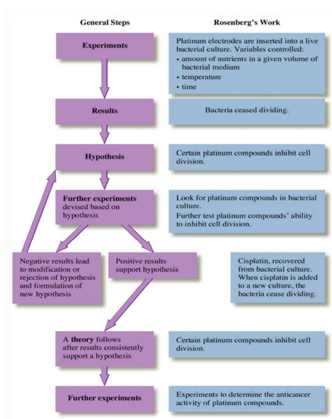
Chapter 1

Chemical Foundations

Steps in the Scientific Method

- 1. Observations
  - quantitative
  - qualitative
- 2. Formulating hypotheses
  - possible explanation for the observation
- 3. Performing experiments
  - gathering new information to decide whether the hypothesis is valid

Figure 1.8: A representation of the scientific method.



Outcomes Over the Long-Term

- Theory (Model)
  - A set of tested hypotheses that give an overall explanation of some natural phenomenon.
- Natural Law
  - The same observation applies to many different systems
- Example - Law of Conservation of Mass

Law vs. Theory

- A law summarizes what happens;
- A theory (model) is an attempt to explain why it happens.

Nature of Measurement

- Measurement - quantitative observation consisting of 2 parts
  - Part 1 - number
  - Part 2 - scale (unit)
- Examples:
  - 20 grams
  - 6.63×10^{-34} Joule seconds

International System (le Système International)

- Based on metric system and units derived from metric system.

Table 1.2
SI Base Units

Quantity	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	cd

TABLE 1.2 The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

TABLE 1.3
Some Examples of Commonly Used Units

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

TABLE 1.4
English–Metric Equivalents

Length	1 m = 1.094 yd 2.54 cm = 1 in
Mass	1 kg = 2.205 lb 453.6 g = 1 lb
Volume	1 L = 1.06 qt 1 ft^3 = 28.32 L

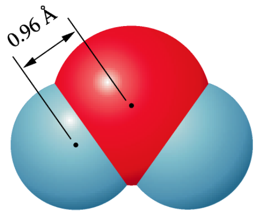
Dimensional Analysis

Proper use of "unit factors" leads to proper units in your answer.

A 1.300 Kg infant weighs how many pounds?

$$1.300\ \cancel{Kg} \times \frac{2.20\ lbs}{\cancel{Kg}} = 2.86\ lbs$$

1.76 Water consists of molecules (groups of atoms). A water molecule has two hydrogen atoms, each connected to an oxygen atom. The distance between any one hydrogen atom and the oxygen atom is 0.96 Å. What is this distance in millimeters?



1 Å = 10^{-10} m, so

(0.96 Å) (10^{-10} m/ Å) = 0.96 x 10^{-10} m

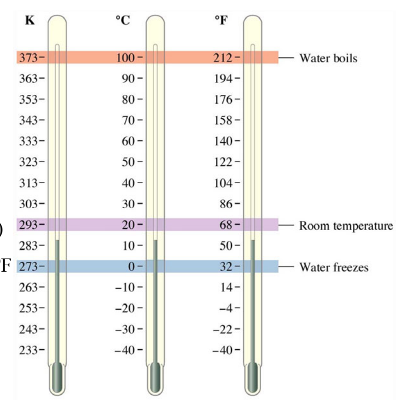
(0.96 x 10^{-10} m)(10^3 mm/m) =

0.96 x 10^{-7} mm or 9.6 x 10^{-8} mm

Temperature

- Celsius scale = °C
- Kelvin scale = K
- Fahrenheit scale = °F

Figure 1.23: Comparison of temperature scales.

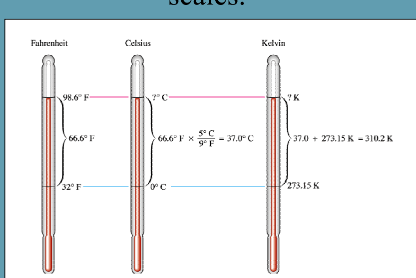


^{O}C = 5/9(^{O}F - 32 ^{O}F)

^{O}F = 9/5 (^{O}C) + 32 ^{O}F

K = ^{O}C (1 K/ ^{O}C) + 273.13 K

Figure 1.12: Normal body temperature on the Fahrenheit, Celsius, and Kelvin scales.



Derived Units

- Units which are composed of two or more single dimension units (mass, time, length)

Examples: area units (length x length) or 1.0 m^2

volume units (length x length x length) or 1 cm^3

density units (mass/unit volume) or g/cm^3

Figure 1.6: Measurement of volume

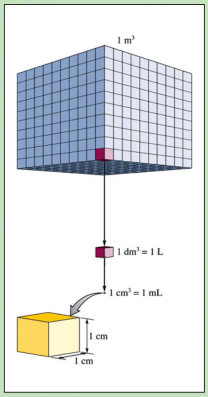


Figure 1.7: Common types of laboratory equipment used to measure liquid volume.

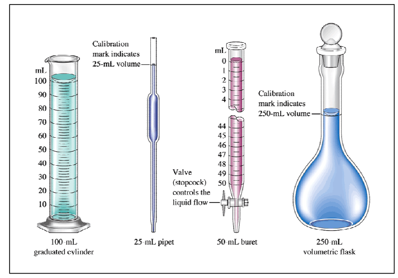


Figure 1.24: Some laboratory glassware. Photo courtesy of James Scherer.

Figure 1.25: The relative densities of copper and mercury. Photo courtesy of James Scherer.

Density = mass/unit volume

$d = m/V$

Figure 1.26: The relative densities of some liquids. Photo courtesy of American Color.

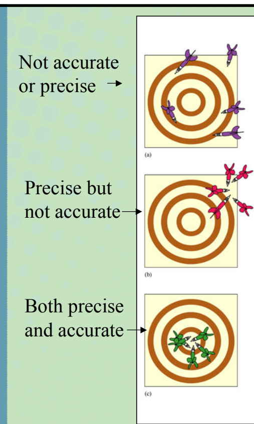
Uncertainty in Measurement

- A digit that must be estimated is called uncertain. A measurement always has some degree of uncertainty.

Precision and Accuracy

- Accuracy refers to the agreement of a particular value with the true value.
- Precision refers to the degree of agreement among several elements of the same quantity.

Figure 1.10: The results of several dart throws show the difference between *precise* and *accurate*.



Types of Error

- Random Error (Indeterminate Error)
  - measurement has an equal probability of being high or low.
- Systematic Error (Determinate Error)
  - Occurs in the same direction each time (high or low), often resulting from poor technique.

Types of Error

Why is it important to determine if an error is random or systematic?

1. For a measurement, we can estimate its mean value and uncertainty only when the error is random. In many cases, we cannot decrease random error – but we can estimate it.
2. Often we can correct systematic error. To correct a systematic error in a measurement, we must fix the cause of the error or compensate for it in some way

Figure 1.20: Precision of measurement with a centimeter.

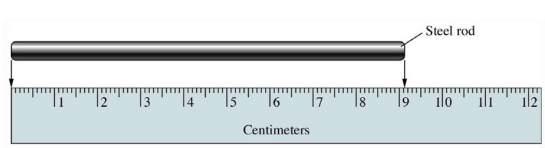
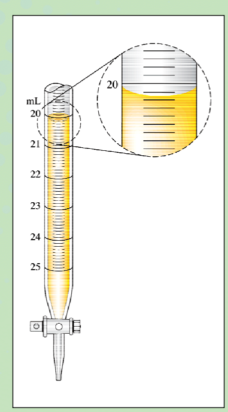


Figure 1.9: Measurement of volume using a buret. The volume is read at the bottom of the liquid curve (called the *meniscus*).



Rules for Counting Significant Figures - Overview

- 1. Nonzero integers
- 2. Zeros
  - leading zeros
  - captive zeros
  - trailing zeros
- 3. Exact numbers

Rules for Counting Significant Figures - Details

- Nonzero integers always count as significant figures.

- 3456 has 4 significant figs.

Rules for Counting Significant Figures - Details

- Zeros
  - Leading zeros do not count as significant figures.

- 0.0486 has 3 significant figs.

Rules for Counting Significant Figures - Details

- Zeros
  - Captive zeros always count as significant figures.

- 16.07 has 4 significant figs.

Rules for Counting Significant Figures - Details

- **Zeros**
  - Trailing zeros are significant only if the number contains a decimal point.

- 9.300 has 4 significant figs.

Rules for Counting Significant Figures - Details

- Exact numbers have an infinite number of significant figures.

- 1 inch = 2.54 cm, exactly

Rules for Significant Figures in Mathematical Operations

- Multiplication and Division: # sign. figs in the result equals the number in the least precise measurement used in the calculation.

- $6.38 \times 2.0 = 12.76 \rightarrow 13$ (2 sign. figs)

Rules for Significant Figures in Mathematical Operations

- Addition and Subtraction: # sig figs in the result is determined using the number of decimal places in the least precise measurement.

- $6.8 + 11.934 = 22.4896 \rightarrow 22.5$ (3 sign. figs)

Figure 1.9: Laboratory balances: A modern single-pan balance. Photo courtesy of American Color.

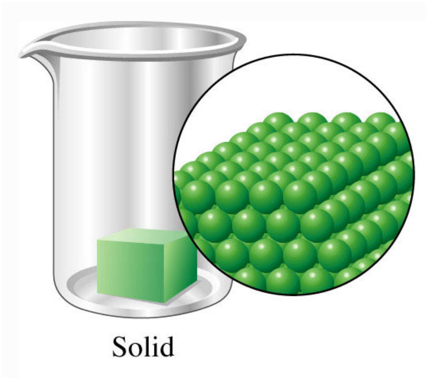
Figure 1.21: Significant figures and calculators. Photo courtesy of American Color.

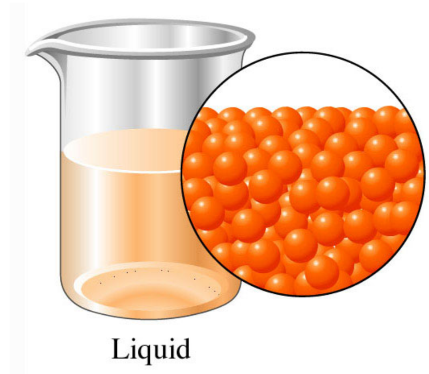
Retain all the significant figures during your calculations. Round to the correct number of significant figures at the end of your calculations.

Matter: Anything occupying space and having mass.

Classification of Matter

- Three States of Matter:
  - Solid: rigid - fixed volume and shape
  - Liquid: definite volume but assumes the shape of its container
  - Gas: no fixed volume or shape - assumes the shape of its container





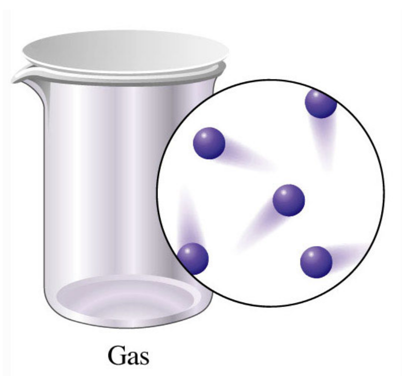


Table 1.1
Characteristics of the States of Matter

State of Matter	Fluidity or Rigidity	Compressibility
Solid	Rigid	Very low
Liquid	Fluid	Very low
Gas	Fluid	High

Pure Substances

- Compound: A substance with a constant composition that can be broken down into elements by chemical processes.
- Element: A substance that cannot be decomposed into simpler substances by chemical means.

Figure 1.14: Elements: sulfur, arsenic, iodine, magnesium, bismuth, mercury. Photo courtesy of American Color.

The element mercury (top left) combines with the element iodine (top right) to form the compound mercuric iodide (bottom). This is an example of a chemical change.

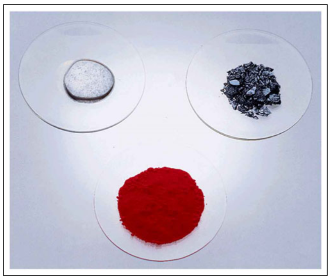
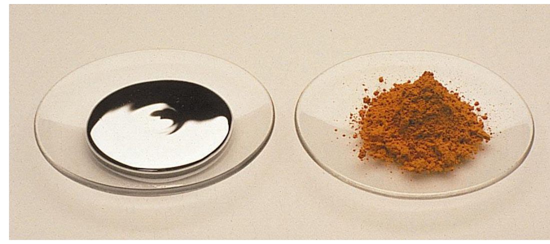


Figure 1.10: Heating mercury metal in air: Mercury metal reacts with Oxygen. Photo courtesy of James Scherer.



$$2\,Hg\,(l) + O_2\,(g) \rightarrow 2\,HgO\,(s)$$

mercury, oxygen, Mercury (II) oxide

Figure 1.10: Heating Mercury in air: Mercury oxide decomposes. Photo courtesy of James Scherer.

$$2\,HgO\,(s) \xrightarrow{heat} 2\,Hg\,(l) + O_2\,(g)$$

Types of Mixtures

- Mixtures have variable composition.
- A homogeneous mixture is a solution (for example, vinegar)
- A heterogeneous mixture is, to the naked eye, clearly not uniform (for example, a bottle of ranch dressing)

Figure 1.3: Sand on a beach looks uniform from a distance, but up close the irregular sand grains are visible.

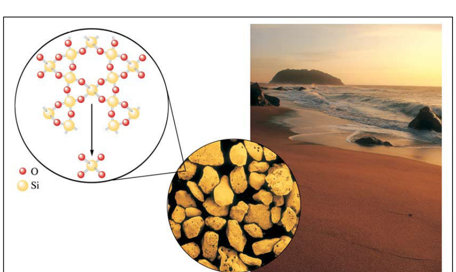


Figure 1.15: A mixture of potassium dichromate and iron fillings. Photo courtesy of James Scherer.

Figure 1.15: A magnet separates the iron filling from the mixture. Photo courtesy of James Scherer.

Mixtures of Pure Substances

- Can be isolated by separation methods:
- Chromatography
- Filtration
- Distillation

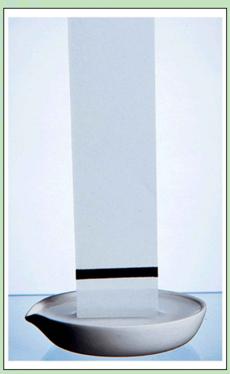


Figure 1.15a: Paper chromatography of ink. (a) A line of the mixture to be separated is placed at one end of a sheet of porous paper.

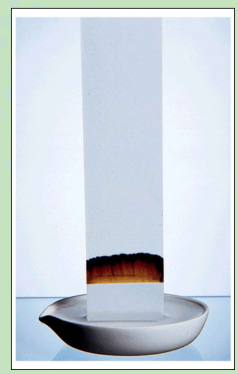


Figure 1.15b: Paper chromatography of ink. (b) The paper acts as a wick to draw up the liquid.

Figure 1.15c: Paper chromatography of ink. (c) The component with the weakest attraction for the paper travels faster than the components that cling to the paper.

Figure 1.14: Simple laboratory distillation apparatus.

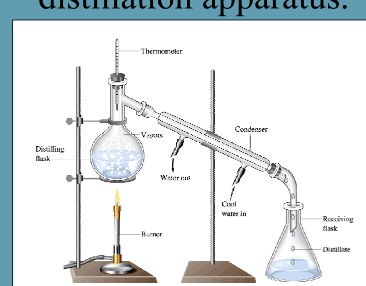
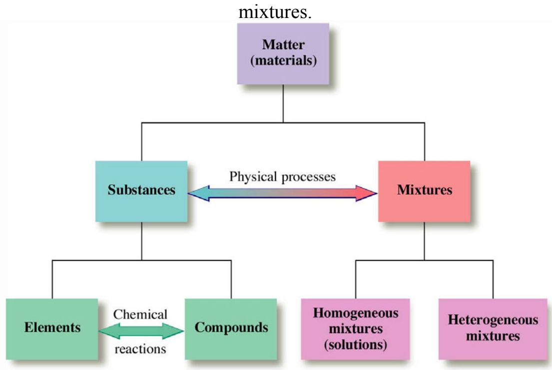


Figure 1.16: Relationships among elements, compounds, and mixtures.



CONCEPT CHECK 1.1

Matter can be represented as being composed of individual units. For example, the smallest individual unit of matter can be represented as a single circle, •, and chemical combinations of these units of matter as connected circles, ••, with each element represented by a different color. Using this model, place the appropriate label—element, compound, or mixture—on each container.