Chapter 1

Essential Ideas

Figure 1.1 Chemical substances and processes are essential for our existence, providing sustenance, keeping us clean and healthy, fabricating electronic devices, enabling transportation, and much more. (credit "left": modification of work by "vxla"/Flickr; credit "left middle": modification of work by "the Italian voice"/Flickr; credit "right middle": modification of work by Jason Trim; credit "right": modification of work by "gosheshe"/Flickr)

Chapter Outline

1.1 Chemistry in Context
1.2 Measurements
1.3 Measurement Uncertainty, Accuracy, and Precision
1.4 Mathematical Treatment of Measurement Results

Introduction

Your alarm goes off and, after hitting "snooze" once or twice, you pry yourself out of bed. You make a cup of coffee to help you get going, and then you shower, get dressed, eat breakfast, and check your phone for messages. On your way to school, you stop to fill your car's gas tank, almost making you late for the first day of chemistry class. As you find a seat in the classroom, you read the question projected on the screen: "Welcome to class! Why should we study chemistry?"Do you have an answer? You may be studying chemistry because it fulfills an academic requirement, but if you consider your daily activities, you might find chemistry interesting for other reasons. Most everything you do and encounter during your day involves

chemistry. Making coffee, cooking eggs, and toasting bread involve chemistry. The products you use—like soap and shampoo, the fabrics you wear, the electronics that keep you connected to your world, the gasoline that propels your car—all of these and more involve chemical substances and processes. Whether you are aware or not, chemistry is part of your everyday world. In this course, you will learn many of the essential principles underlying the chemistry of modern-day life.

1.1 Chemistry in Context

By the end of this module, you will be able to:

- Outline the historical development of chemistry
- Provide examples of the importance of chemistry in everyday life
- Describe the scientific method
- Differentiate among hypotheses, theories, and laws
- Provide examples illustrating macroscopic, microscopic, and symbolic domains

Throughout human history, people have tried to convert matter into more useful forms. Our Stone Age ancestors chipped pieces of flint into useful tools and carved wood into statues and toys. These endeavors involved changing the shape of a substance without changing the substance itself. But as our knowledge increased, humans began to change the composition of the substances as well—clay was converted into pottery, hides were cured to make garments, copper ores were transformed into copper tools and weapons, and grain was made into bread.

Humans began to practice chemistry when they learned to control fire and use it to cook, make pottery, and smelt metals. Subsequently, they began to separate and use specific components of matter. A variety of drugs such as aloe, myrrh, and opium were isolated from plants. Dyes, such as indigo and Tyrian purple, were extracted from plant and animal matter. Metals were combined to form alloys—for example, copper and tin were mixed together to make bronze—and more elaborate smelting techniques produced iron. Alkalis were extracted from ashes, and soaps were prepared by combining these alkalis with fats. Alcohol was produced by fermentation and purified by distillation.

Attempts to understand the behavior of matter extend back for more than 2500 years. As early as the sixth century BC, Greek philosophers discussed a system in which water was the basis of all things. You may have heard of the Greek postulate that matter consists of four elements: earth, air, fire, and water. Subsequently, an amalgamation of chemical technologies and philosophical speculations were spread from Egypt, China, and the eastern Mediterranean by alchemists, who endeavored to transform “base metals” such as lead into “noble metals” like gold, and to create elixirs to cure disease and extend life (**Figure 1.2**).

Figure 1.2 This portrayal shows an alchemist's workshop circa 1580. Although alchemy made some useful contributions to how to manipulate matter, it was not scientific by modern standards. (credit: Chemical Heritage Foundation)

From alchemy came the historical progressions that led to modern chemistry: the isolation of drugs from natural sources, metallurgy, and the dye industry. Today, chemistry continues to deepen our understanding and improve our ability to harness and control the behavior of matter. This effort has been so successful that many people do not realize either the central position of chemistry among the sciences or the importance and universality of chemistry in daily life.

Chemistry: The Central Science

Chemistry is sometimes referred to as "the central science" due to its interconnectedness with a vast array of other STEM disciplines (STEM stands for areas of study in the science, technology, engineering, and math fields). Chemistry and the language of chemists play vital roles in biology, medicine, materials science, forensics, environmental science, and many other fields (**Figure 1.3**). The basic principles of physics are essential for understanding many aspects of chemistry, and there is extensive overlap between many subdisciplines within the two fields, such as chemical physics and nuclear chemistry. Mathematics, computer science, and information theory provide important tools that help us calculate, interpret, describe, and generally make sense of the chemical world. Biology and chemistry converge in biochemistry, which is crucial to understanding the many complex factors and processes that keep living organisms (such as us) 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.

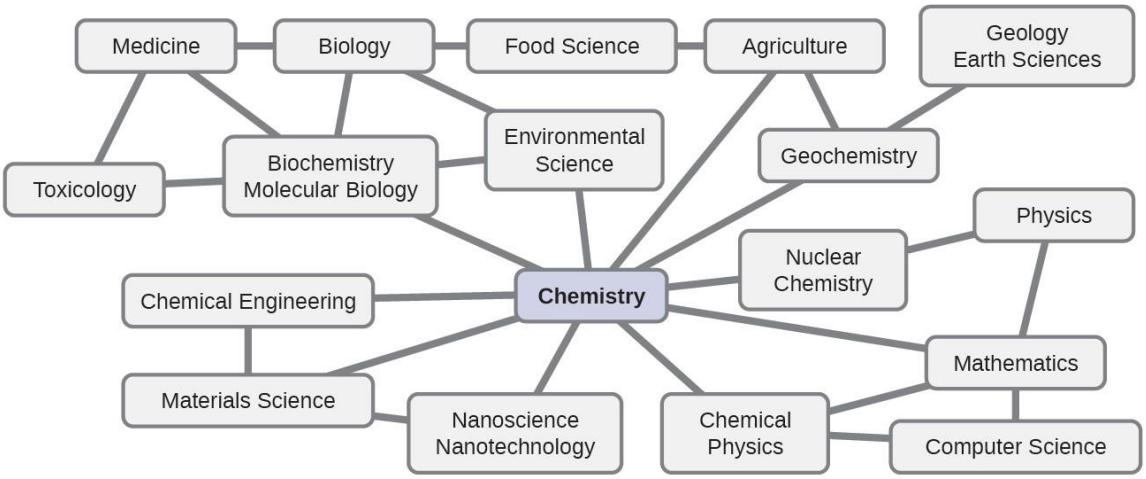


Figure 1.3 Knowledge of chemistry is central to understanding a wide range of scientific disciplines. This diagram shows just some of the interrelationships between chemistry and other fields.

What are some changes in matter that are essential to daily life? Digesting and assimilating food, synthesizing polymers that are used to make clothing, containers, cookware, and credit cards, and refining crude oil into gasoline and other products are just a few examples. As you proceed through this course, you will discover many different examples of changes in the composition and structure of matter, how to classify these changes and how they occurred, their causes, the changes in energy that accompany them, and the principles and laws involved. As you learn about these things, you will be learning **chemistry**, the study of the composition, properties, and interactions of matter. The practice of chemistry is not limited to chemistry books or laboratories: It happens whenever someone is involved in changes in matter or in conditions that may lead to such changes.

The Scientific Method

Chemistry is a science based on observation and experimentation. Doing chemistry involves attempting to answer questions and explain observations in terms of the laws and theories of chemistry, using procedures that are accepted by the scientific community. There is no single route to answering a question or explaining an observation, but there is an aspect common to every approach: Each uses knowledge based on experiments that can be reproduced to verify the results. Some routes involve a **hypothesis**, a tentative

explanation of observations that acts as a guide for gathering and checking information. We test a hypothesis by experimentation, calculation, and/or comparison with the experiments of others and then refine it as needed. Some hypotheses are attempts to explain the behavior that is summarized in laws. The **laws** of science summarize a vast number of experimental observations and describe or predict some facet of the natural world. If such a hypothesis turns out to be capable of explaining a large body of experimental data, it can reach the status of a theory. Scientific **theories** are well-substantiated, comprehensive, testable explanations of particular aspects of nature. Theories are accepted because they provide satisfactory explanations, but they can be modified if new data become available. The path of discovery that leads from question and observation to law or hypothesis to theory, combined with experimental verification of the hypothesis and any necessary modification of the theory, is called the **scientific method** (Figure 1.4).

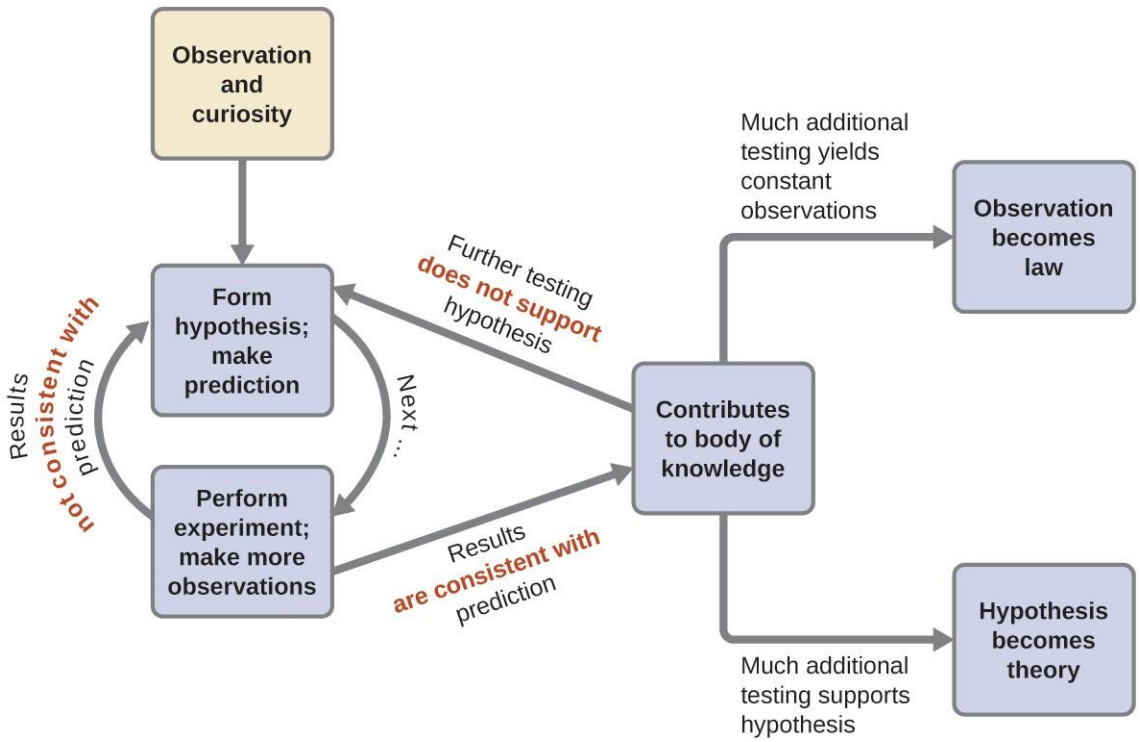


Figure 1.4 The scientific method follows a process similar to the one shown in this diagram. All the key components are shown, in roughly the right order. Scientific progress is seldom neat and clean: It requires open inquiry and the reworking of questions and ideas in response to findings.

The Domains of Chemistry

Chemists study and describe the behavior of matter and energy in three different domains: macroscopic, microscopic, and symbolic. These domains provide different ways of considering and describing chemical behavior.

Macro is a Greek word that means "large." The **macroscopic domain** is familiar to us: It is the realm of

everyday things that are large enough to be sensed directly by human sight or touch. In daily life, this includes the food you eat and the breeze you feel on your face. The macroscopic domain includes everyday and laboratory chemistry, where we observe and measure physical and chemical properties, or changes such as density, solubility, and flammability.

The **microscopic domain** of chemistry is almost always visited in the imagination. *Micro* also comes from Greek and means "small." Some aspects of the microscopic domains are visible through a microscope, such as a magnified image of graphite or bacteria. Viruses, for instance, are too small to be seen with the naked eye, but when we're suffering from a cold, we're reminded of how real they are.

However, most of the subjects in the microscopic domain of chemistry—such as atoms and molecules—are too small to be seen even with standard microscopes and often must be pictured in the mind. Other components of the microscopic domain include ions and electrons, protons and neutrons, and chemical bonds, each of which is far too small to see. This domain includes the individual metal atoms in a wire, the ions that compose a salt crystal, the changes in individual molecules that result in a color change, the conversion of nutrient molecules into tissue and energy, and the evolution of heat as bonds that hold atoms together are created.

The **symbolic domain** contains the specialized language used to represent components of the macroscopic and microscopic domains. Chemical symbols (such as those used in the periodic table), chemical formulas, and chemical equations are part of the symbolic domain, as are graphs and drawings. We can also consider calculations as part of the symbolic domain. These symbols play an important role in chemistry because they help interpret the behavior of the macroscopic domain in terms of the components of the microscopic domain. One of the challenges for students learning chemistry is recognizing that the same symbols can represent different things in the macroscopic and microscopic domains, and one of the features that makes chemistry fascinating is the use of a domain that must be imagined to explain behavior in a domain that can be observed.

A helpful way to understand the three domains is via the essential and ubiquitous substance of water. That water is a liquid at moderate temperatures, will freeze to form a solid at lower temperatures, and boil to form a gas at higher temperatures (**Figure 1.5**) are macroscopic observations. But some properties of water fall into the microscopic domain—what we cannot observe with the naked eye. The description of water as comprised of two hydrogen atoms and one oxygen atom, and the explanation of freezing and boiling in terms of attractions between these molecules, is within the microscopic arena. The formula H_2O, which can describe water at either the macroscopic or microscopic levels, is an example of the symbolic domain. The abbreviations (*g*) for gas, (*s*) for solid, and (*l*) for liquid are also symbolic.

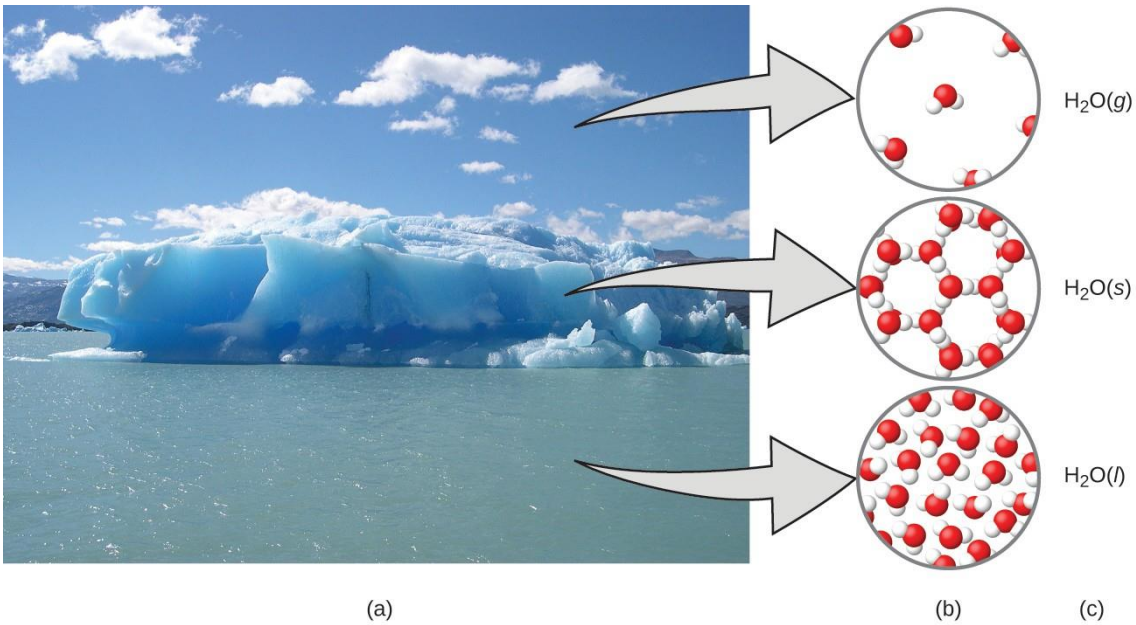


Figure 1.5 (a) Moisture in the air, icebergs, and the ocean represent water in the macroscopic domain. (b) At the molecular level (microscopic domain), gas molecules are far apart and disorganized, solid water molecules are close together and organized, and liquid molecules are close together and disorganized. (c) The formula H_2O symbolizes water, and (*g*), (*s*), and (*l*) symbolize its phases. Note that clouds are actually comprised of either very small liquid water droplets or solid water crystals; gaseous water in our atmosphere is not visible to the naked eye, although it may be sensed as humidity. (credit a: modification of work by "Gorkaazk"/Wikimedia Commons)

1.2 Measurements

By the end of this section, you will be able to:

- Explain the process of measurement
- Identify the three basic parts of a quantity
- Describe the properties and units of length, mass, volume, density, temperature, and time
- Perform basic unit calculations and conversions in the metric and other unit systems

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

scientific notation.

(Scientific notation is also known as exponential notation; a review of this topic can be found in **Appendix B**.) For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as 2.98×10^5 kg. The mass of the average mosquito is about 0.0000025 kilograms, which can be written as 2.5×10^{-6}kg.

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

We usually report the results of scientific measurements in SI units, an updated version of the metric system, using the units listed in **Table 1.2**. Other units can be derived from these base units. The standards for these units are fixed by international agreement, and they are called the **International System of Units** or **SI Units** (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964.

Base Units of the SI System

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

Table 1.2

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

10^3 m). The prefixes used and the powers to which 10 are raised are listed in **Table 1.3**.

Common Unit Prefixes

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

Table 1.3

Link to Learning

Need a refresher or more practice with scientific notation? Visit this **site (http://openstaxcollege.org/l/16notation)** to go over the basics of scientific notation.

SI Base Units

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

Length

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

a yard (**Figure 1.23**); one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

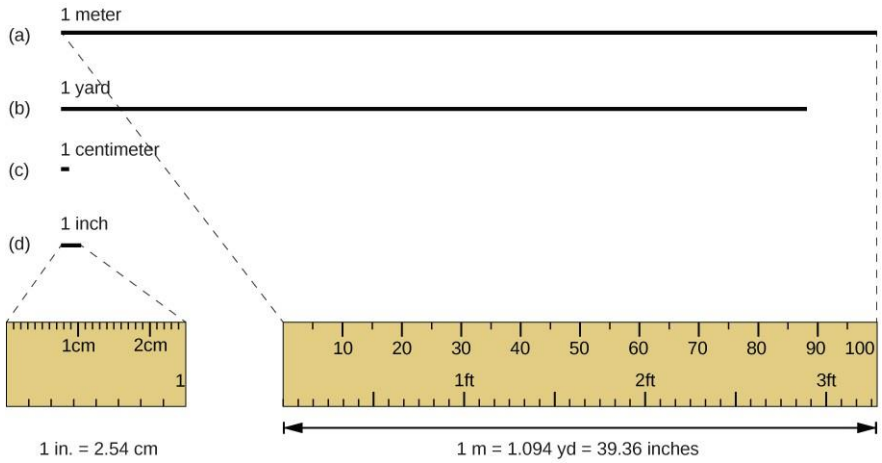


Figure 1.23 The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

Mass

The standard unit of mass in the SI system is the **kilogram (kg)**. A kilogram was originally defined as the mass of a liter of water (a cube of water with an edge length of exactly 0.1 meter). It is now defined by a certain cylinder of platinum-iridium alloy, which is kept in France (**Figure 1.24**). Any object with the same mass as this cylinder is said to have a mass of 1 kilogram. One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (10^{-3} kg).

Figure 1.24 This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology)

Temperature

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

Time

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

Derived SI Units

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

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (**Figure 1.25**). The standard volume is a **cubic meter (m^3)**, a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

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

A **cubic centimeter (cm^3)** is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for **c**ubic **c**entimeter) is often used by health professionals. A cubic centimeter is also called a **milliliter (mL)** and is 1/1000 of a liter.

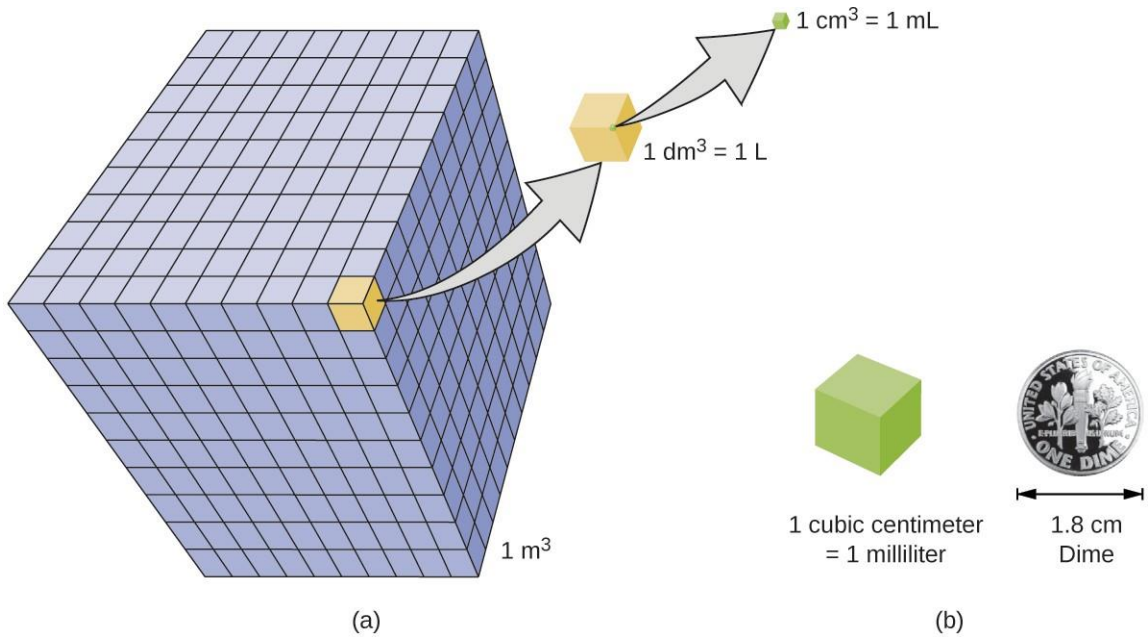


Figure 1.25 (a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

Density

We use the mass and volume of a substance to determine its density. Thus, the units of density are defined by the base units of mass and length. The **density** of a substance is the ratio of the mass of a sample of the substance to its volume. The SI unit for density is the kilogram per cubic meter (kg/m^3). For many situations, however, this as an inconvenient unit, and we often use grams per cubic centimeter (g/cm^3) for the densities of solids and liquids, and grams per liter (g/L) for gases. Although there are exceptions, most liquids and solids have densities that range from about 0.7 g/cm^3 (the density of gasoline) to 19 g/cm^3 (the density of gold). The density of air is about 1.2 g/L. **Table 1.4** shows the densities of some common substances.

Densities of Common Substances

Solids	Liquids	Gases (at 25 °C and 1 atm)
ice (at 0 °C) 0.92 g/cm^3	water 1.0 g/cm^3	dry air 1.20 g/L
oak (wood) 0.60–0.90 g/cm^3	ethanol 0.79 g/cm^3	oxygen 1.31 g/L
iron 7.9 g/cm^3	acetone 0.79 g/cm^3	nitrogen 1.14 g/L
copper 9.0 g/cm^3	glycerin 1.26 g/cm^3	carbon dioxide 1.80 g/L

lead 11.3 g/cm^3	olive oil 0.92 g/cm^3	helium 0.16 g/L
silver 10.5 g/cm^3	gasoline 0.70–0.77 g/cm^3	neon 0.83 g/L
gold 19.3 g/cm^3	mercury 13.6 g/cm^3	radon 9.1 g/L

Table 1.4

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

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

Example 1.1

Calculation of Density: Gold—in bricks, bars, and coins—has been a form of currency for centuries. In order to swindle people into paying for a brick of gold without actually investing in a brick of gold, people have considered filling the centers of hollow gold bricks with lead to fool buyers into thinking that the entire brick is gold. It does not work: Lead is a dense substance, but its density is not as great as that of gold, 19.3 g/cm3. What is the density of lead if a cube of lead has an edge length of 2.00 cm and a mass of 90.7 g?

Solution:

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

$$volume\ of\ lead\ cube = 2.00\ cm * 2.00\ cm * 2.00\ cm = 8.00 cm^3$$

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

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

Check Your Learning

(a) To three decimal places, what is the volume of a cube (cm3) with an edge length of 0.843 cm?

(b) If the cube in part (a) is copper and has a mass of 5.34 g, what is the density of copper to two decimal places?

Answer: (a) 0.599 cm3; (b) 8.91 g/cm3

Link to Learning

To learn more about the relationship between mass, volume, and density, use this **interactive simulator (http://openstaxcollege.org/l/16phetmasvolden)** to explore the density of different materials, like wood, ice, brick, and aluminum.

Example 1.2

Using Displacement of Water to Determine Density: This **PhET simulation (http://openstaxcollege.org/l/16phetmasvolden)** illustrates another way to determine density, using displacement of water. Determine the density of the red and yellow blocks.

Solution:

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

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

$$density = \frac{mass}{volume} = \frac{5.00\ Kg}{1.25\ L} = 4.00\frac{Kg}{L}$$

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

$$density = \frac{mass}{volume} = \frac{5.00\ Kg}{10.00\ L} = 0.500\frac{Kg}{L}$$

Check Your Learning

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

Answer: 2.00 kg/L

1.5 Measurement Uncertainty, Accuracy, and Precision

By the end of this section, you will be able to:

- Define accuracy and precision
- Distinguish exact and uncertain numbers
- Correctly represent uncertainty in quantities using significant figures
- Apply proper rounding rules to computed quantities

Counting is the only type of measurement that is free from uncertainty, provided the number of objects being counted does not change while the counting process is underway. The result of such a counting measurement is an example of an **exact number**. If we count eggs in a carton, we know *exactly* how many eggs the carton contains. The numbers of defined quantities are also exact. By definition, 1 foot is exactly 12 inches, 1 inch

is exactly 2.54 centimeters, and 1 gram is exactly 0.001 kilogram. Quantities derived from measurements other than counting, however, are uncertain to varying extents due to practical limitations of the measurement process used.

Significant Figures in Measurement

The numbers of measured quantities, unlike defined or directly counted quantities, are not exact. To measure the volume of liquid in a graduated cylinder, you should make a reading at the bottom of the meniscus, the lowest point on the curved surface of the liquid.

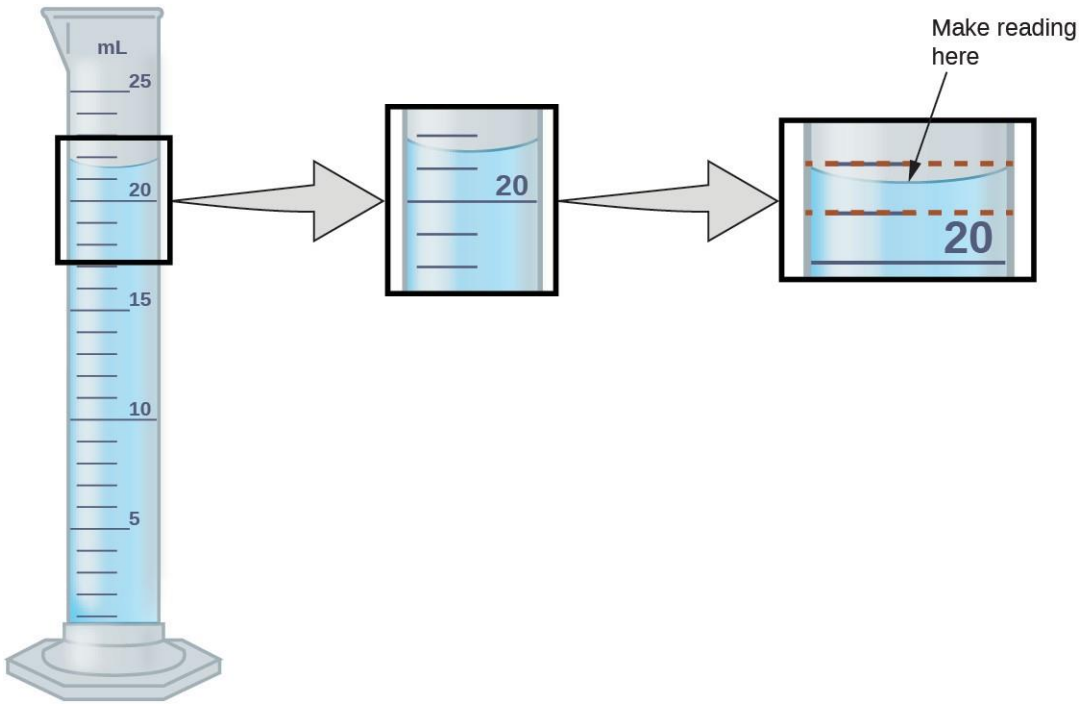


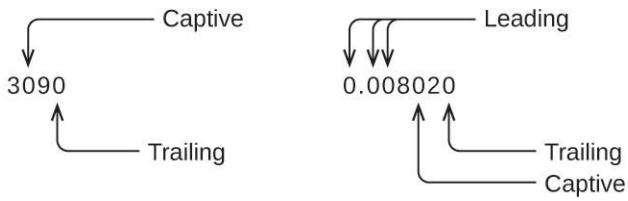
Figure 1.26 To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

Refer to the illustration in **Figure 1.26**. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable estimate of the liquid's volume would be 21.6 mL. In the number 21.6, then, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain. In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in

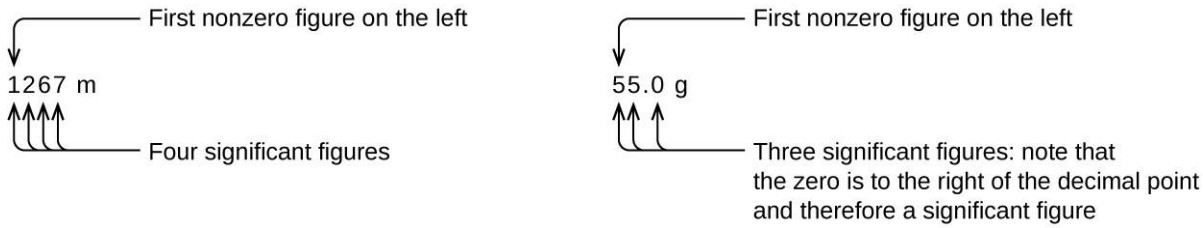
this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

This concept holds true for all measurements, even if you do not actively make an estimate. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter weighs *about* 6.72 grams, with a nominal uncertainty in the measurement of ± 0.01 gram. If we weigh the quarter on a more sensitive balance, we may find that its mass is 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram. Every measurement has some **uncertainty**, which depends on the device used (and the user's ability). All of the digits in a measurement, including the uncertain last digit, are called **significant figures** or **significant digits**. Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows "120," then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

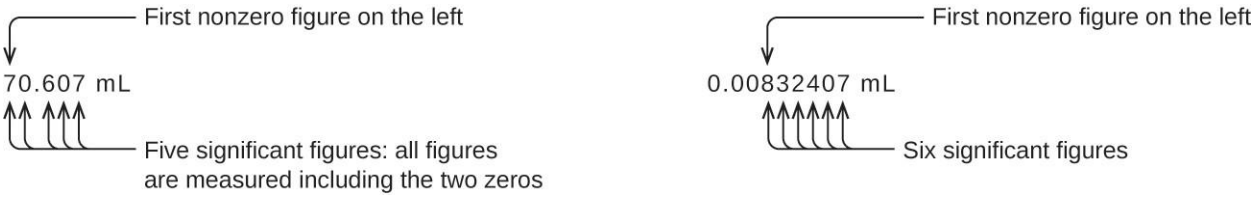
Whenever you make a measurement properly, all the digits in the result are significant. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Well, for starters, all nonzero digits are significant, and it is only zeros that require some thought. We will use the terms "leading," "trailing," and "captive" for the zeros and will consider how to deal with them.



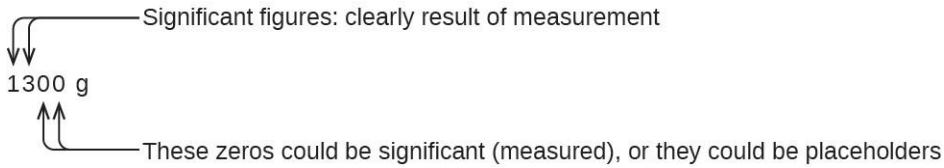
Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point.



Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.



The leading zeros in this example are not significant. We could use exponential notation (as described in Appendix and express the number as 8.32407×10^{-3}; then the number 8.32407 contains all of the significant figures, and 10^{-3} locates the decimal point. The number of significant figures is uncertain in a number that ends with a zero to the left of the decimal point location. The zeros in the measurement 1,300 grams could be significant or they could simply indicate where the decimal point is located. The ambiguity can be resolved with the use of exponential notation: 1.3×10^3 (two significant figures), 1.30×10^3 (three significant figures, if the tens place was measured), or 1.300×10^3 (four significant figures, if the ones place was also measured). In cases where only the decimal-formatted number is



available, it is prudent to assume that all trailing zeros are not significant.

When determining significant figures, be sure to pay attention to reported values and think about the measurement and significant figures in terms of what is reasonable or likely when evaluating whether the value makes sense. For example, the official January 2014 census reported the resident population of the US as 317,297,725. Do you think the US population was correctly determined to the reported nine significant figures, that is, to the exact number of people? People are constantly being born, dying, or moving into or out of the country, and assumptions are made to account for the large number of people who are not actually counted. Because of these uncertainties, it might be more reasonable to expect that we know the population to within perhaps a million or so, in which case the population should be reported as 3.17×10^8 people.

Significant Figures in Calculations

A second important principle of uncertainty is that results calculated from a measurement are at least as uncertain as the measurement itself. We must take the uncertainty in our measurements into account to avoid misrepresenting the uncertainty in calculated results. One way to do this is to report the result of a calculation with the correct number of significant figures, which is determined by the following three rules for **rounding** numbers:

1. When we add or subtract numbers, we should round the result to the same number of decimal places as the number with the least number of decimal places (the least precise value in terms of addition and subtraction).
2. When we multiply or divide numbers, we should round the result to the same number of digits as the number with the least number of significant figures (the least precise value in terms of multiplication and division).
3. If the digit to be dropped (the one immediately to the right of the digit to be retained) is less than 5, we "round down" and leave the retained digit unchanged; if it is more than 5, we "round up" and increase the retained digit by 1; if the dropped digit *is* 5, we round up or down, whichever yields an even value for the retained digit. (The last part of this rule may strike you

as a bit odd, but it's based on reliable statistics and is aimed at avoiding any bias when dropping the digit "5," since it is equally close to both possible values of the retained digit.)

The following examples illustrate the application of this rule in rounding a few different numbers to three significant figures:

- 0.028675 rounds "up" to 0.0287 (the dropped digit, 7, is greater than 5)
- 18.3384 rounds "down" to 18.3 (the dropped digit, 3, is less than 5)
- 6.8752 rounds "up" to 6.88 (the dropped digit is 5, and the retained digit is even)
- 92.85 rounds "down" to 92.8 (the dropped digit is 5 and the retained digit is even) Let's work through these rules with a few example.

92.85 rounds "down" to 92.8 (the dropped digit is 5, and the retained digit is
even) Let's work through these rules with a few examples.

Example 1.3

Rounding Numbers: Round the following to the indicated number of significant figures:

(a) 31.57 (to two significant figures)
(b) 8.1649 (to three significant figures)
(c) 0.051065 (to four significant figures)
(d) 0.90275 (to four significant figures)

Solution:

(a) 31.57 rounds "up" to 32 (the dropped digit is 5, and the retained digit is even)
(b) 8.1649 rounds "down" to 8.16 (the dropped digit, 4, is less than 5)
(c) 0.051065 rounds "down" to 0.05106 (the dropped digit is 5, and the retained digit is even)
(d) 0.90275 rounds "up" to 0.9028 (the dropped digit is 5, and the retained digit is even)

Check Your Learning

Round the following to the indicated number of significant figures:

(a) 0.424 (to two significant figures)
(b) 0.0038661 (to three significant figures)
(c) 421.25 (to four significant figures)
(d) 28,683.5 (to five significant figures)

Answer: (a) 0.42; (b) 0.00387; (c) 421.2; (d) 28,684

Example 1.4

Addition and Subtraction with Significant Figures: Rule: When we add or subtract numbers, we should round the result to the same number of decimal places as the number with the least number of decimal places (i.e., the least precise value in terms of addition and subtraction).

(a) Add 1.0023 g and 4.383 g.

(b) Subtract 421.23 g from 486 g.

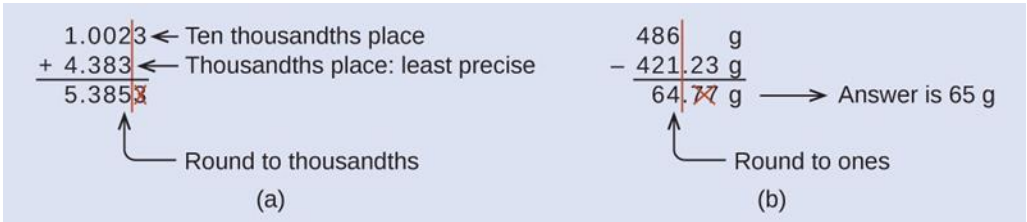
Solution:

$(a)\ 1.0023g + 4.383\ g = 5.3853g$

Answer: 5.385 g (round to the thousandths place; three decimal places)

$(b)\ 486\ g - 421.23\ g = 64.77g$

Answer: 65 g (round to the ones place; no decimal places)



Check Your Learning

(a) Add 2.334 mL and 0.31 mL.

(b) Subtract 55.8752 m from 56.533 m.

Answer: (a) 2.64 mL; (b) 0.658 m

Example 1.5

Multiplication and Division with Significant Figures: Rule: When we multiply or divide numbers, we should round the result to the same number of digits as the number with the least number of significant figures (the least precise value in terms of multiplication and division).

(a) Multiply 0.6238 cm by 6.6 cm.

(b) Divide 421.23 g by 486 mL.

Solution:

$(a) 0.6238cm * 6.6\ cm = 4.11708\ cm^2 \rightarrow result\ is\ 4.1\ cm^2 (round\ to\ two\ significant\ figures)$

four significant figures $*$ two significant figures $\rightarrow$ two significant figures answer

$$\frac{421.23\ g}{486\ mL} = 0.86728 \ldots\ldots \frac{g}{ml} \rightarrow resul\ is\ 0.867 \frac{g}{mL} (\ round\ to\ three\ significant\ figures)$$

$$(b) \frac{significant\ figures}{three\ significant\ figures} \rightarrow three\ significant\ figures\ answer$$

Check your learning

(a) Multiply 2.334 cm and 0.320 cm

(b) Divide 55.8752 m by 56.53 s

Answer: (a) 0.747 cm^2 (b) 0.9884 m/s

In the midst of all these technicalities, it is important to keep in mind the reason why we use significant

figures and rounding rules—to correctly represent the certainty of the values we report and to ensure that a calculated result is not represented as being more certain than the least certain value used in the calculation.

Example 1.6

Calculation with Significant Figures: One common bathtub is 13.44 dm long, 5.920 dm wide, and 2.54 dm deep. Assume that the tub is rectangular and calculate its approximate volume in liters.

Solution:

$$V = l * w * d$$
$$= 13.44\ dm * 5.920\ dm * 2.54\ dm$$
$$= 202.09459 \ldots\ dm^3\ (value\ from\ calculator)$$
$$= 202\ dm^3 or\ 202\ L\ (answer\ rounded\ to\ three\ significant\ figures)$$

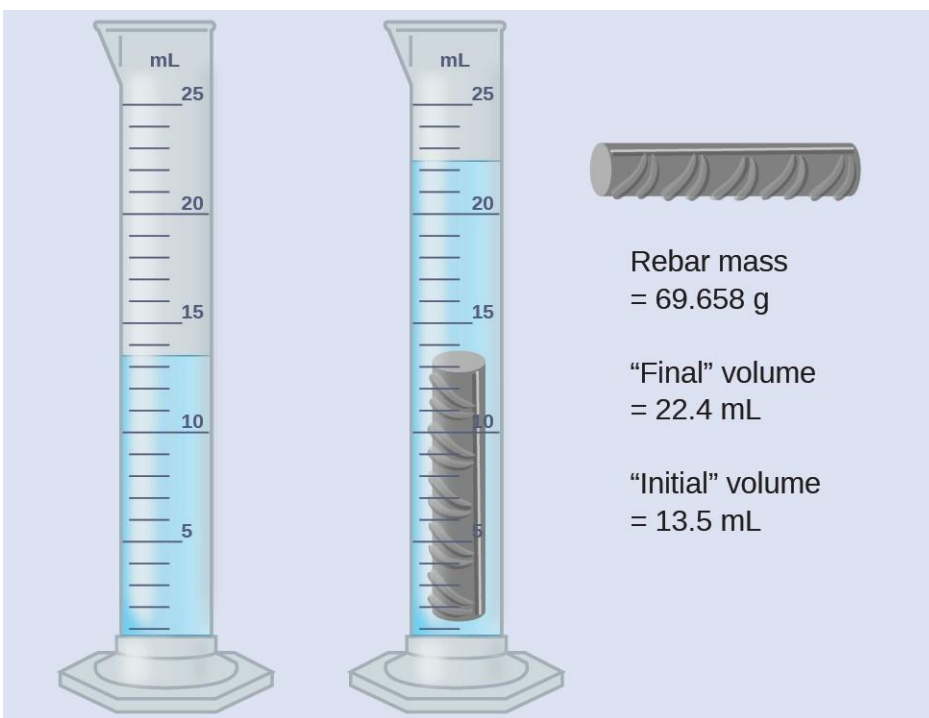
Check Your Learning

What is the density of a liquid with a mass of 31.1415 g and a volume of 30.13 cm3?

Answer: 1.034 g/mL

Example 1.7

Experimental Determination of Density Using Water Displacement: A piece of rebar is weighed and then submerged in a graduated cylinder partially filled with water, with results as shown.



(a) Use these values to determine the density of this piece of rebar.

(b) Rebar is mostly iron. Does your result in (a) support this statement? How?

Solution:

The volume of the piece of rebar is equal to the volume of the water displaced:

$$\text{volume} = 22.4\ \text{mL} - 13.5\ \text{mL} = 8.9\ \text{mL} = 8.9\ \text{cm}$$

Rounded to the nearest 0.1mL per the rule for addition (rounded to the nearest 0.1 mL per the rule for addition and subtraction). The density is the mass to volume ratio;

$$density = \frac{mass}{volume} = \frac{69.658\ g}{8.9cm^3} = 7.8\frac{g}{cm^3}$$

(rounded to two significant figures, per the rule for multiplication and division)

From **Table 1.4**, the density of iron is 7.9 g/cm3, very close to that of rebar, which lends some support to the fact that rebar is mostly iron.

Check Your Learning

An irregularly shaped piece of a shiny yellowish material is weighed and then submerged in a graduated cylinder, with results as shown.

(a) Use these values to determine the density of this material.

(b) Do you have any reasonable guesses as to the identity of this material? Explain your reasoning.

Answer: (a) 19 g/cm3; (b) It is likely gold; the right appearance for gold and very close to the density given for gold in **Table 1.4**

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to know both the **precision** and the **accuracy** of their results. Measurements are said to be precise if they yield very similar results when repeated in the same manner. A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value. These characterizations can be extended to other contexts, such as the results of an archery competition (**Figure 1.27**).

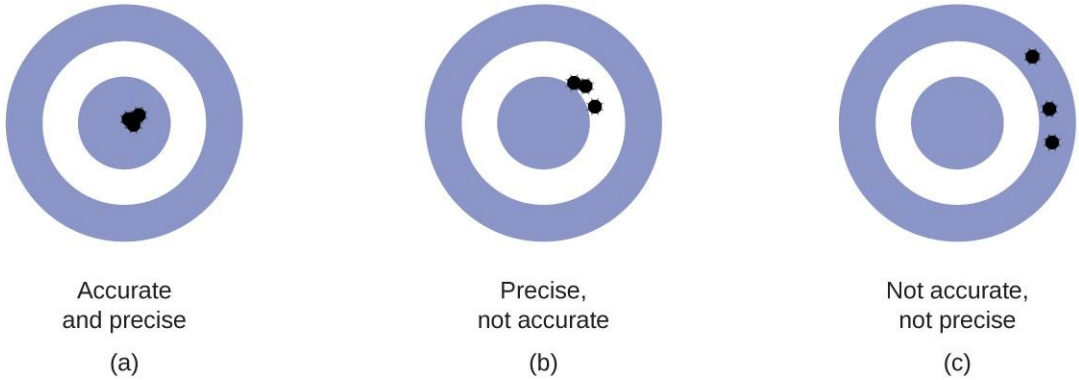


Figure 1.27 (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Suppose a quality control chemist at a pharmaceutical company is tasked with checking the accuracy and precision of three different machines that are meant to dispense 10 ounces (296 mL) of cough syrup into storage bottles. She proceeds to use each machine to fill five bottles and then carefully determines the actual volume dispensed, obtaining the results tabulated in **Table 1.5**.

Volume (mL) of Cough Medicine Delivered by 10-oz (296 mL) Dispensers

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

Table 1.5

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

1.6 Mathematical Treatment of Measurement Results

By the end of this section, you will be able to:

- Explain the dimensional analysis (factor label) approach to mathematical calculations involving quantities
- Use dimensional analysis to carry out unit conversions for a given property and computations involving two or more properties

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

$$Speed = \frac{distance}{time}$$

An Olympic-quality sprinter can run 100 m in approximately 10 s, corresponding to an average speed of

$$\frac{100\ m}{10s} = 10\frac{m}{s}$$

Note that this simple arithmetic involves dividing the numbers of each measured quantity to yield the number of the computed quantity (100/10 = 10) *and likewise* dividing the units of each measured quantity to yield the unit of the computed quantity (m/s = m/s). Now, consider using this same relation to predict the time required for a person running at this speed to travel a distance of 25 m. The same relation between the three properties is used, but in this case, the two quantities provided are a speed (10 m/s) and a distance (25 m). To yield the sought property, time, the equation must be rearranged appropriatel

$$time = \frac{distance}{speed}$$

The time can then be computed as:

$$\frac{25}{10\frac{m}{s}} = 25s$$

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

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

Conversion Factors and Dimensional Analysis

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

$$\frac{2.54\ cm}{1\ in}\ (2.54cm = 1\ in.)\ or\ 2.54\frac{cm}{in}$$

Several other commonly used conversion factors are given in **Table 1.6**.

Common Conversion Factors

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

Table 1.6

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

$$34\ in * \frac{2.54\ cm}{1\ in} = 86\ cm$$

Since this simple arithmetic involves *quantities*, the premise of dimensional analysis requires that we multiply both *numbers and units*. The numbers of these two quantities are multiplied to yield the number of the product quantity, 86, whereas the units are multiplied to yield $\frac{in*cm}{in}$. Just as for numbers, a ratio of identical units is also numerically equal to one, $\frac{in}{in} = 1$, and the unit product thus simplifies to *cm*. (When identical units divide to yield a factor of 1, they are said to "cancel.") Using dimensional analysis, we can determine that a unit conversion factor has been set up correctly by checking to confirm that the original unit will cancel, and the result will contain the sought (converted) unit.

Example 1.8

Using a Unit Conversion Factor: The mass of a competition frisbee is 125 g. Convert its mass to ounces using the unit conversion factor derived from the relationship 1 oz = 28.349 g (**Table 1.6**).

Solution:

If we have the conversion factor, we can determine the mass in kilograms using an equation similar the one used for converting length from inches to centimeters.

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

We write the unit conversion factor in its two forms:

$$\frac{1\ oz}{28.349\ g}\ and\ \frac{28.349\ g}{1\ oz}$$

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

$$X_{oz} = 125g * \frac{1oz}{28.349\ g}$$

$$= \frac{125\ oz}{28.349}$$

$$= 4.41\ oz\ (\ three\ significant\ figures)$$

Check Your Learning

Convert a volume of 9.345 qt to liters.

Answer: 8.844 L

Assessment 1.1:

Historically, vitamin C was used for preventing and treating scurvy. It supports one's immune system and helps the body use the iron one gets from food. One's body also uses it to make collagen, a springy type of connective tissue that makes up parts of the body and helps heal wounds. And it's an antioxidant that helps protect cells from damage. (From *Web* MD). A tablet of Vitamin C weighs 500 *mg*. Assume 3 sf.

(a). Calculate the mass of the tablet in *g*.

(b). From a 2.00 lb batch of powdered Vitamin C, how many 500 mg tablets can be made?

Useful information:

1 *lb* = 453.6 *g*

Answer:

(a). Strategy: Use the conversion factor, given below, for *mg* to *g*. Conversion factor:

$$\frac{1\,g}{1000\,mg}$$

$$Mass\ in\ g = 500\ mg \text{ x } \frac{1\ g}{1000\ mg} = 0.500\ g$$

(b). Strategy: Convert *lb* to *g* and then calculate the number of tablets.Conversion factor:

$$\frac{453.6\ g}{1\ lb}$$

$$Number\ of\ tablets = 2.00\ lb \text{ x } \frac{453.6\ g}{1\ lb} \text{ x } \frac{1\ tablet}{0.500\ g} = 1{,}814\ tablet$$

In the scientific notation; the answer is1.81 x 103 tablets (3 sf)

Assessment 1.2:

A label from a carton of orange juice is given above. It states that one serving, 12 fluid ounces (355 mL, the volume of a can of soda), provides 108 mg Vitamin C, which is 120% of the recommended daily allowance (RDA), based on a 2000 calorie diet.

(a). How many dietary calories are present in one serving?

(b). What is the the recommended daily allowance (RDA) of Vitamin C? Useful information:

1 *g* carbohydrate contains = 4.0 *dietary calories.*

Answer:

(a). Strategy: Use the conversion factor, given below, for *g* to *dietary calories*. Conversion factor:

$$\frac{4.0\ dietary\ calorie}{1\ g\ carbohydrate}$$

$$Dietary\ calories = 33\ g\ carbohydrate\ \text{x}\ \frac{4.0\ dietary\ calorie}{1\ g\ carbohydrate} = 132\ dietary\ calorie$$

The answer in scientific notation is 1.3 x 10^2 *g* (2 sf).

(b). Strategy: Assume that the RDA of Vitamin C is *x mg*, and set up an equation for calculating a percentage.

$$\frac{108\ mg}{x\ mg}\text{x}100 = 120$$

$$\text{Thus,}\quad x = \frac{108\text{x}100}{120} = 90$$

The RDA of Vitamin C is 90 mg.

Assessment 1.3:

During a visit to a doctor, one's height and mass are important measurements. In SI units, the height and mass of an individual are 1.82 *m* and 80.0 *kg*, respectively. Convert these measurements that we commonly use, (a). *feet* and *inches* (for height) and (b). *lb* (for mass).

Useful information:

1 *inch* = 2.54 *cm*

1 *lb* = 453.6 *g*

Answer:

(a). Strategy: For height, convert *m* to *cm*, then cm to inches, then *inches* to *feet*.

Conversion factors:

$$\frac{100\ cm}{1\ m} \qquad \frac{1\ inch}{2.54\ cm} \qquad \frac{12\ inch}{1\ ft}$$

Step 1:

$$Height\ in\ inch = 1.82\ m\ \text{x}\ \frac{100\ cm}{1\ m}\ \text{x}\ \frac{1\ inch}{2.54\ cm} = 71.65\ inch$$

Step 2:

$$Height\ in\ ft = 71.65\,inch \text{ x } \frac{1\ ft}{12\ inch} = 5.97\ ft \ \ (3\ sf)$$

Recognizing that 5 *ft* = 60 *inch (exactly)*, the answer can also be expressed as 5 *ft* 11.65 *inch*, or 5 *ft* 11.7 *inch (3 sf)*

(b). Strategy:

The conversion factors are:

$$\frac{453.6\ g}{1\,lb} \qquad \frac{1\,kg}{1000\ g}$$

$$mass\ in\ lb = 80.0\ kg \text{ x } \frac{1000\ g}{1\,kg} \text{ x } \frac{1\,lb}{453.6\ g} = 176\ lb \ \ (3\text{ sf})$$

Assessment 1.4:

Cholesterol is a waxy, odorless substance made by the liver that is an essential part of cell walls and nerves. Cholesterol plays an important role in body functions such as digestion and hormone production. In addition to being produced by the body, cholesterol comes from animal foods that we eat. Excessive cholesterol in the blood causes an increase in particles called LDL (or "bad" cholesterol), which increases the buildup of plaque in the artery walls and leads to atherosclerosis. (*Web* MD)

An individual has a healthy cholesterol level, 185 *mg*/*dL*. She donates 1.00 *pint* of blood. Calculate the mass of cholesterol in this sample.

Useful information:

1 *pint* = 473.18 *mL* 1 *dL* = 1/10 *L*

Answer:

Strategy: Convert *pint* to *mL*, and then use the conversion factor for *dL*.

Conversion factors:

$$\frac{10\ \mathrm{d}L}{1\,L} \qquad \frac{1L}{1000\ \mathrm{m}L} \qquad \frac{473.18\ mL}{1\ p\mathrm{int}}$$

$$mass\ of\ cholesterol = 1.00\ p\mathrm{int} \text{ x } \frac{473.18\ mL}{1\ p\mathrm{int}} \text{ x } \frac{1\,L}{1000\ mL} \text{ x } \frac{10\ \mathrm{d}L}{1\,L} \text{ x } \frac{185\ mg\ cholesterol}{1\,dL} = 875\ mg$$

Assessment 1.5:

An estimated 50 million Americans use acetaminophen each week to treat conditions such as pain, fever and aches and pains associated with cold and flu symptoms. To help encourage the safe use of acetaminophen, the makers of

TYLENOL® have lowered the maximum daily dose for single-ingredient Extra Strength TYLENOL® (acetaminophen) products sold in the U.S. to 3,000 mg per day for an adult. The dosage is dependent on the age and weight for children. (Adapted from Tylenol's we site).

An adult student uses Nyquil whose label states that a dose cup (30 mL) contains 650 mg acetaminophen. How many dose cups of this brand of Nyquil can an adult take safely in one day? Assume 2 sf.

Caution: This problem is not meant to be used as medical advice. Medications differ in amounts of acetaminophen. Please consult a medical practitioner prior to using any medicine.

Answer:

Strategy: Calculate the number of dose cups that correspond to 3,000 mg acetaminophen.

Conversion factor:

$$\frac{1\,dose\,cup}{650\,mg\,acetomenophen}$$

$$dose\,cup = 3{,}000\,mg\,acetomenophen \text{ x } \frac{1\,dose\,cup}{650\,mg\,acetomenophen} = 4.6\,dose\,cup$$

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

Example 1.9

Computing Quantities from Measurement Results and Known Mathematical Relations: What is the density of common antifreeze in units of g/mL? A 4.00-qt sample of the antifreeze weighs 9.26 lb.

Solution:

Since $density = \frac{mass}{volume}$, we need to divide the mass in grams by the volume in milliliters. In general: the number of units of B = the number of units of A × unit conversion factor. The necessary conversion factors are given in **Table 1.6**: 1 lb = 453.59 g; 1 L = 1.0567 qt; 1 L = 1,000 mL. We can convert mass from pounds to grams in one step

:$9.26lb * \frac{453.59\,g}{1\,lb} = 4.20 * 10^3 g$

We need to use two steps to convert volume from quarts to milliliters.

Step 1. *Convert quarts to liters.*

$$4.00\,qt * \frac{1L}{1.0567\,qt} = 3.78\,L$$

Step 2. *Convert liters to milliliters.*

$$3.78\,L * \frac{1000\,mL}{1\,L} = 3.78 * 10^3\,mL$$

Then,

$$density = \frac{4.20 * 10^3 g}{3.78 * 10^3\ mL} = 1.11 \frac{g}{mL}$$

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

$$\frac{9.26\ lb}{4.00\ qt} * \frac{453.59 g}{1\ lb} * \frac{1.0567\ qt}{1\ L} * \frac{1 L}{1000\ ml} = \frac{1.11 g}{mL}$$

Check Your Learning

What is the volume in liters of 1.000 oz, given that 1 L = 1.0567 qt and 1 qt = 32 oz (exactly)?

Answer: 2.956×10^{-2} L

Example 1.10

Computing Quantities from Measurement Results and Known Mathematical Relations: While being driven from Philadelphia to Atlanta, a distance of about 1250 km, a 2014 Lamborghini Aventador Roadster uses 213 L gasoline.

(a) What (average) fuel economy, in miles per gallon, did the Roadster get during this trip?

(b) If gasoline costs $3.80 per gallon, what was the fuel cost for this trip?

Solution:

(a) We first convert distance from kilometers to miles:

$$1250\ km * \frac{0.62137\ mi}{1\ Km} = 777\ mi$$

And then convert volume from liters to gallons:

$$213\ L * \frac{1.0567\ qt}{1L} * \frac{1\ gal}{4\ qt} = 56.3\ gal$$

Then,

$$average\ mileage = \frac{777\ mi}{56.3\ gal} = 13.8 \frac{miles}{gallon} = 13.8\ mpg$$

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

1250 $\frac{Km}{213\ L} * \frac{0.62137\ mi}{1\ Km} * \frac{1\ L}{1.0567\ qt} * \frac{4\ qt}{1\ gal} = 138. mpg$

(c) Using the previously calculated volume in gallons, we find: $56.3\ gal * \frac{\$\ 3.80}{1\ gal} = \$\ 214$

Check Your Learning

A Toyota Prius Hybrid uses 59.7 L gasoline to drive from San Francisco to Seattle, a distance of 1300 km (two significant digits).

(a) What (average) fuel economy, in miles per gallon, did the Prius get during this trip?

(b) If gasoline costs $3.90 per gallon, what was the fuel cost for this trip?

Answer: (a) 51 mpg; (b) $62

Assessment 1.6:

Saline, a solution of salt (sodium chloride) in water is used to clean wounds and remove contact lenses. It is prepared by dissolving sodium chloride in water, 0.90% mass/volume. Calculate the mass of sodium chloride needed to prepare 1.00 *gallon* of saline.

Useful information:

$$1 gallon = 3.7854\ L$$

Answer:

Strategy: Saline contains sodium chloride, 0.90% mass/volume; thus, 0.90 g sodium chloride is dissolved in 100.0 *mL* of water. Convert *mL* to *gallon*.

Conversion factors:

$$\frac{1000\ mL}{1\ L} \qquad \frac{3.785\ L}{1\ gallon}$$

$$Mass\ of\ sodium\ chloride = \frac{0.90\ g\ sodium\ chloride}{100.0\ mL} \text{x} \frac{1000\ mL}{1\ L} \text{x} \frac{3.7854\ \text{L}}{1\ gallon} \text{x} 1.00\ gallon = 34\ g\ (2sf)$$

Assessment 1.7:

A label for Fish Oil supplements reads, in part: "Supportive but not conclusive research shows that the consumption of Omega-3 fatty acids may reduce the risk of coronary heart disease" One capsule contains 1,200 mg of Omega-3 fatty acids. An individual takes one such capsule. How much salmon (in *ounce*) would the individual have to consume to get an equivalent amount of Omega-3 fatty acids, given that it is present in the amount of 1.9 g per 100.0 g of salmon?

Useful information:

1 *ounce* = 28.35 g

Answer:

Strategy: Convert Omega-3 fatty acids *from mg* to *g*, then calculate mass of salmon in *g*, and convert to *ounce*. Conversion factors:

$$\frac{1\ g}{1000\ mg} \qquad \frac{1\ ounce}{28.35\ g}$$

$$Mass\ of\ salmon = 1{,}200\ mg\ Omega-3 \text{ x } \frac{1\ g\ Omega-3}{1{,}000\ mg\ Omega-3} \text{ x } \frac{100.0\ g\ Salmon}{1.9\ g\ Omega-3} \text{ x } \frac{1\ ounce\ Salmon}{28.35\ g\ Salmon}$$

$$= 2.2\ ounce\ salmon\ (2sf)$$

Assessment 1.8:

Dimensional analysis is important in everyday life, as illustrated below. For ear infections for children two months to six years of age, the recommended dose is 85 *mg* of amoxicillin per 1.00 *kg* of body mass per day. How much amoxicillin should a five year old child weighing 36.0 *lbs* (having been diagnosed with an ear infection) take per day?

Answer:

Strategy: Convert the body mass from *lb* to *g;* then *g* to *kg*; then calculate the mass of amoxicillin. The conversion factors are:

$$\frac{453.6\ g}{1\ lb} \qquad \frac{1\ kg}{1000\ g} \qquad \frac{85\ mg\ amoxicillin}{1\ kg\ body\ mass}$$

The three conversion factors can be used sequentially:

$$Mass\ of\ amoxicillin = 36.0\ lb\ body\ mass \text{ x } \frac{453.6\ g}{1\ lb} \text{ x } \frac{1\ kg}{1000\ g} \text{ x } \frac{85\ mg\ amoxicillin}{1\ kg\ body\ mass} = 1388\ mg\ amoxicillin$$

The answer is reported to 2 sf, 1.4 x 10^3 mg.

Conversion of Temperature Units

We use the word temperature to refer to the hotness or coldness of a substance. One way we measure a change in temperature is to use the fact that most substances expand when their temperature increases and contract when their temperature decreases. The mercury or alcohol in a common glass thermometer changes its volume as the temperature changes. Because the volume of the liquid changes more than the volume of the glass, we can see the liquid expand when it gets warmer and contract when it gets cooler.

To mark a scale on a thermometer, we need a set of reference values: Two of the most commonly used are the freezing and boiling temperatures of water at a specified atmospheric pressure. On the Celsius scale, 0 °C is defined as the freezing temperature of water and 100 °C as the boiling temperature of water. The space between the two temperatures is divided into 100 equal intervals, which we call degrees. On the **Fahrenheit** scale, the freezing point of water is defined as 32 °F and the boiling temperature as 212 °F. The space between these two points on a Fahrenheit thermometer is divided into 180 equal parts (degrees).

Defining the Celsius and Fahrenheit temperature scales as described in the previous paragraph results in a slightly more complex relationship between temperature values on these two scales than for different units of measure for other properties. Most measurement units for a given property are directly proportional to one another (y = mx). Using familiar length units as one example:

$$length\ in\ feet = \left(\frac{1\ ft}{12\ in}\right) * length\ in\ inches$$

where y = length in feet, x = length in inches, and the proportionality constant, m, is the conversion factor. The Celsius

and Fahrenheit temperature scales, however, do not share a common zero point, and so the relationship between these two scales is a linear one rather than a proportional one (y = mx + b). Consequently, converting a temperature from one of these scales into the other requires more than simple multiplication by a conversion factor, m, it also must take into account differences in the scales' zero points (b). The linear equation relating Celsius and Fahrenheit temperatures is easily derived from the two temperatures used to define each scale. Representing the Celsius temperature as *x* and the Fahrenheit temperature as *y*, the slope, *m*, is computed to be:

$$m = \frac{\Delta y}{\Delta x} = \frac{212°\text{F} - 32°\text{F}}{100°\text{C} - 0°\text{C}} = \frac{180°\text{F}}{100°\text{C}} = \frac{9\ °\text{F}}{5\ °\text{C}}$$

The y-intercept of the equation, *b*, is then calculated using either of the equivalent temperature pairs, (100 °C, 212 °F) or (0 °C, 32 °F), as:

$$b = y - mx = 32°\text{F}\ \frac{9°\text{F}}{5°\text{C}} * 0°\text{C} = 32°\text{F}$$

The equation relating the temperature scales is then:

$$T_{°\text{F}} = \left(\frac{9°\text{F}}{5°\text{C}} * T_{°\text{C}}\right) + 32°\text{C}$$

An abbreviated form of this equation that omits the measurement units is: $T_{°\text{F}} = \left(\frac{9}{5} * T_{°\text{F}}\right)$ Rearrangement of this equation yields the form useful for converting from Fahrenheit to Celsius:

$$T_{°\text{C}} = \left(\frac{5}{9} * T_{°\text{F}}\right) - 32°\text{C}$$

As mentioned earlier in this chapter, the SI unit of temperature is the kelvin (K). Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale in which 0 (zero) K corresponds to the lowest temperature that can theoretically be achieved. The early 19th-century discovery of the relationship between a gas's volume and temperature suggested that the volume of a gas would be zero at −273.15 °C. In 1848, British physicist William Thompson, who later adopted the title of Lord Kelvin, proposed an absolute temperature scale based on this concept (further treatment of this topic is provided in this text's chapter on gases).

The freezing temperature of water on this scale is 273.15 K and its boiling temperature 373.15 K. Notice the numerical difference in these two reference temperatures is 100, the same as for the Celsius scale, and so the linear relation between these two temperature scales will exhibit a slope of $1\ \frac{K}{C}$. Following the same approach, the equations for converting between the kelvin and Celsius temperature scales are derived to be:

$$T_{\text{K}} = T_{°\text{C}} + 273.15$$

$$T_{°\text{C}} = T_{\text{K}} - 273.15$$

The 273.15 in these equations has been determined experimentally, so it is not exact. **Figure 1.28** shows the relationship among the three temperature scales. Recall that we do not use the degree sign with temperatures on the Kelvin scale.

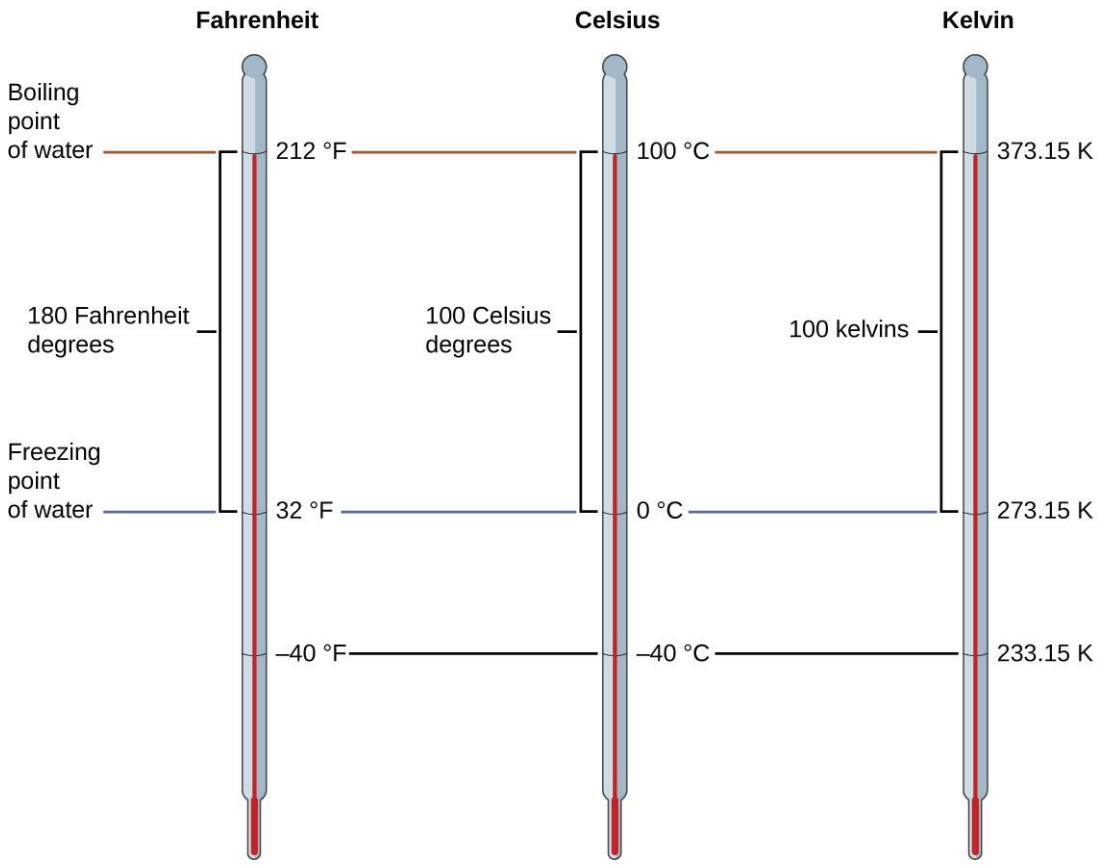


Figure 1.28 The Fahrenheit, Celsius, and kelvin temperature scales are compared.

Although the kelvin (absolute) temperature scale is the official SI temperature scale, Celsius is commonly used in many scientific contexts and is the scale of choice for nonscience contexts in almost all areas of the world. Very few countries (the U.S. and its territories, the Bahamas, Belize, Cayman Islands, and Palau) still use Fahrenheit for weather, medicine, and cooking.

Example 1.11

Conversion from Celsius: Normal body temperature has been commonly accepted as 37.0 °C (although it varies depending on time of day and method of measurement, as well as among individuals). What is this temperature on the kelvin scale and on the Fahrenheit scale?

Solution:

$$k = ℃ + 273.15 = 37.0 + 273.2 = 310.2\ K$$

$$°\text{F} = \frac{9}{5}℃ + 32.0 = \left(\frac{9}{5} * 37.0\right) + 32.0 = 66.6 + 32.0 = 98.6°\text{F}$$

Check Your Learning

Convert 80.92 °C to K and °F.

Answer: 354.07 K, 177.7 °F

Example 1.12

Conversion from Fahrenheit: Baking a ready-made pizza calls for an oven temperature of 450 °F. If you are in Europe, and your oven thermometer uses the Celsius scale, what is the setting? What is the kelvin temperature?

Solution:

$$°\mathrm{C} = \frac{5}{9}\,(°\mathrm{F} - 32) = \frac{5}{9}\,(450 - 32) = \frac{5}{9} * 418 = 232°\mathrm{C} \;\rightarrow set\ oven\ to\ 230°\mathrm{C}\ (\,two\ significant\ figures)$$

$$k = \ °\mathrm{C} + 273.15 = 230 + 273 = 503k \rightarrow 5.0 * 10^2 K\ (two\ significant\ figures)$$

Check your learning

Convert 50 °F to °C and K

Answer: 10 °C, 280 K

Assessment 1.9:

Guidelines for storing eggs safely for 3-5 weeks are: "Store at 4.0ºC; do not freeze". Calculate this temperature in ºF.

Answer:

Strategy: Convert the temperature from ºC to ºF.

$$^{o}\mathrm{F} = \frac{9}{5}\,^{o}\mathrm{C} + 32 = \frac{9}{5}\ \mathrm{x}\ 4.0 + 32 = 39^{o}\mathrm{F}$$

Assessment 1.10 (Summary Problem):

The transportation of gasoline is a way of life. So are oil spills, which are very harmful for the marine ecosystem. Interestingly, a measurement of the thickness of oil can be used to determine the volume of oil spilled. As to what this entails for the entity that made the spill is beyond the scope of this problem!

A tanker spills 25 gallons of oil over an area of 1.0 square miles. Determine the thickness of the oil film in nm.

Useful conversions:

$$1\ gallon = 3.7854\ L$$

$$1L = 1000\ mL$$

$$1\ mL = 1\ cm^3$$

$$1\ mile = 1.6093\ km$$

Answer:

Strategy: Volume = Area x Thickness. Thus, calculate the volume and area separately. Since we want the final

answer in *nm*, it is convenient to work in units of *m*.

The conversion factors are:

$$\frac{3.7854\,L}{1\,gallon} \qquad \frac{1000\,mL}{1\,L} \qquad \frac{1\,cm^3}{1\,\text{m}L} \qquad \frac{1\,m}{100\,cm} \qquad \frac{1\,\text{n}m}{10^{-9}\,m}$$

$$Volume = 25\,gallon \text{ x } \frac{3.7854\,L}{1\,gallon} \text{ x } \frac{1000\,mL}{1\,L} \text{ x } \frac{1\,cm^3}{1\,mL} \text{ x } \frac{1\,m^3}{(100\,cm)^3} = 0.094635\,m^3$$

We will keep extra digits and round at the very end.

$$Area = 1.0\,mile^2 \text{ x } \left(\frac{1.6093\,km}{1\,mile}\right)^2 \text{ x} \left(\frac{1000\,m}{1\,km}\right)^2 = 2.58985 \text{ x } 10^6\,m^2$$

$$Thickness = \left(\frac{Volume}{Area}\right) = \left(\frac{0.094635\,m^3}{2.58985 \text{ x } 10^6\,m^2}\right) = 3.7 \text{ x } 10^{-8}\,m \text{ (2sf)}$$

Now, convert to *nm*

$$Thickness = 3.7 \text{ x } 10^{-8}\,m \text{ x } \left(\frac{1\,nm}{10^{-9}\,m}\right) = 37\,nm$$

Key Terms

accuracy how closely a measurement aligns with a correct value

atom smallest particle of an element that can enter into a chemical combination

Celsius (°C) unit of temperature; water freezes at 0 °C and boils at 100 °C on this scale

chemistry study of the composition, properties, and interactions of matter

compound pure substance that can be decomposed into two or more elements **cubic centimeter (cm^3 or cc)** volume of a cube with an edge length of exactly 1 cm **cubic meter (m^3)** SI unit of volume

density ratio of mass to volume for a substance or object

dimensional analysis (also, factor-label method) versatile mathematical approach that can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities

element substance that is composed of a single type of atom; a substance that cannot be decomposed by a chemical change

exact number number derived by counting or by definition

Fahrenheit unit of temperature; water freezes at 32 °F and boils at 212 °F on this scale

hypothesis tentative explanation of observations that acts as a guide for gathering and checking information

kelvin (K) SI unit of temperature; 273.15 K = 0 ºC

kilogram (kg) standard SI unit of mass; 1 kg = approximately 2.2 pounds

law statement that summarizes a vast number of experimental observations, and describes or predicts some aspect of the natural world

length measure of one dimension of an object

liter (L) (also, cubic decimeter) unit of volume; 1 L = 1,000 cm^3

mass fundamental property indicating amount of matter

meter (m) standard metric and SI unit of length; 1 m = approximately 1.094 yards

microscopic domain realm of things that are much too small to be sensed directly

milliliter (mL) 1/1,000 of a liter; equal to 1 cm^3

precision how closely a measurement matches the same measurement when repeated

macroscopic domain realm of everyday things that are large enough to sense directly by human sight and touch

rounding procedure used to ensure that calculated results properly reflect the uncertainty in the measurements used in the calculation

scientific method path of discovery that leads from question and observation to law or hypothesis to theory, combined with experimental verification of the hypothesis and any necessary modification of the theory

second (s) SI unit of time

SI units (International System of Units) standards fixed by international agreement in the International System of Units (*Le Système International d'Unités*)

significant figures (also, significant digits) all of the measured digits in a determination, including the uncertain last digit

symbolic domain specialized language used to represent components of the macroscopic and microscopic domains, such as chemical symbols, chemical formulas, chemical equations, graphs, drawings, and calculations

theory well-substantiated, comprehensive, testable explanation of a particular aspect of nature

uncertainty estimate of amount by which measurement differs from true value

unit standard of comparison for measurements

unit conversion factor ratio of equivalent quantities expressed with different units; used to convert from one unit to a different unit

weight force that gravity exerts on an object

Key Equations

- $density = \frac{mass}{volume}$
- $T_{°C} = \left(\frac{5}{9} * T_{°F}\right) - 32$
- $T_{°F} = \left(\frac{9}{5} * °C\right) + 32$
- $T_K = °C + 273.15$
- $T_{°C} = K - 273.15$

Summary

1.1 Chemistry in Context

Chemistry deals with the composition, structure, and properties of matter, and the ways by which various forms of matter may be interconverted. Thus, it occupies a central place in the study and practice of science and technology. Chemists use the scientific method to perform experiments, pose hypotheses, and formulate laws and develop theories, so that they can better understand the behavior of the natural world. To do so, they operate in the macroscopic, microscopic, and symbolic domains. Chemists measure, analyze, purify, and synthesize a wide variety of substances that are important to our lives.

1.2 Measurements

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

1.3 Measurement Uncertainty, Accuracy, and Precision

Quantities can be exact or measured. Measured quantities have an associated uncertainty that is represented by the number of significant figures in the measurement. The uncertainty of a calculated value depends on the uncertainties in the values used in the calculation and is reflected in how the value is rounded. Measured values can be accurate (close to the true value) and/or precise (showing little variation when measured repeatedly).

1.4 Mathematical Treatment of Measurement Results

Measurements are made using a variety of units. It is often useful or necessary to convert a measured quantity from one unit into another. These conversions are accomplished using unit conversion factors, which are derived by simple applications of a mathematical approach called the factor-label method or dimensional analysis. This strategy is also employed to calculate sought quantities using measured quantities and appropriate mathematical relations.

Exercises

1.1 Chemistry in Context

1. Explain how you could experimentally determine whether the outside temperature is higher or lower than 0 °C (32 °F) without using a thermometer.

2. Identify each of the following statements as being most similar to a hypothesis, a law, or a theory. Explain your reasoning.

(a) Falling barometric pressure precedes the onset of bad weather.

(b) All life on earth has evolved from a common, primitive organism through the process of natural selection.

(c) My truck's gas mileage has dropped significantly, probably because it's due for a tune-up.

3. Identify each of the following statements as being most similar to a hypothesis, a law, or a theory. Explain your reasoning.

(a) The pressure of a sample of gas is directly proportional to the temperature of the gas.

(b) Matter consists of tiny particles that can combine in specific ratios to form substances with specific properties.

(c) At a higher temperature, solids (such as salt or sugar) will dissolve better in water.

4. Identify each of the underlined items as a part of either the macroscopic domain, the microscopic domain, or the symbolic domain of chemistry. For any in the symbolic domain, indicate whether they are symbols for a macroscopic or a microscopic feature.

(a) The mass of a <u>lead pipe</u> is 14 lb.

(b) The mass of a certain <u>chlorine atom</u> is 35 amu.

(c) A bottle with a label that reads <u>Al</u> contains aluminum metal.

(d) <u>Al</u> is the symbol for an aluminum atom.

5. Identify each of the underlined items as a part of either the macroscopic domain, the microscopic domain, or the symbolic domain of chemistry. For those in the symbolic domain, indicate whether they are symbols for a macroscopic or a microscopic feature.

(a) A certain molecule contains one <u>H</u> atom and one Cl atom.

(b) <u>Copper wire</u> has a density of about 8 g/cm^3.

(c) The bottle contains 15 grams of <u>Ni powder</u>.

(d) A <u>sulfur molecule</u> is composed of eight sulfur atoms.

6. According to one theory, the pressure of a gas increases as its volume decreases because the molecules in the gas have to move a shorter distance to hit the walls of the container. Does this theory follow a macroscopic or microscopic description of chemical behavior? Explain your answer.

7. The amount of heat required to melt 2 lbs of ice is twice the amount of heat required to melt 1 lb of ice. Is this observation a macroscopic or microscopic description of chemical behavior? Explain your answer.

1.2 Measurements

8. Is one liter about an ounce, a pint, a quart, or a gallon?

9. Is a meter about an inch, a foot, a yard, or a mile?

10. Indicate the SI base units or derived units that are appropriate for the following measurements:

(a) the length of a marathon race (26 miles 385 yards)

(b) the mass of an automobile

(c) the volume of a swimming pool

(d) the speed of an airplane

(e) the density of gold

(f) the area of a football field

(g) the maximum temperature at the South Pole on April 1, 1913

11. Indicate the SI base units or derived units that are appropriate for the following measurements:

(a) the mass of the moon

(b) the distance from Dallas to Oklahoma City

(c) the speed of sound

(d) the density of air

(e) the temperature at which alcohol boils

(f) the area of the state of Delaware

(g) the volume of a flu shot or a measles vaccination

12. Give the name and symbol of the prefixes used with SI units to indicate multiplication by the following exact quantities.

(a) 10^3

(b) 10^{-2}

(c) 0.1

(d) 10^{-3}

(e) 1,000,000

(f) 0.000001

13. Give the name of the prefix and the quantity indicated by the following symbols that are used with SI base units.

(a) c

(b) d

(c) G

(d) k

(e) m

(f) n

(g) p

(h) T

14. A large piece of jewelry has a mass of 132.6 g. A graduated cylinder initially contains 48.6 mL water. When the jewelry is submerged in the graduated cylinder, the total volume increases to 61.2 mL.

(a) Determine the density of this piece of jewelry.

(b) Assuming that the jewelry is made from only one substance, what substance is it likely to be? Explain.

15. Visit this **PhET density simulation (http://openstaxcollege.org/l/16phetmasvolden)** and select the Same Volume Blocks.

(a) What are the mass, volume, and density of the yellow block?

(b) What are the mass, volume and density of the red block?

(c) List the block colors in order from smallest to largest mass.

(d) List the block colors in order from lowest to highest density.

(e) How are mass and density related for blocks of the same volume?

16. Visit this **PhET density simulation (http://openstaxcollege.org/l/16phetmasvolden)** and select

Custom Blocks and then My Block.

(a) Enter mass and volume values for the block such that the mass in kg is *less than* the volume in L. What does the block do? Why? Is this always the case when mass < volume?

(b) Enter mass and volume values for the block such that the mass in kg is *more than* the volume in L. What does the block do? Why? Is this always the case when mass > volume?

(c) How would (a) and (b) be different if the liquid in the tank were ethanol instead of water?

(d) How would (a) and (b) be different if the liquid in the tank were mercury instead of water?

17. Visit this **PhET density simulation (http://openstaxcollege.org/l/16phetmasvolden)** and select Mystery Blocks.

(a) Pick one of the Mystery Blocks and determine its mass, volume, density, and its likely identity.

(b) Pick a different Mystery Block and determine its mass, volume, density, and its likely identity.

(c) Order the Mystery Blocks from least dense to most dense. Explain.

1.3 Measurement Uncertainty, Accuracy, and Precision

18. Express each of the following numbers in scientific notation with correct significant figures: (a) 711.0

(b) 0.239

(c) 90743

(d) 134.2

(e) 0.05499

(f) 10000.0

(g) 0.000000738592

19. Express each of the following numbers in exponential notation with correct significant figures:

(a) 704

(b) 0.03344

(c) 547.9

(d) 22086

(e) 1000.00

(f) 0.0000000651

(g) 0.007157

20. Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

(a) the number of eggs in a basket

(b) the mass of a dozen eggs

(c) the number of gallons of gasoline necessary to fill an automobile gas tank

(d) the number of cm in 2 m

(e) the mass of a textbook

(f) the time required to drive from San Francisco to Kansas City at an average speed of 53 mi/h

21. Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

(a) the number of seconds in an hour

(b) the number of pages in this book

(c) the number of grams in your weight

(d) the number of grams in 3 kilograms

(e) the volume of water you drink in one day

(f) the distance from San Francisco to Kansas City

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

(a) 38.7 g

(b) 2×10^{18} m

(c) 3,486,002 kg

(d) 9.74150×10^{-4} J

(e) 0.0613 cm^3

(f) 17.0 kg

(g) 0.01400 g/mL

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

(a) 53 cm

(b) 2.05×10^{8} m

(c) 86,002 J

(d) 9.740×10^{4} m/s

(e) 10.0613 m^3

(f) 0.17 g/mL

(g) 0.88400 s

24. The following quantities were reported on the labels of commercial products. Determine the number of significant figures in each.

(a) 0.0055 g active ingredients

(b) 12 tablets

(c) 3% hydrogen peroxide

(d) 5.5 ounces

(e) 473 mL

(f) 1.75% bismuth

(g) 0.001% phosphoric acid

(h) 99.80% inert ingredients

25. Round off each of the following numbers to two significant figures: (a) 0.436

(b) 9.000

(c) 27.2

(d) 135

(e) 1.497×10^{-3}

(f) 0.445

26. Round off each of the following numbers to two significant figures:

(a) 517

(b) 86.3

(c) 6.382×10^{3}

(d) 5.0008

(e) 22.497

(f) 0.885

27. Perform the following calculations and report each answer with the correct number of significant figures.

(a) 628×342

(b) $(5.63 \times 10^{2}) \times (7.4 \times 10^{3})$

(c) $\frac{28.0}{13.483}$

(d) 8119×0.000023

(e) 14.98 + 27,340 + 84.7593

(f) 42.7 + 0.259

28. Perform the following calculations and report each answer with the correct number of significant figures. (a) 62.8×34

(b) 0.147 + 0.0066 + 0.012

(c) $38 \times 95 \times 1.792$

(d) 15 – 0.15 – 0.6155

(e) $8.78 * \frac{0.0500}{0.478}$

(f) 140 + 7.68 + 0.014

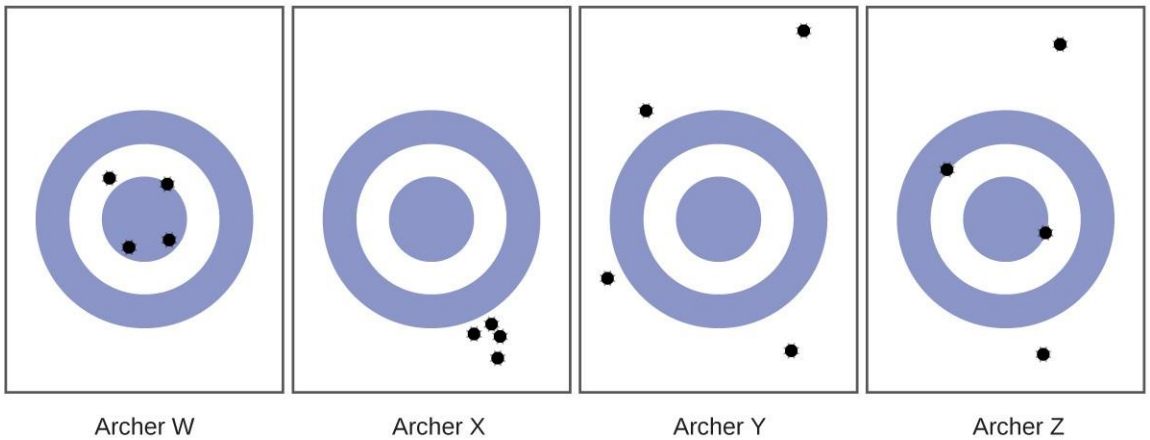
(g) 28.7 – 0.0483

(h) $\frac{(88.5-87.57)}{45.13}$

29. Consider the results of the archery contest shown in this figure.

(a) Which archer is most precise?

(b) Which archer is most accurate?



(c) Who is both least precise and least accurate?

30. Classify the following sets of measurements as accurate, precise, both, or neither.

(a) Checking for consistency in the weight of chocolate chip cookies: 17.27 g, 13.05 g, 19.46 g, 16.92 g

(b) Testing the volume of a batch of 25-mL pipettes: 27.02 mL,26.99 mL, 26.97 mL, 27.01 mL

(c) Determining the purity of gold: 99.9999%, 99.9998%, 99.9998%, 99.9999%

1.4 Mathematical Treatment of Measurement Results

31. Write conversion factors (as ratios) for the number of:

(a) yards in 1 meter

(b) liters in 1 liquid quart

(c) pounds in 1 kilogram

32. Write conversion factors (as ratios) for the number of:

(a) kilometers in 1 mile

(b) liters in 1 cubic foot

(c) grams in 1 ounce

33. The label on a soft drink bottle gives the volume in two units: 2.0 L and 67.6 fl oz. Use this information to derive a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

34. The label on a box of cereal gives the mass of cereal in two units: 978 grams and 34.5 oz. Use this information to find a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

35. Soccer is played with a round ball having a circumference between 27 and 28 in. and a weight between 14 and 16 oz. What are these specifications in units of centimeters and grams?

36. A woman's basketball has a circumference between 28.5 and 29.0 inches and a maximum weight of 20 ounces (two significant figures). What are these specifications in units of centimeters and grams?

37. How many milliliters of a soft drink are contained in a 12.0-oz can?

38. A barrel of oil is exactly 42 gal. How many liters of oil are in a barrel?

39. The diameter of a red blood cell is about 3×10^{-4} in. What is its diameter in centimeters?

40. The distance between the centers of the two oxygen atoms in an oxygen molecule is 1.21×10^{-8} cm. What is this distance in inches?

41. Is a 197-lb weight lifter light enough to compete in a class limited to those weighing 90 kg or less?

42. A very good 197-lb weight lifter lifted 192 kg in a move called the clean and jerk. What was the mass of the weight lifted in pounds?

43. Many medical laboratory tests are run using 5.0 μL blood serum. What is this volume in milliliters?

44. If an aspirin tablet contains 325 mg aspirin, how many grams of aspirin does it contain?

45. Use scientific (exponential) notation to express the following quantities in terms of the SI base units in **Table 1.3**:

(a) 0.13 g

(b) 232 Gg

(c) 5.23 pm

(d) 86.3 mg

(e) 37.6 cm

(f) 54 μm

(g) 1 Ts

(h) 27 ps

(i) 0.15 mK

46. Complete the following conversions between SI units.

(a) 612 g =_____mg (b) 8.160 m =_____cm

(c) 3779 μg =_____g

(d) 781 mL =_____L

(e) 4.18 kg =_____g

(f) 27.8 m =_____km

(g) 0.13 mL =_____L

(h) 1738 km =_____m

(i) 1.9 Gg =_____g

47. Gasoline is sold by the liter in many countries. How many liters are required to fill a 12.0-gal gas tank?

48. Milk is sold by the liter in many countries. What is the volume of exactly 1/2 gal of milk in liters?

49. A long ton is defined as exactly 2240 lb. What is this mass in kilograms?

50. Make the conversion indicated in each of the following:

(a) the men's world record long jump, 29 ft 4¼ in., to meters

(b) the greatest depth of the ocean, about 6.5 mi, to kilometers

(c) the area of the state of Oregon, 96,981 mi^2, to square kilometers

(d) the volume of 1 gill (exactly 4 oz) to milliliters

(e) the estimated volume of the oceans, 330,000,000 mi^3, to cubic kilometers.

(f) the mass of a 3525-lb car to kilograms

(g) the mass of a 2.3-oz egg to grams

51. Make the conversion indicated in each of the following:

(a) the length of a soccer field, 120 m (three significant figures), to feet

(b) the height of Mt. Kilimanjaro, at 19,565 ft the highest mountain in Africa, to kilometers

(c) the area of an 8.5 × 11-inch sheet of paper in cm^2

(d) the displacement volume of an automobile engine, 161 $in.^3$, to liters

(e) the estimated mass of the atmosphere, 5.6×10^{15} tons, to kilograms

(f) the mass of a bushel of rye, 32.0 lb, to kilograms

(g) the mass of a 5.00-grain aspirin tablet to milligrams (1 grain = 0.00229 oz)

52. Many chemistry conferences have held a 50-Trillion Angstrom Run (two significant figures). How long is this run in kilometers and in miles? (1 Å = 1×10^{-10} m)

53. A chemist's 50-Trillion Angstrom Run (see **Exercise 1.78**) would be an archeologist's 10,900 cubit run. How long is one cubit in meters and in feet? (1 Å = 1×10^{-8} cm)

54. The gas tank of a certain luxury automobile holds 22.3 gallons according to the owner's manual. If the density of gasoline is 0.8206 g/mL, determine the mass in kilograms and pounds of the fuel in a full tank.

55. As an instructor is preparing for an experiment, he requires 225 g phosphoric acid. The only container readily available is a 150-mL Erlenmeyer flask. Is it large enough to contain the acid, whose density is 1.83 g/mL?

56. To prepare for a laboratory period, a student lab assistant needs 125 g of a compound. A bottle containing 1/4 lb is available. Did the student have enough of the compound?

57. A chemistry student is 159 cm tall and weighs 45.8 kg. What is her height in inches and weight in pounds?

58. In a recent Grand Prix, the winner completed the race with an average speed of 229.8 km/h. What was his speed in miles per hour, meters per second, and feet per second?

59. Solve these problems about lumber dimensions.

(a) To describe to a European how houses are constructed in the US, the dimensions of "two-by-four" lumber must be converted into metric units. The thickness × width × length dimensions are 1.50 in. × 3.50 in. × 8.00 ft in the US. What are the dimensions in cm × cm × m?

(b) This lumber can be used as vertical studs, which are typically placed 16.0 in. apart. What is that distance in centimeters?

60. The mercury content of a stream was believed to be above the minimum considered safe—1 part per billion (ppb) by weight. An analysis indicated that the concentration was 0.68 parts per billion. What quantity of mercury in grams was present in 15.0 L of the water, the density of which is 0.998 g/ml? ($1 ppb\ Hg = \frac{1\ ngHg}{1\ g\ water}$)

61. Calculate the density of aluminum if 27.6 cm^3 has a mass of 74.6 g.

62. Osmium is one of the densest elements known. What is its density if 2.72 g has a volume of 0.121 cm^3?

63. Calculate these masses.

(a) What is the mass of 6.00 cm^3 of mercury, density = 13.5939 g/cm^3?

(b) What is the mass of 25.0 mL octane, density = 0.702 g/cm^3?

64. Calculate these masses.

(a) What is the mass of 4.00 cm^3 of sodium, density = 0.97 g/cm^3 ?

(b) What is the mass of 125 mL gaseous chlorine, density = 3.16 g/L?

65. Calculate these volumes.

(a) What is the volume of 25 g iodine, density = 4.93 g/cm^3?

(b) What is the volume of 3.28 g gaseous hydrogen, density = 0.089 g/L?

66. Calculate these volumes.

(a) What is the volume of 11.3 g graphite, density = 2.25 g/cm^3?

(b) What is the volume of 39.657 g bromine, density = 2.928 g/cm^3?

67. Convert the boiling temperature of gold, 2966 °C, into degrees Fahrenheit and kelvin.

68. Convert the temperature of scalding water, 54 °C, into degrees Fahrenheit and kelvin.

69. Convert the temperature of the coldest area in a freezer, −10 °F, to degrees Celsius and kelvin.

70. Convert the temperature of dry ice, −77 °C, into degrees Fahrenheit and kelvin.

71. Convert the boiling temperature of liquid ammonia, −28.1 °F, into degrees Celsius and kelvin.

72. The label on a pressurized can of spray disinfectant warns against heating the can above 130 °F. What are the corresponding temperatures on the Celsius and kelvin temperature scales?

73. The weather in Europe was unusually warm during the summer of 1995. The TV news reported temperatures as high as 45 °C. What was the temperature on the Fahrenheit scale?