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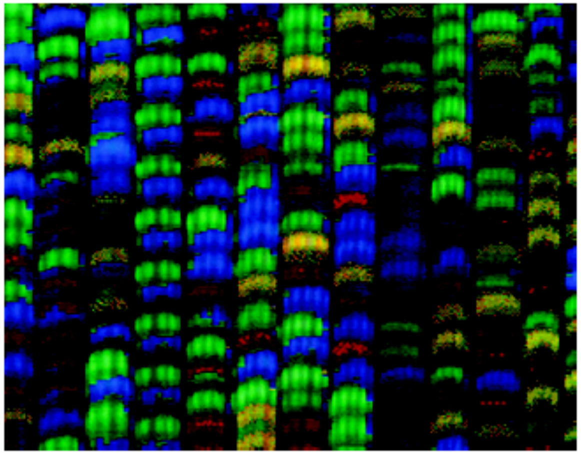
Chemistry:
The Study of Change

Chapter 1

Chemistry: A Science for the 21st Century

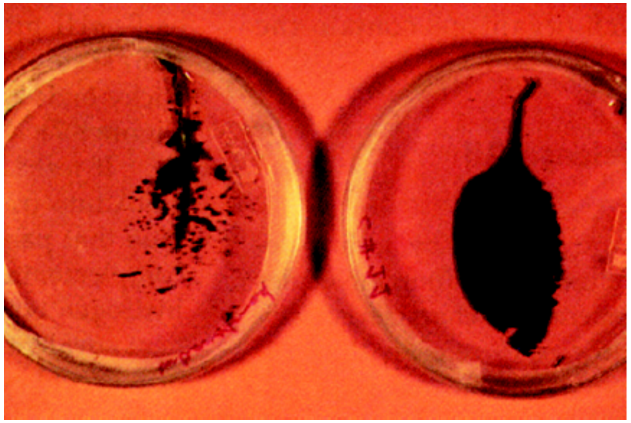
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



- Energy and the Environment
  - Fossil fuels
  - Solar energy
  - Nuclear energy

Chemistry: A Science for the 21st Century

- Materials and Technology
  - Polymers, ceramics, liquid crystals
  - Room-temperature superconductors?
  - Molecular computing?

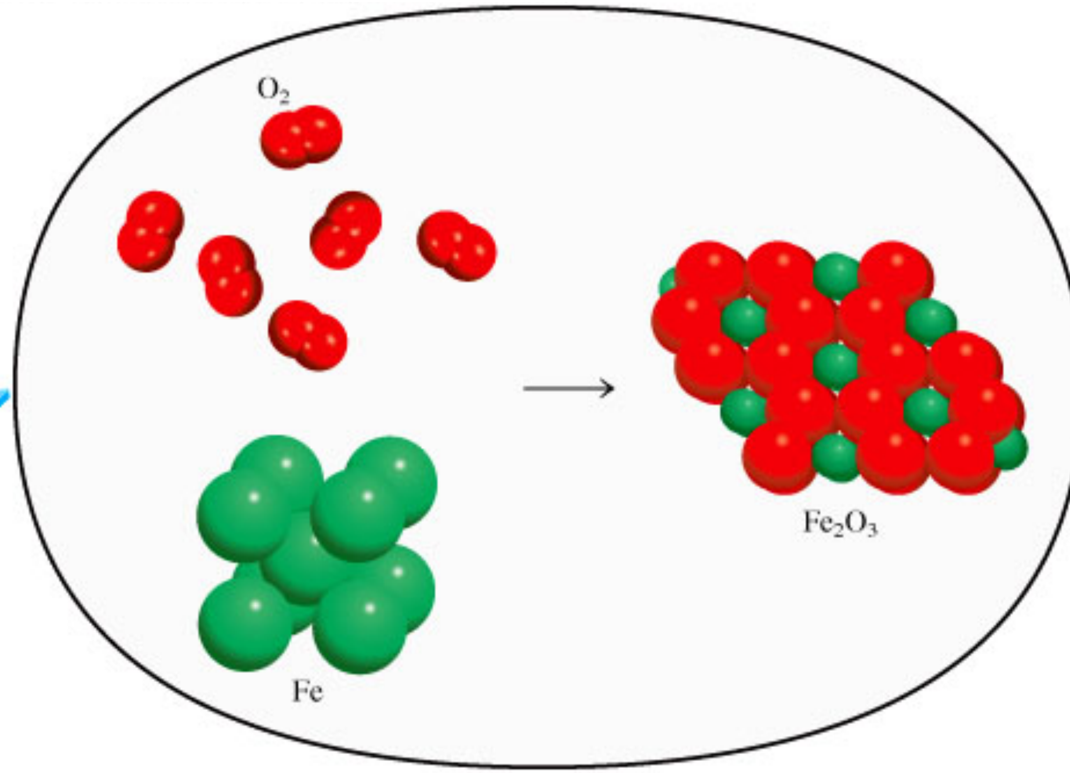


- Food and Agriculture
  - Genetically modified crops
  - "Natural" pesticides
  - Specialized fertilizers

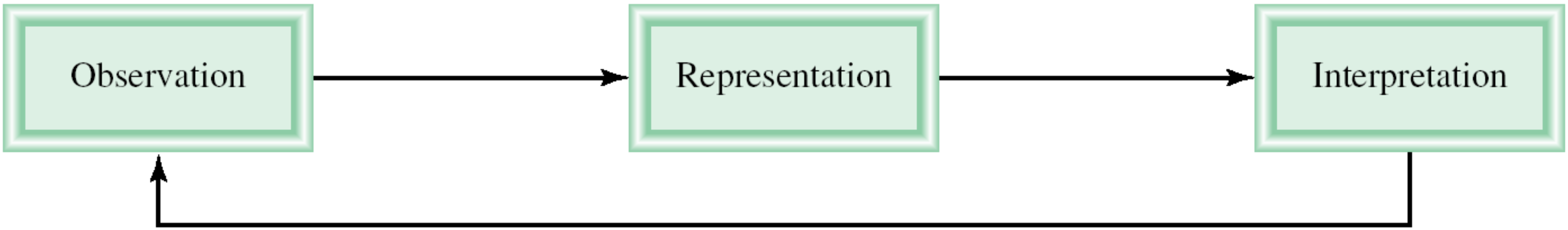
The Study of Chemistry

Macroscopic

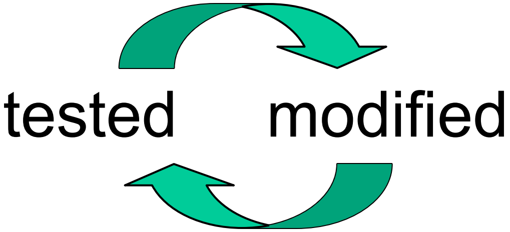
Microscopic



The ***scientific method*** is a systematic approach to research



A ***hypothesis*** is a tentative explanation for a set of observations

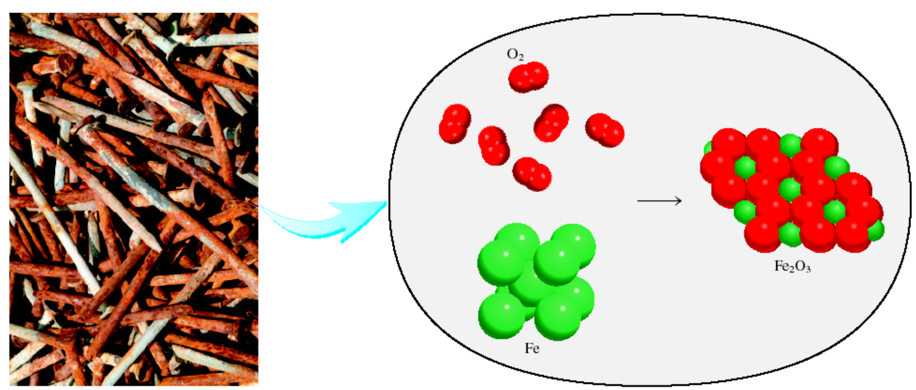


A ***law*** is a concise statement of a relationship between phenomena that is always the same under the same conditions.

Force = mass x acceleration

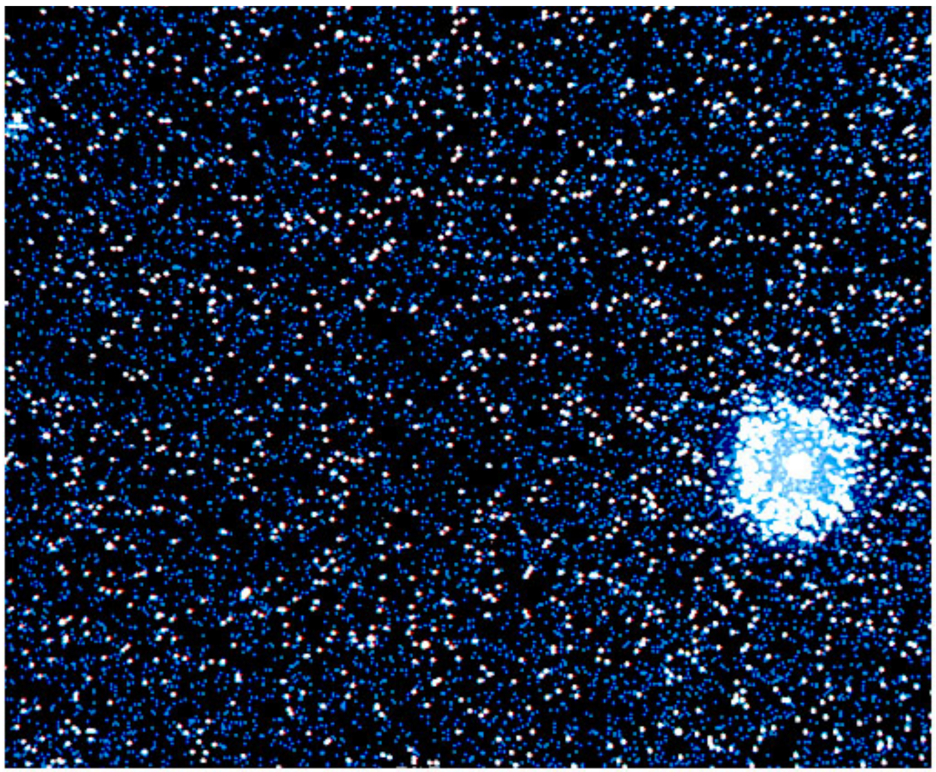
A ***theory*** is a unifying principle that explains a body of facts and/or those laws that are based on them.

Atomic Theory



Chemistry In Action:
Primordial Helium and the Big Bang Theory

In 1940 George Gamow ***hypothesized*** that the universe began with a gigantic explosion or big bang.



Experimental Support

- expanding universe
- cosmic background radiation
- primordial helium

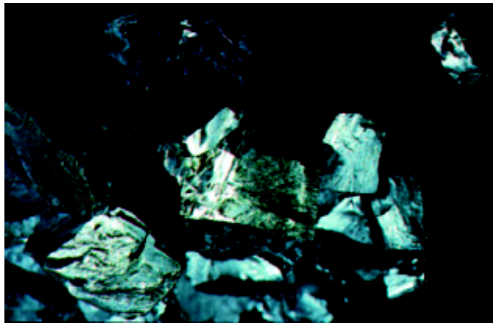
Chemistry is the study of matter and the changes it undergoes

Matter is anything that occupies space and has mass.

A ***substance*** is a form of matter that has a definite composition and distinct properties.

liquid nitrogen

gold ingots



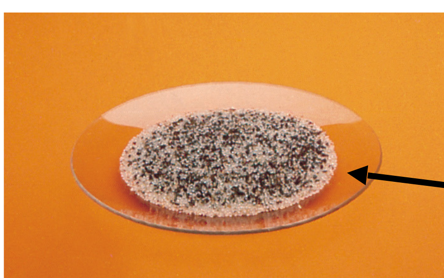
silicon crystals

A ***mixture*** is a combination of two or more substances in which the substances retain their distinct identities.

1. ***Homogenous mixture*** – composition of the mixture is the same throughout.

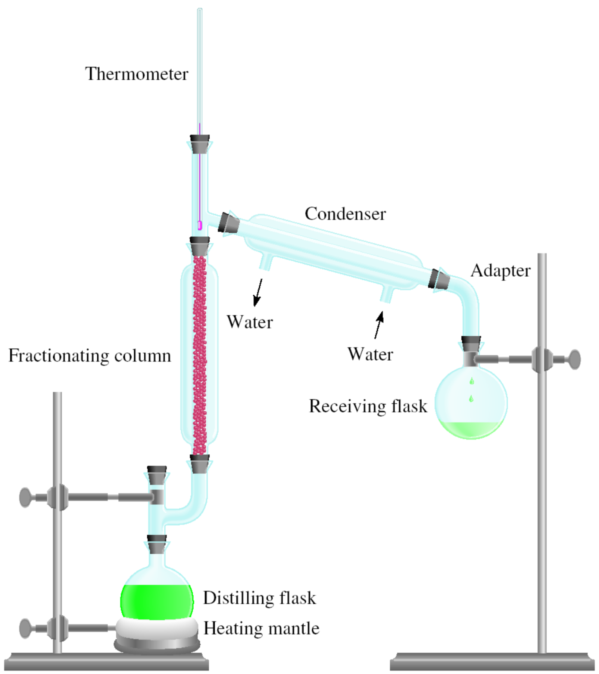
soft drink, milk, solder

2. ***Heterogeneous mixture*** – composition is not uniform throughout.



cement,
iron filings in sand

Physical means can be used to separate a mixture into its pure components.



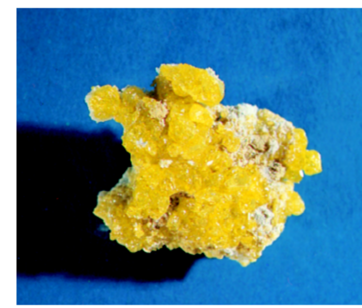
distillation

magnet

An ***element*** is a substance that cannot be separated into simpler substances by ***chemical means***.

- 114 elements have been identified
  - 82 elements occur naturally on Earth

    gold, aluminum, lead, oxygen, carbon, sulfur



  - 32 elements ha ntists

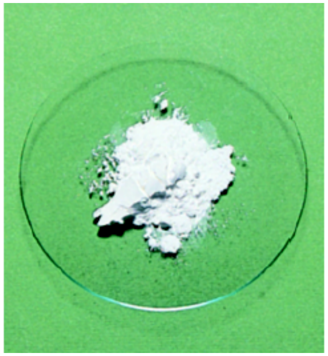
    technetium, americium, seaborgium

TABLE 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

A ***compound*** is a substance composed of atoms of two or more elements chemically united in fixed proportions.

Compounds can only be separated into their pure components (elements) by ***chemical*** means.

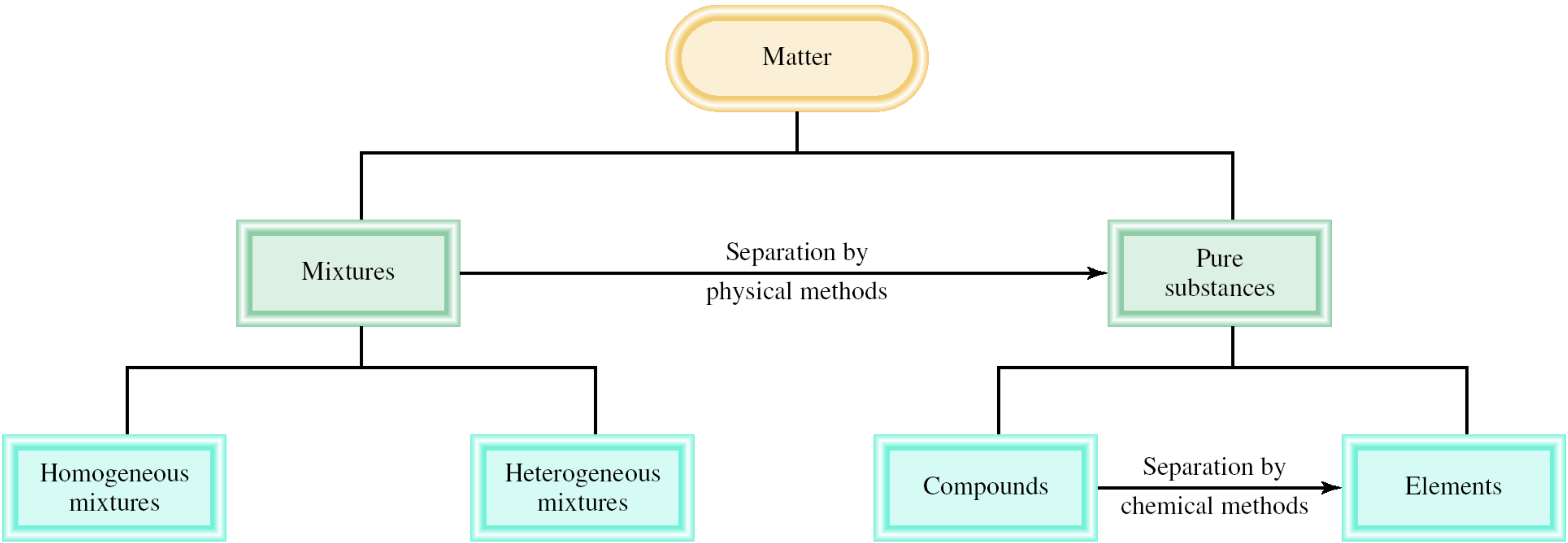


lithium fluoride

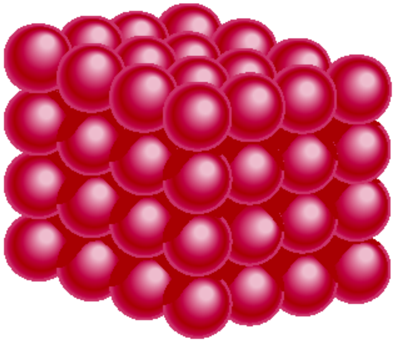
quartz

dry ice – carbon dioxide

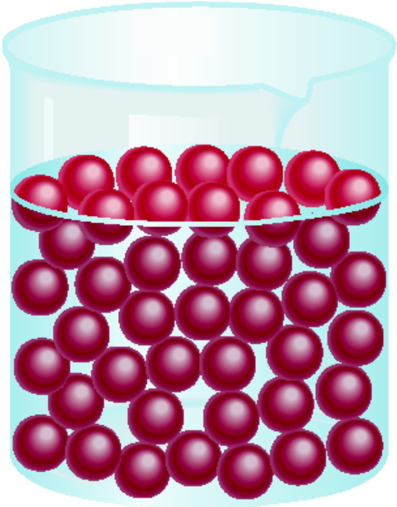
Classifications of Matter



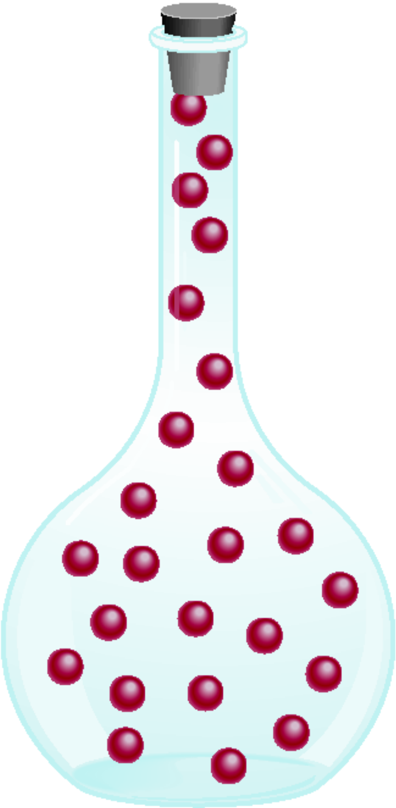
A Comparison: The Three States of Matter



Solid

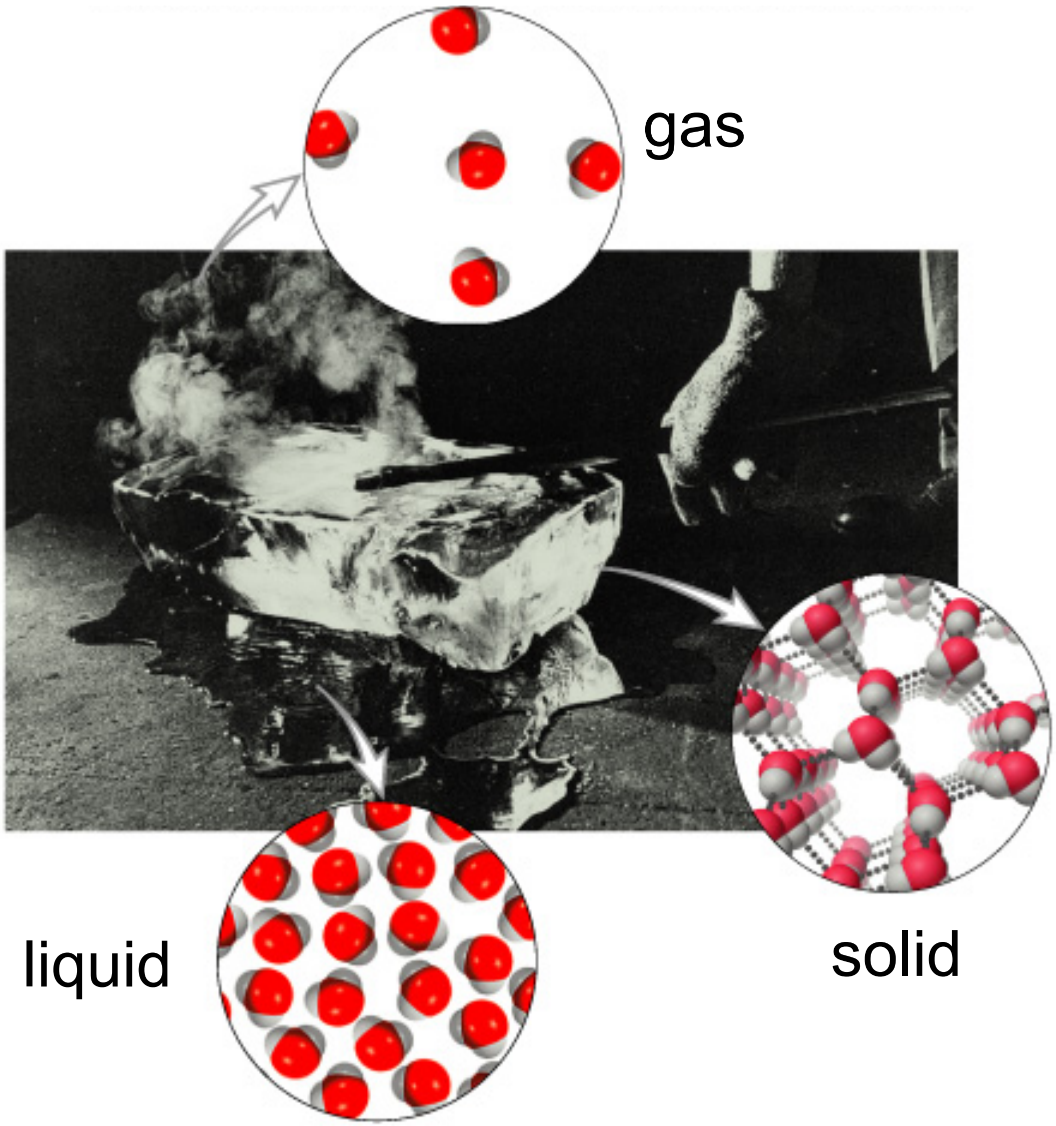


Liquid



Gas

The Three States of Matter: Effect of a Hot Poker on a Block of Ice



Types of Changes

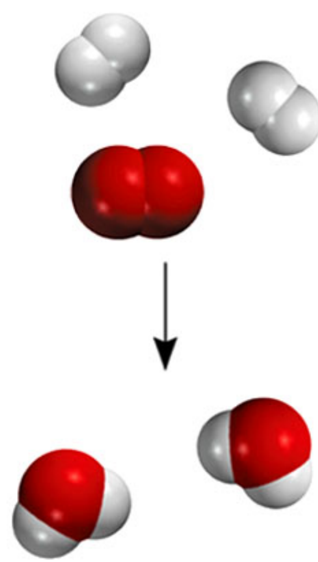
A ***physical change*** does not alter the composition or identity of a substance.

ice melting

sugar dissolving in water

A ***chemical change*** alters the composition or identity of the substance(s) involved.

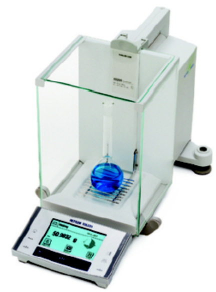
hydrogen burns in air to form water

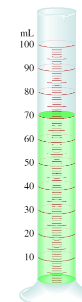


Extensive and Intensive Properties

An ***extensive property*** of a material depends upon how much matter is is being considered.

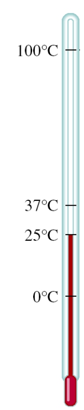
- mass
- length
- volume





An ***intensive property*** of a material does not depend upon how much matter is is being considered.

- density
- temperature
- color



Matter - anything that occupies space and has ***mass***.

mass – measure of the quantity of matter

SI unit of mass is the ***kilogram*** (kg)

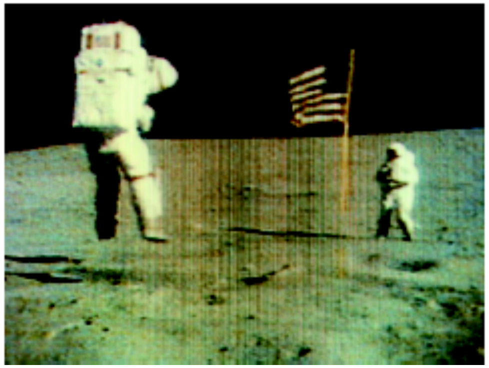
1 kg = 1000 g = 1×10^3 g

weight – force that gravity exerts on an object

weight = *c* x mass

on earth, *c* = 1.0

on moon, *c* ~ 0.1



A 1 kg bar will weigh

1 kg on earth

0.1 kg on moon

International System of Units (SI)

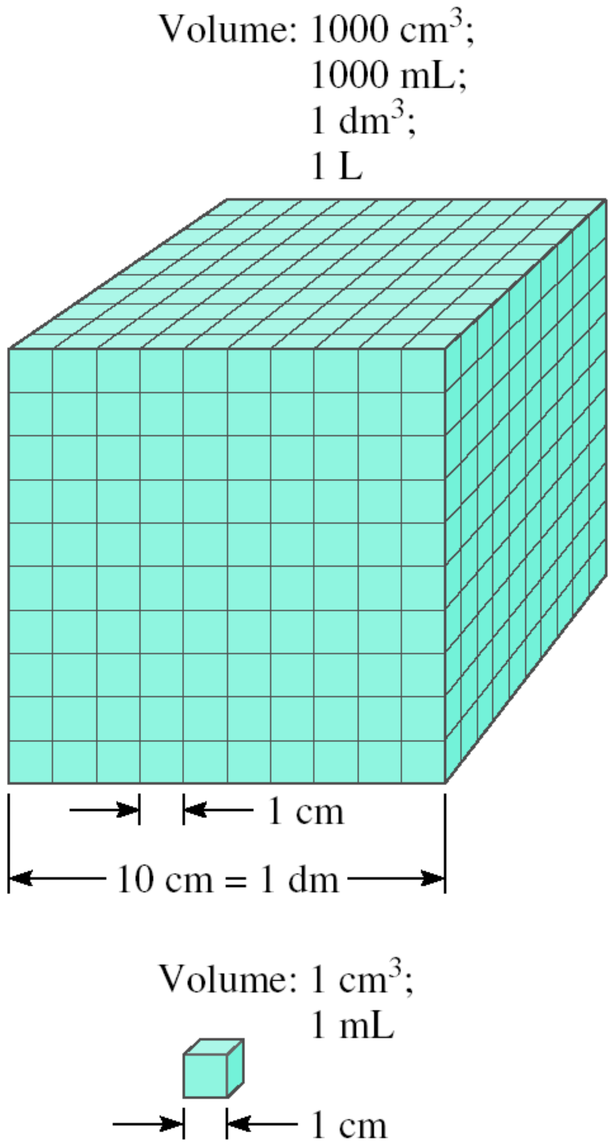
TABLE 1.2 **SI Base Units**

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

TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m

Volume – SI derived unit for volume is cubic meter (m^3)



$$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$

$$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$

$$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$$

$$\boxed{1\ mL = 1\ cm^3}$$

Volumetric flask

Density – SI derived unit for density is kg/m^3

$$1\ g/cm^3 = 1\ g/mL = 1000\ kg/m^3$$

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{m}{V}$$

A piece of platinum metal with a density of $21.5\ g/cm^3$ has a volume of $4.49\ cm^3$. What is its mass?

$$d = \frac{m}{V}$$

$$m = d \times V = 21.5\ g/\cancel{cm^3} \times 4.49\ \cancel{cm^3} = 96.5\ g$$

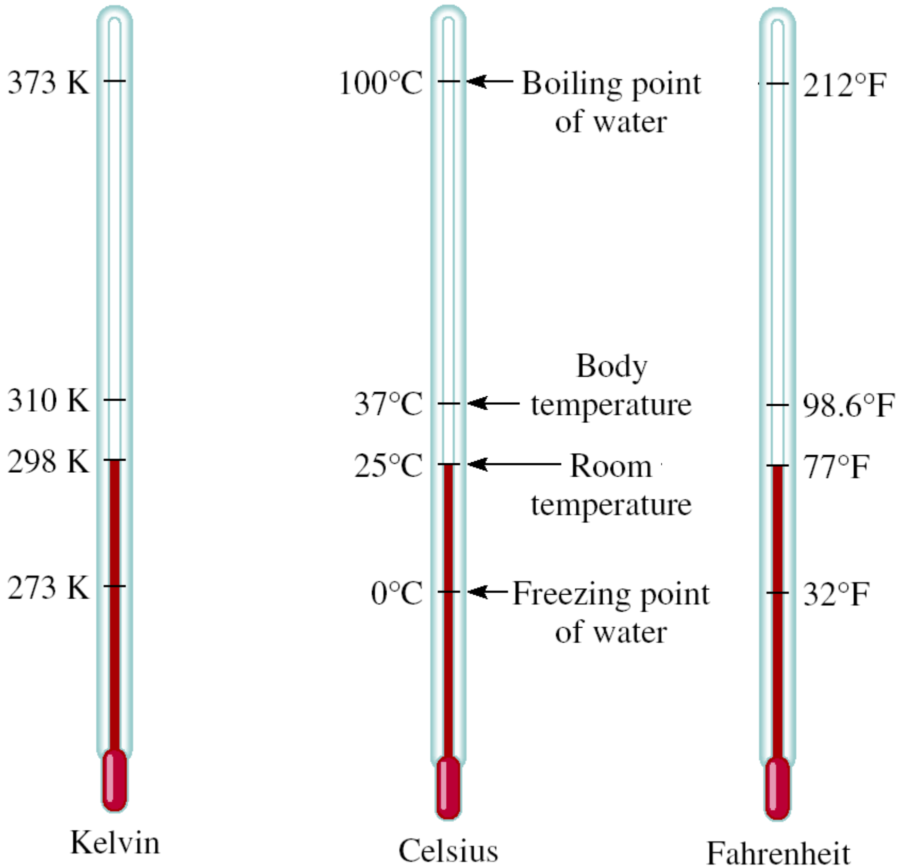
TABLE 1.4

Densities of Some Substances at 25°C

Substance	Density (g/cm^3)
Air*	0.001
Ethanol	0.79
Water	1.00
Mercury	13.6
Table salt	2.2
Iron	7.9
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.

A Comparison of Temperature Scales



$$K = {}^{0}C + 273.15$$

$$273\ K = 0\ {}^{0}C$$
$$373\ K = 100\ {}^{0}C$$

$${}^{0}F = \frac{9}{5} \times {}^{0}C + 32$$

$$32\ {}^{0}F = 0\ {}^{0}C$$
$$212\ {}^{0}F = 100\ {}^{0}C$$

Convert 172.9 ^{0}F to degrees Celsius.

$$^0F = \frac{9}{5} \times {}^0C + 32$$

$$^0F - 32 = \frac{9}{5} \times {}^0C$$

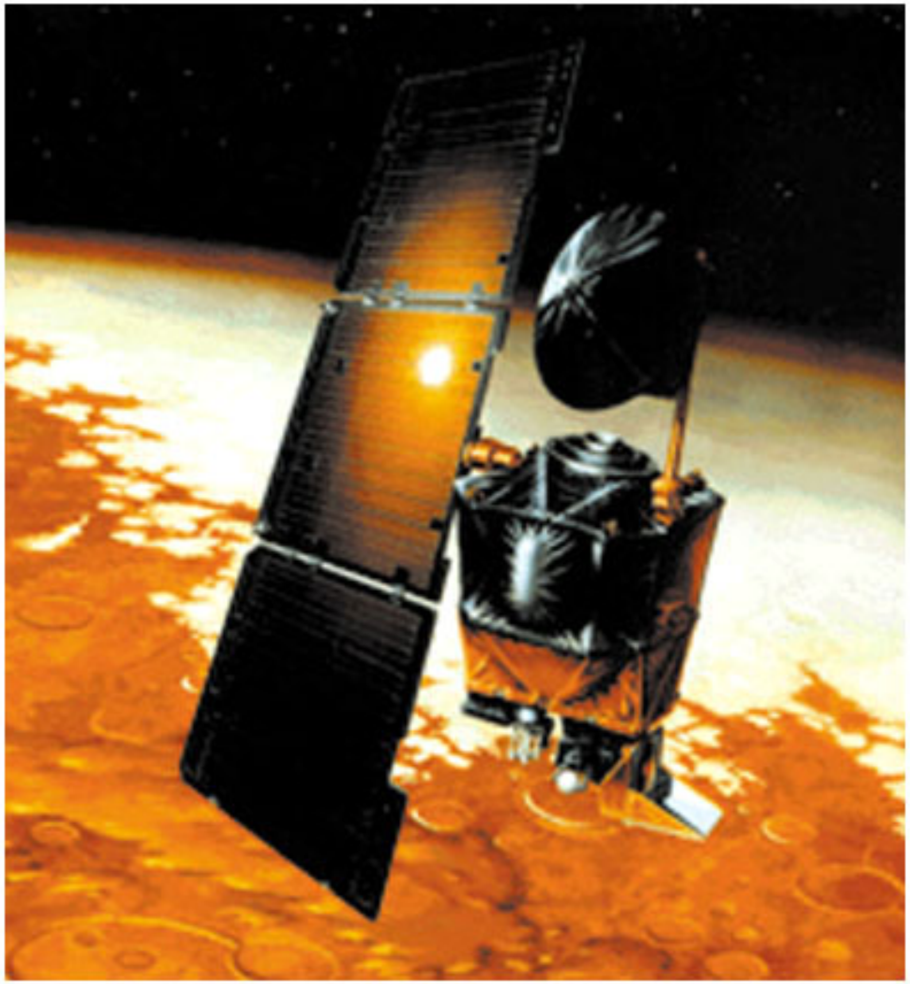
$$\frac{5}{9} \times ({}^0F - 32) = {}^0C$$

$$^0C = \frac{5}{9} \times ({}^0F - 32)$$

$$^0C = \frac{5}{9} \times (172.9 - 32) = 78.3$$

Chemistry In Action

On 9/23/99, \$125,000,000 Mars Climate Orbiter entered Mar's atmosphere 100 km (62 miles) lower than planned and was destroyed by heat.



1 lb $\neq$ 1 N

1 lb = 4.45 N

"This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time."

Scientific Notation

The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

6.022×10^{23}

The mass of a single carbon atom in grams:

0.0000000000000000000000199

1.99×10^{-23}

$\boxed{N \times 10^{n}}$

N is a number between 1 and 10

n is a positive or negative integer

Scientific Notation

568.762

← move decimal left

n > 0

$568.762 = 5.68762 \times 10^{2}$

0.00000772

→ move decimal right

n < 0

$0.00000772 = 7.72 \times 10^{-6}$

Addition or Subtraction

1. Write each quantity with the same exponent *n*
2. Combine N_1 and N_2
3. The exponent, *n*, remains the same

$4.31 \times 10^{4} + 3.9 \times 10^{3} =$

$4.31 \times 10^{4} + 0.39 \times 10^{4} =$

4.70×10^{4}

Scientific Notation

Multiplication

1. Multiply N_1 and N_2
2. Add exponents n_1 and n_2

$$(4.0 \times 10^{-5}) \times (7.0 \times 10^{3}) =$$
$$(4.0 \times 7.0) \times (10^{-5+3}) =$$
$$28 \times 10^{-2} =$$
$$2.8 \times 10^{-1}$$

Division

1. Divide N_1 and N_2
2. Subtract exponents n_1 and n_2

$$8.5 \times 10^{4} \div 5.0 \times 10^{9} =$$
$$(8.5 \div 5.0) \times 10^{4-9} =$$
$$1.7 \times 10^{-5}$$

Significant Figures

- Any digit that is not zero is significant

  1.234 kg 4 significant figures

- Zeros between nonzero digits are significant

  606 m 3 significant figures

- Zeros to the left of the first nonzero digit are **not** significant

  0.08 L 1 significant figure

- If a number is greater than 1, then all zeros to the right of the decimal point are significant

  2.0 mg 2 significant figures

- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

  0.00420 g 3 significant figures

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

24 mL	2 significant figures
3001 g	4 significant figures
0.0320 m^3	3 significant figures
6.4 x 10^4 molecules	2 significant figures
560 kg	2 significant figures

Significant Figures

Addition or Subtraction

The answer cannot have more digits to the right of the decimal point than any of the original numbers.

$$\begin{array}{r} 89.332 \\ +1.1 \\ \hline 90.432 \end{array}$$

$+1.1 \longleftarrow$ one significant figure after decimal point

$90.432 \longleftarrow$ round off to 90.4

$$\begin{array}{r} 3.70 \\ -2.9133 \\ \hline 0.7867 \end{array}$$

$3.70 \longleftarrow$ two significant figures after decimal point

$0.7867 \longleftarrow$ round off to 0.79

Significant Figures

Multiplication or Division

The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures

$$4.51 \times 3.6666 = 16.536366 = 16.5$$

3 sig figs

round to
3 sig figs

$$6.8 \div 112.04 = 0.0606926 = 0.061$$

2 sig figs

round to
2 sig figs

Significant Figures

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures

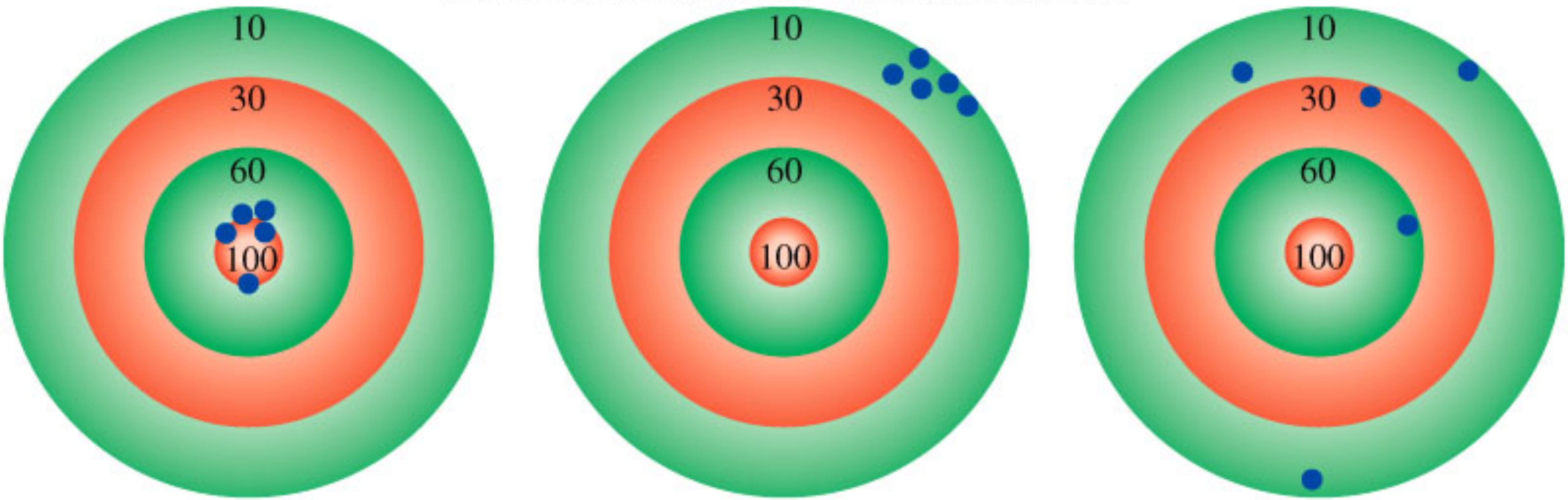
The average of three measured lengths; 6.64, 6.68 and 6.70?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 = \cancel{7}$$

Because 3 is an ***exact number***

Accuracy – how close a measurement is to the *true* value

Precision – how close a set of measurements are to each other



accurate
&
precise

precise
but
not accurate

not accurate
&
not precise

Dimensional Analysis Method of Solving Problems

1. Determine which unit conversion factor(s) are needed
2. Carry units through calculation
3. If all units cancel except for the ***desired unit(s)***, then the problem was solved correctly.

given quantity x conversion factor = desired quantity

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

Dimensional Analysis Method of Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

$$1.63\,\cancel{L} \times \frac{1000\ mL}{1\cancel{L}} = 1630\ mL$$

$$1.63\ L \times \frac{1L}{1000\ mL} = 0.001630\ \frac{L^2}{mL}$$

The speed of sound in air is about 343 m/s. What is this speed in miles per hour?

conversion units

meters to miles

seconds to hours

1 mi = 1609 m 1 min = 60 s 1 hour = 60 min

$$343 \frac{\cancel{m}}{\cancel{s}} \times \frac{1 \text{ mi}}{1609 \cancel{m}} \times \frac{60 \cancel{s}}{1 \cancel{min}} \times \frac{60 \cancel{min}}{1 \text{ hour}} = 767 \frac{\text{mi}}{\text{hour}}$$