2.1 Physical Quantities

Physical properties such as height, volume, and temperature that can be measured are called **physical quantities**. Both a number and a unit of defined size is required to describe physical quantity.

Number → 61.2 ← Unit

61.2 kilograms

- A number without a unit is meaningless.
- To avoid confusion, scientists have agreed on a standard set of units.
- Scientists use SI or the closely related metric units.

TABLE 2.1 Some SI and Metric Units and Their Equivalents

QUANTITY	SI UNIT (SYMBOL)	METRIC UNIT (SYMBOL)	EQUIVALENTS
Mass	Kilogram (kg)	Gram (g)	1 kg = 1000 g = 2.205 lb
Length	Meter (m)	Meter (m)	1 m = 3.280 ft
Volume	Cubic meter (m^3)	Liter (L)	1 m^3 = 1000 L = 264.2 gal
Temperature	Kelvin (K)	Celsius degree (°C)	1 K = 1°C 1°C = 1.8°F
Time	Second (s)	Second (s)	—

Frequently used prefixes are shown below.

TABLE 2.2 Some Prefixes for Multiples of Metric and SI Units

PREFIX	SYMBOL	BASE UNIT MULTIPLIED BY*	EXAMPLE
mega	M	1,000,000 = 10^6	1 megameter (Mm) = 10^6 m
kilo	k	1,000 = 10^3	1 kilogram (kg) = 10^3 g
hecto	h	100 = 10^2	1 hectogram (hg) = 100 g
deka	da	10 = 10^1	1 dekaliter (daL) = 10 L
deci	d	0.1 = 10^{-1}	1 deciliter (dL) = 0.1 L
centi	c	0.01 = 10^{-2}	1 centimeter (cm) = 0.01 cm
milli	m	0.001 = 10^{-3}	1 milligram (mg) = 0.001 g
micro	μ	0.000 001 = 10^{-6}	1 micrometer (μm) = 10^{-6} m
nano	n	0.000 000 001 = 10^{-9}	1 nanogram (ng) = 10^{-9} g
pico	p	0.000 000 000 001 = 10^{-12}	1 picogram (pg) = 10^{-12} g

The scientific notation method of writing large and small numbers (for example, 10^6 for 1,000,000) is explained in Section 2.5.

2.2 Measuring Mass

- **Mass** is a measure of the amount of matter in an object. Mass does not depend on location.
- **Weight** is a measure of the gravitational force acting on an object. Weight depends on location.
- A scale responds to weight.
- At the same location, two objects with identical masses have identical weights.
- The mass of an object can be determined by comparing the weight of the object to the weight of a reference standard of known mass.

Relationships between metric units of mass and the mass units commonly used in the United States are shown below.

TABLE 2.3 Units of Mass

UNIT	EQUIVALENT	UNIT	EQUIVALENT
1 kilogram (kg)	= 1000 grams = 2.205 pounds	1 ton	= 2000 pounds = 907.03 kilograms
1 gram (g)	= 0.001 kilogram = 1000 milligrams = 0.035 27 ounce	1 pound (lb)	= 16 ounces = 0.454 kilogram = 454 grams
1 milligram (mg)	= 0.001 gram = 1000 micrograms	1 ounce (oz)	= 0.028 35 kilogram = 28.35 grams = 28,350 milligrams
1 microgram (μg)	= 0.000 001 gram = 0.001 milligram		

2.3 Measuring Length and Volume

- The meter (m) is the standard measure of length or distance in both the SI and the metric system.
- Volume is the amount of space occupied by an object. A volume can be described as a length^3.
- The SI unit for volume is the cubic meter (m^3).

Relationships between metric units of length and volume and the length and volume units commonly used in the United States are shown below and on the next slide.

TABLE 2.4 Units of Length

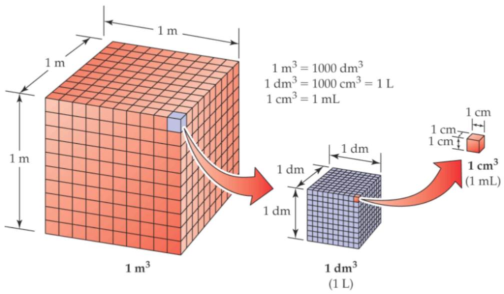
UNIT	EQUIVALENT	UNIT	EQUIVALENT
1 kilometer (km)	= 1000 meters = 0.6214 mile	1 mile (mi)	= 1.609 kilometers = 1609 meters
1 meter (m)	= 100 centimeters = 1000 millimeters = 1.0936 yards = 39.37 inches	1 yard (yd)	= 0.9144 meter = 91.44 centimeters
		1 foot (ft)	= 0.3048 meter = 30.48 centimeters
1 centimeter (cm)	= 0.01 meter = 10 millimeters = 0.3937 inch	1 inch (in.)	= 2.54 centimeters = 25.4 millimeters
1 millimeter (mm)	= 0.001 meter = 0.1 centimeter		

A m^3 is the volume of a cube 1 m or 10 dm on edge. Each m^3 contains $(10\ dm)^3 = 1000\ dm^3$ or liters. Each liter or $dm^3 = (10cm)^3 = 1000\ cm^3$ or milliliters. Thus, there are 1000 mL in a liter and 1000 L in a m^3.

TABLE 2.5 Units of Volume

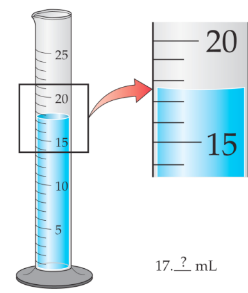
UNIT	EQUIVALENT	UNIT	EQUIVALENT
1 cubic meter (m^3)	= 1000 liters = 264.2 gallons	1 gallon (gal)	= 3.7854 liters
		1 quart (qt)	= 0.9464 liter = 946.4 milliliters
1 liter (L)	= 0.001 cubic meter = 1000 milliliters = 1.057 quarts	1 fluid ounce (fl oz)	= 29.57 milliliters
1 deciliter (dL)	= 0.1 liter = 100 milliliters		
1 milliliter (mL)	= 0.001 liter = 1000 microliters		
1 microliter (μL)	= 0.001 milliliter		

The metric system is based on factors of 10 and is much easier to use than common U.S. units. Does anyone know how many teaspoons are in a gallon?



2.4 Measurement and Significant Figures

- Every experimental measurement has a degree of uncertainty.
- The volume, V, at right is certain in the 10's place, 10mL<V<20mL
- The 1's digit is also certain, 17mL<V<18mL
- A best guess is needed for the tenths place.



- To indicate the precision of a measurement, the value recorded should use all the digits known with certainty, plus one additional estimated digit that usually is considered uncertain by plus or minus 1.
- No further insignificant digits should be recorded.
- The total number of digits used to express such a measurement is called the number of **significant figures**.
- All but one of the significant figures are known with certainty. The last significant figure is only the best possible estimate.

Below are two measurements of the mass of the same object. The same quantity is being described at two different levels of precision or certainty.

Uncertain digit

54.07 g A mass between 54.06 g and 54.08 g (±0.01 g)

Uncertain digit

54.071 38 g A mass between 54.071 37 g and 54.071 39 g (±0.000 01 g)

- When reading a measured value, all nonzero digits should be counted as significant. There is a set of rules for determining if a zero in a measurement is significant or not.
- **RULE 1.** Zeros in the middle of a number are like any other digit; they are always significant. Thus, 94.072 g has five significant figures.
- **RULE 2.** Zeros at the beginning of a number are not significant; they act only to locate the decimal point. Thus, 0.0834 cm has three significant figures, and 0.029 07 mL has four.

- **RULE 3.** Zeros at the end of a number and *after* the decimal point are significant. It is assumed that these zeros would not be shown unless they were significant. 138.200 m has six significant figures. If the value were known to only four significant figures, we would write 138.2 m.
- **RULE 4.** Zeros at the end of a number and *before* an implied decimal point may or may not be significant. We cannot tell whether they are part of the measurement or whether they act only to locate the unwritten but implied decimal point.

2.5 Scientific Notation

- **Scientific notation** is a convenient way to write a very small or a very large number.
- Numbers are written as a product of a number between 1 and 10, times the number 10 raised to power.
- 215 is written in scientific notation as:

$215 = 2.15 \times 100 = 2.15 \times (10 \times 10) = 2.15 \times 10^2$

Two examples of converting standard notation to scientific notation are shown below.

$$215. = 2.15 \times 10^2$$

Decimal point is moved two places to the left, so exponent is 2.

$$0.002\,15 = 2.15 \times \frac{1}{1000} = 2.15 \times \frac{1}{10 \times 10 \times 10} = 2.15 \times \frac{1}{10^3} = 2.15 \times 10^{-3}$$

$$0.002\,15 = 2.15 \times 10^{-3}$$

Decimal point is moved three places to the right, so exponent is −3.

Two examples of converting scientific notation back to standard notation are shown below.

$$3.7962 \times 10^4 = 37{,}962$$

Positive exponent of 4, so decimal point is moved to the right four places.

$$1.56 \times 10^{-8} = 0.000\,000\,015\,6$$

Negative exponent of –8, so decimal point is moved to the left eight places.

- Scientific notation is helpful for indicating how many significant figures are present in a number that has zeros at the end but to the left of a decimal point.
- The distance from the Earth to the Sun is 150,000,000 km. Written in standard notation this number could have anywhere from 2 to 9 significant figures.
- Scientific notation can indicate how many digits are significant. Writing 150,000,000 as 1.5×10^8 indicates 2 and writing it as 1.500×10^8 indicates 4.
- Scientific notation can make doing arithmetic easier. Rules for doing arithmetic with numbers written in scientific notation are reviewed in Appendix A.

2.6 Rounding Off Numbers

- Often when doing arithmetic on a pocket calculator, the answer is displayed with more significant figures than are really justified.
- How do you decide how many digits to keep?
- Simple rules exist to tell you how.

RULE 1. In carrying out a multiplication or division, the answer cannot have more significant figures than either of the original numbers.

Three significant figures

Three significant figures

$$\frac{278\ \text{mi}}{11.70\ \text{gal}} = 23.8\ \text{mi/gal}$$

Four significant figures

RULE 2. In carrying out an addition or subtraction, the answer cannot have more digits after the decimal point than either of the original numbers.

Volume of water at start →	3.18? ?? L	← Two digits after decimal point
Volume of water addded →	+ 0.013 15 L	← Five digits after decimal point
Total volume of water →	3.19? ?? L	← Two digits after decimal point

- Once you decide how many digits to retain, the rules for rounding off numbers are straightforward:
- **RULE 1.** If the first digit you remove is 4 or less, drop it and all following digits. 2.4271 becomes 2.4 when rounded off to two significant figures because the first dropped digit (a 2) is 4 or less.
- **RULE 2.** If the first digit removed is 5 or greater, round up by adding 1 to the last digit kept. 4.5832 is 4.6 when rounded off to 2 significant figures since the first dropped digit (an 8) is 5 or greater.
- If a calculation has several steps, it is best to round off at the end.

2.7 Problem Solving: Converting a Quantity from One Unit to Another

- **Factor-Label Method:** A quantity in one unit is converted to an equivalent quantity in a different unit by using a conversion factor that expresses the relationship between units.

(Starting quantity) x (Conversion factor) = Equivalent quantity

Writing 1 km = 0.6214 mi as a fraction restates it in the form of a conversion factor. This and all other conversion factors are numerically equal to 1.

Conversion factors between kilometers and miles

$$\frac{1\ \text{km}}{0.6214\ \text{mi}} = 1 \quad \text{or} \quad \frac{0.6214\ \text{mi}}{1\ \text{km}} = 1$$

The numerator is equal to the denominator. Multiplying by a conversion factor is equivalent to multiplying by 1 and so causes no change in value.

These two quantities are the same.

$$\frac{1\ \text{km}}{0.6214\ \text{mi}} \quad \text{or} \quad \frac{0.6214\ \text{mi}}{1\ \text{km}}$$

These two quantities are the same.

When solving a problem, the idea is to set up an equation so that all unwanted units cancel, leaving only the desired units.

$$26.22\ \cancel{\text{mi}} \times \frac{1\ \text{km}}{0.6214\ \cancel{\text{mi}}} = 42.20\ \text{km}$$

Starting quantity — Conversion factor — Equivalent quantity

2.8 Problem Solving: Estimating Answers

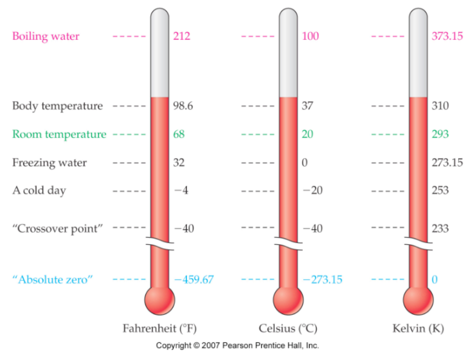
- **STEP 1:** Identify the information given.
- **STEP 2:** Identify the information needed to answer.
- **STEP 3:** Find the relationship(s) between the known information and unknown answer, and plan a series of steps, including conversion factors, for getting from one to the other.
- **STEP 4:** Solve the problem.
- **BALLPARK CHECK:** Make a rough estimate to be sure the value and the units of your calculated answer are reasonable.

2.9 Measuring Temperature

- Temperature is commonly reported either in degrees Fahrenheit (°F) or degrees Celsius (°C).
- The SI unit of temperature is the Kelvin (K).
- 1 Kelvin, no degree, is the same size as 1 °C.
- 0 K is the lowest possible temperature, 0 °C = 273.15 K is the normal freezing point of water. To convert, adjust for the zero offset.
- Temperature in K = Temperature in °C + 273.15
- Temperature in °C = Temperature in K - 273.15

Freezing point of H_2O	Boiling point of H_2O
32°F	212°F
0°C	100°C

212°F - 32°F = 180°F covers the same range of temperature as 100°C - 0°C = 100°C covers. Therefore, a Celsius degree is exactly 180/100 = 1.8 times as large as a Fahrenheit degree. The zeros on the two scales are separated by 32°F.



Fahrenheit, Celsius, and Kelvin temperature scales.

2.10 Energy and Heat

- **Energy:** The capacity to do work or supply heat.
- Energy is measured in SI units by the **Joule** (J); the calorie is another unit often used to measure energy.
- One **calorie** (cal) is the amount of heat necessary to raise the temperature of 1 g of water by 1°C.
- A kilocalorie (kcal) = 1000 cal. A **Calorie**, with a capital C, used by nutritionists, equals 1000 cal.
- An important energy conversion factor is:

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

- Not all substances have their temperatures raised to the same extent when equal amounts of heat energy are added.
- One calorie raises the temperature of 1 g of water by 1°C but raises the temperature of 1 g of iron by 10°C.
- The amount of heat needed to raise the temperature of 1 g of a substance by 1°C is called the **specific heat** of the substance.
- Specific heat is measured in units of cal/g°C

- Knowing the mass and specific heat of a substance makes it possible to calculate how much heat must be added or removed to accomplish a given temperature change.
- (Heat Change) = (Mass) x (Specific Heat) x (Temperature Change)
- Using the symbols Δ for change, H for heat, m for mass, C for specific heat, and T for temperature, a more compact form is:

$$\Delta H = mC\Delta T$$

2.11 Density

Density relates the mass of an object to its volume. Density is usually expressed in units of grams per cubic centimeter (g/cm^3) for solids, and grams per milliliter (g/mL) for liquids.

$$\text{Density} = \frac{\text{Mass (g)}}{\text{Volume (mL or cm}^3\text{)}}$$

- Which is heavier, a ton of feathers or a ton of bricks?
- Which is larger?
- If two objects have the same mass, the one with the higher density will be smaller.

2. 12 Specific Gravity

Specific gravity (sp gr): density of a substance divided by the density of water at the same temperature. Specific gravity is unitless. The density of water is so close to 1 g/mL that the specific gravity of a substance at normal temperature is numerically equal to the density.

$$\text{Specific gravity} = \frac{\text{Density of substance } (\cancel{\text{g/ml}})}{\text{Density of water at the same temperature } (\cancel{\text{g/ml}})}$$