

CHAPTER 1

Measurements

Chapter Outline

- 1.1 THE INTERNATIONAL SYSTEM OF UNITS
- 1.2 UNIT CONVERSIONS
- 1.3 UNCERTAINTY IN MEASUREMENTS
- 1.4 REFERENCES



Measurements, measurements, measurements! They affect your life in many ways each and every day. When cooking, it is very important to carefully measure each ingredient so that your recipe turns out exactly the way you intend. We make time measurements frequently, from cooking times, to estimating travel times and deciding how much time will be required for various tasks. The builders of your house or apartment hopefully made good measurements of the various wooden, plastic, and metal materials that went into its construction. A chemist also must make accurate measurements in the laboratory when conducting his or her experiments. The image above shows some of the common glassware that a chemist frequently uses in measuring volumes of liquids. In this chapter, you will learn about the critical skills of making, reporting, and performing calculations with measurements.

User:Tweenk/Wikimedia Commons. commons.wikimedia.org/wiki/File:Lab_glassware.jpg. CC BY 3.0.

1.1 The International System of Units

Lesson Objectives

- Identify the seven base units of the International System of Units.
- Know the commonly used metric prefixes.
- Convert between the Celsius and Kelvin temperature scales.
- Understand volume and energy as combinations of SI Units.
- Distinguish between mass and weight.

Lesson Vocabulary

- energy
- International System of Units (SI)
- joule
- kinetic energy
- liter
- measurement
- scientific notation
- temperature
- weight

Check Your Understanding

Recalling Prior Knowledge

- Why is the metric system easier to use than the English system of units?
- How is scientific notation used to represent very large or very small numbers?
- What units are used to measure length, mass, and volume in the metric system?

The temperature outside is 52 degrees Fahrenheit. Your height is 67 inches and your weight is 145 pounds. All are examples of measurements. A **measurement** is a quantity that includes both a number and a unit. If someone were to describe the height of a building as 85, that would be meaningless. 85 meters? 85 feet? Without a unit, a measurement does not convey enough information to be useful. In this lesson, we will begin an exploration of the units that are typically used in chemistry.

SI Base Units

All measurements depend on the use of units that are well known and understood. The English system of measurement units (inches, feet, ounces, etc.) is not used in science because of the difficulty in converting from one unit to another. The metric system is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. The **International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units (**Table 1.1**).

TABLE 1.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first five units are frequently encountered in chemistry. The amount of a substance, the mole, will be discussed in detail in a later chapter. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

You can learn more about base units at www.nist.gov/pml/wmd/metric/si-units.cfm .

Metric Prefixes and Scientific Notation

As stated earlier, conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. Prefixes are used to distinguish between units of different size. Listed below (**Table 1.2**) are the most common metric prefixes and their relationship to the central unit, which has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

TABLE 1.2: SI Prefixes

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
giga	G	10^9	1,000,000,000	1 gigameter (Gm) = 10^9 m
mega	M	10^6	1,000,000	1 megameter (Mm) = 10^6 m
kilo	k	10^3	1000	1 kilometer (km) = 1000 m
hecto	h	10^2	100	1 hectometer (hm) = 100 m
deka	da	10^1	10	1 dekameter (dam) = 10 m
		10^0	1	1 meter (m)

TABLE 1.2: (continued)

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
deci	d	10^{-1}	1/10	1 decimeter (dm) = 0.1 m
centi	c	10^{-2}	1/100	1 centimeter (cm) = 0.01 m
milli	m	10^{-3}	1/1000	1 millimeter (mm) = 0.001 m
micro	μ	10^{-6}	1/1,000,000	1 micrometer (μm) = 10^{-6} m
nano	n	10^{-9}	1/1,000,000,000	1 nanometer (nm) = 10^{-9} m
pico	p	10^{-12}	1/1,000,000,000,000	1 picometer (pm) = 10^{-12} m

There are more prefixes, although some of them are rarely used. Have you ever heard of a zeptometer? You can learn more about metric prefixes at www.nist.gov/pml/wmd/metric/prefixes.cfm.

The table above (**Table 1.2**) introduces a very useful tool for working with numbers that are either very large or very small. **Scientific notation** is a way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power. As an example, the distance from Earth to the Sun is about 150,000,000,000 meters—a very large distance indeed. In scientific notation, the distance is written as 1.5×10^{11} m. The coefficient is 1.5 and must be a number greater than or equal to 1 and less than 10. The power of 10, or exponent, is 11. Pictured below are two more examples of scientific notation (**Figure 1.1**). Scientific notation is sometimes referred to as exponential notation.

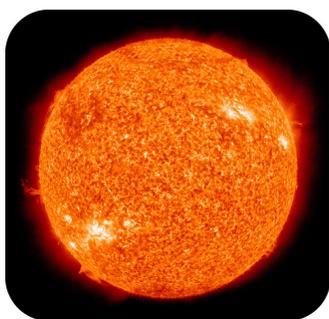


FIGURE 1.1

The sun is very large and very distant, so solar data is better expressed in scientific notation. The mass of the sun is 2.0×10^{30} kg, and its diameter is 1.4×10^9 m.

Very small numbers can also be expressed using scientific notation. The mass of an electron in decimal notation is 0.00000000000000000000000000911 grams. In scientific notation, the mass is expressed as 9.11×10^{-28} g. Notice that the value of the exponent is chosen so that the coefficient is between 1 and 10.

Typical Units in Chemistry

Length and Volume

The SI basic unit of length, or linear measure, is the meter (m). All measurements of length may be made in meters, though the prefixes listed above (**Table 1.2**) will often be more convenient. The width of a room may be expressed

as about 5 meters (m), whereas a large distance such as the distance between New York City and Chicago is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 20 micrometers (μm).

Volume is the amount of space occupied by a sample of matter (**Figure 1.2**). The volume of a regular object can be calculated by multiplying its length by its width by its height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter (m^3), which is the volume occupied by a cube that measures 1 m on each side. This very large volume is not very convenient for typical use in a chemistry laboratory. A **liter (L)** is the volume of a cube that measures 10 cm (1 dm) on each side. A liter is thus equal to both 1000 cm^3 ($10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm}$) and to 1 dm^3 . A smaller unit of volume that is commonly used is the milliliter (mL). A milliliter is the volume of a cube that measures 1 cm on each side. Therefore, a milliliter is equal to a cubic centimeter (cm^3). There are 1000 mL in 1 L, which is the same as saying that there are 1000 cm^3 in 1 dm^3 .

**FIGURE 1.2**

(A) A typical water bottle is 1 liter in volume. (B) These dice measure 1 cm on each side, so each die has a volume of 1 cm^3 or 1 mL. (C) Volume in the laboratory is often measured with graduated cylinders, which come in a variety of sizes.

You can watch a video about measuring volume using graduated cylinders at <http://www.benchfly.com/video/153/how-to-use-graduated-cylinders/> .

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The mass of an object is made in comparison to the standard mass of 1 kilogram. The kilogram was originally defined as the mass of 1 L of liquid water at 4°C (the volume of a liquid changes slightly with temperature). In the laboratory, mass is measured with a balance (**Figure 1.3**), which must be calibrated with a standard mass so that its measurements are accurate.

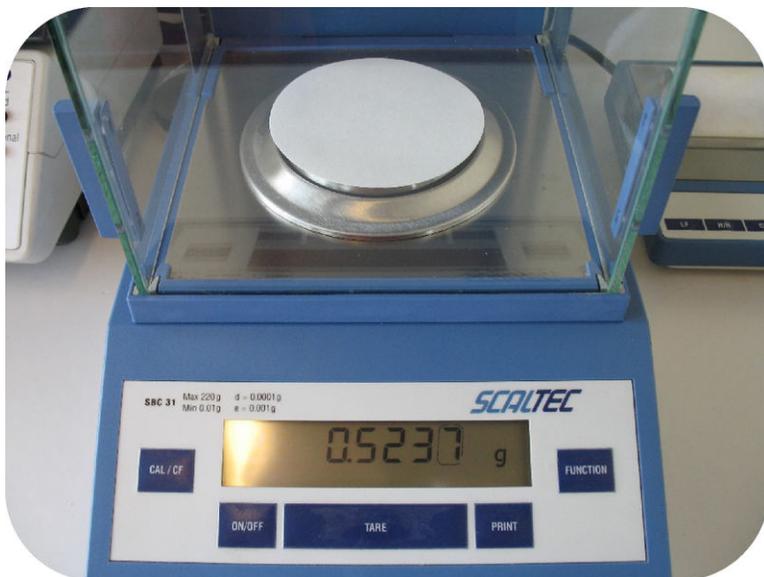
You can watch a short video about using an analytical balance at <http://www.benchfly.com/video/54/how-to-weigh-small-amounts/> .

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are 1000 g in 1 kg. A milligram is 1/1000th of a gram, so there are 1000 mg in 1 g.

Mass is often confused with the term weight. **Weight** is a measure of force that is equal to the gravitational pull on an object. The weight of an object is dependent on its location. On the moon, the force due to gravity is about one sixth that of the gravitational force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale by reading the distance that a certain object pulls down and stretches a spring.

Temperature and Energy

Touch the top of the stove after it has been on and it feels hot. Hold an ice cube in your hand and it feels cold. Why? The particles of matter in a hot object are moving much faster than the particles of matter in a cold object.

**FIGURE 1.3**

An analytical balance in the laboratory takes very sensitive measurements of mass, usually in grams.

An object's **kinetic energy** is the energy due to motion. The particles of matter that make up the hot stove have a greater amount of kinetic energy than those in the ice cube (**Figure 1.4**). **Temperature** is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature is how hot or cold an object is. Temperature determines the direction of heat transfer. When two objects at different temperatures are brought into contact with one another, heat flows from the object at the higher temperature to the object at the lower temperature. This occurs until their temperatures are the same.

**FIGURE 1.4**

The glowing charcoal on the left is composed of particles with a high level of kinetic energy, while the snow and ice on the right are made of particles with much less kinetic energy.

Temperature can be measured with several different scales. The Fahrenheit scale is typically not used for scientific purposes. The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at 0°C and 100°C , respectively. The distance between those two points is divided into 100 equal intervals, each of which is referred to as one degree.

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824-1907). It is based on molecular motion, with the temperature of 0 K, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is 273.15 K, while the boiling point is 373.15 K. As can be seen by the 100 kelvin difference between the two, a change of one degree on the Celsius scale is equivalent to the change of one kelvin on the Kelvin scale. Converting from one scale to another is easy, as you simply add or subtract 273.15 (**Figure 1.5**).

$$^{\circ}\text{C} = \text{K} - 273.15$$

$$\text{K} = ^{\circ}\text{C} + 273.15$$

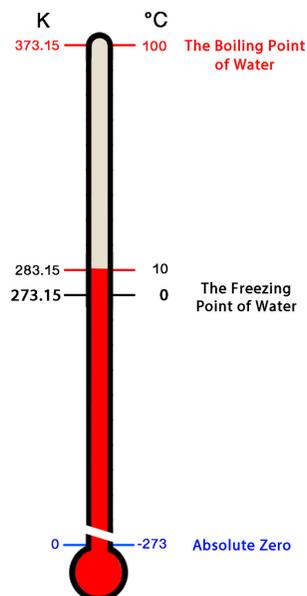


FIGURE 1.5

A comparison of the Kelvin (left) and Celsius (right) temperature scales. The two scales differ from one another by 273.15 degrees.

Energy is defined as the capacity to do work or to produce heat. As discussed previously, kinetic energy is one type of energy and is associated with motion. Another frequently encountered form of energy is potential energy, which is a type of energy that is stored in matter. The **joule (J)** is the SI unit of energy and is named after English physicist James Prescott Joule (1818-1889). In terms of SI base units, a joule is equal to a kilogram times a meter squared divided by a second squared ($\text{kg}\cdot\text{m}^2/\text{s}^2$). A common non-SI unit of energy that is often used is the calorie (cal), which is equal to 4.184 J.

Lesson Summary

- Measurements are critical to any field of science and must consist of a quantity and an appropriate unit. The International System of Units consists of seven base units.
- The metric system utilizes prefixes and powers of 10 to make conversions between units easy.
- Length (m), mass (kg), temperature (K), time (s), and amount (mol) are the base units that are most frequently used in chemistry. Quantities such as volume and energy can be derived from combinations of the base units.

Lesson Review Questions

Reviewing Concepts

1. Give the SI base unit of measurement for each of the following quantities.

- a. mass
 - b. length
 - c. time
 - d. temperature
2. Convert the following numbers into scientific notation.
 - a. 85,000,000
 - b. 0.00019
 3. Put the following into decimal notation.
 - a. 8.72×10^{-8}
 - b. 3×10^4
 4. Place the following units of mass in order from smallest to largest: g, kg, μg , g, pg, Mg, ng, cg, dg.
 5. Explain what is wrong with the following statement: "This rock weighs 8 kilograms."
 6. What is absolute zero on the Celsius temperature scale?

Problems

7. Calculate the volume in mL of a cube that is 2.20 cm on each side.
8. A rectangular solid has a volume of 80 cm^3 . Its length is 2.0 cm and its width is 8.0 cm. What is the height of the solid?
9. Convert the following Celsius temperatures to Kelvin.
 - a. 36°C
 - b. -104°
10. Convert the following Kelvin temperatures to degrees Celsius.
 - a. 188 K
 - b. 631 K
11. Temperature in degrees Fahrenheit can be converted to Celsius by first subtracting 32, then dividing by 1.8. What is the Celsius temperature outside on a warm 88°F day?
12. Two samples of water are at different temperatures. A 2 L sample is at 40°C , while a 1 L sample is at 70°C .
 - a. The particles of which sample have a larger average kinetic energy?
 - b. The water samples are mixed. Assuming no heat loss, what will be the temperature of the 3 L of water?

Further Reading / Supplemental Links

- E.A. Mechtly, *International System of Units: Fundamental Constants and Conversion Factors*. Stipes Pub Llc, 1977.
- SI Metric System: <http://www.simetric.co.uk/sibasis.htm>
- You can do an online metric system crossword puzzle at http://education.jlab.org/sciencecrossword/metricsystem_01.html .
- Take a museum tour of weights and measurements at <http://museum.nist.gov/exhibits/ex1/index.html> .
- You can view a comparison of the sizes of viruses, DNA, and biological molecules, along with information about DNA-based computers, at <http://publications.nigms.nih.gov/chemhealth/cool.htm> .
- Time is standardized by an atomic clock. You can view a short video about the Amazing Atomic Clock at <http://video.pbs.org/video/2167682634> .
- Test your metric skills by playing Metric System Hangman. There are two Hangman games.
 - The first is at http://education.jlab.org/vocabhangman/metric_system_01/8.html .
 - The second is at http://education.jlab.org/vocabhangman/metric_system_02/8.html .

Points to Consider

Conversions between units of the metric system are made easy because they are related by powers of ten and because the prefixes are consistent across various types of measurement (length, volume, mass, etc.).

- What is the mass in grams of a 2.50 kg book?
- What is the length in cm of a field that is 0.65 km?

1.2 Unit Conversions

Lesson Objectives

- Identify and use conversion factors.
- Use the method of dimensional analysis to convert between units.
- Understand density as a physical property of matter.
- Perform calculations with derived units, including density.

Lesson Vocabulary

- conversion factor
- density
- derived unit
- dimensional analysis

Check Your Understanding

Recalling Prior Knowledge

- Why are units required when reporting the results of a measured quantity?
- When a quantity with a large unit (such as km) is changed into a quantity with a smaller unit (such as cm), will the numerical value of the quantity increase or decrease?

When traveling in another country, you may be faced with a unit problem. For example, if you are driving, you may encounter a sign saying that the next town is 30 km away. If your car's odometer measures distances in miles, how far will you need to go to get to that town? In this lesson, you will learn to solve this and other unit-conversion problems with a technique called dimensional analysis.

Conversion Factors

Many quantities can be expressed in several different ways. For example, the English system measurement of 4 cups is also equal to 2 pints, 1 quart, and 1/4 of a gallon.

$$4 \text{ cups} = 2 \text{ pints} = 1 \text{ quart} = 0.25 \text{ gallon}$$

Notice that the numerical component of each quantity is different, while the actual amount of material that it represents is the same. That is because the units are different. We can establish the same set of equalities for the metric system:

$$1 \text{ meter} = 10 \text{ decimeters} = 100 \text{ centimeters} = 1000 \text{ millimeters}$$

The metric system's use of powers of 10 for all conversions makes this quite simple.

Whenever two quantities are equal, a ratio can be written that is numerically equal to 1. Using the metric examples above:

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

The fraction 1 m/100 cm is called a conversion factor. A **conversion factor** is a ratio of equivalent measurements. Because both 1 m and 100 cm represent the exact same length, the value of the conversion factor is 1. The conversion factor is read as "1 meter per 100 centimeters." Other conversion factors from the cup measurement example can be:

$$\frac{4 \text{ cups}}{2 \text{ pints}} = \frac{2 \text{ pints}}{1 \text{ quart}} = \frac{1 \text{ quart}}{0.25 \text{ gallon}} = 1$$

Since the numerator and denominator represent equal quantities in each case, all are valid conversion factors.

Dimensional Analysis

Conversion factors are used in solving problems in which a certain measurement must be expressed with different units. When a given measurement is multiplied by an appropriate conversion factor, the numerical value changes, but the actual size of the quantity measured remains the same. **Dimensional analysis** is a technique that uses the units (dimensions) of the measurement in order to correctly solve problems. Dimensional analysis is best illustrated with an example.

Sample Problem 3.1: Dimensional Analysis

How many seconds are in a day?

Step 1: List the known quantities and plan the problem.

Known

- 1 day = 24 hours
- 1 hour = 60 minutes
- 1 minute = 60 seconds

Unknown

- 1 day = ? seconds

The known quantities above represent the conversion factors that we will use. The first conversion factor will have day in the denominator so that the "day" unit will cancel. The second conversion factor will then have hours in the denominator, while the third conversion factor will have minutes in the denominator. As a result, the unit of the last numerator will be seconds and that will be the units for the answer.

Step 2: Calculate.

$$1 \text{ d} \times \frac{24 \text{ h}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} = 86,400 \text{ s}$$

Applying the first conversion factor, the “d” unit cancels and $1 \times 24 = 24$. Applying the second conversion factor, the “h” unit cancels and $24 \times 60 = 1440$. Applying the third conversion factor, the “min” unit cancels and $1440 \times 60 = 86400$. The unit that remains is “s” for seconds.

Step 3: Think about your result.

A second is a much smaller unit of time than a day, so it makes sense that there are a very large number of seconds in one day.

Practice Problems

1. How many minutes are in a year?
2. How many days are equal to one million seconds?

Dimensional Analysis and the Metric System

The metric system’s many prefixes allow quantities to be expressed in many different units. Dimensional analysis is useful to convert from one metric system unit to another.

Sample Problem 3.2: Metric Unit Conversions

A particular experiment requires 120 mL of a solution. The teacher knows that he will need to make enough solution for 40 experiments to be performed throughout the day. How many liters of solution should he prepare?

Step 1: List the known quantities and plan the problem.

Known

- 1 experiment requires 120 mL of solution
- 1 L = 1000 mL

Unknown

- 40 experiments require? L of solution

Since each experiment requires 120 mL of solution and the teacher needs to prepare enough for 40 experiments, multiply 120 by 40 to get 4800 mL of solution needed. Now you must convert mL to L by using a conversion factor.

Step 2: Calculate.

$$4800 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 4.8 \text{ L}$$

Note that conversion factor is arranged so that the mL unit is in the denominator and thus cancels out, leaving L as the remaining unit in the answer.

Step 3: Think about your result.

A liter is much larger than a milliliter, so it makes sense that the number of liters required is less than the number of milliliters.

Practice Problems

3. Perform the following conversions.

- (a) 0.074 km to m
- (b) 24,600 μg to g
- (c) 1300 ms to s
- (d) 3.8×10^{-5} L to μL

Some metric conversion problems are most easily solved by breaking them down into more than one step. When both the given unit and the desired unit have prefixes, one can first convert to the simple (unprefixed) unit, followed by a conversion to the desired unit. An example will illustrate this method.

Sample Problem 3.3: Two-Step Metric Conversion

Convert 4.3 cm to μm .

Step 1: List the known quantities and plan the problem.

Known

- 1 m = 100 cm
- 1 m = 10^6 μm

Unknown

- 4.3 cm = ? μm

You may need to consult the table in the lesson, "The International System of Units," for the multiplication factor represented by each metric prefix. First convert cm to m; then convert m to μm .

Step 2: Calculate.

$$4.3 \text{ cm} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{10^6 \mu\text{m}}{1 \text{ m}} = 43,000 \mu\text{m}$$

Each conversion factor is written so that unit of the denominator cancels with the unit of the numerator of the previous factor.

Step 3: Think about your result.

A micrometer is a smaller unit of length than a centimeter, so the answer in micrometers is larger than the number of centimeters given.

Practice Problems

4. Perform the following conversions.

- (a) 4.9×10^7 μg to kg
- (b) 84 dm to mm
- (c) 355 nm to cm
- (d) 70.5 ML to mL

Dimensional Analysis and Derived Units

Some units are combinations of SI base units. A **derived unit** is a unit that results from a mathematical combination of SI base units. We have already discussed volume and energy as two examples of derived units. Some others are listed below (**Table 1.3**).

TABLE 1.3: Derived SI Units

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

Using dimensional analysis with derived units requires special care. When units are squared or cubed, as with area or volume, the conversion factors themselves must also be squared. Shown below is the conversion factor for cubic centimeters and cubic meters.

$$\left(\frac{1 \text{ m}}{100 \text{ cm}}\right)^3 = \frac{1 \text{ m}^3}{10^6 \text{ cm}^3} = 1$$

Because a cube has 3 sides, each side is subject to the conversion of 1 m to 100 cm. Since 100 cubed is equal to 1 million (10⁶), there are 10⁶ cm³ in 1 m³. Two convenient volume units are the liter, which is equal to a cubic decimeter, and the milliliter, which is equal to a cubic centimeter. The conversion factor would be:

$$\left(\frac{1 \text{ dm}}{10 \text{ cm}}\right)^3 = \frac{1 \text{ dm}^3}{1000 \text{ cm}^3} = 1$$

There are thus 1000 cm³ in 1 dm³, which is the same thing as saying there are 1000 mL in 1 L (**Figure 1.6**).

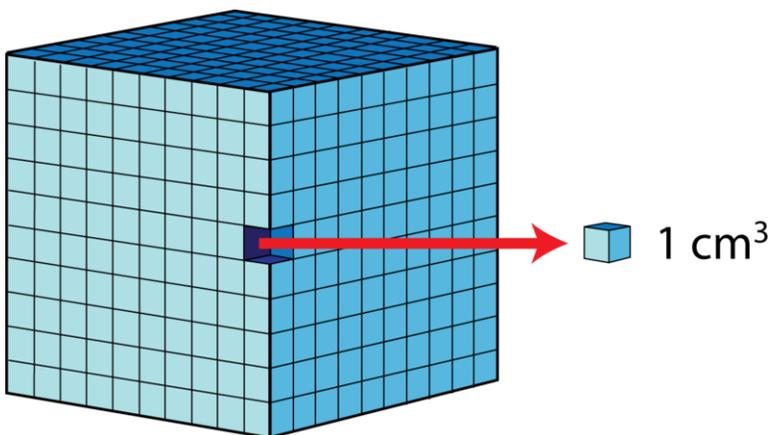


FIGURE 1.6

There are 1000 cm³ in 1 dm³. Since 1 cm³ is equal to 1 mL and 1 dm³ is equal to 1 L, we can say that there are 1000 mL in 1 L.

You can participate in an interactive version of this cube at www.dlt.ncssm.edu/core/Chapter1-Introduction/Chapter1-Animations/M3_DM3_CM3.html .

Sample Problem 3.4: Derived Unit Conversion

Convert 3.6 mm^3 to mL.

Step 1: List the known quantities and plan the problem.

Known

- $1 \text{ m} = 1000 \text{ mm}$
- $1 \text{ mL} = 1 \text{ cm}^3$
- $1 \text{ m} = 100 \text{ cm}$

Unknown

- $3.6 \text{ mm}^3 = ? \text{ mL}$

This problem requires multiple steps and the technique for converting with derived units. Simply proceed one step at a time: mm^3 to m^3 to $\text{cm}^3 = \text{mL}$.

Step 2: Calculate.

$$3.6 \text{ mm}^3 \times \left(\frac{1 \text{ m}}{1000 \text{ mm}}\right)^3 \times \left(\frac{100 \text{ cm}}{1 \text{ m}}\right)^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = 0.0036 \text{ mL}$$

Numerically, the steps are to divide 3.6 by 10^9 , followed by multiplying by 10^6 . Alternatively, you can shorten the calculation by one step if you first determine the conversion factor between mm and cm. Using the fact that there are 10 mm in 1 cm, you can avoid the intermediate step of converting to meters.

$$3.6 \text{ mm}^3 \times \left(\frac{1 \text{ cm}}{10 \text{ mm}}\right)^3 \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = 0.0036 \text{ mL}$$

Both conversion methods give the same result of 0.0036 mL.

Step 3: Think about your result.

Cubic millimeters are much smaller than cubic centimeters, so the final answer is much less than the original number of mm^3 .

Practice Problems

5. Perform the following conversions.

- (a) 0.00722 km^2 to m^2
- (b) 129 cm^3 to L
- (c) $4.9 \times 10^5 \mu\text{m}^3$ to mm^3

You can find more help with dimensional analysis by watching this video at www.chemcollective.org/stoich/dimensionalanalysis.php .

You can download an instructional packet that explains dimensional analysis step by step and provides practice problems at https://docs.google.com/open?id=0B_ZuEGrhVEfMS09WeUtOMFNTaFk .

Density

A golf ball and a table tennis ball are about the same size. However, the golf ball is much heavier than the table tennis ball. Now imagine a similar size ball made out of lead. That would be very heavy indeed! What are we comparing? By comparing the mass of an object relative to its size, we are studying a property called density. **Density is the ratio of the mass of an object to its volume.**

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad D = \frac{m}{V}$$

Density is an intensive property, meaning that it does not depend on the amount of material present in the sample. For example, water has a density of 1.0 g/mL. That density is the same whether you have a small glass of water or a swimming pool full of water. Density is a property that is constant for a specific type of matter at a given temperature.

The SI units of density are kilograms per cubic meter (kg/m^3), because the kg and the m are the SI units for mass and length, respectively. Unfortunately, this unit is awkwardly large for everyday usage in the laboratory. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter (g/cm^3). Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L. Densities of some common substances at 20°C are listed below (**Table 1.4**). Since most materials expand as temperature increases, the density of a substance is temperature-dependent and usually decreases as temperature increases.

TABLE 1.4: Densities of Some Common Substances

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084
Ice (0°C)	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water (4°C)	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.92		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

You know from experience that ice floats in water. This is consistent with the values listed above (**Table 1.4**), which show that ice is less dense than liquid water. Alternatively, corn syrup would sink if placed into water because it has a higher density. The video below shows a number of substances arranged into a density column.



MEDIA

Click image to the left or use the URL below.

URL: <http://www.ck12.org/flx/render/embeddedobject/65509>

With gases, a balloon filled with helium floats because helium is less dense than air. However, a balloon filled with carbon dioxide sinks because carbon dioxide is denser than air.

Sample Problem 3.5: Density Calculations

An 18.2 g sample of zinc metal has a volume of 2.55 cm³. Calculate the density of zinc.

Step 1: List the known quantities and plan the problem.

Known

- mass = 18.2 g
- volume = 2.55 cm³

Unknown

- density = ? g/cm³

Use the equation for density, $D = m/V$, to solve the problem.

Step 2: Calculate.

$$D = \frac{m}{V} = \frac{18.2 \text{ g}}{2.55 \text{ cm}^3} = 7.14 \text{ g/cm}^3$$

Step 3: Think about your result.

If 1 cm³ of zinc has a mass of about 7 grams, then a sample that is approximately 2.5 cm³ will have a mass between 2 and 3 times its density, which is consistent with the values given in this problem. Additionally, metals are expected to have a density that is greater than that of water, and zinc's density falls within the range of the other metals listed above (**Table 1.4**).

Since density values are known for many substances, density can be used to determine an unknown mass or an unknown volume. Dimensional analysis will be used to ensure that units cancel appropriately.

Sample Problem 3.6: Using Density to Determine Mass and Volume

1. What is the mass of 2.49 cm³ of aluminum?
2. What is the volume of 50.0 g of aluminum?

Step 1: List the known quantities and plan the problem.

Known

- density = 2.70 g/cm³
- 1. volume = 2.49 cm³
- 2. mass = 50.0 g

Unknown

- 1. mass = ? g
- 2. volume = ? cm³

Use the equation for density, $D = m/V$, and dimensional analysis to solve each problem.

Step 2: Calculate.

$$1. \ 2.49 \text{ cm}^3 \times \frac{2.70 \text{ g}}{1 \text{ cm}^3} = 6.72 \text{ g}$$

$$2. 50.0 \text{ g} \times \frac{1 \text{ cm}^3}{2.70 \text{ g}} = 18.5 \text{ cm}^3$$

In problem one, the mass is equal to the density multiplied by the volume. In problem two, the volume is equal to the mass divided by the density.

Step 3: Think about your results.

Because a 1 cm³ sample of aluminum has a mass of 2.70 g, the mass of a 2.49 cm³ sample should be greater than 2.70 g. The mass of a 50-g block of aluminum is substantially more than the value of its density in g/cm³, so that amount should occupy a volume that is significantly larger than 1 cm³.

Practice Problems

6. A student finds the mass of a “gold” ring to be 41.7 g and its volume to be 3.29 cm³. Calculate the density of the ring. Is it pure gold? (**Table 1.4.**)
7. What is the mass of 125 L of oxygen gas?
8. The density of silver is 10.5 g/cm³. What is the volume of a 13.4 g silver coin?

Finally, conversion problems involving density or other derived units like speed or concentration may involve a separate conversion of each unit. To convert a density in g/cm³ to kg/m³, two steps must be used. One step converts g to kg, while the other converts cm³ to m³. They may be performed in either order.

Sample Problem 3.7: Density Conversions

The average density of the planet Jupiter is 1.33 g/cm³. What is Jupiter’s density in kg/m³?

Step 1: List the known quantities and plan the problem.

Known

- density = 1.33 g/cm³
- 1 kg = 1000 g
- 1 m = 100 cm

Unknown

- density = ? kg/m³

Use separate conversion factors to convert the mass from g to kg and the volume from cm³ to m³.

Step 2: Calculate.

$$\frac{1.33 \text{ g}}{\text{cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left(\frac{100 \text{ cm}}{1 \text{ m}} \right)^3 = 1330 \text{ kg/m}^3$$

Step 3: Think about your result.

Since a cubic meter is so much larger (1 million times) than a cubic centimeter, the density of Jupiter is larger in kg/m³ than in g/cm³.

You can perform a density experiment to identify a mystery object online. Find this simulation at <http://phet.colorado.edu/en/simulation/density> .

Lesson Summary

- Conversion factors are ratios of equivalent quantities expressed in different units. When multiplying by a conversion factor, the numerical value and the unit changes while the actual size of the quantity remains the same.
- Dimensional analysis employs conversion factors to solve problems in which the units are changing. Dimensional analysis can be used to solve metric system conversion problems.
- Density is a derived unit of mass per unit volume and is a physical property of a substance. Density problems can be solved using dimensional analysis.

Lesson Review Questions

Reviewing Concepts

1. What must be true for a ratio of two measurements to be a conversion factor?
2. Which of the following ratios qualify as conversion factors? For the ones that do not, explain why.
 - a. $\frac{10 \text{ pennies}}{1 \text{ dime}}$
 - b. $\frac{3 \text{ dogs}}{\text{several}}$
 - c. $\frac{60 \text{ seconds}}{1 \text{ hour}}$
 - d. $\frac{1 \text{ dozen donuts}}{12 \text{ donuts}}$
3. How do you decide which unit should go in the denominator of a conversion factor?
4. What is a derived unit?
5. Explain what is wrong with this statement: “The density of a heavy bar of pure gold is greater than the density of a small ingot of pure gold.”

Problems

6. Make the following conversions.
 - a. 128 mL to L
 - b. $2.5 \times 10^5 \mu\text{g}$ to g
 - c. 0.481 km to m
 - d. 1890 cm to km
 - e. 6.2×10^{-5} ms to ns
 - f. 75,000 pg to cg
7. Make the following conversions.
 - a. 2800 cm^3 to m^3
 - b. 5.8 g/cm^3 to g/L
 - c. A speed of 60.0 miles per hour to m/s (1 mile = 1608 m)
 - d. A flow rate of 125 mL/min to liters per hour
8. The speed of light is 3.0×10^8 m/s. If the distance from Earth to the Sun is 1.5×10^8 km, how many minutes does it take for light from the Sun to reach Earth?
9. A regular solid has dimensions of 3.20 cm by 4.90 cm by 5.40 cm. The mass of the solid is 235 g. What is its density in g/cm^3 ?

10. What is the mass of a cube of copper that is 1.80 cm on each side? The density of copper is 8.92 g/cm^3 .
11. A balloon is filled with 2300 mL of an unknown gas. The mass of the gas is 3.24 g. Will the balloon float or sink in air?
12. A cube of lead (density = 11.35 g/cm^3) has a mass of 145.7 g. What is the length of each side of the cube in cm?

Further Reading / Supplemental Links

- Barber, Jacqueline; Buegler, Marion; Lowell, Laura, *Discovering Density*. GEMS –Regents of the Univ of CA, 1998.
- Huggins, Shawna, *How to Make a Lava Lamp: Exploring Density*. Blue and Green Wonders Publishing, 2011.
- Fun With Dimensional Analysis: <http://www.alysion.org/dimensional/fun.htm>
- Can you make a golf ball float on water? Find out at www.youtube.com/watch?v=cXLTSLa3yYs .
- Watch the density of water change when heated at <http://www.youtube.com/watch?v=GThJTncPf0> .
- Density differences in liquids and gases can cause convection currents. You can see this at <http://www.youtube.com/watch?v=ovSMAujQbz4> .
- Test your skills with a unit conversion hangman game at http://education.jlab.org/vocabhangman/measuremnt_04/1.html .

Points to Consider

Measurements all must have a certain amount of uncertainty in them, since no measuring tool is 100% accurate. The uncertainty in a measurement must be considered both when reporting measured values and when doing calculations.

- How is the uncertainty in a given measurement indicated in the reported value?
- When a quantity such as density is calculated from two measurements (mass and volume), is it important to measure both accurately or is just one sufficient?
- What are the meanings of the terms precision and accuracy when dealing with measurements?

1.3 Uncertainty in Measurements

Lesson Objectives

- Distinguish between accuracy and precision in measurements.
- Calculate the percent error of a measured quantity.
- Report measured values to the correct number of significant figures based on the measuring tool.
- Perform calculations with measured quantities, rounding the answers to the correct number of significant figures.

Lesson Vocabulary

- accepted value
- accuracy
- error
- experimental value
- percent error
- precision
- significant figures

Check Your Understanding

Recalling Prior Knowledge

- Suppose that a baseball pitcher throws very accurately. What does that mean? Is it different from throwing very precisely?
- When you make a measurement with a specific measuring tool, how well can you read that measurement? Will individual measurements have an affect on quantities that are calculated from those measurements?

When making a measurement, there is always going to be some uncertainty. Some of that uncertainty is related to the reliability of the measuring tool, while some of it is related to the skill of the measurer. When you are performing measurements, you should always strive for the greatest accuracy and precision that you possibly can (**Figure 1.7**).

Accuracy and Precision

In everyday speech, the terms accuracy and precision are frequently used interchangeably. However, their scientific meanings are quite different. **Accuracy** is a measure of how close a measurement is to the correct or accepted value



FIGURE 1.7

Students in a chemistry lab are separating a solution into volumetric flasks, which are each calibrated to hold a specific volume that is known to a high level of accuracy. The long thin stems of these flasks are designed to increase the precision with which a sample of this known volume can be produced.

of the quantity being measured. **Precision** is a measure of how close a series of measurements are to one another. Precise measurements are highly reproducible, even if the measurements are not near the correct value.

Darts thrown at a dartboard are helpful in illustrating accuracy and precision (**Figure 1.8**).

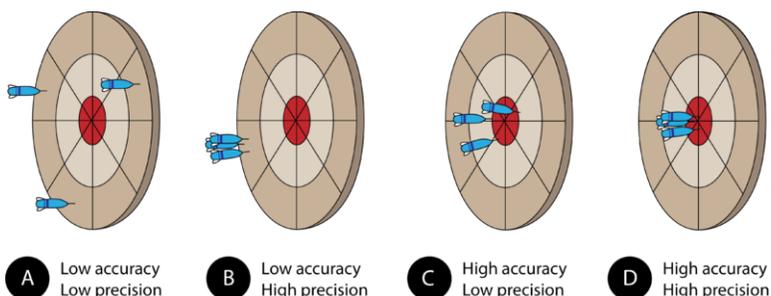


FIGURE 1.8

The distribution of darts on a dartboard shows the difference between accuracy and precision.

Assume that three darts are thrown at the dartboard, with the bulls-eye representing the true, or accepted, value of what is being measured. A dart that hits the bulls-eye is highly accurate, whereas a dart that lands far away from the bulls-eye displays poor accuracy. Pictured above are the four possible outcomes (**Figure 1.8**).

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

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

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

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

Percent Error

An individual measurement may be accurate or inaccurate, depending on how close it is to the true value. Suppose that you are doing an experiment to determine the density of a sample of aluminum metal. The **accepted value** of a measurement is the true or correct value based on general agreement with a reliable reference. For aluminum, the accepted density is 2.70 g/cm^3 . The **experimental value** of a measurement is the value that is measured during the experiment. Suppose that in your experiment you determine an experimental value of 2.42 g/cm^3 for the density of aluminum. The **error** of an experiment is the difference between the experimental and accepted values.

$$\text{Error} = \text{experimental value} - \text{accepted value}$$

If the experimental value is less than the accepted value, the error is negative. If the experimental value is larger than the accepted value, the error is positive. Often, error is reported as the absolute value of the difference in order to avoid the confusion of a negative error. The **percent error** is the absolute value of the error divided by the accepted value and multiplied by 100%.

$$\text{Percent Error} = \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

To calculate the percent error for the aluminum density measurement, we can substitute the given values of 2.45 g/cm^3 for the experimental value and 2.70 g/cm^3 for the accepted value.

$$\text{Percent Error} = \frac{|2.45 \text{ g/cm}^3 - 2.70 \text{ g/cm}^3|}{2.70 \text{ g/cm}^3} \times 100\% = 9.26\%$$

If the experimental value is equal to the accepted value, the percent error is equal to 0. As the accuracy of a measurement decreases, the percent error of that measurement rises.

Significant Figures in Measurements

Uncertainty

Some error or uncertainty always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest 0.1 g, other highly sensitive balances are capable of measuring to the nearest 0.001 g or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Pictured below is an object (indicated by the blue arrow) whose length is being measured by two different rulers (**Figure 1.9**).

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

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

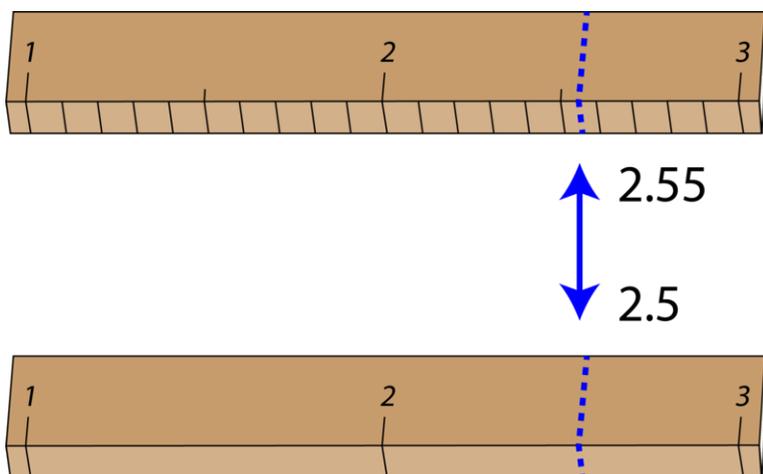


FIGURE 1.9

Again, another measurer may report the length to be 2.54 cm or 2.56 cm. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

Determining Significant Figures

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler example, the bottom ruler gave a length with two significant figures, while the top ruler gave a length with three significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. It would not be correct to report the length as 2.553 cm with either ruler because there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below (**Table 1.5**) details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

TABLE 1.5: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant	A. 237 has three significant figures. B. 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits are always significant.	A. 39,004 has five significant figures. B. 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called left-end zeros. Left-end zeros are never significant.	A. 0.008 has one significant figure. B. 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called right-end zeros. Right-end zeros in a number that lacks a decimal point are not significant.	A. 140 has two significant figures. B. 75,210 has four significant figures.
5. Right-end zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	A. 620.0 has four significant figures. B. 19,000. has five significant figures

It needs to be emphasized that just because a certain digit is not significant does not mean that it is not important or that it can be left out. Though the zero in a measurement of 140 may not be significant, the value cannot simply be reported as 14. An insignificant zero functions as a placeholder for the decimal point. When numbers are written in scientific notation, this becomes more apparent. The measurement 140 can be written as 1.4×10^2 , with two

when we begin looking at how significant figures are dealt with during calculations. Numbers in many conversion factors, especially for simple unit conversions, are also exact quantities and have infinite significant figures. There are exactly 100 centimeters in 1 meter and exactly 60 seconds in 1 minute. Those values are definitions and are not the result of a measurement.

Sample Problem 3.8: Counting Significant Figures

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

1. 19.5 m
2. 0.0051 L
3. 204.80 g
4. 1.90×10^5 s
5. 14 beakers
6. 700 kg

Step 1: Plan the problem.

Follow the rules for counting the number of significant figures in a measurement, paying special attention to the location of zeros in each. Note each rule that applies according to the table above (**Table 1.5**).

Step 2: Solve.

1. three (rule one)
2. two (rule three)
3. five (rules two five)
4. three (rule five)
5. infinite
6. one (rule four)

The 14 beakers is a counted set of items and not a measurement, so it has an infinite number of significant figures.

Practice Problems

1. Count the number of significant figures in each measurement.
 - a. 0.00090 L
 - b. 255 baseballs
 - c. 435,210 m
 - d. 40.1 kg
 - e. 9.026×10^{-6} mm
 - f. 12.40°C

Significant Figures in Calculations

Many reported quantities in science are the result of calculations involving two or more measurements. Density involves mass and volume, both of which are measured quantities. As an example, say that you have a precise balance that gives the mass of a certain object as 21.513 g. However, the volume is measured very roughly by water displacement, using a graduated cylinder that can only be read to the nearest tenth of a milliliter. The volume of the object is determined to be 8.2 mL. On a calculator, the density (mass divided by volume) would come out as

2.623536585 g/mL. Hopefully, it should be apparent that the calculator is giving us far more digits than we actually can be certain of knowing. In fact, the density should be reported as 2.6 g/mL. This is because the result of a calculated answer can be no more precise than the least precise measurement from which it was calculated. Since the volume was known only to two significant figures, the resultant density needs to be rounded to two significant figures.

Rounding

Before we get to the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To round a number, first decide how many significant figures the number should have. Once you know that, round to the correct number of digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped, and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement 207.518 m. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process listed below (**Table 1.6**).

TABLE 1.6: Significant Figure Rounding

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Significant Figures in Addition and Subtraction

Consider two separate mass measurements: 16.7 g and 5.24 g. The first mass measurement (16.7 g) is known only to the tenths place, which is one digit after the decimal point. There is no information about its hundredths place, so that digit cannot be assumed to be zero. The second measurement (5.24 g) is known to the hundredths place, which is two digits after the decimal point.

When these masses are added together, the result on a calculator is $16.7 + 5.24 = 21.94$ g. Reporting the answer as 21.94 g suggests that the sum is known all the way to the hundredths place. However that cannot be true because the hundredths place of the first mass was completely unknown. The calculated answer needs to be rounded in such a way as to reflect the certainty of each of the measured values that contributed to it. *For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the lowest number of decimal places.* The sum of the above masses would be properly rounded to a result of 21.9 g.

When working with whole numbers, pay attention to the last significant digit that is to the left of the decimal point and round your answer to that same point. For example, consider the subtraction problem 78,500 m $-$ 362 m. The calculated result is 78,138 m. However, the first measurement is known only to the hundreds place, as the 5 is the last significant digit. Rounding the result to that same point means that the final calculated value should be reported as 78,100 m.

Significant Figures in Multiplication and Division

The density of a certain object is calculated by dividing the mass by the volume. Suppose that a mass of 37.46 g is divided by a volume of 12.7 cm³. The result on a calculator would be:

$$D = \frac{m}{V} = \frac{37.46 \text{ g}}{12.7 \text{ cm}^3} = 2.94960299 \text{ g/cm}^3$$

The value of the mass measurement has four significant figures, while the value of the volume measurement has only three significant figures. *For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the lowest number of significant figures.* Applying this rule results in a density of 2.95 g/cm³, which has three significant figures—the same as the volume measurement.

Note that the rule for multiplication and division problems is different than the rule for addition and subtraction problems. For multiplication and division, it is the number of significant figures that must be considered. For addition and subtraction, it is the position of the decimal place that determines the correct rounding. Review the sample problem below, paying special attention to this distinction.

Sample Problem 3.9: Significant Figures in Calculations

Perform the following calculations, rounding the answers to the appropriate number of significant figures.

1. 0.048 m × 32.97 m
2. 21.9 g – 19.417 g
3. 14,570 kg ÷ 5.81 L
4. 71.2 cm + 90 cm

Step 1: Plan the problem.

Decide which calculation rule applies. Analyze each of the measured values to determine how many significant figures should be in the result. Perform the calculation and round appropriately. Apply the correct units to the answer. When adding or subtracting, the units in each measurement must be identical and then remain the same in the result. When multiplying or dividing, the units are also multiplied or divided.

Step 2: Calculate.

1. 0.048 m × 32.97 m = 1.6 m² - Round to two significant figures because 0.048 has two.
2. 21.9 g - 19.417 g = 2.5 g - Answer ends at the tenths place because of 21.9.
3. 14,570 kg ÷ 5.81 L = 2510 kg/L - Round to three significant figures because 5.81 has three.
4. 71.2 cm + 90 cm = 160 cm - Answer ends at the tens place because of 90.

Practice Problems

2. Solve each problem, rounding each answer to the correct number of significant figures.
 - (a) 132.3 g ÷ 29.600 mL
 - (b) 3.27 g/cm³ × 0.086 cm³
 - (c) 125 m + 61.3 m + 310 m
 - (d) 3.0 × 10⁴ L – 1244 L
3. A rectangular prism has dimensions of 3.7 cm by 4.81 cm by 1.90 cm. The mass of the prism is 49.72 g. Calculate its density. (Hint: Do the entire calculation without rounding until the final answer in order to reduce “rounding error.”)

Lesson Summary

- Accuracy refers to how close a measured value is to the accepted value, whereas precision indicates how close individual measurements within a set are to each other.
- Percent error is the difference between the experimental and accepted values divided by the accepted value and multiplied by 100.
- The measuring tool dictates how many significant figures can be reported in a measurement. Significant figures include all of the certain digits plus one uncertain digit. A set of rules is followed for determining the number of significant figures in numbers that contain zeros. Counted quantities have infinite significant figures.
- For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the lowest number of decimal places. For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the lowest number of significant figures.

Lesson Review Concepts

Reviewing Concepts

1. The density of a sample of copper metal was determined by three different students (**Table 1.7**). Each performed the measurement three times. Describe the accuracy and precision of each student's measurements. The accepted value for the density of copper is 8.92 g/cm^3 .

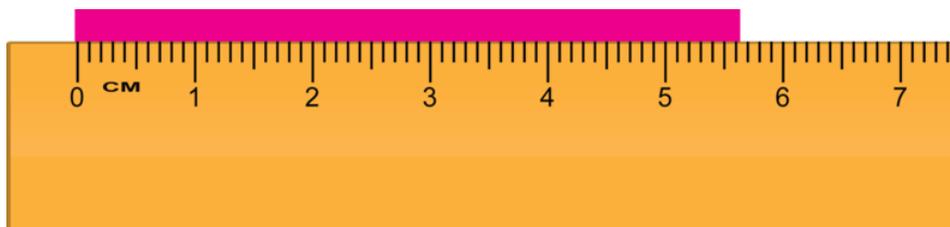
TABLE 1.7: Density of copper (g/cm)

Student	Trial 1	Trial 2	Trial 3
Jane	8.94	8.89	8.91
Justin	8.32	8.31	8.34
Julia	8.64	9.71	9.13

2. What is wrong with the following statement? "My measurement of 8.45 m for the width of the room is very precise."
3. Consider the following 5 mL graduated cylinders, which contain identical quantities of liquid. Which cylinder yields a measurement with a greater number of significant figures? How many significant figures can be reported for each cylinder?



4. Report the length measurement of the pink bar to the correct number of significant figures. Which digits in your measurement are certain? Which are uncertain?



5. How many significant figures are in each of the following measurements?
- 9 potatoes
 - 4.05 cm
 - 0.0061 kg
 - 50 mL
 - $8.00 \times 10^9 \mu\text{g}$
 - 720.00 s
6. Round the measured quantity of 31.0753 g to each of the following amounts of significant figures.
- five
 - four
 - three
 - two
 - one

Problems

7. Kyle measures the mass of a solid sample to be 8.09 g. The accepted value for the mass is 8.42 g. Calculate Kyle's percent error.

8. Jamelle performs three separate determinations of the density of a mineral sample. She gets values of 4.58 g/cm³, 4.79 g/cm³, and 4.55 g/cm³.
- Calculate the average value of the density of the mineral.
 - The deviation of a measured value is defined as the absolute value of the difference between the measured value and the average value: Deviation = |measured value – average value|. Calculate the deviation for each of the three measurements. According to the deviations, which measurement appears to be poorest compared to the others?
 - The average deviation is the sum of all the deviations divided by the total number of measurements. Calculate the average deviation of the three measurements.
 - When average deviation is high, does that indicate good precision or poor precision in the measurements? Explain.
9. Convert a measurement of 2.75 hours to seconds. Are the required conversion factors measured or exact quantities?
10. Convert 466.84 cm to inches, given that 1 inch = 2.54 cm. The conversion factor between centimeters and inches is a measured quantity.
11. Perform the following calculations and round to the correct number of significant figures.
- $78.2 \text{ g} \div 32 \text{ cm}^3$
 - $3.0 \text{ m/s} \times 9.21 \text{ s}$
 - $59 \text{ g} + 4 \text{ g} + 0.79 \text{ g}$
 - $34,000 \text{ km} - 2430 \text{ km}$
 - $(9.59 \text{ g} + 1.098 \text{ g}) \div 2.313 \text{ mL}$

Further Reading / Supplemental Links

- Tutorial on the Use of Significant Figures: <http://www.chem.sc.edu/faculty/morgan/resources/sigfigs/index.html>
- Another tutorial video can be found at <http://www.youtube.com/watch?v=ctj07mSIJ0w> .
- Significant Figures Calculator: <http://calculator.sig-figs.com/>
- Winter, Paul K., *Significant Figures*. International Textbook Company, 1965.

Points to Consider

Measurements will be a constant consideration throughout your study of chemistry. Next you will begin a study of the atom, its component parts, and the evolution of the atomic model.

- Atoms are extremely small and extremely light. How do you think that the mass and size of an atom can be measured? Do you think the accuracy of these measurements has improved over time?

1.4 References

1. Courtesy of NASA/SDO. http://commons.wikimedia.org/wiki/File:The_Sun_by_the_Atmospheric_Imaging_-_Assembly_of_NASA%27s_Solar_Dynamics_Observatory_-_20100819.jpg . Public Domain
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