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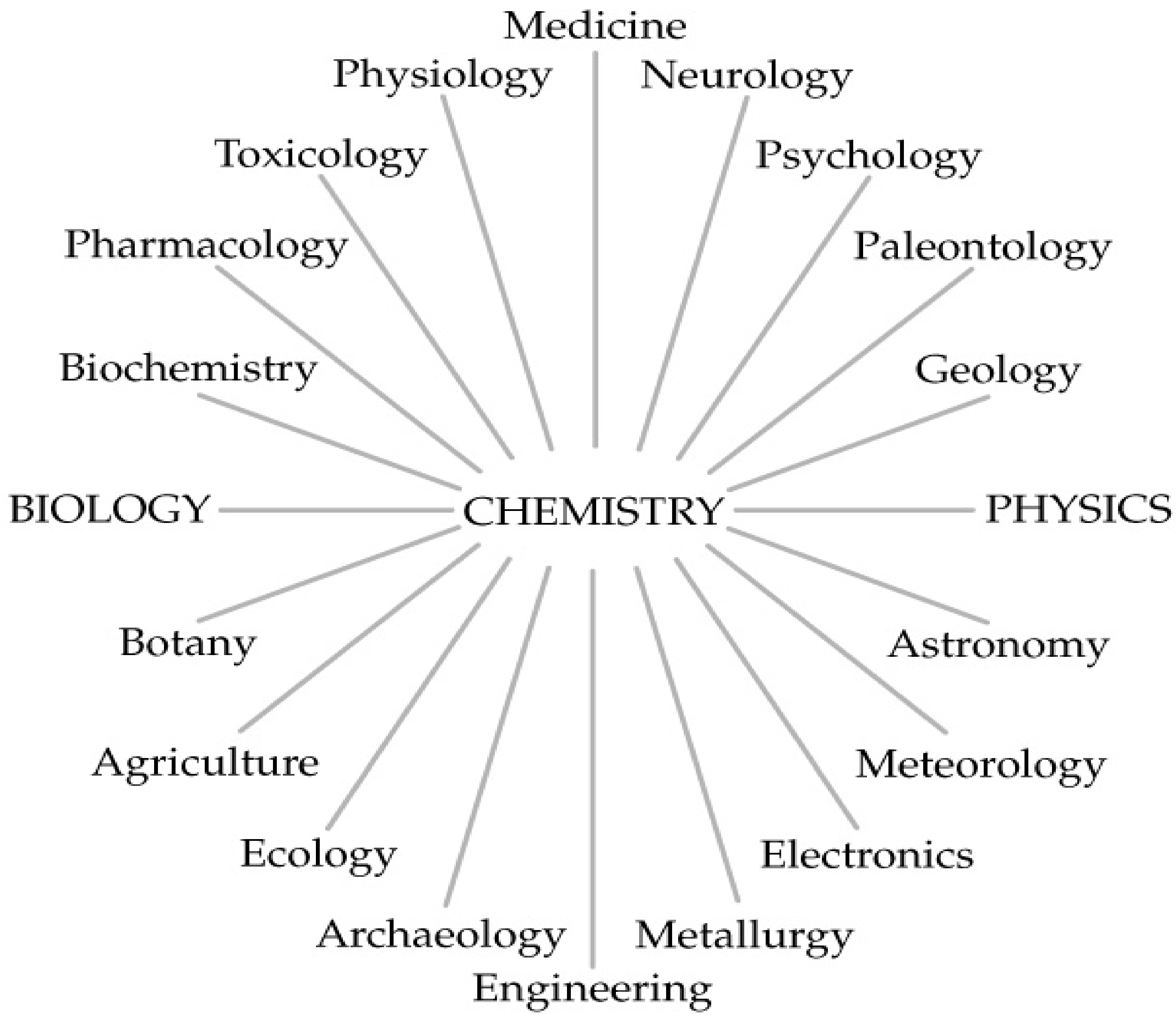
Chemistry:
The Study of Change

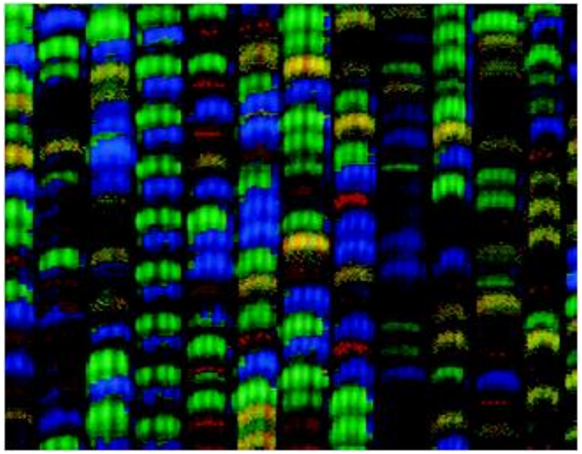
Chapter 1

Chemistry: the central science



Chemistry: A Science for the 21st Century

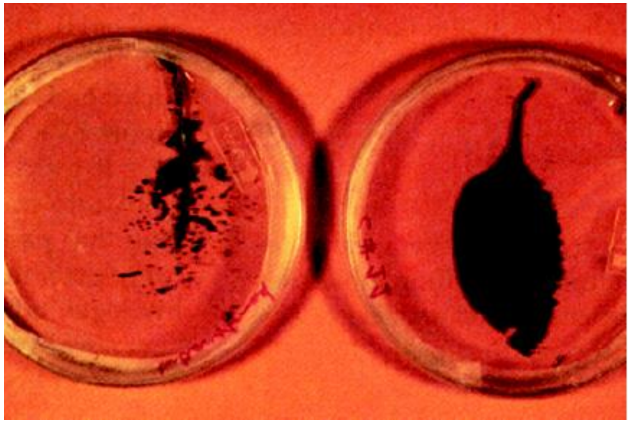
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



- 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?

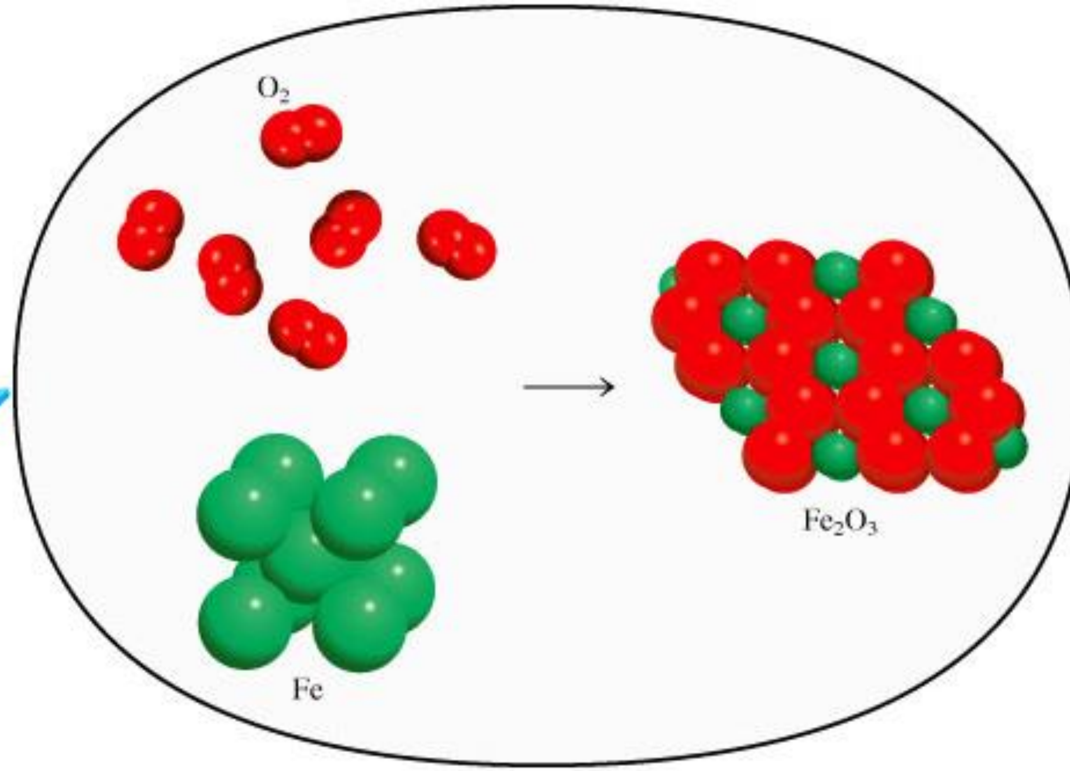


- Food and Agriculture
  - Genetically modified crops
  - “Natural” pesticides
  - Specialized fertilizers

The Study of Chemistry

Macroscopic

Microscopic



Chemistry is the study of matter and the changes it undergoes.

Matter is anything that occupies space and has mass.

Classification of Matter includes, **substances**, **mixtures**, **elements**, **compounds**, **atoms & molecules**.

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


liquid nitrogen


gold ingots


silicon crystals

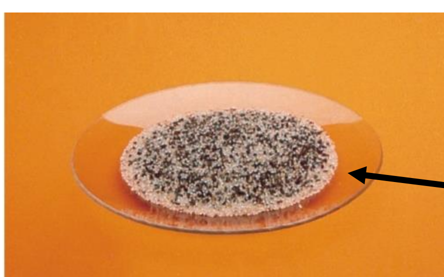
A ***mixture*** is a combination of two or more substances in which the substances retain their distinct identities.

Two types:

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

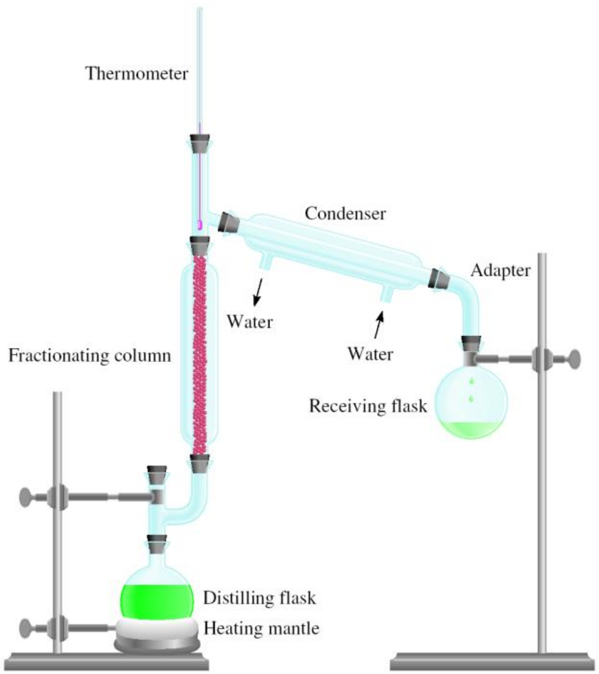
soft drink, milk, solder

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.



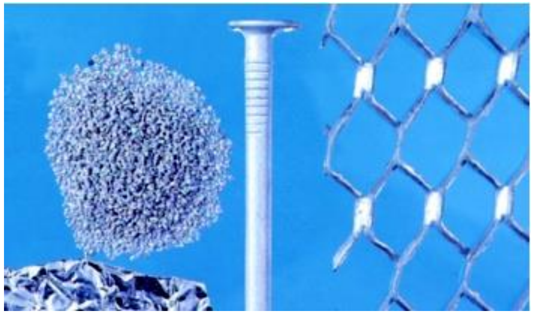
distillation

magnet

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

- **118** elements have been identified
  - **82** elements occur naturally on Earth

    gold, aluminum, lead, oxygen, carbon, sulfur



  - **36** elements have been created by scientists

    technetium, americium, seaborgium

TABLE 1.1 Some Common Elements and Their Symbols

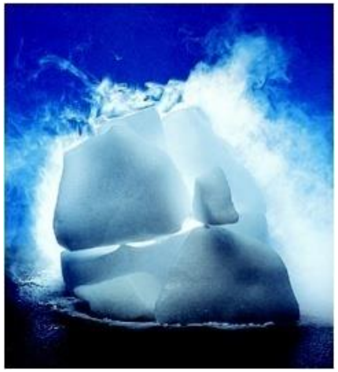
Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	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.

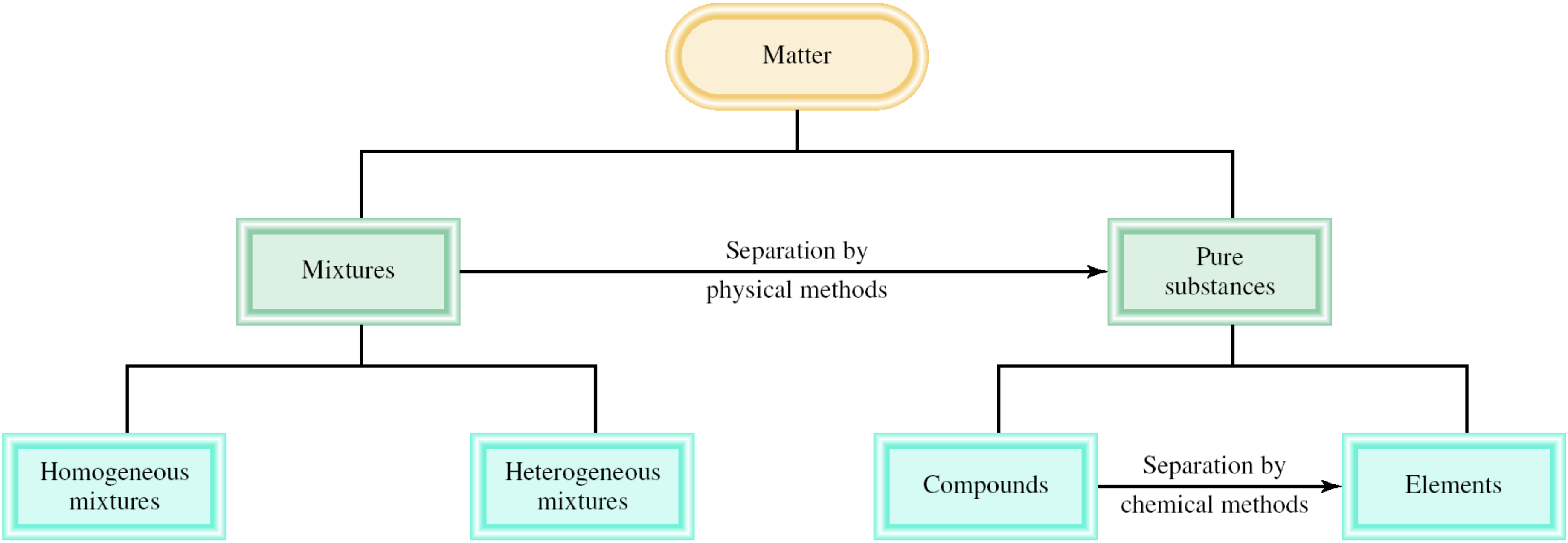
lithium fluoride

quartz

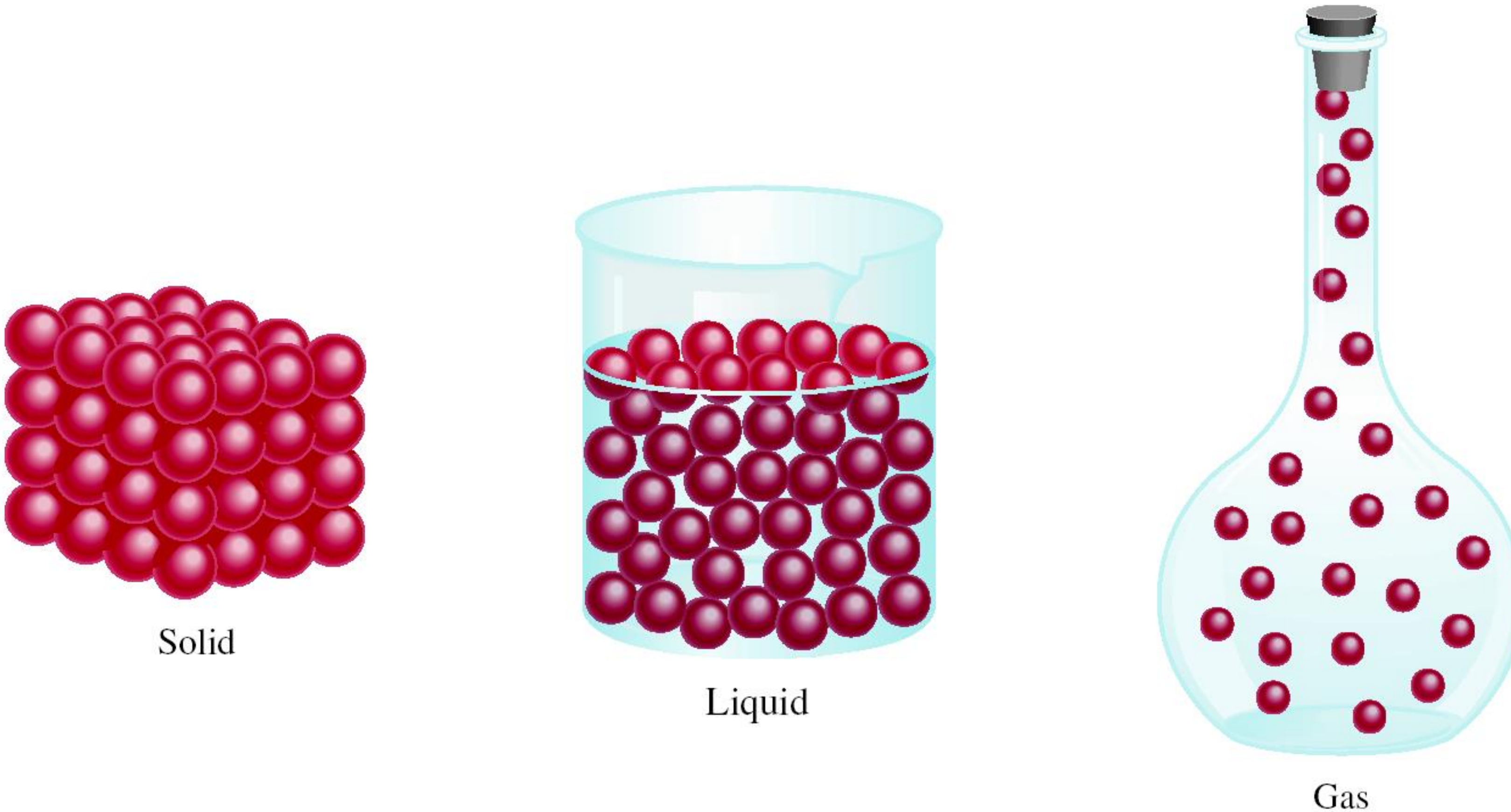


dry ice – carbon dioxide

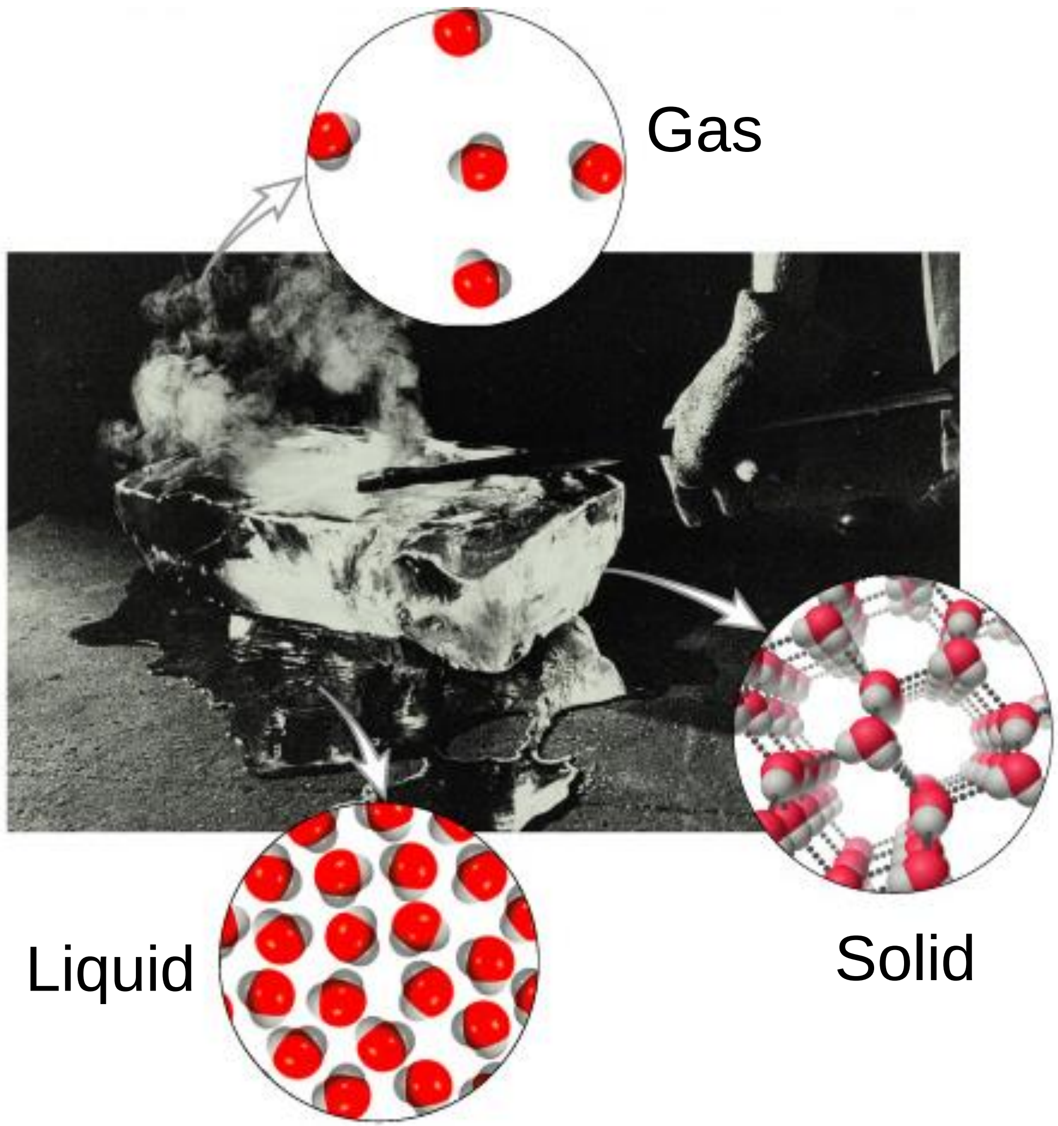
Classifications of Matter



A Comparison: The Three States of Matter



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



Types of Changes

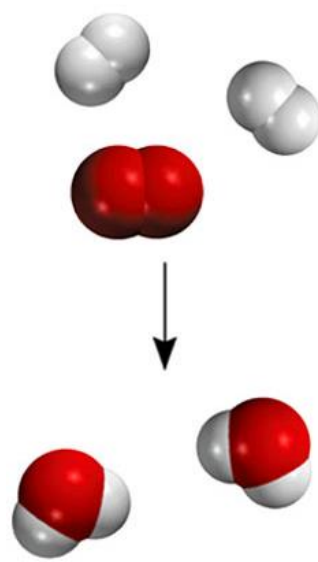
A ***physical change*** does not alter the composition or identity of a substance.

ice melting

sugar dissolving in water

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

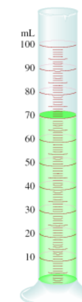
hydrogen burns in air to form water



Extensive and Intensive Properties

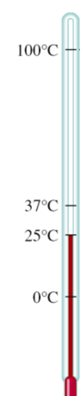
An ***extensive property*** of a material depends upon how much matter is being considered.

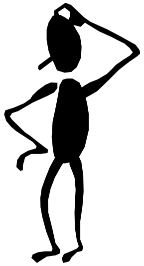
- mass
- length
- volume



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

- density
- temperature
- color





Which one is extensive and which one is intensive property?

Length	Extensive	Colour	Intensive
Weight	Extensive	Surface area	Extensive
Taste	Intensive	Hardness	Intensive
Width	Extensive	Melting point	Intensive
Boiling point	Intensive	Energy content	Extensive
Volume	Extensive	Density	Intensive
Mass	Extensive	Luster	Intensive

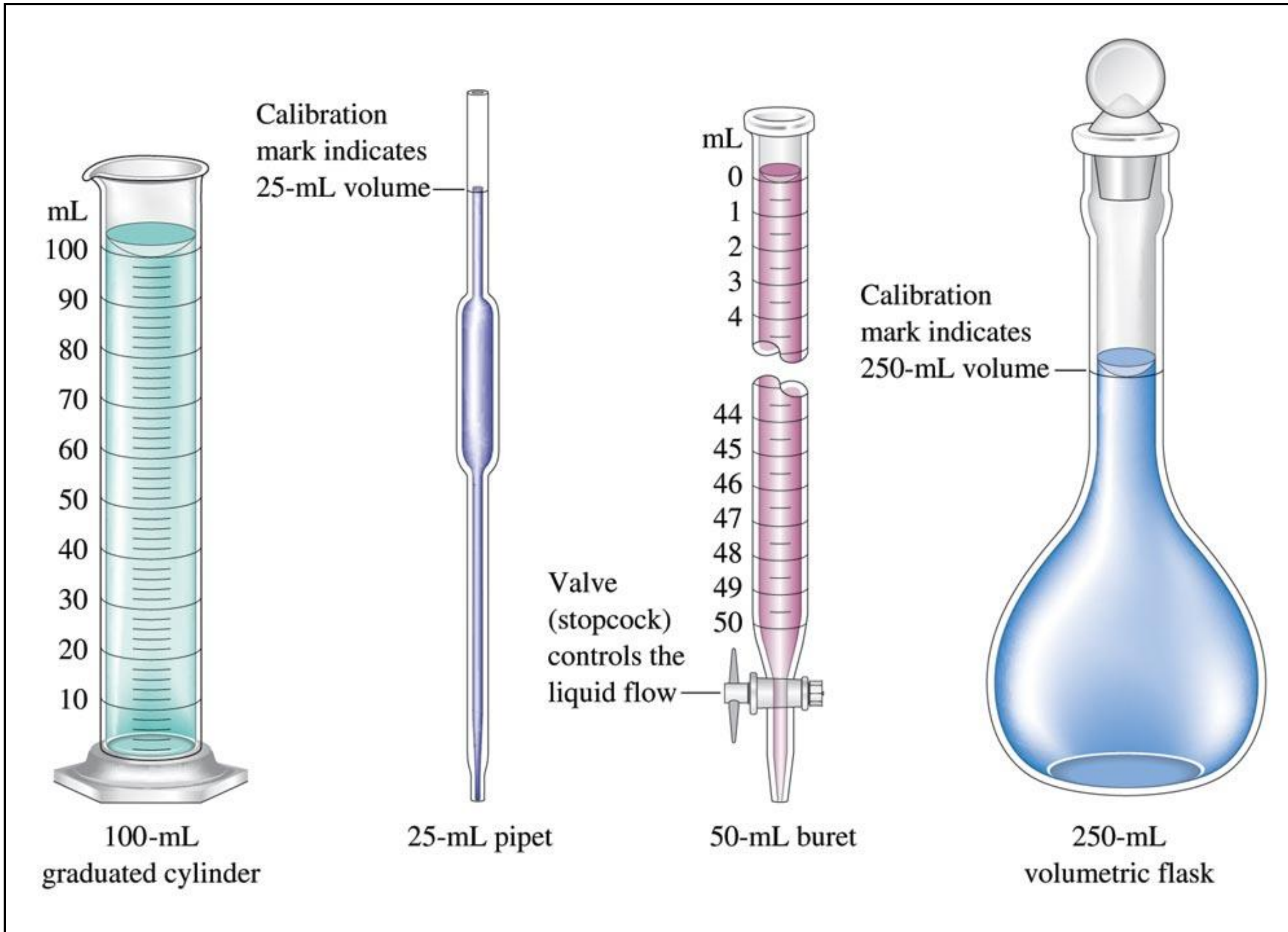
Measurements

- ***Measurement*** – quantitative observation consisting of two parts:
  - Number
  - Scale (unit)

- Examples:
  - 20 grams
  - 6.63×10^{-34} joule·seconds

Common Types of Laboratory Equipment Used to Measure Mass & Length

Common Types of Laboratory Equipment Used to Measure Liquid Volume



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

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

mass – measure of the quantity of matter

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

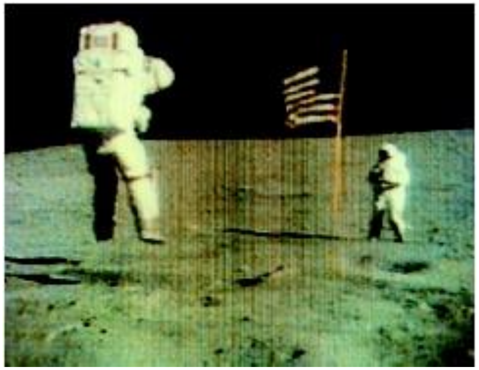
1 kg = 1000 g = 1×10^3 g

weight – force that gravity exerts on an object

weight = g x mass

on earth, $g = 9.81$ m s^{-2}

on moon, g is ~1/6 of earth



A 1 kg bar will weigh

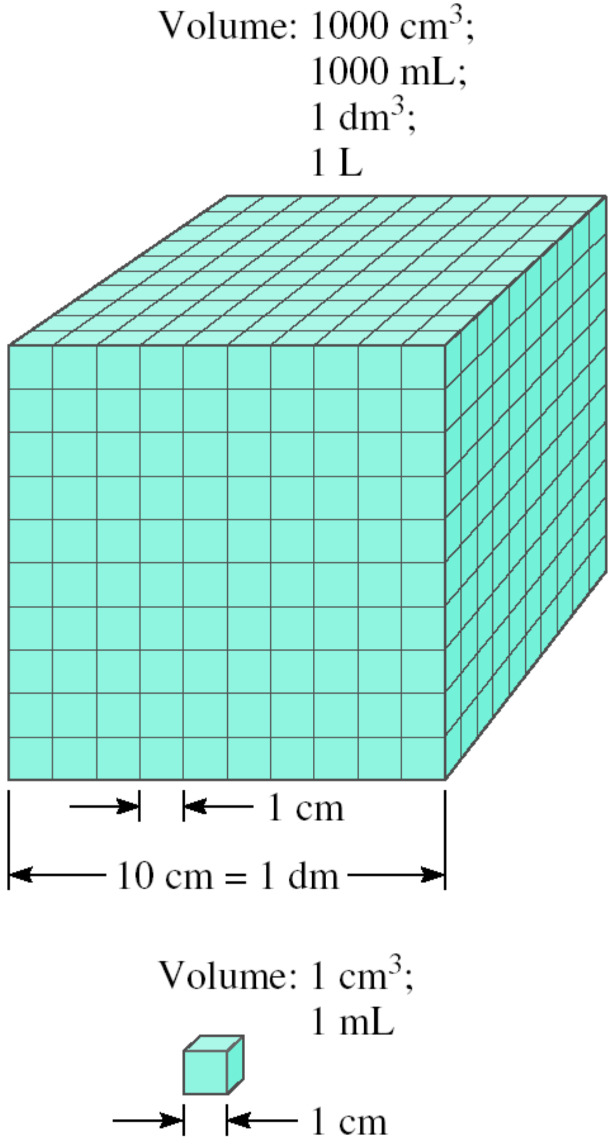
1 kg on earth

0.17 kg on moon

MKr has a mass of 63 kg weighs 618 Newtons !

Taken g =1

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?

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

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

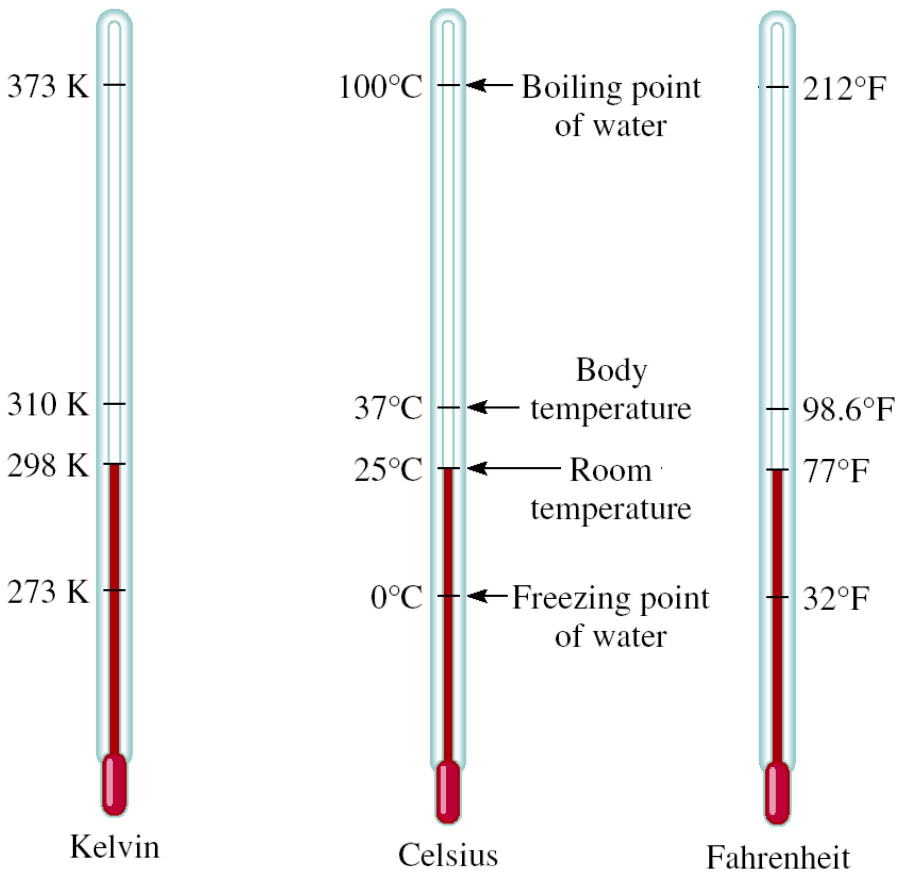
TABLE 1.4

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



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

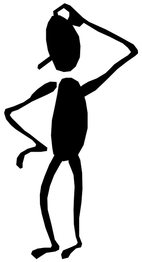
$$273\ K = 0\ {}^{0}C$$
$$373\ K = 100\ {}^{0}C$$

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

$${}^{0}C = ({}^{0}F - 32) \times \frac{100}{180}$$

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

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

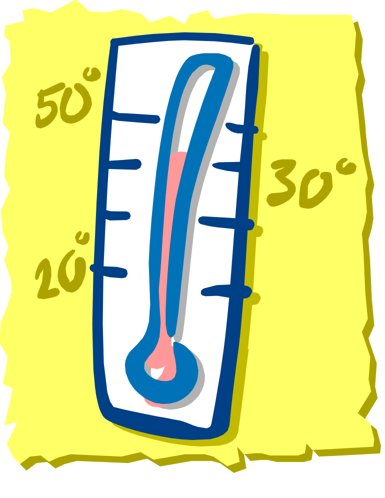


Convert 172.9 ^{0}F to degrees Celsius.

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

$$= \frac{5}{9} \times (172.9 - 32)$$

$$= 78.3$$



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."

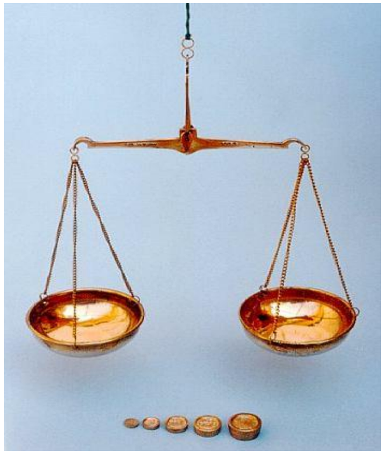
Significant Figures

The meaningful digits in a measured or calculated quantity.

10.5583 g 10.55 g ?

±0.0001 g

Last digit is uncertain.



1.55 kg 1.5583 kg ?

± 0.01 kg

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

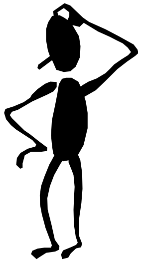
  0.08 L   1 significant figure

- 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

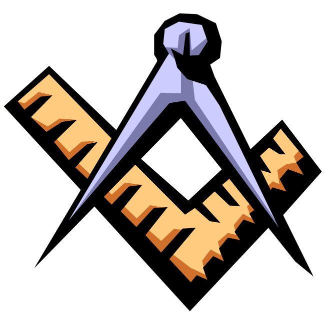
- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

  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 or 3 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.

```
 89.332
  +1.1    ←——  one digit after decimal point
 ------
 90.432   ←——  round off to 90.4


   3.70   ←——  two digits after decimal point
 -2.9133
 -------
 0.7867   ←——  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 x 3.6666 = 16.536366 = 16.5

3 sig figs

round to
3 sig figs

6.8 ÷ 112.04 = 0.0606926 = 0.061

2 sig figs

round to
2 sig figs

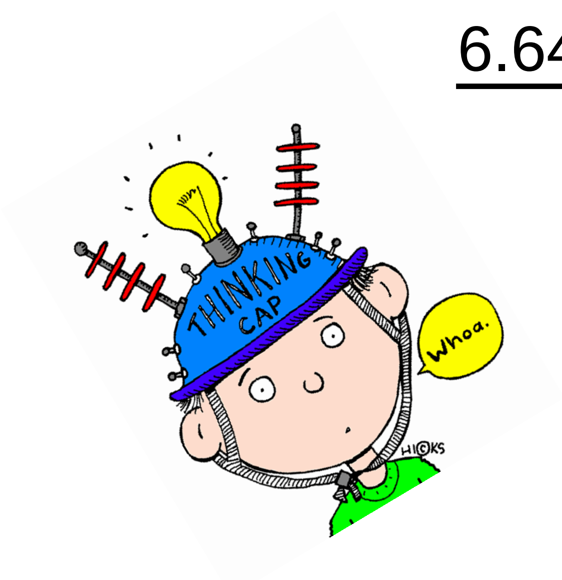
Significant Figures

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?

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

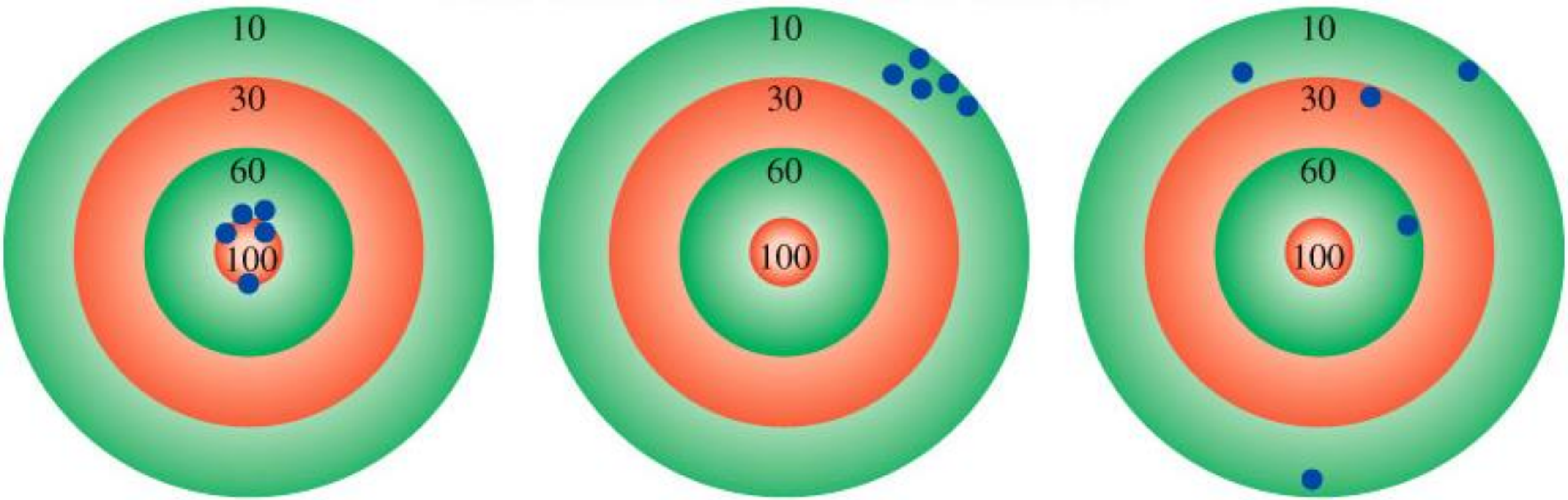


Because 3 is an ***exact number***

Accuracy and Precision

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

Accuracy and Precision

Mass of a copper wire measured by three students.

	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

- True mass = 2.000 g
- Student B's result is more precise than student A, neither set is very accurate.
- Student C's results are most precise and most accurate.
- Highly accurate measurements are usually precise too.
- Highly precise measurements do not necessarily accurate. (*e.g.*, a faulty balance)

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

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

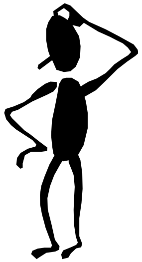
Dimensional Analysis Method of Solving Problems

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

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

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



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

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