For example, when barium hydroxide and sulfuric acid are mixed in aqueous solution, the neutralization reaction is:

$$Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ H_2O(\ell)$$

Example 7.4

Consider the following acid-base reactions. For each, write a balanced molecular equation, a complete ionic equation, and a net ionic equation.

(a) potassium hydroxide and hydrochloric acid

(b) lithium hydroxide and sulfuric acid

(c) ammonium hydroxide and nitric acid

Solutions

(a) *Molecular Equation:* $KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(\ell)$

Complete Ionic Equation:

$K^+(aq) + OH^-(aq) + H^+(aq) + Cl^-(aq) \rightarrow K^+(aq) + Cl^-(aq) + H_2O(\ell)$

Net Ionic Equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

(b) *Molecular Equation:* $2\ LiOH(aq) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + 2\ H_2O(\ell)$

Complete Ionic Equation:

$2\ Li^+(aq) + 2\ OH^-(aq) + 2\ H^+(aq) + SO_4^{2-}(aq) \rightarrow Li_2SO_4(aq) + 2\ H_2O(\ell)$

Net Ionic Equation:

$2\ H^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(\ell)$ simplifies to $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

(c) *Molecular Equation:* $NH_4OH(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq) + H_2O(\ell)$

Complete Ionic Equation:

$NH_4^+(aq) + OH^-(aq) + H^+(aq) + NO_3^-(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq) + H_2O(\ell)$

Net Ionic Equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

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7.5 Oxidation-Reduction Reactions

The term oxidation was first used to describe reactions in which metals react with oxygen in air to produce metal oxides. When iron is exposed to air in the presence of water, for example, the iron turns to rust—an iron oxide. When exposed to air, aluminum metal develops a continuous, transparent layer of aluminum oxide on its surface.

$$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$$

The term reduction originally referred to the decrease in mass observed when a metal oxide was heated with a substance such as hydrogen gas in a process used to extract metals from ores. When solid copper(I) oxide is heated with hydrogen, for example, the mass of the original sample decreases, or is *reduced*, because pure copper forms and *oxygen atoms are lost* in the form of water vapor.

$$Cu_2O(s) + H_2(g) \rightarrow 2\ Cu(s) + H_2O(g)$$

In modern chemistry, the term **oxidation** has been broadened to include any process by which an atom loses electrons. Similarly, the term **reduction** now refers to any process by which an atom gains electrons. An oxidation-reduction reaction, or a **redox** reaction, is one in which there is gain or loss of electrons in one or more of the reactants.

Identifying gains and losses of electrons begins with the **oxidation state** of each atom in a chemical reaction. The oxidation state is the "charge" that an atom appears to have when in a compound. The rules for assigning oxidation states are as follows:

1. Atoms in substances containing only one element have a zero oxidation state.
   - *H_2, N_2, O_2, O_3, P_4, and S_8 contain atoms with zero oxidation states.*

2. Monatomic ions have an oxidation state that is the same as their charge.
   - *F^- has an oxidation state of −1. O^{2-} has an oxidation state of −2.*

3. First assign an oxidation state to each hydrogen atom and then to each oxygen atom. Then continue with the other atoms in the compound.

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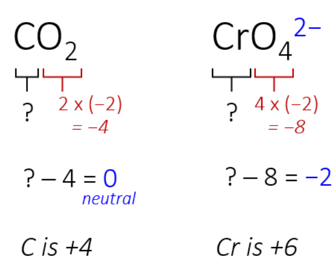
(a) Hydrogen has an oxidation state of +1 when combined with a nonmetal.

- *In the compound HF, the hydrogen atom has an oxidation state of +1. Fluoride has an oxidation state of −1.*

(b) Oxygen almost always has an oxidation state of −*2.*

- *In the compound H_2O, each hydrogen atom has an oxidation state of +1 and the oxygen atom has an oxidation state of −2.*
- *One exception is hydrogen peroxide, H_2O_2, in which each hydrogen atom has an oxidation state of +1 and each oxygen atom has an oxidation state of −1.*

4. The sum of the oxidation states of a species must equal the overall charge of that species.
    - *Carbon dioxide, CO_2, has two oxygen atoms, each with an oxidation state of −2. In order to have a neutral molecule, the oxidation state of carbon must be +4.*
    - *Chromate, CrO_4^{2-}, contains four oxygen atoms, each with an oxidation state of −2 for a total of −8. Chromium must therefore have an oxidation state of +6. When added, these leave −2 for the overall charge of the ion.*



Example 7.

Assign an oxidation state to *each* atom in the following compounds.

(a) carbon monoxide (b) iron(III) sulfate (c) zinc chloride

Solutions

(a) CO is a neutral molecule. The oxygen has an oxidation state of −2 and the carbon has an oxidation state of +2.

(b) $Fe_2(SO_4)_3$ is a neutral compound. The name of the compound states that iron has an oxidation state of +3. Oxygen has an oxidation state of −2. The oxidation state for the three sulfur atoms

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must be determined. Two iron atoms give a total of +6. Twelve oxygen atoms give a total of −24. The compound is neutral, so the total must be zero. Let x = total sulfur charge:

$$\text{iron} + \text{oxygen} + \text{sulfur} = 0$$

$$2(+3) + 12(-2) + x = 0$$

$$x = +18$$

There are three sulfur atoms in the compound, so each must have a charge of +18/3 = +6.

(c) $ZnCl_2$ is a neutral compound. Each chlorine atom has an oxidation state of −1. In order for the compound to be neutral, zinc must have an oxidation state of +2.

Test Yourself

Assign an oxidation state to *<u>each</u>* atom in the following compounds.

(a) cobalt(III) chloride	(b) $KMnO_4$	(c) silver sulfate
(d) potassium selenide	(e) OsO_4	(f) magnesium dichromate

Answers

(a) Co: +3 Cl: −1	(b) K: +1 Mn: +7 O: −2	(c) Ag: +1 S:+6 O: −2
(d) K: +1 Se: −2	(e) Os: +8 O: −2	(f) Mg: +2 Cr:+6 O: −2

Redox Reactions

In an oxidation-reduction reaction, one or more of the atoms on the reactant side of the equation undergoes a change in its oxidation state as a result of the transfer of electrons. **The loss of electrons is called oxidation.** The electron must have a place to go, and some other species in the reaction must absorb, or gain the electron. **The gain of electrons is called reduction.**

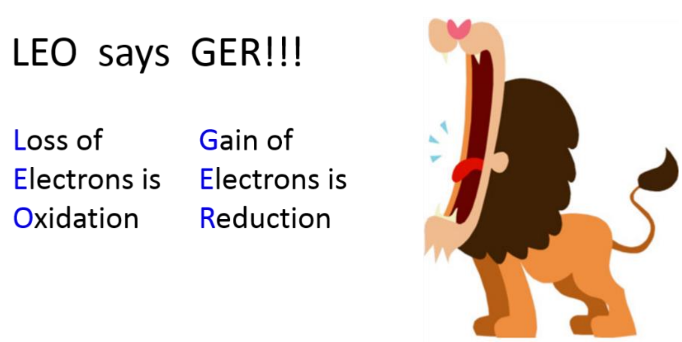

Figure 7.6 An easy way to remember what happens during oxidation and reduction is LEO says GER! *Image Credit: Wikimedia Commons.*

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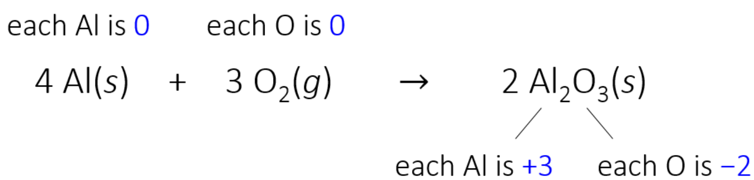
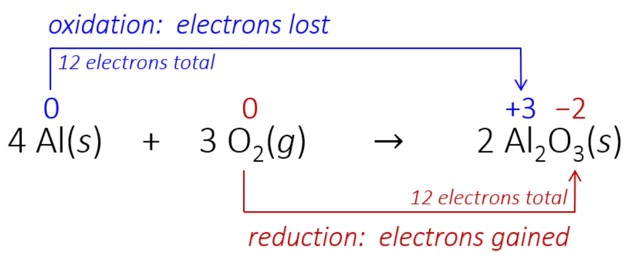
Consider the oxidation reaction presented at the beginning of this section:

$$4\ \mathrm{Al}(s) + 3\ \mathrm{O_2}(g) \rightarrow 2\ \mathrm{Al_2O_3}(s)$$

Aluminum is an element, so Al(s) has an oxidation state of zero. Aluminum reacts with oxygen gas, also an element. Each atom in $O_2(g)$ therefore has an oxidation state of zero. The product, $Al_2O_3(s)$ has an oxidation state of +3 for the aluminum atoms and −2 for the oxygen atoms.



When this reaction takes place, each Al(s) atom loses three electrons to become Al^{3+}. For the four aluminum atoms, a total of 12 electrons are lost. Each oxygen atom in $O_2(g)$ gains two electrons to become O^{2-}. There are six oxygen atoms in the reactants, so a total of 12 electrons are gained.



It is crucial to remember that when an atom loses and electron, they must have a place to go. In the reaction above, the 12 electrons lost from aluminum are taken up by the oxygen. This means that, for every oxidation, there *must* be an associated reduction. Oxidation and reduction are coupled processes.

Example 7.6

Write the molecular equation for each oxidation-reduction reaction described. List the oxidation state for each atom in the reactants and for each atom in the products.

(a) Barium solid reacts with oxygen gas to produce solid barium oxide.

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(b) Zinc solid reacts with aqueous HCl to produce zinc chloride and hydrogen gas.

(c) Carbon solid reacts with elemental sulfur to produce liquid carbon disulfide.

Solutions

(a) 2 Ba(*s*) + O_2(*g*) → 2 BaO(*s*)

Reactants: Ba is 0 O is 0 Products: Ba is +2 O is −2

(b) Zn(*s*) + 2 HCl(*aq*) → $ZnCl_2$(*aq*) + H_2(*g*)

Reactants: Zn is 0 H is +1 Cl is −1 Products: Zn is +2 H is 0

(c) 4 C(*s*) + S_8(*s*) → 4 CS_2(*ℓ*)

Reactants: C is 0 S is 0 Products: C is +4 S is −2

Oxidizing Agents and Reducing Agents

In the reaction between solid aluminum and oxygen gas, the aluminum atoms in the reactants lose electrons, and they are transferred to oxygen atoms. The aluminum that loses the electrons therefore *causes* oxygen to be reduced. Al(*s*) is termed the **reducing agent**. A reducing agent is the reactant that gets oxidized during a redox reaction. Similarly, the oxygen atoms in the reactants pick up electrons. These oxygen atoms therefore *cause* oxidation and O_2(*g*) is called the **oxidizing agent**. An oxidizing agent is the reactant that gets reduced during a redox reaction.

Al gives up electrons and causes reduction.

Al(s) is the reducing agent.

$$\overset{0}{4\,Al(s)} \;+\; \overset{0}{3\,O_2(g)} \;\rightarrow\; 2\,\overset{+3}{Al_2}\overset{-2}{O_3}(s)$$

O_2 is the oxidizing agent.

O takes electrons and causes oxidation.

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Example 7.7

Write the molecular equation for each oxidation-reduction reaction described. Identify both the oxidizing agent and the reducing agent.

(a) Copper(I) oxide reacts with hydrogen gas to produce copper solid and water vapor.

(b) Methane gas, $CH_4(g)$, and oxygen gas react to produce CO_2 gas and water vapor.

(c) $Zn(s)$ and aqueous iron(II) nitrate react to produce aqueous zinc nitrate and iron solid.

Solutions

(a) $Cu_2O(s) + H_2(g) \rightarrow 2\ Cu(s) + H_2O(g)$ Ox. Agent: $Cu_2O(s)$ Red. Agent: $H_2(g)$

(b) $CH_4(g) + 2\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$ Ox. Agent: $O_2(s)$ Red. Agent: $CH_4(g)$

(c) $Zn(s) + Fe(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Fe(s)$

Ox. Agent: $Fe(NO_3)_2(aq)$ Red. Agent: $Zn(s)$

Earlier in this chapter, a table of solubility rules was given to help predict the products of precipitation reactions. The products of oxidation-reduction reactions can also be predicted given a bit of extra information. In order to predict if electrons will be transferred between one metal and another, an **activity series** can be used (Table 7.2). The metals in an activity series are listed according to reactivity in an oxidation reaction. The metals at the top of the series are most reactive and readily give up an electron to become a positive ion. In Chapter 4, the formation of ions and the prediction of ion charge from position on the periodic table was discussed (Figure 4.12). Recall that alkali metals readily form 1+ ions and alkaline earths readily form 2+ ions. This means that these metals are easily oxidized – the loss of electrons occurs spontaneously making these metals especially active. Alkali metals such as lithium and potassium appear at the top of the activity series.

The metals at the bottom of the series are relatively unreactive and prefer to hold on to electrons and stay in solid, elemental form. Precious metals such as gold, platinum, and silver are examples of metals that prefer to remain in the pure, solid state. These metals are difficult to oxidize and appear at the bottom of the activity series.

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If a very reactive metal such a zinc comes into contact with the *cation of a less reactive metal* such as platinum, electrons will be transferred from the zinc (an oxidation) and to the platinum (a reduction) and the reaction will occur spontaneously.

$$Zn(s) + Pt(NO_3)_2(aq) \rightarrow Pt(s) + Zn(NO_3)_2(aq)$$

Without spectator ions, the net reaction is:

$$Zn(s) + Pt^{2+}(aq) \rightarrow Pt(s) + Zn^{2+}$$

By contrast, if a relatively unreactive metal such at copper comes into contact with the *cation of a more reactive metal* such as Co^{2+}, the copper will keep its electrons and remain as the elemental solid. Cobalt prefers to be in the 2+ form and will not accept electrons from copper.

$$Cu(s) + Co(NO_3)_2(aq) \nrightarrow \textit{no reaction}$$

When a reaction contains a solid metal and another metal cation, consult the activity table to determine whether it will occur spontaneously.

Table 7.2 Activity Series of Selected Elements

Metal	Reactivity	Halogen
Li(*s*)	most reactive	F_2
K(s)	↓	Cl_2
Mg(s)	↓	Br_2
Al(*s*)	↓	I_2
Zn(*s*)	reactivity	
Ni(s)	decreases	
Pb(*s*)	↓	
Cu(*s*)	↓	
Ag(*s*)	↓	
Pt(*s*)	↓	
Au(*s*)	least reactive	

**These are activities at standard laboratory conditions.*

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The halogens can also be ordered according to their reactivity and ability to replace other halogens. For instance, the element fluorine, $F_2(g)$, is the most reactive of the halogens. A fluorine atom will readily replace an atom of any halogen below it. For example, aqueous sodium bromide will react with fluorine gas to produce aqueous sodium fluoride and bromine gas.

$$NaBr(aq) + F_2(g) \rightarrow NaCl(aq) + Br_2(g)$$

If aqueous sodium fluoride were to be mixed with chlorine gas, no reaction would occur. Chlorine is below fluorine on the activity table and will not replace it.

$$NaF(aq) + Cl_2(g) \nrightarrow \textit{no reaction}$$

Example 7.8

Determine whether or not each reaction will occur spontaneously. If a reaction does occur, complete and balance the equation. Write out both the molecular and net ionic equations.

(a) $Zn(s) + Pb(NO_3)_2(aq) \rightarrow$

(b) $Ni(s) + Co(NO_3)_2(aq) \rightarrow$

(c) $Pb(s) + Pt(NO_3)_2(aq) \rightarrow$

Solutions

(a) $Zn(s)$ is higher than $Pb(s)$ on the reactivity table. Zn prefers to be Zn^{2+} and Pb prefers to be $Pb(s)$. A spontaneous reaction occurs.

$$Zn(s) + Pb(NO_3)_2(aq) \rightarrow Pb(s) + Zn(NO_3)_2(aq)$$

$$Zn(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Zn^{2+}(aq)$$

(b) $Ni(s)$ is lower than $Co(s)$ on the reactivity table. Ni prefers to be $Ni(s)$. Co prefers to be Co^{2+} in this case. No reaction occurs.

(c) $Pb(s)$ is higher than $Pt(s)$ on the reactivity table. In relative terms, Pb prefers to be Pb^{2+} and Pt prefers to be $Pt(s)$. A spontaneous reaction occurs.

$$Pb(s) + Pt(NO_3)_2(aq) \rightarrow Pt(s) + Pb(NO_3)_2(aq)$$

$$Pb(s) + Pt^{2+}(aq) \rightarrow Pt(s) + Pb^{2+}(aq)$$

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Example 7.9

Determine whether or not the following reactions will occur. Write the balanced equation for the reaction that do occur.

(a) $CaCl_2(aq) + I_2(g) \rightarrow$ (b) $Br_2(g) + KI(aq) \rightarrow$ (c) $AlF_3(aq) + Cl_2(g) \rightarrow$

Solutions

(a) Iodine is lower than chlorine on the activity table on the reactivity table. No reaction.

(b) Bromine is higher than iodine on the activity table, so a reaction will occur.

$Br_2(g) + KI(aq) \rightarrow KBr(aq) + I_2(g)$

(c) Chlorine is lower than fluorine on the activity table on the reactivity table. No reaction.

7.6 Classifying Chemical Reactions

The chemical reactions introduced in this chapter are only a tiny sampling of the infinite number of chemical reactions possible. How do chemists cope with this overwhelming diversity? How do they predict which compounds will react with one another and what products will be formed? The key to success is to find useful ways to categorize reactions. Familiarity with a few basic types of reactions will help you to predict the products that form when certain kinds of compounds or elements come in contact.

Most chemical reactions can be classified into one of five groups:

1. Synthesis: $A + B \rightarrow AB$
2. Decomposition: $AB \rightarrow A + B$
3. Double Displacement: $AB + CD \rightarrow AD + CB$
4. Single Replacement: $A + CD \rightarrow AD + C$
5. Combustion: $A + O_2 \rightarrow A_xO_y$

   $C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$

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Synthesis

In a synthesis reaction, two or more substances react to form a more complex substance. The reaction can be represented in a general manner by:

$$A + B \rightarrow AB$$

The reactant substances may be solid, liquid, gaseous, or aqueous. For instance, solid potassium combines with bromine gas to produce solid potassium bromide.

$$2\ K(s) + Br_2(g) \rightarrow 2\ KBr(s)$$

Some other examples of synthesis reactions are:

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

$$8\ Fe(s) + S_8(s) \rightarrow 8\ FeS(s)$$

$$Li_2CO_3(s) + H_2O(\ell) + CO_2(g) \rightarrow 2\ LiHCO_3(aq)$$

$$6\ CO_2(g) + 6\ H_2O(\ell) + \mathit{light} \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$$

The last reaction is **photosynthesis** and occurs when green plants use sunlight to produce food such as glucose.

Decomposition

Decomposition reaction are easy to spot because a complex initial substance decomposes into simpler substances.

$$AB \rightarrow A + B$$

For instance, solid potassium chlorate, $KClO_3(s)$, decomposes into potassium chloride and oxygen gas.

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

Some other examples of decomposition reactions are:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

$$H_2O_2(aq) \rightarrow H_2(g) + O_2(g)$$

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$$4\ C_3H_5(NO_3)_3(\ell) \rightarrow 6\ N_2(g) + O_2(g) + 12\ CO_2(g) + 10\ H_2O(\ell)$$

The last reaction is the decomposition of nitroglycerin, an explosive process that releases a large amount of energy.

Double Displacement

A double displacement reaction occurs when positive and negative ions of two different ionic compounds exchange places to form two new compounds.

$$AB + CD \rightarrow AD + CB$$

In this example, the cations are A and C while the anions are B and D. When they swap, cation A must pair with anion D and cation C must pair with anion B. The driving force for double displacement reactions is usually the formation of a stable solid, gas, or liquid water.

During double displacement reactions the ion charges do not change. If a reactant in a double displacement contains Fe^{3+}, for example, one of the products will have Fe^{3+} in it as well. The precipitation reactions presented in Section 7.3 and the acid-base neutralization reactions presented in Section 7.4 are double displacement reactions. Double displacement reactions are typically written as molecular equations unless otherwise specified.

Examples of some double displacement reactions are:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(\ell) + NaCl(aq)$$

$$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow 2\ KNO_3(aq) + PbI_2(s)$$

$$NaCN(aq) + HBr(aq) \rightarrow NaBr(aq) + HCN(g)$$

$$Na_2S(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2S(g)$$

Single Replacement

In a single replacement reaction, one element replaces a similar element in a compound. The general form of single replacement reaction is

$$A + CD \rightarrow AD + C$$

Oxidation-reduction reactions involving solid metals are often single replacement reactions. For example, solid magnesium metal reacts with aqueous copper(II) nitrate to produce solid copper and aqueous magnesium nitrate.

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$$

Halogen replacement reactions are also classified as single replacement. Chlorine gas reacts with aqueous sodium bromide to produce aqueous sodium chloride and bromine gas.

$$Cl_2(g) + 2\ NaBr(aq) \rightarrow 2\ NaCl\ (aq) + Br_2(g)$$

Examples of some other single replacement reactions are:

$$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

$$2\ Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2\ Fe(s)$$

$$2\ Na(s) + 2\ H_2O(\ell) \rightarrow 2\ NaCl(aq) + H_2(g)$$

Combustion

Combustion is a special type of redox reaction as discussed at the beginning of Section 7.5. If the reactants contain only carbon, hydrogen, and oxygen, a combustion reaction can be represented by the general equations:

$$A + O_2 \rightarrow A_xO_y$$

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

The A in the first equation represents an element that burns in oxygen gas to product an oxide. In the second equation, the combustion of an organic substance containing only carbon, hydrogen, and oxygen is shown. The products of this type of combustion are carbon dioxide and water.

Combustion reactions are often challenging to balance due to the fact that there may be odd *vs.* even numbers of oxygen atoms on different sides of the equation to start.

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Examples of some balanced combustion reactions are:

$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$$

$$P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$$

$$Ti(s) + O_2(g) \rightarrow TiO_2(s)$$

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

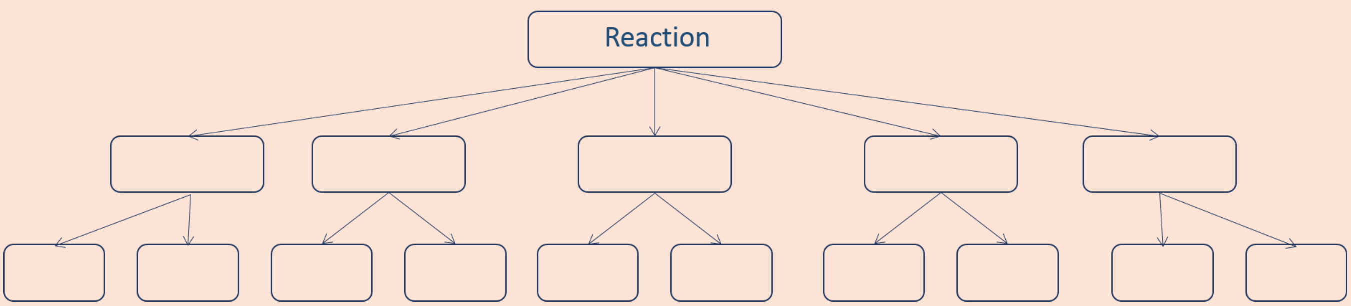
$$C_2H_5OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(g)$$

$$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$$

Classifying Reactions – Create a Decision Tree

Draw a decision tree for classifying chemical reactions.

1. Write the reaction and identify its components.
2. Place the reaction into one of the five categories listed at the beginning of Section 7.6.
3. Within each classification, are there any special reaction subgroups that you can identify?



Keep adding layers, arrows, and boxes!

Use the decision tree you've created to work through the example reactions given in the chapter and in the practice problems. Revise your decision tree until it works for all cases!

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Chapter 7 Practice Problems

7.1 Representing Chemical Reactions

1. For each of the following reactions, use the given information and find the missing coefficient.
   (a) $C_3H_8(g)$ + ____ $O_2(g)$ → 3 $CO_2(g)$ + 4 $H_2O(g)$
   (b) 2 $Al(s)$ + ____ $H_2SO_4(aq)$ → $Al_2(SO_4)_3(aq)$ + 3 $H_2(g)$
   (c) $BaO_2(s)$ + $H_2SO_4(aq)$ → ____ $BaSO_4(s)$ + $H_2O_2(aq)$
   (d) 2 $C_6H_6(\ell)$ + 15 $O_2(g)$ → ____ $CO_2(g)$ + 6 $H_2O(g)$

2. For each of the following reactions, use the given information and find the missing coefficient.
   (a) $C_8H_{18}(\ell)$ + ____ $O_2(g)$ → 16 $CO_2(g)$ + 18 $H_2O(\ell)$
   (b) $V_2O_5(s)$ + 2 $H_2(g)$ → $V_2O_3(s)$ + ____ $H_2O(\ell)$
   (c) $Na_3PO_4(s)$ + ____ $HCl(aq)$ → 3 $NaCl(aq)$ + $H_3PO_4(aq)$
   (d) $C_6H_{12}O_6(aq)$ + ____ $O_2(g)$ → 6 $CO_2(g)$ + 6 $H_2O(\ell)$

3. Write the balanced chemical equation for each of the following reactions:
   (a) Butane gas, $C_4H_{10}(g)$, reacts with oxygen during a combustion reaction to produce gaseous carbon dioxide and liquid water.
   (b) Zinc solid reacts with hydrochloric acid to produce zinc chloride and hydrogen gas.
   (c) Lithium oxide reacts with liquid water to produce aqueous lithium hydroxide.
   (d) Silver iodide and sodium sulfide react to form silver sulfide and sodium iodide.

4. Write the balanced chemical equation for each of the following reactions:
   (a) Nitric oxide gas, $NO(g)$, reacts with carbon monoxide gas to produce nitrogen gas and carbon dioxide gas.
   (b) Carbon monoxide gas and hydrogen gas react to form octane, $C_8H_{18}(\ell)$ and liquid water.
   (c) Titanium(IV) chloride reacts with liquid water to produce solid titanium(IV) oxide and hydrochloric acid.
   (d) Ammonia gas, $NH_3(g)$, reacts with oxygen gas to form nitric oxide gas, $NO(g)$, and water vapor.

5. Balance each of the following chemical equations:
   (a) $As(s)$ + $NaOH(aq)$ → $Na_3AsO_3(g)$ + $H_2(g)$
   (b) $CO_2(g)$ + $H_2O(\ell)$ → $C_6H_{12}O_6(aq)$ + $O_2(g)$
   (c) $C_2H_6O(\ell)$ + $O_2(g)$ → $CO_2(g)$ + $H_2O(g)$
   (d) $LiAlH_4(s)$ + $AlCl_3(s)$ → $AlH_3(s)$ + $LiCl(s)$

6. Balance each of the following chemical equations:

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(a) $Ba_3N_2(s)$ + $H_2O(\ell)$ → $Ba(OH)_2(aq)$ + $NH_3(aq)$
(b) $FeS(s)$ + $O_2(g)$ → $Fe_2O_3(s)$ + $SO_2(g)$
(c) $PCl_5(s)$ + $H_2O(\ell)$ → $H_3PO_4(aq)$ + $HCl(aq)$
(d) $KClO_3(s)$ → $KCl(s)$ + $O_2(g)$

7. Balance each of the following chemical equations:
(a) $SnO_2(s)$ + $H_2(g)$ → $Sn(s)$ + $H_2O(\ell)$
(b) $KOH(aq)$ + $H_2SO_4(aq)$ → $K_2SO_4(aq)$ + $H_2O(\ell)$
(c) $C_6H_{14}O(\ell)$ + $O_2(g)$ → $CO_2(g)$ + $H_2O(g)$
(d) $B_2Br_6(aq)$ + $HNO_3(aq)$ → $B(NO_3)_3(aq)$ + $HBr(aq)$

8. Balance each of the following chemical equations:
(a) $H_2(g)$ + $N_2(g)$ → $NH_3(g)$
(b) $AsH_3(g)$ + $O_2(g)$ → $As_2O_3(s)$ + $H_2O(\ell)$
(c) $GeCl_4(s)$ + $H_2S(g)$ → $GeS_2(s)$ + $HCl(aq)$
(d) $Hg(OH)_2(s)$ + $H_3PO_4(aq)$ → $Hg_3(PO_4)_2(aq)$ + $H_2O(\ell)$

9. Balance the following chemical equation:
The combustion of AlC_3H_9 produces solid aluminum oxide, carbon dioxide gas, and liquid water.

10. Balance the following chemical equation:
The reaction between liquid $C_4H_{10}O_2$ and fluorine gas produces C_2F_6 gas, hydrofluoric acid gas, and oxygen difluoride gas.

11. *Challenge Problem – Attempt this only after you have completed all the practice problems for this chapter*. Balance the following chemical equation:
In aqueous solution, copper(I) sulfide reacts with nitric acid to yield copper(II) nitrate, copper(II) sulfate, nitrogen dioxide gas, and liquid water.

12. *Challenge Problem – Attempt this only after you have completed all the practice problems for this chapter*. Balance the following chemical equation:
$CuSCN(aq)$ + $KIO_3(aq)$ + $HCl(aq)$ → $CuSO_4(aq)$ + $KCl(aq)$ + $HCN(aq)$ + $ICl(aq)$ + $H_2O(\ell)$

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7.2 Reactions in Aqueous Solution

13. Write the balanced molecular, complete ionic, and net ionic equations for each of the following reactions.
    (a) $Ag(s) + H_2SO_4(aq) \rightarrow Ag_2SO_4(s) + H_2(g)$
    (b) $CaCl_2(aq) + H_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + HCl(aq)$
    (c) $K_2CrO_4(aq) + SrI_2(aq) \rightarrow SrCrO_4(s) + KI(aq)$
    (d) $NH_4Br(aq) + Pb(NO_3)_2(aq) \rightarrow NH_4NO_3(aq) + PbBr_2(s)$

14. Write the balanced molecular, complete ionic, and net ionic equations for each of the following reactions.
    (a) $Ba(OH)_2(aq) + MgSO_4(aq) \rightarrow BaSO_4(s) + Mg(OH)_2(aq)$
    (b) $CaCl_2(aq) + Na_3PO_4(aq) \rightarrow Ca_3(PO_4)_2(s) + NaCl(aq)$
    (c) $I_2(g) + NaBr(aq) \rightarrow NaI(aq) + Br_2(g)$
    (d) $AgNO_3(aq) + AlCl_3(aq) \rightarrow AgCl(s) + Al(NO_3)_3(aq)$

15. Write the balanced molecular, complete ionic, and net ionic equations for each of the following reactions. Assume that the species are aqueous unless otherwise noted.
    (a) Silver nitrate reacts with sodium bromide to form solid silver bromide and $NaNO_3$.
    (b) Sodium carbonate and copper(II) chloride react to form sodium chloride and copper(II) carbonate.
    (c) Ammonium chloride and potassium hydroxide react to form ammonium hydroxide and potassium chloride
    (d) Calcium chloride and potassium carbonate react to form solid calcium carbonate and potassium chloride.

16. Write the balanced molecular, complete ionic, and net ionic equations for each of the following reactions.
    (a) Nitric acid reacts with potassium hydroxide to form potassium nitrate and water.
    (b) Silver nitrate and copper solid react to from copper(II) nitrate and silver solid.
    (c) Sodium chloride and bromine liquid react to form sodium bromide and chlorine gas.
    (d) Sodium sulfide reacts with hydrochloric acid to form aqueous sodium chloride and hydrogen sulfide gas.

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7.3 Precipitation Reactions

17. Use the solubility rules in Table 7.1 to complete the reactions given below. Write the balanced molecular equation and net ionic equation for each precipitation reaction. If a precipitation does not occur, simply write the balanced molecular equation.
   (a) $Pb(NO_3)_2(aq)$ + $NaI(aq)$ →
   (b) $AgNO_3(aq)$ + $MgCl_2(aq)$ →
   (c) $Na_3PO_4(aq)$ + $(NH_4)_2S(aq)$ →
   (d) $HNO_3(aq)$ + $NaOH(aq)$ →

18. Use the solubility rules in Table 7.1 to complete the reactions given below. Write the balanced molecular equation and net ionic equation for each precipitation reaction. If a precipitation does not occur, simply write the balanced molecular equation.
   (a) $K_2CO_3(aq)$ + $CaI_2(aq)$ →
   (b) $HCl(aq)$ + $KOH(aq)$ →
   (c) $(NH_4)_3PO_4(aq)$ + $Mg(OH)_2(aq)$ →
   (d) $Ba(NO_3)_2(aq)$ + $(NH_4)_2SO_4(aq)$ →

19. Use the solubility rules in Table 7.1 to determine whether or not a reaction will take place in each of the aqueous mixtures given. If a precipitation reaction takes place, write the balanced molecular equation and net ionic equation.
   (a) lead(II) nitrate and sodium carbonate
   (b) potassium chloride and silver nitrate
   (c) calcium bromide and ammonium hydroxide
   (d) silver acetate and sodium sulfate

20. Use the solubility rules in Table 7.1 to complete the reactions given below. Write the balanced molecular equation and net ionic equation for each precipitation reaction. If a precipitation does not occur, simply write the balanced molecular equation. Assume that the species are aqueous unless otherwise noted.
   (a) ammonium chloride and lead(II) acetate
   (b) magnesium hydroxide solid and nitric acid
   (c) copper(II) nitrate and potassium hydroxide
   (d) magnesium bromide and iron(II) acetate

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7.4 Acid-Base Reactions

21. For the following aqueous acid-base combinations, predict the products of a neutralization reaction. Write a balanced molecular equation, complete ionic equation, and net ionic equation for each.
    (a) HCl and KOH
    (b) H_2SO_4 and KOH
    (c) H_3PO_4 and $Ni(OH)_2$

22. For the following aqueous acid-base combinations, predict the products of a neutralization reaction. Write a balanced molecular equation, complete ionic equation, and net ionic equation for each.
    (a) HBr and $Fe(OH)_3$
    (b) HNO_2 and $Al(OH)_3$
    (c) $HClO_3$ and $Mg(OH)_2$

23. Complete and balance the following neutralization reactions. Write a balanced molecular equation, complete ionic equation, and net ionic equation for each.
    (a) $HI(aq) + KOH(aq) \rightarrow$
    (b) $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow$

24. Complete and balance the following neutralization reactions. Write a balanced molecular equation, complete ionic equation, and net ionic equation for each.
    (a) $HNO_3(aq) + Fe(OH)_3(s) \rightarrow$
    (b) $H_3PO_4(aq) + CsOH(aq) \rightarrow$

25. Write the complete and net ionic equations for the neutralization reaction between $HClO_3(aq)$ and $Mg(OH)_2(s)$.

26. Write the complete and net ionic equations for the neutralization reaction between $H_2C_2O_4(s)$ and $KOH(aq)$.

27. Explain why the net ionic equation for the neutralization reaction between $HCl(aq)$ and $KOH(aq)$ is the same as the net ionic equation for the neutralization reaction between $HNO_3(aq)$ and RbOH.

28. Compare the net ionic equation for the neutralization reaction between $HCl(aq)$ and $KOH(aq)$ is different from the net ionic equation for the neutralization reaction between $H_2SO_4(aq)$ and $Ba(OH)_2$.

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29. Write the complete and net ionic equations for the neutralization reaction between HCl(aq) and KOH(*aq*) using the hydronium ion in place of H^+. What difference does it make when using the hydronium ion?

30. Write the complete and net ionic equations for the neutralization reaction between HClO3(aq) and $Zn(OH)_2(s)$ using the hydronium ion in place of H^+. Assume the salt is soluble. What difference does it make when using the hydronium ion?

7.5 Oxidation-Reduction Reactions

31. Assign an oxidation state to *each* atom in the following compounds.

(a) K_2O (b) $ZnCl_2$ (c) MnO_2
(d) $Tc(CO_3)_2$ (e) WO_3 (f) $CaCrO_4$

32. Assign an oxidation state to *each* atom in the following compounds.

(a) In_2O_3 (b) $KMnO_4$ (c) CCl_4
(d) SO_2 (e) $PtCl_4$ (f) P_4O_{10}

33. Assign an oxidation state to *each* atom in the following ions.

(a) nitrate (b) chromate (c) acetate
(d) $Cr_2O_7^{2-}$ (e) $C_2O_4^{2-}$ (f) $Cr_2O_7^{2-}$

34. Assign an oxidation state to *each* atom in the following ions.

(a) carbonate (b) ammonium (c) phosphate
(d) $Se_2O_3^{2-}$ (e) $RuCl_4^{2-}$ (f) ClO_4^-

35. Use the activity series in Table 7.2 to determine whether or not a reaction occurs. If a reaction occurs, complete the equation, balance, and write the net ionic equation.

(a) $K(s) + Al(NO_3)_3(aq) \rightarrow$
(b) $NaNO_3(aq) + Ag(CH_3OO)(aq) \rightarrow$
(c) $Li(s) + Pb(CH_3OO)_2(aq) \rightarrow$
(d) $Ag(s) + Pt(NO_3)_2(aq) \rightarrow$

36. Use the activity series in Table 7.2 to determine whether or not a reaction occurs. If a reaction occurs, complete the equation, balance, and write the net ionic equation.

(a) $Co(s) + ZnCl_2(aq) \rightarrow$
(b) $Ni(NO_3)_2(aq) + Co(s) \rightarrow$
(c) $Pb(s) + Cu(NO_3)_2(aq) \rightarrow$
(d) $AuCl_3(aq) + AgNO_3(aq) \rightarrow$

7.6 Classifying Chemical Reactions

Use the concepts presented in this chapter to complete and balance each of the following reactions. Assume standard laboratory conditions and include the phases of all reaction components. Classify each reaction according to the groups presented in Section 7.6.

35.
- (a) $Zn + Fe(NO_3)_2 \rightarrow$
- (b) $CH_4 + O_2 \rightarrow$
- (c) $AgNO_3 + MgCl_2 \rightarrow$
- (d) $CaO + CO_2 \rightarrow$
- (e) $F_2 + FeI_3 \rightarrow$
- (f) $Li + MgSO_4 \rightarrow$
- (g) $NaBr + Cl_2 \rightarrow$
- (h) $Sn + H_2SO_4 \rightarrow$
- (i) $Al + NiBr_2 \rightarrow$
- (j) $H_2 + S_8 + 0_2 \rightarrow H_2SO_4$
- (k) $HI + Br_2 \rightarrow$
- (l) $FeCl_2 + Br_2 \rightarrow$
- (m) $Li + N_2 \rightarrow$
- (n) $AgNO_3 + MgCl_2 \rightarrow$
- (o) $Fe(NO_3)_3 + Al \rightarrow$
- (p) $Zn + Fe_3(PO_4)_2 \rightarrow$
- (q) $Ag + HNO_3 \rightarrow$
- (r) $C_8H_{18} + O_2 \rightarrow$
- (s) $NaI + Cl_2 \rightarrow$
- (t) $AgCl + Au \rightarrow$

36.
- (a) $Pt + H_3PO_4 \rightarrow$
- (b) $Li + H_2O \rightarrow$
- (c) $Zn(NO_3)_2 + NaOH \rightarrow$
- (d) $HCl + Na_2S \rightarrow$
- (e) $Zn + H_2SO_4 \rightarrow$
- (f) $Ca(OH)_2 + HNO_3 \rightarrow$
- (g) $Na_2CO_3 + Sr(NO_2)_2 \rightarrow$
- (h) $Pb(NO_3)2 + KBr \rightarrow$
- (i) $CH_3CH_2CH_2OH + O_2 \rightarrow$
- (j) $K_2O + MgCO_3 \rightarrow$
- (k) $Sn(OH)_2 + FeBr_3 \rightarrow$
- (l) $CsNO_3 + KCl \rightarrow$
- (m) $Pb(NO_3)_2 + KBr \rightarrow$
- (n) $K_3PO_4 + SrCl_2 \rightarrow$
- (o) $C_5H_{12} + O_2 \rightarrow$
- (p) $4\ C_3H_5(NO_3)_3(\ell) \rightarrow N_2 + O_2 + CO_2 + H_2O$
- (q) $KCl + Ag_2SO_4 \rightarrow$
- (r) $KNO_3 + Li_2CO_3 \rightarrow$
- (s) $C_6H_6 + O_2 \rightarrow$
- (t) $KOH + AgNO_3 \rightarrow$

Review: How to Succeed in Chemistry

37. How is your study plan working? Are you on track to earn the grade you want for the course? Schedule an appointment with your instructor if you need help getting back on track or have other concerns.

38. Look up the date of each of your final exams. Make a study schedule for all your classes so that you are prepared when final exam week arrives!

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Chapter 8
Stoichiometry

Quantities are important in science, especially in chemistry. It is important to make accurate measurements of a variety of quantities when performing experiments. However, it is also important to be able to relate one measured quantity to another, unmeasured quantity. In this chapter, we will consider how to relate quantities in a chemical reaction.

When baking, the correct ratio of ingredients is very important. Just like a baker would follow a recipe when making this pound cake, chemists refer to chemical equations for the ratio of chemicals needed to make a product. *Image credit: Wikimedia Commons*

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LEARNING OBJECTIVES

1. Define *stoichiometry.*
2. Relate quantities in a balanced chemical reaction on a molecular basis.
3. Relate quantities of atoms within a chemical formula.
4. Balance a chemical equation in terms of moles.
5. Use a balanced equation to construct conversion factors in terms of moles.
6. Perform mole/mole and mole/mass conversion using balanced chemical equations.
7. Define and determine theoretical yields, actual yields, and percent yields.
8. Find the limiting reagent for a chemical reaction and perform calculations based upon that limiting reagent.

8.1 Stoichiometry

Consider a classic recipe for pound cake: 1 pound of eggs, 1 pound of butter, 1 pound of flour, and 1 pound of sugar. (That's why it's called "pound cake.") If you have 4 pounds of butter, you would need 4 pounds each of sugar, flour, and eggs.

Now suppose you have 1.00 g H_2. If the chemical reaction proceeds according to the balanced chemical equation,

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(\ell)$$

then what mass of oxygen do you need to make water?

Curiously, this chemical reaction question is very similar to the pound cake question. Both of them involve relating a quantity of one substance to a quantity of another substance. The relating of one chemical substance to another using a balanced chemical reaction is called **stoichiometry** (pronounced STOY-kee-OM-ah-tree). Using stoichiometry is a fundamental skill in chemistry; it greatly broadens your ability to predict what will occur and, more importantly, how much of a product is produced in a chemical reaction.

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Consider a more complicated example. A recipe for pancakes calls for 2 cups (c) of pancake mix, 1 egg, and 1/2 c of milk. We can write this in the form of a balanced equation:

2 c mix + 1 egg + 1/2 c milk → 1 batch of pancakes

Image by Michael Stern
Wikimedia Commons

If you have 9 c of pancake mix, how many eggs and how much milk do you need? It might take a little bit of work, but eventually you will find you need 4½ eggs and 2¼ c milk.

How can this be formalized? A conversion factor can be made using our original recipe and that conversion factor can be used to convert from a quantity of one substance to a quantity of another substance, similar to the way we constructed a conversion factor between feet and yards in Chapter 2. Because one recipe's worth of pancakes requires 2 c of pancake mix, 1 egg, and 1/2 c of milk, the following mathematical relationships can be derived to relate these quantities:

2 c pancake mix ≡ 1 egg ≡ 1/2 c milk

≡ is the mathematical symbol for "is equivalent to." This does not mean that 2 c of pancake mix equal 1 egg. However, *as far as this recipe is concerned*, these are the equivalent quantities needed for a single recipe of pancakes. So, any possible quantities of two or more ingredients must have the same numerical ratio as the quantities in the equivalence if the recipe is being followed. If the ingredients are not in that numerical ratio there will be some ingredients left over when other ingredients have run out.

These equivalences can be dealt with in the same way as unit conversions, by making conversion factors that essentially equal 1. For example, to determine how many eggs we need for 9 c of pancake mix, two conversion factors can be constructed that relate eggs to pancake mix.

$$\frac{1\ egg}{2\ c\ pancake\ mix} \quad or \quad \frac{2\ c\ pancake\ mix}{1\ egg}$$

These conversion factors are equivalent to 1 because the recipe relates the two equivalent quantities. In order to determine which conversion factor is needed, the rules of dimensional

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analysis from Chapter 2 can be applied. Starting with the initial quantity and multiplying by the conversion factor, the "*cups of pancake mix"* cancels, leaving the units of "*eggs".*

$$9\ \cancel{\text{c pancake mix}} \times \frac{1\ \text{egg}}{2\ \cancel{\text{c pancake mix}}} = 4\frac{1}{2}\ \text{eggs}$$

This is the formal, mathematical way of getting our amounts to mix with 9 c of pancake mix. A similar conversion factor for the amount of milk can be used.

$$9\ \cancel{\text{c pancake mix}} \times \frac{\frac{1}{2}\ \text{c milk}}{2\ \cancel{\text{c pancake mix}}} = 2\frac{1}{4}\ \text{c milk}$$

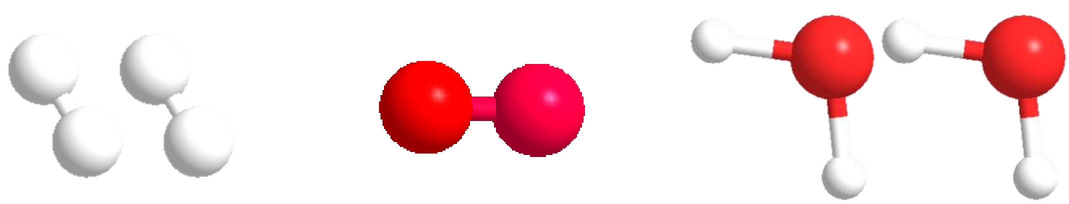
Again, units cancel, and new units are introduced.

A balanced chemical equation is nothing more than *a recipe for a chemical reaction*. The difference is that a balanced chemical equation is written in terms of formula units such as atoms and molecules, not cups, pounds, and eggs.

For example, consider the following chemical equation:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(\ell)$$

This can be interpreted literally as "two hydrogen molecules react with one oxygen molecule to make two water molecules." That interpretation directly leads to some equivalences, just as the pancake recipe did.



2 H_2 molecules ≡ 1 O_2 molecule ≡ 2 H_2O molecules

These equivalences allow us to use the coefficients from a balanced equation to construct conversion factors. The ≡ symbol means that the factors are equivalent to each other.

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$$\frac{2 \text{ molecules } H_2}{1 \text{ molecule } O_2} \equiv \frac{2 \text{ molecules } H_2}{2 \text{ molecules } H_2O} \equiv \frac{1 \text{ molecule } O_2}{2 \text{ molecules } H_2O}$$

These conversions factors can be used to relate quantities of one substance to quantities of another. For example, suppose you need to know how many molecules of oxygen are needed to react with **16 molecules of H_2**. In a similar manner to converting units, start with the given quantity and use the appropriate conversion factor.

$$16 \cancel{\text{ molecules } H_2} \times \frac{1 \text{ molecule } O_2}{2 \cancel{\text{ molecules } H_2}} = 8 \text{ molecules } O_2$$

Note how the unit *molecules H_2* cancels algebraically, just as any unit does in a conversion like this. The conversion factor came directly from the coefficients in the balanced chemical equation.

Example 8.1

How many molecules of SO_3 are needed to react with 144 formula units of Fe_2O_3 given the following balanced chemical equation? (Remember, since Fe_2O_3 is ionic, its chemical formula expresses one formula unit, not one molecule.)

$$Fe_2O_3(s) + 3\ SO_3(g) \rightarrow Fe_2(SO_4)_3(s)$$

Solution

Use the balanced chemical equation to construct a conversion factor between Fe_2O_3 and SO_3. The number of formula units of Fe_2O_3 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO_3 go on the top. The conversion factor is:

$$\frac{3 \text{ molecules } SO_3}{1 \text{ formula unit } Fe_2O_3}$$

Starting with the given amount and applying the conversion factor, the result is

$$144 \cancel{\text{ formula units } Fe_2O_3} \times \frac{3 \text{ molecules } SO_3}{1 \cancel{\text{ formula unit } Fe_2O_3}} = 432 \text{ molecules } SO_3$$

432 molecules of SO_3 are needed to react with 144 formula units of Fe_2O_3.

Test Yourself

How many molecules of H_2 are needed to react with 29 molecules of N_2 to make ammonia if the balanced chemical equation is $N_2 + 3\ H_2 \rightarrow 2\ NH_3$? *Answer* 87 molecules

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Example 8.2

How many molecules of NH_3 can you make if you have 228 atoms of H?

Solution

From the formula, we know that one molecule of NH_3 has three H atoms. Use that fact as a conversion factor:

$$228\ \cancel{\text{atoms}} \times \frac{1\ \text{molecule NH}_3}{1\ \cancel{\text{atoms}}} = 76\ \text{molecules NH}_3$$

Test Yourself How many formula units of $Fe_2(SO_4)_3$ can you make from 777 atoms of S?

Answer 259 formula units

8.2 The Mole in Chemical Reactions

Consider this balanced chemical equation:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

This is interpreted as "two molecules of hydrogen react with one molecule of oxygen to make two molecules of water." The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:

$$100\ H_2(g) + 50\ O_2(g) \rightarrow 100\ H_2O(g)$$

This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:

$$500\ H_2(g) + 250\ O_2(g) \rightarrow 500\ H_2O(g)$$

Again, this is not conventional, but it is still balanced. Suppose a much larger number is used:

$$12.044 \times 10^{23}\ H_2(g)\ +\ 6.022 \times 10^{23}\ O_2(g)\ \rightarrow\ 12.044 \times 10^{23}\ H_2O(g)$$

These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro's number, while the second number is Avogadro's number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

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Notice, this is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. This chemical equation can be read as "two moles of hydrogen react with one mole of oxygen to make two moles of water." **All balanced chemical reactions are balanced in terms of moles**.

Example 8.3

Interpret this balanced chemical equation in terms of moles.

$$P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$$

Answer

The coefficients represent the number of moles involved in the reaction , not individual molecules. This would be described as "one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide."

Test Yourself Interpret this balanced chemical equation in terms of moles.

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

Answer

One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

In Section 8.1, the chemical equation was introduced as a simple recipe for a chemical reaction. As such, chemical equations also give equivalences between the reactants and the products. However, now it can be said that *these equivalences are expressed in terms of moles*. Consider the balanced chemical equation

$$2\ H_2(g)\ +\ O_2(g)\ \rightarrow\ 2\ H_2O(g)$$

This chemical reaction gives us the following equivalences:

$$2\ moles\ H_2\ \equiv\ 1\ mole\ O_2\ \equiv\ 2\ moles\ H_2O$$

Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance.

$$\frac{2\ moles\ H_2}{1\ mole\ O_2} = \frac{2\ moles\ H_2}{2\ moles\ H_2O} = \frac{1\ mole\ O_2}{2\ moles\ H_2O}$$

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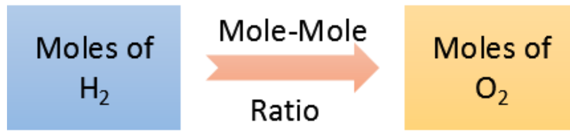
If, for example, if a chemist needs to know how many moles of oxygen will react with 17.6 moles of hydrogen, a conversion factor between 2 moles of H_2 and 1 mole of O_2 should be used to convert from moles of H_2 to moles of O_2 according to the concept map:



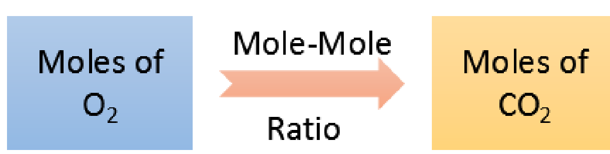
$$17.6\ \cancel{\text{mol H}_2} \times \frac{1\ \text{mol O}_2}{2\ \cancel{\text{mol H}_2}} = 8.80\ \text{mol O}_2$$

Note how the unit *mol H_2* cancels, and mol O_2 is the new unit introduced. This is an example of a **mole-mole calculation**, where you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the chemical equation is straightforward, but you may have to work with more complex chemical equations!

Consider the balanced equation below that represents the combustion of propane in oxygen to produce carbon dioxide and water. Check the equation below to make sure all elements are balanced.

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$

In the previous problem, the amount of the other *reactant* used in a reaction was calculated. The balanced equation can also be used to predict how much *product* will be formed from a given reactant. Consider if you have 6.25 moles O_2 and want to determine the amount of CO_2 produced.



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$$6.25\ \cancel{\text{mol O}_2} \times \frac{3\ \text{mol CO}_2}{5\ \cancel{\text{mol O}_2}} = 3.75\ \text{mol CO}_2$$

The calculation indicates that 3.75 mol CO_2 will be produced given that the C_3H_8 does not run out before the reaction is complete. This is called the **theoretical yield** because it is the amount that should be collected if all of the O_2 reacts.

Example 8.4

For the balanced chemical equation:

$$2\ C_4H_{10}(g) + 13\ O_2\ (g) \rightarrow 8\ CO_2(g) + 10\ H_2O(\ell)$$

if 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution
In order to relate the amount of oxygen to an amount of carbon dioxide, the equivalence between these two substances is needed. According to the balanced chemical equation, the equivalence is *13 mol O_2 ≡ 8 mol CO_2.*
This equivalence is used to construct the proper conversion factor. Start with the given amount and apply the conversion factor:

$$154\ \cancel{\text{mol O}_2} \times \frac{8\ \text{mol CO}_2}{13\ \cancel{\text{mol O}_2}} = 94.8\ \text{mol CO}_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they don't contribute to the number of significant figures in the final answer.

Test Yourself
Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?
Answer 118 mol H_2O

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or liters—but moles. Any stoichiometry problem will require the chemicals involved to be in units of moles at some point.

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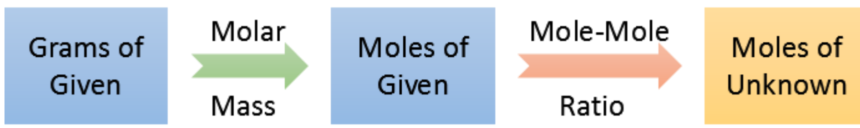
8.3 Mole-Mole and Mole-Mass Conversions

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall that the molar mass can be determined from a chemical formula and used as a conversion factor. The molar mass can be added as another step in a calculation to make a **mole-mass calculation**, where a given number of moles of a substance is used as the starting point to calculate the mass of another substance involved in the chemical equation. Study the steps in the concept map below.



For example, suppose we have the balanced chemical equation

$$2\ \mathrm{Al}(s) + 3\ \mathrm{Cl_2}(g) \rightarrow 2\ \mathrm{AlCl_3}(s)$$

If there is a sample of 123.2 g of Cl_2, how can the number of moles of $AlCl_3$ be determined when the reaction is complete? First and foremost, ***chemical equations are not balanced in terms of grams; they are balanced in terms of moles***. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of $AlCl_3$, first convert the given amount of Cl_2 into moles. This type of calculation was first introduced in Chapter 6 by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. This conversion factor must be inverted so that the units cancel properly:

$$123.2\ \cancel{\mathrm{g\ Cl_2}} \times \frac{1\ \mathrm{mol\ Cl_2}}{70.90\ \cancel{\mathrm{g\ Cl_2}}} = 1.738\ \mathrm{mol\ Cl_2}$$

Now that the quantity is in moles, the balanced chemical equation can be used to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of $AlCl_3$. The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2\ \mathrm{mol\ AlCl_3}}{3\ \mathrm{mol\ Cl_2}}$$

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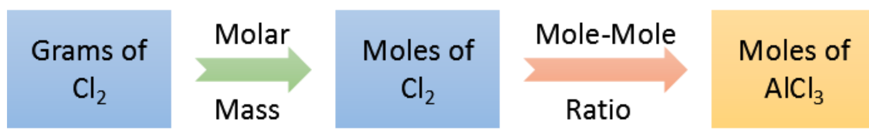
Using this conversion factor with the molar quantity calculated above, the moles of $AlCl_3$ can be calculated.

$$1.738\ \cancel{\text{mol Cl}_2} \times \frac{2\ \text{mol AlCl}_3}{3\ \cancel{\text{mol Cl}_2}} = 1.159\ \text{mol AlCl}_3$$

This means that 1.159 mol of $AlCl_3$ will be produced if 123.2 g of Cl_2 is reacted with excess aluminum.

In the last example, the calculation was done in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line, as shown below.



$$123.2\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{2\ \text{mol AlCl}_3}{3\ \cancel{\text{mol Cl}_2}} = 1.159\ \text{mol AlCl}_3$$

This is the preferred method because it minimizes the number of steps where an answer is rounded since the numbers can be input into the calculator all at once. The units still cancel appropriately, and the same numerical answer is achieved in the end.

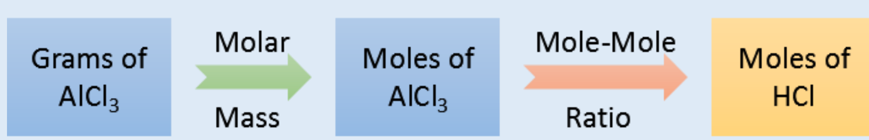
Example 8.5

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2\ AlCl_3 + 3\ H_2O(\ell) \rightarrow Al_2O_3 + 6\ HCl(g)$$

Solution

This can be done in two steps. First, convert the mass of $AlCl_3$ to moles, and then use the balanced chemical equation to find the number of moles of HCl formed.



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$$249\ \cancel{\text{g AlCl}_3} \times \frac{1\ \cancel{\text{mol AlCl}_3}}{133.33\ \cancel{\text{g AlCl}_3}} \times \frac{6\ \text{mol HCl}}{2\ \cancel{\text{mol AlCl}_3}} = 5.60\ \text{mol HCl}$$

Test Yourself
How many moles of Al_2O_3 will be produced when 23.9 g of H_2O are reacted according to the chemical equation above?
Answer 0.442 mol Al_2O_3

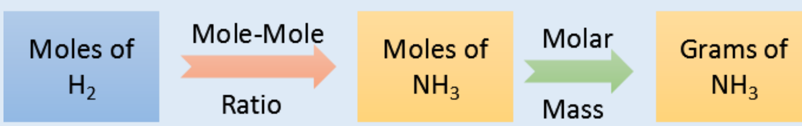
A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams, as shown in Example 8.6. The steps are the same but are performed in reverse order.

Example 8.6

How many grams of NH_3 will be produced when 33.9 mol of H_2 are reacted according to this chemical equation?

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

Solution
This can be performed in two steps. First, use the balanced chemical equation to convert to moles of another substance, and then use its molar mass to determine the mass of the final substance.



$$33.9\ \cancel{\text{mol H}_2} \times \frac{2\ \cancel{\text{mol NH}_3}}{3\ \cancel{\text{mol H}_2}} \times \frac{17.03\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}} = 385\ \text{g NH}_3$$

Test Yourself
How many grams of N_2 are needed to produce 2.17 mol of NH_3 when reacted according to this chemical equation?

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

Answer 30.4 g
Note: Here the calculation goes from a product to a reactant, showing that mole-mass problems can begin and end with <u>any substance</u> in the chemical equation.

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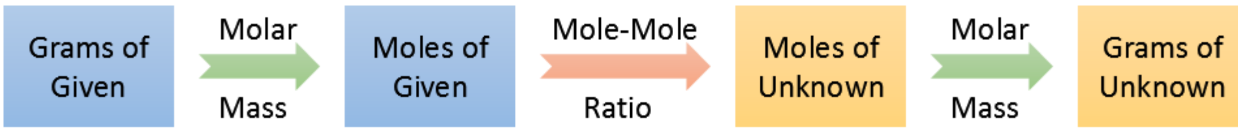
It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with the mass of some substance and end with the mass of another substance related by the same chemical equation. As shown in the concept map below, the calculation must proceed by working through the mole.

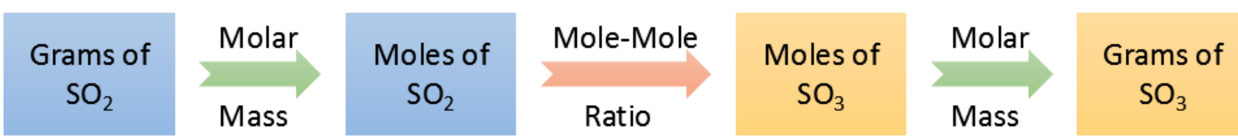


For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to **moles**. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO_3 that can be produced by the reaction of 45.3 g of SO_2 and excess O_2:

$$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$$

This is a three-step calculation, as shown by the concept map below.



Despite involving three steps, this problem should be set up using one long sequence of conversion factors and input into your calculator as a single problem.

$$45.3\ \cancel{g\,K} \times \frac{1\ \cancel{mol\ SO_2}}{64.06\ \cancel{g\ SO_2}} \times \frac{2\ \cancel{mol\ SO_3}}{2\ \cancel{mol\ SO_2}} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

Note how the initial and all of the intermediate units cancel, leaving grams of SO_3, which is the final answer.

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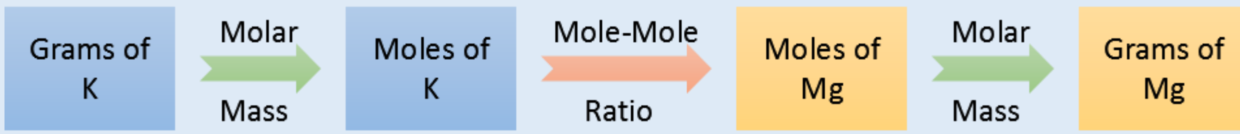
Example 8.7

What mass of Mg will be produced when 86.4 g of K are reacted?

$$MgCl_2(s) + 2\ K(s) \rightarrow Mg(s) + 2\ KCl(s)$$

Solution

Follow the steps outlined in the concept map:



$$86.4\ \cancel{g\ K} \times \frac{1\ \cancel{mol\ K}}{39.09\ \cancel{g\ K}} \times \frac{1\ \cancel{mol\ Mg}}{2\ \cancel{mol\ K}} \times \frac{24.31\ g\ Mg}{1\ \cancel{mol\ Mg}} = 26.9\ g\ Mg$$

Test Yourself

What mass of H_2 will be produced when 122 g of Zn are reacted?

$$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

Answer 3.77 g

8.4 Yields

In previous calculations that involving balanced chemical equations, two assumptions were made: (1) the reaction goes exactly as written, and (2) the reaction proceeds completely. In reality, side reactions may occur that make some chemical reactions rather messy. For example, in the actual combustion of some carbon-containing compounds, such as a log in a fireplace, some CO is produced as well as CO_2. This is why carbon monoxide detectors should be placed next to fireplaces.

In many calculations, we assume that side reactions are minimal. The second assumption, that the reaction proceeds completely, is a bigger concern because in practice, no reaction proceeds to 100% yield. By calculating an amount of product assuming that all the reactant reacts, the **theoretical yield** is determined.

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Figure 8.1 Combustion of a log involves side reactions that produce carbon monoxide. The presence of soot at the end shows that the reaction does not proceed completely. *Image credit: Wikimedia Commons.*

In many cases, this is not what really happens; sometimes much less product is made during the course of a chemical reaction. For example, when combustion engines work hard, fuel used to power the engine is not completely burned. When you see black soot coming from exhaust, the engine is not burning its fuel efficiently.

Figure 8.2 At a tractor pull, the diesel fueling the engine is not completely burned, so some of the output is soot rather than the normal products of combustion, water and carbon dioxide. *Image credit: [Principles of General Chemistry](). *

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The amount that is actually produced in a reaction is called the **actual yield**. By definition, the actual yield is less than or equal to the theoretical yield. If it is not, then an error has been made or a product has not been purified completely. For example, a product crystalized from out of solution in water might not be completely dried so the measured mass of the product is actually the mass of the product plus some water. Both theoretical yields and actual yields are expressed in units of moles or grams. It is also common to see something called a percent yield. The **percent yield** is a comparison between the actual yield and the theoretical yield and is defined as

$$Percent\ Yield = \frac{actual\ yield}{theoretical\ yield} \times 100\%$$

It does not matter whether the actual and theoretical yields are expressed in moles or grams, as long as they are expressed in the same units. However, the percent yield always has units of percent. Proper percent yields are between 0% and 100%— if percent yield is greater than 100%, an error has been made.

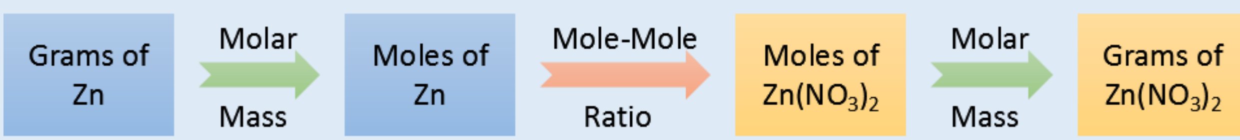
Example 8.8

A chemist reacts 30.5 g of Zn with nitric acid and evaporates the remaining water to obtain 65.2 g of $Zn(NO_3)_2$. What are the theoretical yield, the actual yield, and the percent yield?

$$Zn(s) + 2\ HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$$

Solution

A mass-mass calculation can be performed to determine the theoretical yield.



$$30.5\ \cancel{g\ Zn} \times \frac{1\ \cancel{mol\ Zn}}{65.39\ \cancel{g\ Zn}} \times \frac{1\ \cancel{mol\ Zn(NO_3)_2}}{1\ \cancel{mol\ Zn}} \times \frac{189.41\ g\ Zn(NO_3)_2}{1\ \cancel{mol\ Zn(NO_3)_2}} = 88.3\ g\ Zn(NO_3)_2$$

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Thus, the theoretical yield is 88.3 g of $Zn(NO_3)_2$. The actual yield is the amount that was actually made, which was 65.2 g of $Zn(NO_3)_2$. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100:

$$\frac{65.2\text{ g }\ Zn(NO_3)_2}{88.3\text{ g }\ Zn(NO_3)_2} \times 100\% = 73.8\%$$

Test Yourself

A synthesis produced 2.05 g of NH_3 from 16.5 g of N_2. What is the theoretical yield and the percent yield?

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

Answer theoretical yield = 20.1 g; percent yield = 10.2%

Chemistry is Everywhere: Actual Yields in Drug Synthesis

Many drugs are the product of several steps of chemical synthesis. Each step typically occurs with less than 100% yield, so the overall percent yield may be very small. The general rule is that the overall percent yield is the product of the percent yields of the individual synthesis steps. To calculate the overall percent yield first divide the individual percent yields by 100%, multiply them all together, then multiply by 100% to get the overall percent yield. For example, if a two-step process has a percent yield of 40.% for the first step and 60.% for the second step, the overall percent yield is 0.40 x 0.60 x 100% = 24%. For a drug synthesis that has many steps, the overall percent yield can be very tiny, which is one factor in the huge cost of some drugs. For example, if a 10-step synthesis has a percent yield of 90% for each step, the overall yield for the entire synthesis is only 35%. Many scientists work every day trying to improve percent yields of the steps in the synthesis to decrease costs, improve profits, and minimize waste.

Purifications of complex molecules into pure active pharmaceutical substances are subject to percent yields. Consider the purification of albuterol. Albuterol ($C_{13}H_{21}NO_2$) is a drug (delivered *via* inhaler) used to treat asthma, bronchitis, and other obstructive pulmonary diseases.

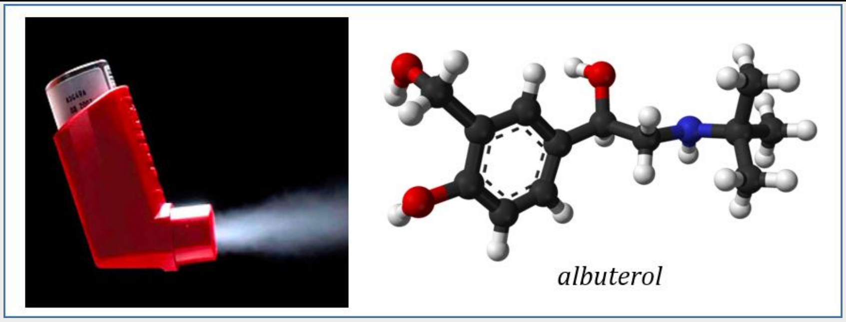

Image credit: Teva Pharmaceuticals and Wikimedia Commons.

Albuterol is synthesized from norepinephrine, a naturally occurring hormone and neurotransmitter. Its initial synthesis makes very impure albuterol that is purified in five chemical steps. Each step has its own percent yield and the overall percent yield must be calculated by combining the information from each step.

Step 1	impure albuterol	→	intermediate A	70% yield
Step 2	intermediate A	→	intermediate B	100% yield
Step 3	intermediate B	→	intermediate C	40% yield
Step 4	intermediate C	→	intermediate D	71% yield
Step 5	intermediate D	→	drug product	35% yield

$$\text{Overall Yield} = 70\% \times 100\% \times 40\% \times 71\% \times 35\%$$
$$= 0.70 \times 1.00 \times 0.40 \times 0.71 \times 0.35 = 0.070 \times 100\% = 7.0\%$$

That is, only about *one-fourteenth* of the original material was turned into the purified drug. The multiple steps and low yields during the industrial synthesis of drug products is one reason for the high cost of many pharmaceuticals.

8.5 Limiting Reagents

One additional assumption we have made about chemical reactions—in addition to the assumption that reactions proceed all the way to completion—is that all the reactants are present in the proper quantities to react to products. This is not always the case.

Consider if you are making lunch for friends and bought a 5-pack of hot dogs and a 4-pack of buns. How many people will be able to enjoy their hot dog on a bun? An easy way to solve this problem is to lay out all of the ingredients, as Figure 8.3 shows. It immediately becomes obvious that one hot dog will be left over.

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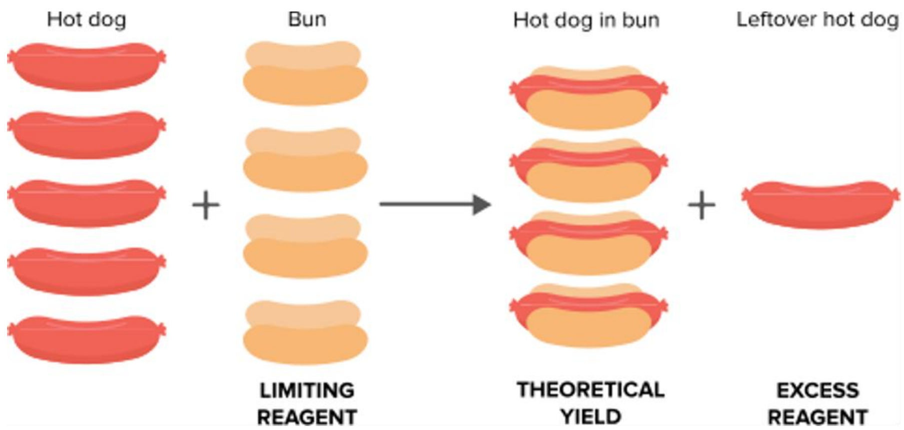

Figure 8.3 Assembling hots dogs in buns is an example of a limiting reagent problem. It is rarely the case that the buns and the hot dogs run out at the same time! *Image credit: Khan Academy.*

When an ingredient or a chemical is left over at the end of a process, we cannot use it to determine the amount of a product (here, hot dogs in buns) that will be formed. Instead, the ingredient that runs out first – the buns – is used to determine the amount of product that is actually formed.

Figure 8.4 explores the same idea with molecules instead of food. Here we are taking hydrogen molecules and oxygen molecules (left) to make water molecules (right) according to the following reaction:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(\ell)$$

However, there are not enough oxygen atoms to use up all of the hydrogen atoms. The oxygen atoms run out and cannot be used to make more than four water molecules. Thus the process stops when the reaction runs out of oxygen atoms.

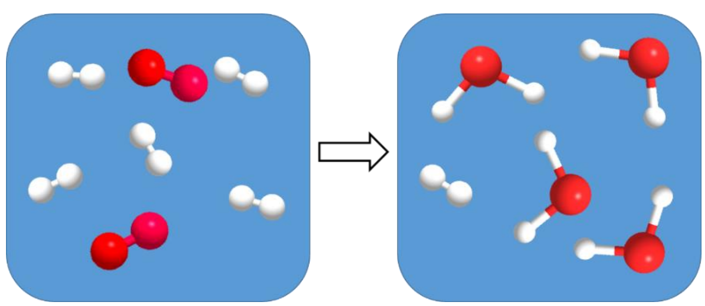

Figure 8.4 In this scenario for making water molecules, oxygen (red) runs out before hydrogen (white), making oxygen the limiting reagent.

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A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the **limiting reagent**; any remaining reactants are called **excess reagents**. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reagent and which is in excess. Visually, the limiting reactant can be determined by grouping together atoms in the correct ratio to form a product molecule. In Figure 8.5, after all of the atoms have been grouped to make four complete sets of H_2O, one H_2 molecule has not been used. It is, therefore, the excess reagent.

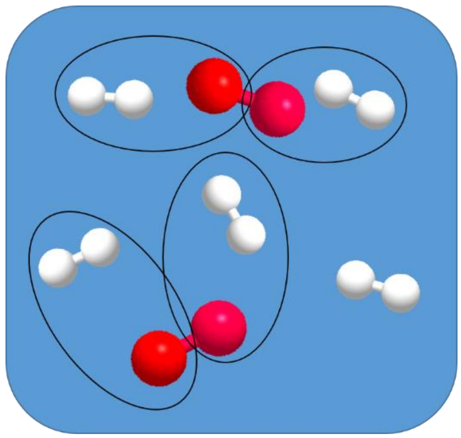
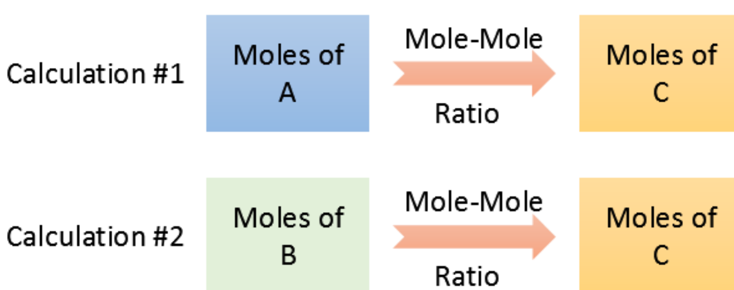


Figure 8.5 By grouping together two hydrogen atoms (white) with every one oxygen atom (red), four H_2O molecules are made. One H_2 molecule is left over, making it the excess reagent.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: *whichever reactant gives the lesser amount of product is the limiting reagent*. To determine which reactant is limiting, first predict the theoretical amount of *one product* from the given amounts of *each reactant*. Whichever reactant gives the least amount of that product is the limiting reagent. Consider the concept map below for a general reaction A + B → C + D where moles of each reactant are given.



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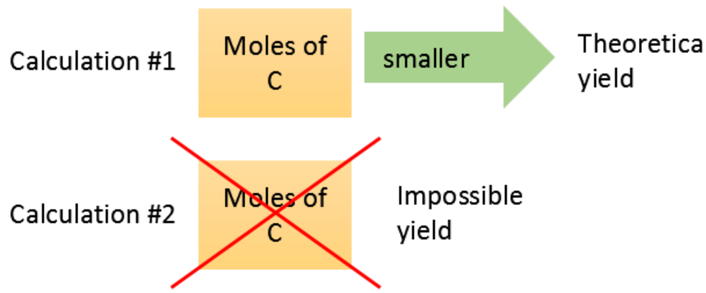
Note that two amounts for product C were calculated! Be careful: this does not mean that both amounts will be produced. As shown below, the smaller number represents the theoretical yield, and the larger amount is impossible to produce because you will run out of reactant A first.



It does not matter which product we use, as long as we use the same one each time. We could have converted to product D and still arrived at the conclusion that reactant A is the limiting reagent. Furthermore, it does not matter whether we determine the number of moles or grams of that product. For example, consider the following reaction:

$$4\ As(s) + 3\ O_2(g) \rightarrow 2\ As_2O_3(s)$$

Suppose we start a reaction with 50.0 g of As and 50.0 g of O_2. Which one is the limiting reagent? We need to perform two mole-mass calculations, each assuming that the reactant reacts completely. Then we compare the amount of the product produced by each and determine which is less.

The concept map for this problem involves one additional step: converting the initial grams to moles. We will still proceed with two calculations, one for each reactant, as shown in the concept map below. Calculation #1 represents a reaction that uses up all of the arsenic. Calculation #2 represents a reaction that uses up all of the oxygen.

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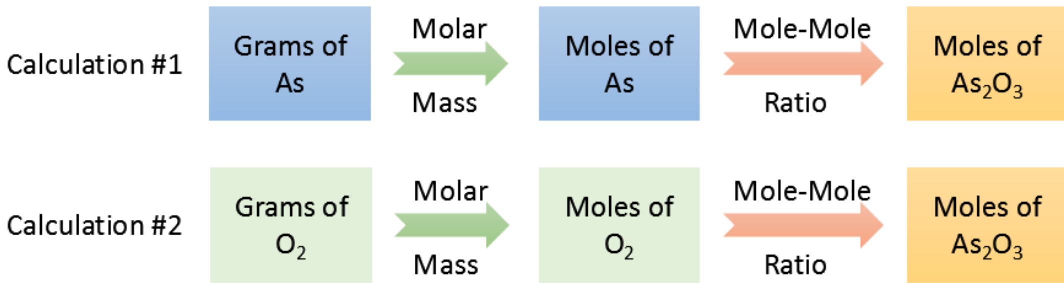

The calculations are as follows:

$$\text{Calculation \#1} \quad 50.0\ \text{g}\ \cancel{\text{As}} \times \frac{1\ \cancel{\text{mol As}}}{74.92\ \cancel{\text{g As}}} \times \frac{2\ \text{mol As}_2\text{O}_3}{4\ \cancel{\text{mol As}}} = 0.334\ \text{mol As}_2\text{O}_3$$

$$\text{Calculation \#2} \quad 50.0\ \text{g}\ \cancel{\text{O}_2} \times \frac{1\ \cancel{\text{mol O}_2}}{32.00\ \cancel{\text{g As}}} \times \frac{2\ \text{mol As}_2\text{O}_3}{3\ \cancel{\text{mol O}_2}} = 1.04\ \text{mol As}_2\text{O}_3$$

Comparing these two answers, it is clear that 0.334 mol of As_2O_3 is less than 1.04 mol of As_2O_3, so arsenic is the limiting reagent. If this reaction is performed under these initial conditions, the arsenic will run out before the oxygen runs out. We say that the oxygen is "in excess."
Often you will be asked what *mass* of product is theoretically formed. To answer this, one additional step must be performed. Making sure to use the smallest of the values from the two calculations above, the moles can be converted to grams.

$$0.334\ \cancel{\text{mol As}_2\text{O}_3} \times \frac{197.84\ \text{g As}_2\text{O}_3}{1\ \cancel{\text{mol As}_2\text{O}_3}} = 66.1\ \text{g As}_2\text{O}_3$$

Thus, 66.1 g As_2O_3 are theoretically formed. In summary, the steps that you will follow for a limiting reactant problem are:

1. Determine the number of moles of each reactant.
2. Convert to moles of the same product using stoichiometry.
3. Determine which reactant produced the **least** amount of product to identify the limiting reactant.
4. Convert the least number of moles of product to mass of product.

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The limiting reagent not only determines how much of each product is made, but also determines how much of the other reactants are used. By subtracting how much of the reactant was used up from the amount that was originally there, you can determine how much of the other reactants remain unreacted. This assumes that none of the limiting reagents are left over or are unreacted.

Example 8.9

In the reaction between 5.00 g of Rb and 3.44 g of $MgCl_2$ the limiting reagent is Rb. What mass of $MgCl_2$ is leftover?

$$2\ Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2\ RbCl(s)$$

Solution

To determine how much of the excess reagent is left, we have to do a mass-mass calculation to determine what mass of $MgCl_2$ reacted with the 5.00 g of Rb.

$$5.00\ \cancel{\text{g Rb}} \times \frac{1\ \cancel{\text{mol Rb}}}{85.47\ \cancel{\text{g Rb}}} \times \frac{1\ \cancel{\text{mol MgCl}_2}}{2\ \cancel{\text{mol Rb}}} \times \frac{95.21\ \text{g MgCl}_2}{1\ \cancel{\text{mol MgCl}_2}} = 2.78\ \text{g MgCl}_2\ \text{reacted}$$

Then subtract the amount reacted from the original amount. Because the reaction began with 3.44 g of $MgCl_2$, there are: 3.44 g $MgCl_2$ − 2.78 g $MgCl_2$ reacted = 0.66 g $MgCl_2$ left

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Chapter 8 Practice Problems

8.1 Stoichiometry

1. Think back to the pound cake recipe. What possible conversion factors can you construct relating the components of the recipe?

2. Think back to the pancake recipe. What possible conversion factors can you construct relating the components of the recipe?

3. What are all the conversion factors that can be constructed from the balanced chemical reaction: $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(\ell)$?

4. What are all the conversion factors that can be constructed from the balanced chemical reaction: $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$?

8.2 The Mole in Chemical Reactions

5. Balance the chemical equation: $Na(s) + H_2O(\ell) \rightarrow NaOH(aq) + H_2(g)$
   How many molecules of H_2 are produced when 332 atoms of Na react?

6. Balance the chemical equation: $S(s) + O_2(g) \rightarrow SO_3(g)$
   How many molecules of O_2 are needed when 38 atoms of S react?

7. For the balanced chemical equation:
   $6\ H^+(aq) + 2\ MnO_4^-(aq) + 5\ H_2O_2(\ell) \rightarrow 2\ Mn^{2+}(aq) + 5\ O_2(g) + 8\ H_2O(\ell)$
   how many molecules of H_2O are produced when 75 molecules of H_2O_2 react?

8. For the balanced chemical reaction: $2\ C_6H_6(\ell) + 15\ O_2(g) \rightarrow 12\ CO_2(g) + 6\ H_2O(\ell)$
   how many molecules of CO_2 are produced when 56 molecules of C_6H_6 react?

9. Given the balanced chemical equation: $Fe_2O_3(s) + 3\ SO_3(g) \rightarrow Fe_2(SO_4)_3(aq)$
   how many molecules of $Fe_2(SO_4)_3$ are produced if 321 atoms of S are reacted?

10. For the balanced chemical equation: $CuO(s) + H_2S(g) \rightarrow CuS(s) + H_2O(\ell)$
    how many molecules of CuS are formed if 9,044 atoms of H react?

11. For the balanced chemical equation: $Fe_2O_3(s) + 3\ SO_3(g) \rightarrow Fe_2(SO_4)_3(\ell)$
    suppose we need to make 145,000 molecules of $Fe_2(SO_4)_3$. How many molecules of SO_3 do we need?

12. One way to make sulfur hexafluoride is to react thioformaldehyde, CH_2S, with elemental fluorine: $CH_2S(g) + 6\ F_2(g) \rightarrow CF_4(g) + 2\ HF(g) + SF_6(g)$
    If 45,750 molecules of SF_6 are needed, how many molecules of F_2 are required?

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13. Construct the three independent conversion factors possible for these two reactions:
$2\ H_2 + O_2 \rightarrow 2\ H_2O$ *and* $H_2 + O_2 \rightarrow H_2O_2$
Why are the ratios between H_2 and O_2 different?

14. Construct the three independent conversion factors possible for these two reactions:
$2\ Na + Cl_2 \rightarrow 2\ NaCl$ *and* $4\ Na + 2Cl_2 \rightarrow 4\ NaCl$
What similarities, if any, exist in the conversion factors from these two reactions?

15. Express in mole terms what this chemical equation means: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

16. Express in mole terms what this chemical equation means:
$Na_2CO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2O(\ell) + CO_2(aq)$

17. How many molecules of each substance are involved in the equation in Exercise 15 if it is interpreted in terms of moles?

18. How many molecules of each substance are involved in the equation in Exercise 16 if it is interpreted in terms of moles?

19. For the chemical equation: $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 4\ CO_2(g) + 6\ H_2O(\ell)$, what equivalences can you write in terms of moles? Use the $\Leftrightarrow$ sign.

20. For the chemical equation: $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$
what equivalences can you write in terms of moles? Use the $\Leftrightarrow$ sign.

21. Write the balanced chemical reaction for the combustion of C_5H_{12} (the products are CO_2 and H_2O) and determine how many moles of H_2O are formed when 5.8 mol of O_2 are reacted.

22. Write the balanced chemical reaction for the formation of $Fe_2(SO_4)_3$ from Fe_2O_3 and SO_3 and determine how many moles of $Fe_2(SO_4)_3$ are formed when 12.7 mol of SO_3 are reacted.

23. For the balanced chemical equation:
$3\ Cu(s) + 2\ NO_3^-(aq) + 8\ H^+(aq) \rightarrow 3\ Cu^{2+}(aq) + 4\ H_2O(\ell) + 2N\ O(g)$
how many moles of Cu^{2+} are formed when 55.7 mol of H^+ are reacted?

24. For the balanced chemical equation: $Al(s) + 3\ Ag^+(aq) \rightarrow Al^{3+}(aq) + 3\ Ag(s)$
how many moles of Ag are produced when 0.661 mol of Al are reacted?

25. For the balanced chemical reaction: $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(\ell)$, how many moles of H_2O are produced when 0.669 mol of NH_3 react?

26. For the balanced chemical reaction:
$4\ NaOH(aq) + 2\ S(s) + 3\ O_2(g) \rightarrow 2\ Na_2SO_4(aq) + 2\ H_2O(\ell)$, how many moles of Na_2SO_4 are formed when 1.22 mol of O_2 react?

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27. For the balanced chemical reaction: $4\ KO_2(s) + 2\ CO_2(g) \rightarrow 2\ K_2CO_3(s) + 3\ O_2(g)$ determine the number of moles of each of the products that are formed when 6.88 mol of KO_2 react.

28. For the balanced chemical reaction: $2\ AlCl_3(s) + 3\ H_2O(\ell) \rightarrow Al_2O_3sg) + 6\ HCl(g)$ determine the number of moles of each of the products that are formed when 0.0552 mol of $AlCl_3$ react.

8.3 Mole-mass and Mass-mass Calculations

29. What mass of CO_2 is produced by the combustion of 1.00 mol of CH_4?
$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(\ell)$

30. What mass of H_2O is produced by the combustion of 1.00 mol of CH_4?
$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(\ell)$

31. What mass of HgO is required to produce 0.692 mol of O_2?
$2\ HgO(s) \rightarrow 2\ Hg(\ell) + O_2(g)$

32. What mass of $NaHCO_3$ is needed to produce 2.659 mol of CO_2?
$2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(\ell) + CO_2(g)$

33. How many moles of Al can be produced from 10.87 g of Ag?
$Al(NO_3)_3(aq) + 3\ Ag(s) \rightarrow Al(s) + 3\ AgNO_3(aq)$

34. How many moles of HCl can be produced from 0.226 g of $SOCl_2$?
$SOCl_2(\ell) + H_2O(\ell) \rightarrow SO_2(g) + 2\ HCl(g)$

35. How many moles of O_2 are needed to prepare 1.00 g of $Ca(NO_3)_2$?
$Ca(s) + N_2(g) + 3\ O_2(g) \rightarrow Ca(NO_3)_2(s)$

36. How many moles of C_2H_5OH are needed to generate 106.7 g of H_2O?
$C_2H_5OH(\ell) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 3\ H_2O(\ell)$

37. What mass of O_2 can be generated by the decomposition of 100.0 g of $NaClO_3$?
$2\ NaClO_3(s) \rightarrow 2\ NaCl(s) + 3\ O_2(g)$

38. What mass of Li_2O is needed to react with 1,060 g of CO_2?
$Li_2O(aq) + CO_2(g) \rightarrow Li_2CO_3(aq)$

39. What mass of Fe_2O_3 must be reacted to generate 324 g of Al_2O_3?
$Fe_2O_3(s) + 2\ Al(s) \rightarrow 2\ Fe(s) + Al_2O_3(\mathrm{s})$

40. What mass of Fe is generated when 100.0 g of Al are reacted?
$Fe_2O_3(s) + 2\ Al(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$

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41. What mass of MnO_2 is produced when 445 g of H_2O are reacted?
$H_2O(\ell) + 2\ MnO_4^-(aq) + Br^-(aq) \rightarrow BrO_3^-(aq) + 2\ MnO_2(s) + 2\ OH^-(aq)$

42. What mass of $PbSO_4$ is produced when 29.6 g of H_2SO_4 are reacted?
$Pb(s) + PbO_2(s) + 2\ H_2SO_4(aq) \rightarrow 2\ PbSO_4(s) + 2\ H_2O(\ell)$

43. If 83.9 g of ZnO are formed, what mass of Mn_2O_3 is formed with it?
$Zn(s) + 2\ MnO_2(s) \rightarrow ZnO(s) + Mn_2O_3(s)$

44. If 14.7 g of NO_2 are reacted, what mass of H_2O is reacted with it?
$3\ NO_2(g) + H_2O(\ell) \rightarrow 2\ HNO_3(aq) + NO(g)$

45. If 88.4 g of CH_2S are reacted, what mass of HF is produced?
$CH_2S(g) + 6\ F_2\ (g) \rightarrow CF_4(g) + 2\ HF(g) + SF_6(g)$

46. If 100.0 g of Cl_2 are needed, what mass of NaOCl must be reacted?
$NaOCl(aq) + HCl(aq) \rightarrow NaOH(aq) + Cl_2(g)$

8.4 Yields

47. What is the difference between the theoretical yield and the actual yield?

48. What is the difference between the actual yield and the percent yield?

49. A chemist isolates 2.675 g of SiF_4 after reacting 2.339 g of SiO_2 with HF. What are the theoretical yield and the actual yield?

$SiO_2(s) + 4\ HF(g) \rightarrow SiF_4(g) + 2\ H_2O(\ell)$

50. A pharmaceutical chemist synthesizes aspirin, $C_9H_8O_4$, according to this chemical equation. If 12.66 g of $C_7H_6O_3$ are reacted and 12.03 g of aspirin are isolated, what are the theoretical yield and the actual yield? What is the percent yield?

$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + HC_2H_3O_2$

51. A chemist decomposes 1.006 g of $NaHCO_3$ and obtains 0.0334 g of Na_2CO_3. What are the theoretical yield and the actual yield? What is the percent yield?

$2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(\ell) + CO_2(g)$

52. A chemist combusts a 3.009 g sample of C_5H_{12} and obtains 3.774 g of H_2O. What are the theoretical yield and the actual yield? What is the percent yield?

$C_5H_{12}(\ell) + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(\ell)$

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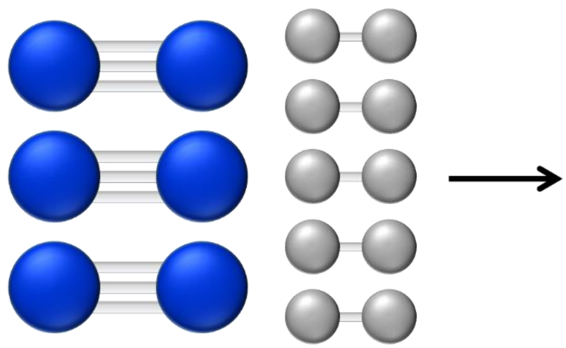
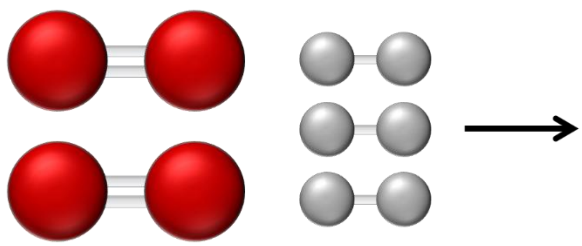
8.5 Limiting Reagents

53. The box below shows molecules of nitrogen (N_2, blue spheres) and of hydrogen (H_2, gray spheres) that will react to produce ammonia, NH_3. Which is the limiting reagent? How many ammonia molecules are made? How many N atoms are left over? How many H atoms are left over?



54. The box below shows molecules of oxygen (O_2, red spheres) and of hydrogen (H_2, gray spheres) that will react to produce water, H_2O. Which is the limiting reagent? How many water molecules are made? How many oxygen atoms are left over? How many hydrogen atoms are left over?



55. Given the statement "20.0 g of methane is burned in excess oxygen," what is the limiting reagent?

56. Given the statement "the metal is heated in the presence of excess hydrogen," what is the limiting reagent despite not specifying any quantity of reactant?

8.6 Excess Reagents

57. Acetylene (C_2H_2) is formed by reacting 7.08 g of C and 4.92 g of H_2.

$$2\,C(s) + H_2(g) \rightarrow C_2H_2(g)$$

What is the limiting reagent? How much of the other reactant is in excess?

58. Ethane (C_2H_6) is formed by reacting 7.08 g of C and 4.92 g of hydrogen gas.

$$2\,C(s) + 3\,H_2(g) \rightarrow C_2H_6(g)$$

What is the limiting reagent? How much of the other reactant is in excess?

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59. Given 35.6 g P_4O_6 and 4.77 g H_2O, what is the limiting reagent for the reaction below, and how much of the other reactant is in excess?

$$P_4O_6(s) + 6\ H_2O(\ell) \rightarrow 4\ H_3PO_4(aq)$$

60. Given 377 g NO_2 and 244 g H_2O, what is the limiting reagent for the reaction below, and how much of the other reactant is in excess?

$$3\ NO_2(g) + H_2O(\ell) \rightarrow 2\ HNO_3(aq) + NO(g)$$

61. To form the precipitate $PbCl_2$, 2.88 g of NaCl and 7.21 g of $Pb(NO_3)_2$ are mixed in solution. How much precipitate is formed? How much of which reactant is in excess?

62. In a neutralization reaction, 18.06 g of KOH are reacted with 13.43 g of HNO_3. What mass of H_2O is produced, and what mass of which reactant is in excess?

Review: How to Succeed in Chemistry

63. Identify the chapters and sections that are difficult for you. Make sure to spend some extra review time on these topics and see a chemistry tutor if necessary.

64. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

65. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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Chapter 9
Aqueous Solutions

LEARNING OBJECTIVES

1. Explain the nature of aqueous solutions.
2. Identify how water molecules interact in solutions.
3. Distinguish between qualitative and quantitative descriptions of solutions.
4. Calculate concentrations of solutions in molarity.
5. Calculate amounts and volumes used in dilutions.
6. Describe the procedure for diluting a solution.
7. Use concentration units such as molarity as conversion factors in calculations.
8. Describe a titration experiment and explain the purpose of indicators.
9. Perform titration calculations.

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The Dead Sea is an example of a concentrated aqueous solution. It has such a high concentration of salt that people are able to float on its surface. *Image credit: Wikimedia Commons.*

Mixtures with water are present all around us, from the salt in the ocean, to the sugar in our blood. In this chapter the interaction of water with other water molecules and dissolved substances will be explored. Also, the amount of a substance in water will be calculated to determine concentration.

9.1 Forming Mixtures with Water

In Chapter 3, matter was classified as either a pure substance or a mixture. Distilled water is a pure substance, and the only interactions in a sample of distilled water are those between water molecules. In Figure 9.1, water molecules can be seen at the particle level being held together by electrostatic attractions (dotted lines) between the hydrogen and oxygen atoms on adjacent molecules. Because the interactions between water molecules are continually breaking and reforming, liquid water does not have a single fixed structure.

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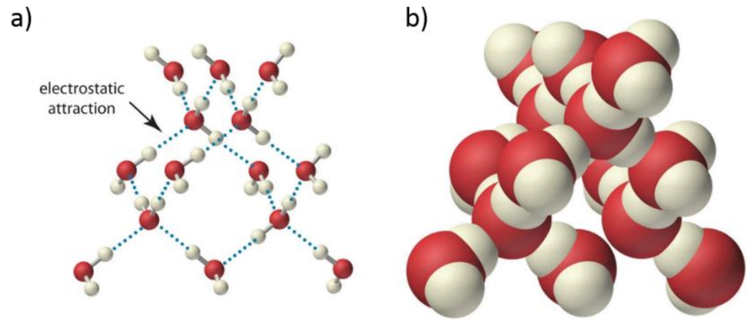

Figure 9.1 Two views of water molecules are shown: (a) a ball-and-stick structure and (b) a space-filling model. *Image credit: [Principles of General Chemistry.]*

The electrostatic attraction between molecules arise because each H_2O molecule has partial charges due to how the electrons are distributed. Figure 9.2 shows that oxygen has a partial negative (δ^-) charge since electrons prefer to be closer to the oxygen atom, and each hydrogen has a partial positive (δ^+) charge since the electrons are being pulled away from hydrogen. Thus in a sample of many H_2O molecules, neighboring molecules stick together because the partially negative-charged oxygen atom of one molecule and the partially positive-charged hydrogen atom on another molecule acts like glue.

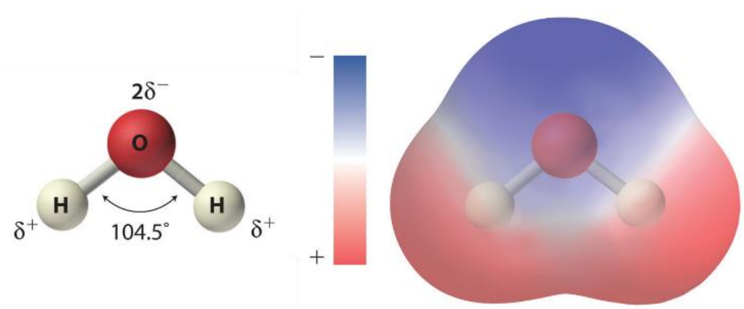

Figure 9.2 Because oxygen attracts electrons to itself more than hydrogen, a partial negative charge (δ^-) is present on oxygen and a partial positive charge (δ^+) is present on hydrogen. *Image credit: [Principles of General Chemistry].*

The partial charges on atoms in water molecules have a significant effect when it comes to forming mixtures. When other substances form a homogeneous mixture with water, this is called an **aqueous solution**. The major component of a solution is called the **solvent**. The minor component

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of a solution is called the **solute**. The terms major and minor refer to whichever component has the greater presence by mass or by moles. Sometimes this becomes confusing, especially with substances with very different molar masses. However, in this chapter only solutions for which the major component and the minor component are obvious will be considered. For most solutions the solvent is usually a liquid. For aqueous solutions, the solvent is generally considered to be water. Usually scientists say that the solute is placed in the solvent; for example, a solution where toluene is the solute and benzene is the solvent is said to have toluene dissolved in benzene.

Example 9.1

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Answer
Either by mass or by moles, the obvious minor component is sucrose, so it is the solute. Water—the majority component—is the solvent. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Test Yourself
A solution is made by dissolving 3.33 g of HCl(*g*) in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.
Answer The solute is HCl(*g*) and the solvent is $CH_3OH(\ell)$.

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water; soda water is a solution of gaseous CO_2 in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent.

Consider the atomic-level view of an NaCl solution in Figure 9.3. Notice how the partial charges on a molecule of water interact with the charge of Na^+ and Cl^- ions. An ionic solid such as sodium chloride dissolves in water because of the electrostatic attraction between the cations (Na^+) and the partially negatively charged oxygen atoms of water molecules, and between the anions (Cl^-) and the partially positively charged hydrogen atoms of water.

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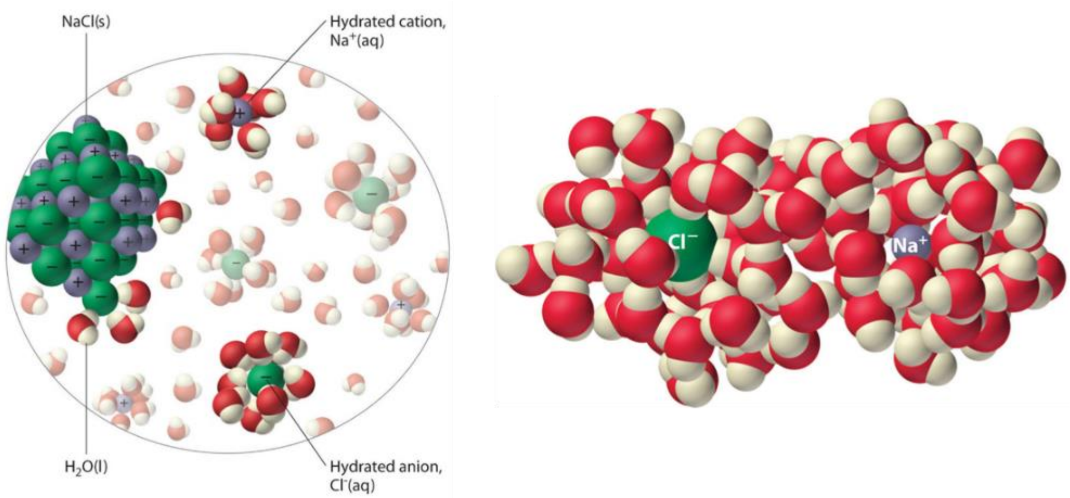

Figure 9.3 Water molecules in a solution with NaCl orient themselves according to the partial charges on water. Negative chloride ions align with the (δ^+) hydrogen atoms. Positive sodium ions align with the (δ^-) oxygen atoms. *Image credit: Principles of General Chemistry.*

Recall from Chapter 7 that when ionic substances dissolve in water, solvated ions are formed as ionic bonds are broken. For example, $BaCl_2$ dissolves in water according to the following ionic equation:

$$BaCl_2(s) \rightarrow Ba^{2+}(aq) + 2\ Cl^-(aq)$$

The notation (aq) indicates that the dissolution is taking place in water as opposed to some other solvent.

If an ionic substance contains a polyatomic ion, the polyatomic ion will not break apart in water because it contains covalent bonds. Note in the following example that the NO_3^- ion is not broken apart.

$$Ba(NO_3)_2(s) \rightarrow Ba^{2+}(aq) + 2\ NO_3^-(aq)$$

The above balanced equations and similar equations with other ionic compounds will be important later in this chapter because stoichiometry will be applied to reactions that occur in water.

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9.2 Concentration and Molarity

One important concept of solutions is in defining how much solute is dissolved in a given amount of solvent. This concept is called **concentration**. Various words are used to describe the relative amounts of solute. The word **dilute** describes a solution with very little dissolved solute. The term **concentrated** describes a solution that has a relatively large amount of solute per unit volume. One problem is that these terms are qualitative. Rather than qualitative terms, a quantitative method to express the amount of solute in a solution is needed; that is, specific units of concentration must be defined. In this section, several common and useful units of concentration will be introduced.

Molarity (M) is defined as the number of moles of solute divided by the number of liters of solution and has units of mol/L (also called "molar, with a unit symbol of M).

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

As with any mathematical equation, if you know any two quantities, you can calculate the third, unknown, quantity. For example, suppose you have 0.500 L of solution that has 0.24 mol of NaOH dissolved in it. The concentration of the solution can be calculated as follows:

$$\text{Molairty} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{0.24\text{ mol NaOH}}{0.500\text{ L}} = 0.48\text{ M NaOH}$$

The concentration of the solution is 0.48 M, which is spoken as "zero point four eight molar." If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L? First, convert the mass of solute to moles using the molar mass of HCl:

$$22.4\ \cancel{\text{g HCl}} \times \frac{1\text{ mol HCl}}{36.45\ \cancel{\text{g HCl}}} = 0.615\text{ mol HCl}$$

Now we can use the definition of molarity to determine a concentration:

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$$\text{Molarity} = \frac{0.615 \text{ mol NaOH}}{1.56 \text{ L}} = 0.394 \text{ M NaOH}$$

When working with concentrations in chemistry, a wide range of numbers will be encountered. Molarity can be very small with some cellular concentrations as small as 10^{-6} M, or it can be greater than 1, as is the case with concentrated acids such as 12 M HCl.

Example 9.2

What is the molarity when 32.7 g of NaOH are dissolved to make 445 mL of solution?

Solution
To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$445 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 0.445 \text{ L}$$

Now we convert the amount of solute to moles, using the molar mass of NaOH:

$$32.7 \cancel{\text{g NaOH}} \times \frac{1 \text{ mol NaOH}}{40.0 \cancel{\text{g NaOH}}} = 0.818 \text{ mol NaOH}$$

Now we can use the definition of molarity to determine the molar concentration:

$$\text{Molarity} = \frac{0.818 \text{ mol NaOH}}{0.445 \text{ L}} = 1.84 \text{ M NaOH}$$

Test Yourself
What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution? *Answer* 1.57 M

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Concentration can be used as a conversion factor between the amount/moles of solute and the amount/volume of solution. For example, suppose you are asked how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, the second expression for the concentration can be used as a conversion factor between moles and liters, where there is an understood 1 in the denominator of the conversion factor.

$$0.108\ \cancel{\text{L solution}} \times \frac{0.887\ \text{mol NaCl}}{\cancel{\text{L solution}}} = 0.0958\ \text{mol NaCl}$$

Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can also be used as a conversion factor. Consider if you wish to make a 1.2 M NaCl solution and start with 0.54 mol NaCl. What volume of solution will be made

$$0.54\ \cancel{\text{mol NaCl}} \times \frac{\text{L solution}}{1.2\ \cancel{\text{mol NaCl}}} = 0.45\ \text{L solution}$$

In many instances you will need to use the molarity of a solution to solve for either the volume or the number of moles. In both cases, molarity can be used as a conversion factor following the guidelines of dimensional analysis. Consider the red wine in Figure 9.4, which is a solution of 2.1 M ethanol (C_2H_5OH).

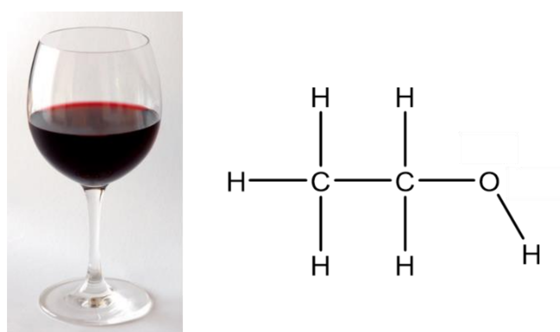

Figure 9.4 Red wine contains the molecule ethanol, CH_3CH_2OH, with a concentration of about 2.1 M. *Image credit: Wikimedia Commons by Andre Karwath.*

Suppose you want to know how many moles of ethanol are in a 175 mL glass of red wine, which is 2.1 M ethanol. Molarity should be used as a conversion factor, but first mL must be converted to L.

$$175\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{2.1\ \text{mol ethanol}}{\cancel{\text{L}}} = 0.37\ \text{mol ethanol}$$

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Example 9.3

How many moles of solute are present in 108 μL of a 0.887 M NaCl solution?

Solution

Use dimensional analysis with molarity as a conversion factor. First, convert μL to L.

$$108\ \cancel{\mu L} \times \frac{1\ \cancel{L}}{10^6\ \cancel{\mu L}} \times \frac{0.887\ \text{mol NaCl}}{\cancel{L}} = 9.58 \times 10^{-5} \text{mol NaCl}$$

Test Yourself How many moles of solute are present in 225 mL of a 1.44 M $CaCl_2$ solution?

Answer 0.324 mol

Example 9.4

What volume (in mL) of a 2.33 M $NaNO_3$ solution is needed to obtain 0.222 mol of solute?

Solution

Dimensional analysis can be used, and since molarity is our conversion factor, don't begin with it, but rather with the number of moles.

$$0.222\ \cancel{\text{mol NaNO}_3} \times \frac{1\ \text{L}}{2.33\ \cancel{\text{mol NaNO}_3}} \times \frac{1000\ \text{mL}}{1\ \text{L}} = 95.3\ \text{mL g NaNO}_3$$

Test Yourself What volume of a 0.570 M K_2SO_4 solution is needed to obtain 0.872 mol of solute?

Answer 1.53 L

Example 9.5

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH to convert to mass:

$$0.765\ \cancel{\text{L solution}} \times \frac{1.93\ \cancel{\text{mol NaOH}}}{\cancel{\text{L solution}}} \times \frac{40.00\ \text{g NaOH}}{1\ \cancel{\text{mol NaOH}}} = 59.1\ \text{g NaOH}$$

Test Yourself What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?

Answer 6.12 g

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9.3 The Preparation of Solutions

To prepare a solution that contains a specified concentration of a substance, it is necessary to dissolve the desired number of moles of solute in enough solvent to give the desired final volume of solution. Figure 9.5 illustrates the procedure to prepare a solution of cobalt (II) chloride dihydrate in ethanol. Example 9.6 calculates the concentration of this solution.

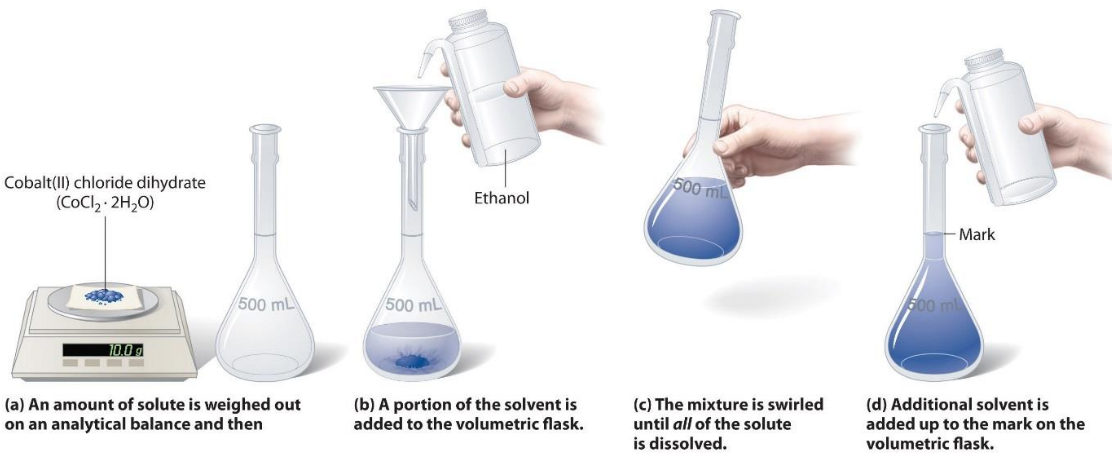

Figure 9.5 In order to prepare a solution with a specified concentration, one must (a) weigh out the calculated amount, (b) add a small amount of solvent, (c) swirl the solution to dissolve completely, and (d) add more solvent to give the appropriate final volume. *Image credit: Principles of General Chemistry.*

Note that the volume of the *solvent* is not specified; only the volume of *solution* is known. Because the solute occupies space in the solution, the volume of the solvent needed is always *less* than the desired volume of solution. Unfortunately, the volumes of different substances in a solution are generally not additive, so you cannot make 500mL of solution by dissolving 8mL of salt in 492 mL of water. Some of the salt "fits in the gaps between the water molecules" so the resulting total volume is smaller than expected. Similarly, if the desired volume were 1.00 L, it would be incorrect to add 1.00 L of water to 342 g of sucrose because that would produce more than 1.00 L of solution. As shown in Figure 9.6 for some substances this effect can be significant, especially for concentrated solutions.

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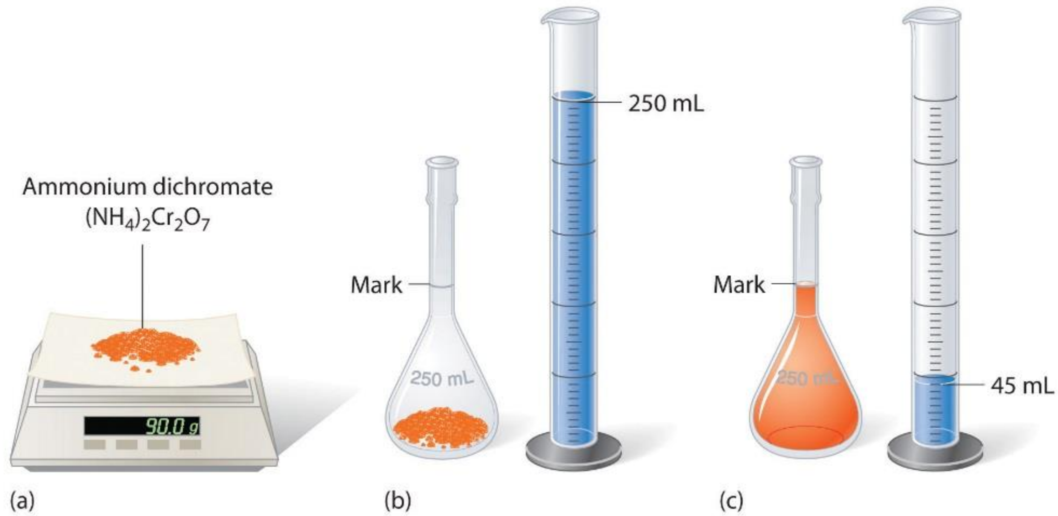

Figure 9.6 In order to prepare an aqueous solution with a specified molarity, one will (a) weigh out the calculated amount of solute, (b) add water according to the procedure in Figure 9.5, (c) note that the volume of the solution (orange) is not the same as the volume of the water added. *Image credit: Principles of General Chemistry.*

In Figure 9.6 it becomes obvious that the solute occupies space in the solution, so less than 250 mL of water are needed to make 250 mL of solution.

Example 9.6

The solution in Figure 9.5 contains 10.0 g of cobalt(II) chloride dihydrate, $CoCl_2{\cdot}2H_2O$, in enough ethanol to make 500.0 mL of solution. What is the molar concentration of $CoCl_2{\cdot}2H_2O$?

Solution

In order to find the molarity, first find the amount of solute in units of moles and the volume of solution in units of liters. Find the number of moles of $CoCl_2{\cdot}2H_2O$ using its molar mass as a conversion factor. The chemical formula shows that this compound is a hydrate, which means that two H_2O molecules are **added** to the mass of $CoCl_2$.

$$10.0\ \cancel{\mathrm{g\ CoCl_2 \cdot 2H_2O}} \times \frac{1\ \mathrm{mol}}{165.87\ \cancel{\mathrm{g\ CoCl_2 \cdot 2H_2O}}} = 0.0603\ \mathrm{mol\ CoCl_2 \cdot 2H_2O}$$

Then find the volume of solution in liters:

$$500.0\ \cancel{\mathrm{mL}} \times \frac{1\ \mathrm{L}}{1000\ \cancel{\mathrm{mL}}} = 0.5000\ \mathrm{L}$$

Calculate the molarity of the solution by dividing the number of moles of solute by the volume of the solution in liters.

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$$\text{Molarity} = \frac{0.0603 \text{ mol } \mathrm{CoCl_2 \cdot 2H_2O}}{0.5000 \text{ L}} = 0.121\ M\ \mathrm{CoCl_2 \cdot 2H_2O}$$

Test Yourself
The solution shown in Figure 9.6 contains 90.0 g of $(NH_4)_2Cr_2O_7$ in enough water to give a final volume of exactly 250.0 mL. What is the molar concentration of ammonium dichromate?
Answer 1.43 M $(NH_4)_2Cr_2O_7$

For solutions of ionic compounds, the molar concentration of each ion may be different from the molar concentration of the overall compound. For example, consider the dissolution of 1M NaCl in water: $NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$. The solution could also be described as a solution of 1 M $Na^+(aq)$ and 1 M $Cl^-(aq)$ because there is one Na^+ ion and one Cl^- ion per formula unit of the salt.

Now consider a larger salt that contains a polyatomic ion. Ammonium dichromate, $(NH_4)_2Cr_2O_7$, is an ionic compound that contains two NH_4^+ ions and one $Cr_2O_7^{2-}$ ion per formula unit. Like most other water-soluble ionic compounds, it is a strong electrolyte that dissociates in aqueous solution to give aqueous NH_4^+ and $Cr_2O_7^{2-}$ ions:

$$(NH_4)_2Cr_2O_7(s) \rightarrow 2\ NH_4^+(aq) + Cr_2O_7^{2-}(aq)$$

Thus 1 mole of ammonium dichromate formula units dissolves in water to produce two moles of ammonium, NH_4^+, cations and one mole of dichromate, $Cr_2O_7^{2-}$, anions (Figure 9.7).

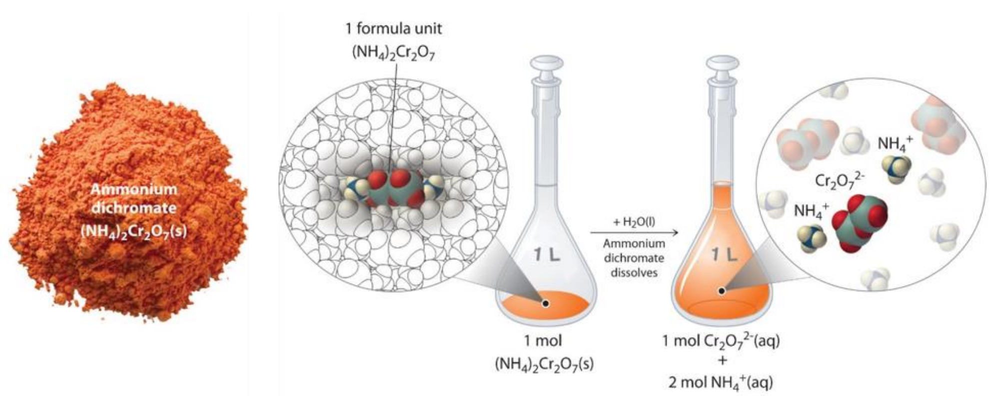

Figure 9.7 Dissolving 1 mol of $(NH_4)_2Cr_2O_7$ produces a solution that contains 1 mol of $Cr_2O_7^{2-}$ ions and 2 mol of NH_4^+ ions. *Image credit: Principles of General Chemistry.*

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When a chemical reaction using a solution of a salt such as ammonium dichromate is carried out, the concentration of each ion present in the solution must be considered. If a solution contains 1.43 M $(NH_4)_2Cr_2O_7$, the concentration of dichromate, $Cr_2O_7^{2-}$, must also be 1.43 M because there is one $Cr_2O_7^{2-}$ ion per formula unit. Ammonium, NH_4^+, ions are present in a 2:1 ratio compared to dichromate. The concentration of dichromate ion in solution can be calculated by considering the stoichiometry of the compound:

$$1.43 \text{ M } (\mathrm{NH_4})_2\mathrm{Cr_2O_7} = \frac{1.43 \text{ mol } \cancel{(\mathrm{NH_4})_2\mathrm{Cr_2O_7}}}{1 \text{ L solution}} \times \frac{1 \text{ mol } \mathrm{Cr_2O_7^{2-}}}{1 \text{ mol } \cancel{(\mathrm{NH_4})_2\mathrm{Cr_2O_7}}}$$

$$= \frac{1.43 \text{ mol } \mathrm{Cr_2O_7^{2-}}}{1 \text{ L solution}} = 1.43 \text{ M } \mathrm{Cr_2O_7^{2-}}$$

The concentration of ammonium ion in solution can be calculated in a similar way:

$$1.43 \text{ M } (\mathrm{NH_4})_2\mathrm{Cr_2O_7} = \frac{1.43 \text{ mol } \cancel{(\mathrm{NH_4})_2\mathrm{Cr_2O_7}}}{1 \text{ L solution}} \times \frac{2 \text{ mol } \mathrm{NH_4^+}}{1 \text{ mol } \cancel{(\mathrm{NH_4})_2\mathrm{Cr_2O_7}}}$$

$$= \frac{2.86 \text{ mol } \mathrm{NH_4^+}}{1 \text{ L solution}} = 2.86 \text{ M } \mathrm{NH_4^+}$$

Each formula unit of $(NH_4)_2Cr_2O_7$ produces *three* ions when dissolved in water (one dichromate ion and two ammonium ions), so the *total* concentration of ions in the solution is 1.43 M for dichromate and 2 x 1.43 M for ammonium resulting in a total ion concentration of 4.29 M.

9.4 Dilutions

In the lab, chemists and biologists often need to use a concentrated or stock solution to make a less concentrated, or **dilute**, solution. The concentration of a solution can be lowered by the addition of extra solvent. Recall the definition of molarity:

$$Molarity = \frac{moles\ of\ solute}{liters\ of\ solution} = \frac{mol}{L}$$

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Dimensional analysis can be used to quickly and conveniently calculate the concentration of a dilute solution. For instance, if a 15.8 M stock solution of nitric acid, HNO_3, is available in the lab, (Figure 9.8) and a student needs a dilute solution such as 500.0 mL of a 0.632 M nitric acid solution, the number of moles required can be determined as follows:

Figure 9.8 Nitric acid, HNO_3, is often purchased in its most concentrated form (15.8 M) and then later diluted to the desired concentration. *Image credit: Wikimedia Commons.*

In the lab, chemists and biologists often need to use a concentrated or stock solution to make a less concentrated, or **dilute**, solution. The concentration of a solution can be lowered by the addition of extra solvent. Recall the definition of molarity:

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$$500.0\ \text{mL of } 0.632\ \text{M HNO}_3 = 500.0\ \text{mL of } 0.632\ \frac{\text{mol}}{\text{L}}\ \text{HNO}_3 = 0.5000\ \text{L of } 0.632\frac{\text{mol}}{\text{L}}\ \text{HNO}_3$$

$$0.5000\ \text{L of } 0.632\frac{\text{mol}}{\text{L}}\ \text{HNO}_3 = 0.5000\ \text{L} \times 0.632\frac{\text{mol}}{\text{L}}\ \text{HNO}_3 = 0.0316\ \text{mol HNO}_3$$

The student needs 0.0316 moles of HNO_3 and must obtain that number of moles of HNO_3 from the 15.8 M stock solution of HNO_3. The amount of stock solution needed is therefore:

$$0.0316\ \cancel{\text{mol HNO}_3} \times \frac{1\ \text{L solution}}{15.8\ \cancel{\text{mol HNO}_3}} = 0.00200\ \text{L} \times \frac{1000\ \text{mL}}{1\ \text{L}} = 20.0\ \text{mL}$$

A volume of 20.0 mL of the 15.8 M stock HNO_3 solution must be transferred to a 500.0 mL (the final volume needed) volumetric flask. Water is the solvent and is added until the flask is filled exactly to the mark on the neck. The volumetric flask now contains a total volume of 500.0 mL, and in that 500.0 mL there are 0.0316 moles of HNO_3 dissolved (Figure 9.9).

It is a good idea to double-check the concentration just to be sure that the correct volumes and concentrations were used in the calculation. The 500.0 mL (or 0.5000 L) volumetric flask contains 0.0316 moles of HNO_3. The concentration of the solution in the flask is therefore:

$$[\text{HNO}_3] = \frac{moles\ of\ solute}{liters\ of\ solution} = \frac{0.0316\ \text{mol HNO}_3}{0.5000\ \text{L solution}} = 0.0632\frac{\text{mol}}{\text{L}} = 0.0632\ \text{M}$$

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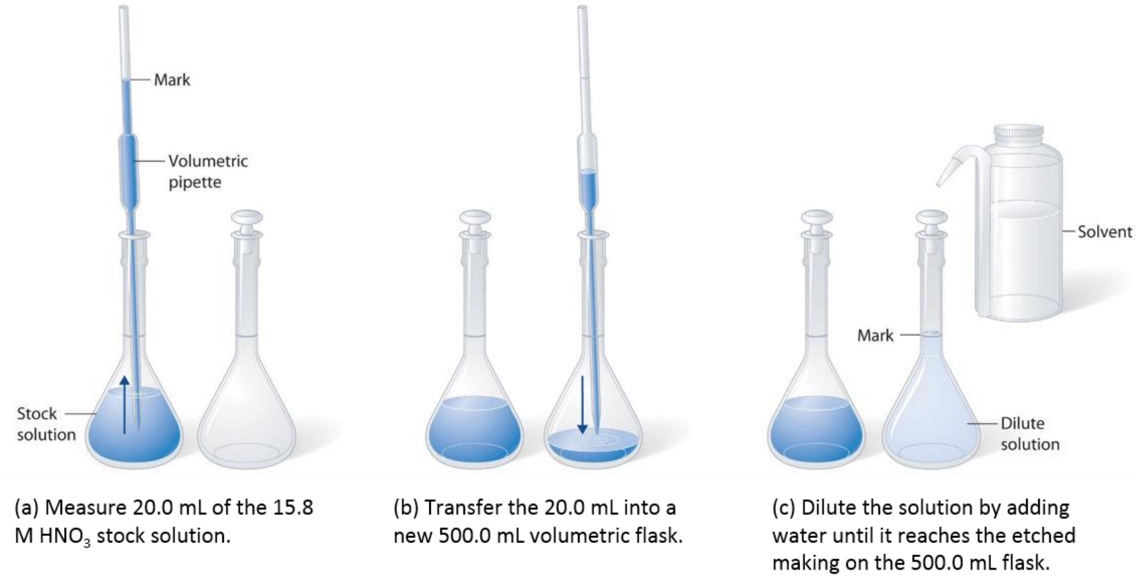

Figure 9.9 The steps used to make 500.0 mL of a 0.0632 M solution of nitric acid from a stock solution of 15.8 M nitric acid. *Image credit: Principles of General Chemistry*.

Many students use a common formula for dilution: $M_1V_1 = M_2V_2$. Note that this formula is *only* for dilution and *does not apply* to every calculation involving molarities and volumes. To avoid errors in this course and in general chemistry, it is best to *use dimensional analysis* for all calculations.

Do not use $M_1V_1 = M_2V_2$.

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Example 9.7

A laboratory technician needs to make 150.0 mL of 0.575 M NaOH. On the shelf in the lab, there is a stock solution of 2.50 M NaOH. Use dimensional analysis to solve the problem.

Solution

The number of moles of NaOH required can be calculated:

$$150.0\text{ mL of }0.575\text{ M NaOH} \rightarrow 150.0\text{ mL of }0.575\,\frac{\text{mol}}{\text{L}}\text{ NaOH} \rightarrow 0.1500\text{ L of }0.575\,\frac{\text{mol}}{\text{L}}\text{ NaOH}$$

$$0.1500\text{ L of }0.575\,\frac{\text{mol}}{\text{L}}\text{ NaOH} = 0.1500\text{ L} \times 0.575\,\frac{\text{mol}}{\text{L}}\text{ NaOH} = 0.08625\text{ mol NaOH}$$

To obtain 0.08625 mol of NaOH from a 2.50 M (2.50 mol/L) stock solution, what volume is necessary?

$$0.08625\,\cancel{\text{mol NaOH}} \times \frac{1\text{ L solution}}{2.50\,\cancel{\text{mol NaOH}}} = 0.0345\text{ L} \times \frac{1000\text{ mL}}{1\text{ L}} = 34.5\text{ mL}$$

The final volume needs to be 150.0 mL, so a 150.0 mL volumetric flask must be used. 34.5 mL of 2.50 M NaOH should be transferred to the 150.0 mL volumetric flask. The flask should then be filled to the mark on the neck with water, the solvent. The solution in the flask will contain 0.08625 moles of NaOH and a total of 150.0 mL of solution. Double-checking the concentration:

$$[\text{NaOH}] = \frac{moles\ of\ solute}{liters\ of\ solution} = \frac{0.08625\text{ mol HNO}_3}{0.1500\text{ L solution}} = 0.575\,\frac{\text{mol}}{\text{L}} = 0.575\text{ M}$$

Test Yourself

To run an assay in the biology lab, a student needs 25.0 mL of a 0.175 M buffer solution. The stock solution for the buffer is 1.00 M. What volume of the buffer stock solution must be used? How should the dilute solution be made?

Answer

4.38 mL of the 1.00 M buffer stock solution should be transferred to a 25.0 mL volumetric flask. The flask should be filled to the mark on the neck. The resulting 25.0 mL solution will have a concentration of 0.175 M.

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Chemistry is Everywhere: Preparing IV Solutions

In a hospital emergency room, a physician orders an intravenous (IV) delivery of 100 mL of 0.070 M KCl for a patient suffering from hypokalemia (low potassium levels). Does an aide run to a supply cabinet and take out an IV bag containing this 0.070 M potassium chloride?

Not likely. It is more probable that the aide must make the proper solution from a sterile *stock solution*, of potassium chloride. A syringe must be used to draw up a certain volume of stock solution and inject it into an IV bag so along with an appropriate amount of water such that the final concentration in the IV bag ends up being 0.070 M KCl.

A dilution calculation must be done quickly and accurately.

If the stock solution is 0.200 M KCl and the patient needs 100 mL of a 0.070 M KCl IV solution, the volume of solution needed can be calculated.

The patient needs 100 mL of 0.070 mol/L KCl:

$$100 \text{ mL} \times 0.070 \text{ mol/L KCl} = 0.100 \text{ L} \times 0.070 \text{ mol/L KCl} = 0.0070 \text{ mol KCl needed}$$

The stock solution available is 0.200 mol/L KCl. Start with the moles required.

$$0.0070 \text{ mol KCl} \times (1 \text{ L}/0.200 \text{ mol}) = 0.035 \text{ L} = 35. \text{ mL of the stock solution is required}$$

In order to get 100 mL of the solution for the IV, 35. mL of sterile stock KCl must be transferred to the IV bag. Then 65. mL of sterile water must be added to the IV bag. This will produce a solution that is 100 mL total with a concentration of 0.070 mol/L.

Medical and pharmaceutical personnel are constantly dealing with dosages that require concentration measurements and dilutions. It is an important responsibility: calculating the *wrong* dose can be useless, harmful, or even fatal!

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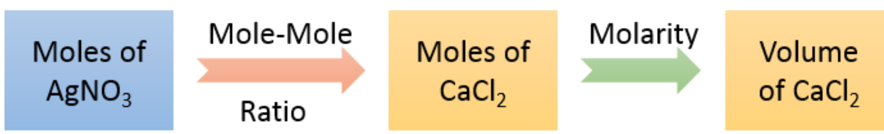
9.5 Concentrations as Conversion Factors

In Section 9.2, molarity was used as a conversion factor to determine the amount of moles or the volume of a solution. More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, consider the following chemical reaction:

$$2\ AgNO_3\ (aq) + CaCl_2\ (aq) \rightarrow 2\ AgCl\ (s) + Ca(NO_3)_2\ (aq)$$

If you wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, stoichiometry is required. According to the concept map below, first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution.



$$1.25\ \cancel{\text{mol}\ AgNO_3} \times \frac{1\ \cancel{\text{mol}\ CaCl_2}}{2\ \cancel{\text{mol}\ AgNO_3}} \times \frac{1\ \text{L solution}}{0.555\ \cancel{\text{mol}\ CaCl_2}} = 1.13\ \text{L}\ \ CaCl_2$$

This can be extended by starting with the mass of reactant instead of moles in Example 9.8.

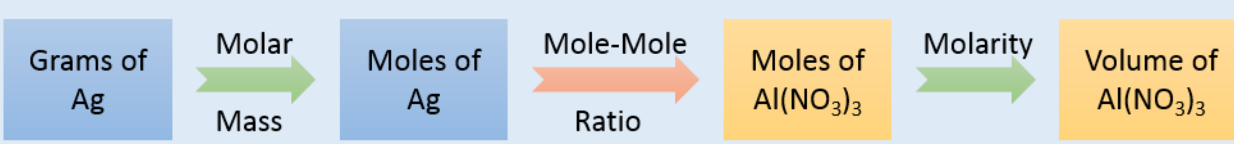
Example 9.8

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\ Ag(s) + Al(NO_3)_3(aq) \rightarrow 3\ AgNO_3(aq) + Al(s)$$

Solution

First convert the mass of Ag to moles before using the balanced chemical equation, and then use the definition of molarity as a conversion factor:



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$$3.66\ \cancel{\text{g Ag}} \times \frac{1\ \cancel{\text{mol Ag}}}{107.97\ \cancel{\text{g Ag}}} \times \frac{1\ \cancel{\text{mol Al(NO}_3)_3}}{3\ \cancel{\text{mol Ag}}} \times \frac{1\ \text{L solution}}{0.0995\ \cancel{\text{mol Al(NO}_3)_3}} = 0.114\ \text{L}\ \ \text{Al(NO}_3)_3$$

Test Yourself
What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$H_2C_2O_4(s) + 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(\ell)$$

Answer 0.777 L

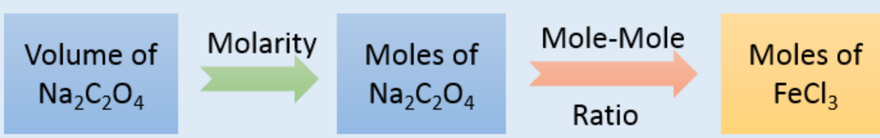
To extend your skills even further, recognize that quantities of one solution can be related to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, you can determine how much of another solution will be needed.

Example 9.9

A student finds that 9.04 mL of 0.1074 M $Na_2C_2O_4$ solution is required to react all of the $FeCl_3$ in 10.0 mL of a $FeCl_3$ solution. What was the concentration of the $FeCl_3$ in the original solution? The balanced chemical equation is as follows:

$$2\ FeCl_3\ (aq) + 3\ Na_2C_2O_4(aq) \rightarrow Fe_2(C_2O_4)_3(s) + 6\ NaCl(aq)$$

Solution
First determine the number of moles of $Na_2C_2O_4$ that reacted using the concentration of the solution as a conversion factor. Then, use the balanced chemical equation to determine the number of moles of $FeCl_3$.



$$9.04\ \cancel{\text{mL NaC}_2\text{O}_4} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.1074\ \cancel{\text{mol NaC}_2\text{O}_4}}{1\ \cancel{\text{L}}} \times \frac{2\ \cancel{\text{mol FeCl}_3}}{3\ \cancel{\text{mol NaC}_2\text{O}_4}} = 6.47 \times 10^{-4}\ \text{mol FeCl}_3$$

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Next, determine the concentration of $FeCl_3$ in the original solution. Converting 10.00 mL into liters (0.01000 L), you can use the definition of molarity directly:

$$Molarity = \frac{moles}{liters} = \frac{6.47 \times 10^{-4}\ \text{mol FeCl}_3}{0.01000\ \text{L}} = 0.0647\ \frac{\text{mol}}{\text{L}}\text{FeCl}_3 = 0.0647\ \text{M FeCl}_3$$

Test Yourself
A student uses 54.06 mL of 0.0987 M KOH to complete the reaction with 25.00 mL of H_3PO_4. What is the concentration of H_3PO_4?

$$H_3PO_4(aq) + 3\ KOH(aq) \rightarrow K_3PO_4(aq) + 3\ H_2O(\ell)$$

Answer 0.0711 M

9.6 Reactions in Solution Titrations

In Chapter 7 neutralization reactions, or reactions between acids and bases, were discussed. In this chapter neutralization reactions will be used to quantitatively determine amounts of a reactant. The experimental technique to do so is called a **titration**. Although a titration can be performed with almost any chemical reaction for which the balanced chemical equation is known, here, only titrations that involve acid-base reactions will be considered.

In a titration, one reagent has a known concentration or amount, while the other reagent has an unknown concentration or amount. Typically, the known reagent (the **titrant**) is added to a buret, as the chemist in Figure 9.10(a) shows. The unknown amount of substance (the **analyte**) is dissolved in water and placed under the buret. Figure 9.10(b) shows the buret markings that are used to determine the volume of titrant that has been added to the analyte.

When the reaction is complete, it is said to be at the **equivalence point** and the moles of acid equals the moles of base. Chemists need to know when the equivalence point is reached so that the sample is not over-titrated. To help see the equivalence point for a reaction, a chemical **indicator** is added to the analyte, such as phenolphthalein used in Figure 9.11 to give a pink color.

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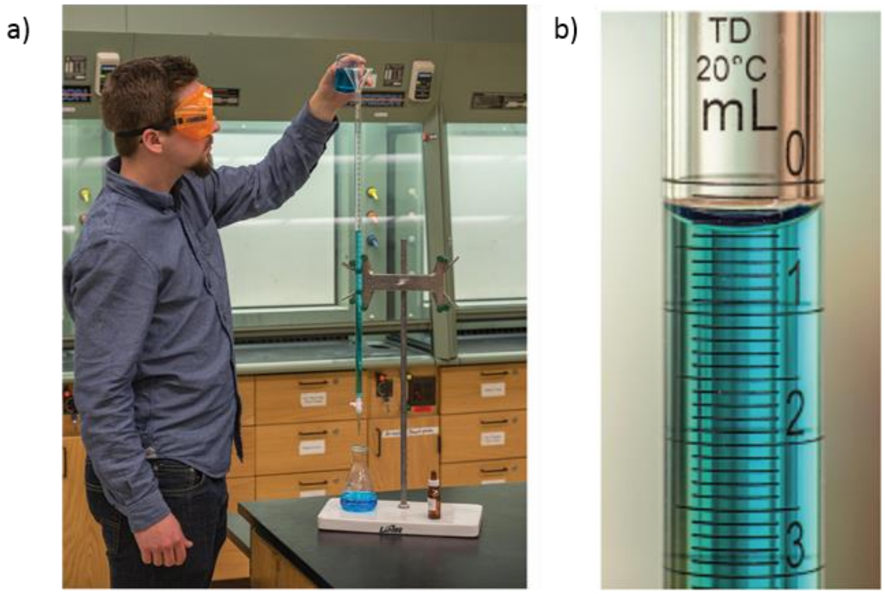

Figure 9.10 In (a), a chemist adds liquid to a buret. This solution will be the titrant. The flask below the buret contains the analyte with unknown concentration. The volume markings on the buret (b) are used to determine the amount of liquid dispensed during the experiment. *Image credit: OpenStax Chemistry.*

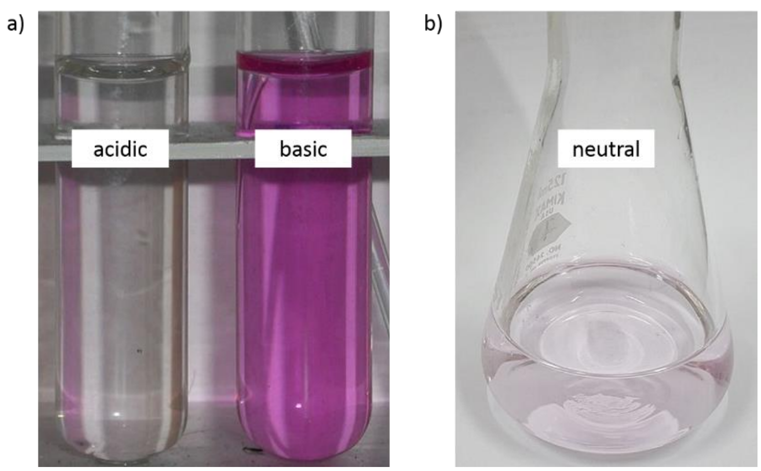

Figure 9.11 When titrating an acid with a base, the indicator phenolphthalein (a) first appears clear and then turns pink as base is added. (b) The equivalence point is reached when a pale pink is formed. *Image credit: Wikimedia Commons by Kengskn.*

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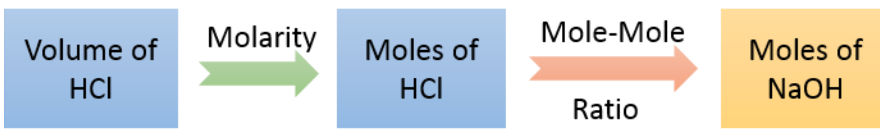
The data acquired from a titration can be used to determine the unknown amount of analyte. For example, consider if you are given an aqueous solution of NaOH but do not know the molarity. You can use the reaction between HCl and NaOH as shown below to find the missing molarity.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(\ell)$$

Suppose a student performs and experiment and uses 25.66 mL of 0.1078 M HCl to completely titrate a 150.0 mL sample of the NaOH. To find the molarity of the 150.0 mL sample of NaOH, the student will need to determine the moles of HCl used and then perform stoichiometry to find the moles of NaOH present in the sample. Consider the concept map below.



$$25.66\ \cancel{\text{mL HCl}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.1078\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol HCl}}} = 0.002766\ \text{mol NaOH}$$

To determine the molarity of the solution, divide the number of moles by the number of liters of Solution

$$Molarity = \frac{moles}{liters} = \frac{0.002766\ \text{mol NaOH}}{0.1500\ \text{L}} = 0.01844\ \frac{\text{mol}}{\text{L}}\text{NaOH} = 0.01844\ \text{M NaOH}$$

Example 9.10

What mass of $Ca(OH)_2$ is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO_3? The balanced chemical equation is as follows:

$$2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(\ell)$$

Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

$$44.02\ \cancel{\text{mL HNO}_3} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.0885\ \cancel{\text{mol HNO}_3}}{1\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{mol Ca(OH)}_2}}{2\ \cancel{\text{mol HNO}_3}} \times \frac{74.10\ \cancel{\text{g Ca(OH)}_2}}{1\ \cancel{\text{mol Ca(OH)}_2}} =$$

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$$= 0.144 \text{ g Ca(OH)}_2$$

Test Yourself
What mass of $H_2C_2O_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is:

$$H_2C_2O_4(aq) + 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(\ell)$$

Answer 0.182 g $H_2C_2O_4$

Example 9.10

What mass of $Ca(OH)_2$ is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO_3? The balanced chemical equation is as follows:

$$2\ HNO_3(aq) + Ca(OH)_2(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(\ell)$$

Solution
In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

$$44.02\ \cancel{\text{mL HNO}_3} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{0.0885\ \cancel{\text{mol HNO}_3}}{1\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{mol Ca(OH)}_2}}{2\ \cancel{\text{mol HNO}_3}} \times \frac{74.10\ \cancel{\text{g Ca(OH)}_2}}{1\ \cancel{\text{mol Ca(OH)}_2}} =$$

$$= 0.144 \text{ g Ca(OH)}_2$$

Test Yourself
What mass of $H_2C_2O_4$ is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is:

$$H_2C_2O_4(aq) + 2\ NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2\ H_2O(\ell)$$

Answer 0.182 g $H_2C_2O_4$

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Chapter 9 Practice Problems

9.2 Concentration and Molarity

1. What is the concentration of each species present in the following aqueous solutions?

   (a) 0.489 mol of $NiSO_4$ in 600.0 mL of solution
   (b) mol of magnesium bromide in 500.0 mL of solution
   (c) 0.146 mol of glucose in 800.0 mL of solution
   (d) 0.479 mol of $CeCl_3$ in 700.0 mL of solution

2. What is the concentration of each species present in the following aqueous solutions?

   (a) 0.324 mol of K_2MoO_4 in 250.0 mL of solution
   (b) 0.528 mol of potassium formate in 300.0 mL of solution
   (c) 0.477 mol of $KClO_3$ in 900.0 mL of solution
   (d) 0.378 mol of potassium iodide in 750.0 mL of solution

3. What is the molarity of a solution made by dissolving 13.4 g of $NaNO_3$ in 345 mL of H_2O?

4. What is the molarity of a solution made by dissolving 332 g of $C_6H_{12}O_6$ in 4.66 L of H_2O?

5. What is the molar concentration of each solution?

   (a) 9.8 g of lithium sulfate in 300.0 mL of solution
   (b) 12.4 g of sucrose ($C_{12}H_{22}O_{11}$) in 750.0 mL of solution
   (c) 14.2 g of iron(III) nitrate hexahydrate in 300.0 mL of solution
   (d) 3.59 g of calcium bromide in 250.0 mL of solution

6. What is the molar concentration of each solution?
   (a) 12.8 g of sodium hydrogen sulfate in 400.0 mL of solution
   (b) 67.3 g of potassium hydrogen phosphate in 250.0 mL of solution
   (c) 0.9864 g of barium chloride in 350.0 mL of solution
   (d) 3.097g of tartaric acid ($C_4H_6O_6$) in 250.0 mL of solution

7. Give the concentration of each reactant in the following equations, assuming 20.0 g of each and a solution volume of 250.0 mL for each reactant.
   (a) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow$
   (b) $Ca(OH)_2(aq) + H_3PO_4(aq) \rightarrow$
   (c) $Al(NO_3)_3(aq) + H_2SO_4(aq) \rightarrow$
   (d) $Pb(NO_3)_2(aq) + CuSO_4(aq) \rightarrow$
   (e) $Al(CH_3CO_2)_3(aq) + NaOH(aq) \rightarrow$

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8. How many moles of $MgCl_2$ are present in 0.0331 L of a 2.55 M solution?

9. How many moles of NH_4Br are present in 88.9 mL of a 0.228 M solution?

10. What volume of 0.556 M NaCl is needed to obtain 0.882 mol of NaCl?

11. What volume of 3.99 M H_2SO_4 is needed to obtain 4.61 mol of H_2SO_4?

12. Calculate the number of grams of solute in 1.500 L of each solution.
    (a) 0.2593 M $NaBrO_3$
    (b) 1.592 M KNO_3
    (c) 1.559 M acetic acid
    (d) 0.943 M potassium iodate

13. Calculate the number of grams of solute in 2.000 L of each solution.
    (a) 0.1065 M BaI_2
    (b) 1.135 M Na_2SO_4
    (c) 1.428 M NH_4Br
    (d) 0.889 M sodium acetate

14. What volume of 0.333 M $Al(NO_3)_3$ contains 26.7 g of $Al(NO_3)_3$?

15. What volume of 1.772 M $BaCl_2$ contains 123 g of $BaCl_2$?

16. If all solutions below contain the same solute, which contains the greater mass of solute?
    (a) 1.40 L of a 0.334 M solution or 1.10 L of a 0.420 M solution
    (b) 25.0 mL of a 0.134 M solution or 10.0 mL of a 0.295 M solution
    (c) 250 mL of a 0.489 M solution or 150 mL of a 0.769 M solution

17. Complete the following table for 500 mL of solution.

Compound	**Mass**	**Moles**	**Concentration (M)**
calcium sulfate	4.86		
acetic acid		3.62	
hydrogen iodide dihydrate			1.273
barium bromide	3.92		
glucose			0.983
sodium acetate		2.42	

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18. What are the individual ion concentrations and the total ion concentration in 0.66 M $Mg(NO_3)_2$?

19. What are the individual ion concentrations and the total ion concentration in 1.04 M $Al_2(SO_4)_3$?

20. If the $C_2H_3O_2^-$ ion concentration in a solution is 0.554 M, what is the concentration of $Ca(C_2H_3O_2)_2$?

21. If the Cl^- ion concentration in a solution is 2.61 M, what is the concentration of $FeCl_3$?

22. An experiment required 200.0 mL of a 0.330 M solution of Na_2CrO_4. A stock solution of Na_2CrO_4 containing 20.0% solute by mass with a density of 1.19 g/cm^3 was used to prepare this solution. Describe how to prepare 200.0 mL of a 0.330 M solution of Na_2CrO_4 using the stock solution.

9.3 Concentration The Preparation of Solutions

23. An experiment required 200.0 mL of a 0.330 M solution of Na_2CrO_4. A stock solution of Na_2CrO_4 containing 20.0% solute by mass with a density of 1.19 g/cm^3 was used to prepare this solution. Describe how to prepare 200.0 mL of a 0.330 M solution of Na_2CrO_4 using the stock solution.

24. Calcium hypochlorite [$Ca(ClO)_2$] is an effective disinfectant for clothing and bedding. If a solution has a $Ca(ClO)_2$ concentration of 3.4 g per 100 mL of solution, what is the molarity of hypochlorite?

25. Phenol (C_6H_5OH) is often used as an antiseptic in mouthwashes and throat lozenges. If a mouthwash has a phenol concentration of 1.5 g per 100 mL of solution, what is the molarity of phenol?

26. If a tablet containing 100 mg of caffeine ($C_8H_{10}N_4O_2$) is dissolved in water to give 10.0 oz of solution, what is the molar concentration of caffeine in the solution?

27. A certain drug label carries instructions to add 10.0 mL of sterile water, stating that each milliliter of the resulting solution will contain 0.500 g of medication. If a patient has a prescribed dose of 900.0 mg, how many milliliters of the solution should be administered?

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9.4 Dilutions

28. What is the difference between dilution and concentration?

29. What quantity remains constant when you dilute a solution?

30. A 1.88 M solution of NaCl has an initial volume of 34.5 mL. What is the final concentration of the solution if it is diluted to 134 mL?

31. A 0.664 M solution of NaCl has an initial volume of 2.55 L. What is the final concentration of the solution if it is diluted to 3.88 L?

32. If 1.00 mL of a 2.25 M H_2SO_4 solution needs to be diluted to 1.00 M, what will be its final volume?

33. If 12.00 L of a 6.00 M HNO_3 solution needs to be diluted to 0.750 M, what will be its final volume?

34. If 665 mL of a 0.875 M KBr solution are boiled gently to concentrate the solute to 1.45 M, what will be its final volume?

35. If 1.00 L of an LiOH solution is boiled down to 164 mL and its initial concentration is 0.00555 M, what is its final concentration?

36. How much water must be added to 75.0 mL of 0.332 M $FeCl_3(aq)$ to reduce its concentration to 0.250 M? Assume volumes are additive.

37. How much water must be added to 1.55 L of 1.65 M $Sc(NO_3)_3(aq)$ to reduce its concentration to 1.00 M? Assume volumes are additive.

9.5 Concentrations as Conversion Factors

38. Using concentration as a conversion factor, how many moles of solute are in 3.44 L of 0.753 M $CaCl_2$?

39. Using concentration as a conversion factor, how many moles of solute are in 844 mL of 2.09 M $MgSO_4$?

40. Using concentration as a conversion factor, how many liters are needed to provide 0.822 mol of NaBr from a 0.665 M solution?

41. Using concentration as a conversion factor, how many liters are needed to provide 2.500 mol of $(NH_2)_2CO$ from a 1.087 M solution?

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42. What is the mass of solute in 24.5 mL of 0.755 M $CoCl_2$?

43. What is the mass of solute in 3.81 L of 0.0232 M $Zn(NO_3)_2$?

44. What volume of solution is needed to provide 9.04 g of NiF_2 from a 0.332 M solution?

45. What volume of solution is needed to provide 0.229 g of CH_2O from a 0.00560 M solution?

9.6 Reactions in Solution Titrations

46. What volume of 3.44 M HCl will react with 5.33 mol of $CaCO_3$?
    $2\ HCl(aq) + CaCO_3(aq) \rightarrow CaCl_2(aq) + H_2O(\ell) + CO_2(g)$

47. What volume of 0.779 M NaCl will react with 40.8 mol of $Pb(NO_3)_2$?
    $Pb(NO_3)_2(aq) + 2\ NaCl(aq) \rightarrow PbCl_2(aq) + 2\ NaNO_3(aq)$

48. What volume of 0.905 M H_2SO_4 will react with 26.7 mL of 0.554 M NaOH?
    $H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(\ell)$

49. What volume of 1.500 M Na_2CO_3 will react with 342 mL of 0.733 M H_3PO_4?
    $3\ Na_2CO_3(aq) + 2\ H_3PO_4(aq) \rightarrow 2\ Na_3PO_4(aq) + 3\ H_2O(\ell) + 3\ CO_2(g)$

50. It takes 23.77 mL of 0.1505 M HCl to titrate with 15.00 mL of $Ca(OH)_2$. What is the concentration of $Ca(OH)_2$? You will need to write the balanced chemical equation first.

51. It takes 97.62 mL of 0.0546 M NaOH to titrate a 25.00 mL sample of H_2SO_4. What is the concentration of H_2SO_4? You will need to write the balanced chemical equation first.

52. It takes 4.667 mL of 0.0997 M HNO_3 to dissolve some solid Cu. What mass of Cu can be dissolved?
    $Cu(s) + 4\ HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2\ NO_2(g) + 2\ H_2O(\ell)$

53. It takes 49.08 mL of 0.877 M NH_3 to dissolve some solid AgCl. What mass of AgCl can be dissolved?
    $AgCl(s) + 4\ NH_3(aq) \rightarrow Ag(NH_3)_4Cl(aq)$

54. What mass of 3.00% by mass H_2O_2 is needed to produce 66.3 g of O_2(g)?
    $2\ H_2O_2(aq) \rightarrow 2\ H_2O(\ell) + O_2(g)$

55. A 0.75% by mass solution of Na_2CO_3 is used to precipitate Ca^{2+} ions from solution. What mass of solution is needed to precipitate 40.7 L of solution with a concentration of 0.0225 M $Ca^{2+}(aq)$?
    $Na_2CO_3(aq) + Ca^{2+}(aq) \rightarrow CaCO_3(s) + 2\ Na^+(aq)$

56. Define *titration*.

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57. What is the difference between the *titrant* and the *analyte*?

58. True or false: An acid is always the titrant. Explain your answer.

59. True or false: An analyte is always dissolved before reaction. Explain your answer.

60. If 55.60 mL of 0.2221 M HCl was needed to titrate a sample of NaOH to its equivalence point, what mass of NaOH was present?

61. If 16.33 mL of 0.6664 M KOH was needed to titrate a sample of $HC_2H_3O_2$ to its equivalence point, what mass of $HC_2H_3O_2$ was present?

62. It takes 45.66 mL of 0.1126 M HBr to titrate 25.00 mL of $Ca(OH)_2$ to its equivalence point. What is the original concentration of the $Ca(OH)_2$ solution?

63. It takes 9.77 mL of 0.883 M H_2SO_4 to titrate 15.00 mL of KOH to its equivalence point. What is the original concentration of the KOH solution?

Review: How to Succeed in Chemistry

64. Confirm the dates of all your final exams. Make a study schedule that includes all your classes.

65. How do you plan to study for your chemistry final exam? Make sure you look over lecture notes, worksheets, homework problems, and old quizzes/exams.

66. Identify the chapters and sections that are difficult for you. Make sure to spend some extra review time on these topics and see a chemistry tutor if necessary.

67. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two short-answer exam questions*. You may consult the internet for related problems. Include *worked solutions* to the problems you create.

68. What's important? Try to predict what your professor will ask about this chapter on an exam by creating *two multiple-choice exam questions*. Remember to include *worked solutions* to the problems you create.

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