GENERAL CHEMISTRY CHE 101

The Study of Change

Dos and Don’ts for CHE 101

- Don’t miss the classes
- Don’t be late and interrupt the class
- Bring your personal notebook and scientific calculator
- Cramming wont work in this course
- Try to understand and discuss in groups
- Keep updated and get ready for quizzes
- Ask as much as you can (I don’t mind with stupid questions)
- You are more than welcome if you have suggestions
- Please feel free to contact me (mostly office hours!) or through email (anytime!)
- I promise, the course will be fun if you assist me!

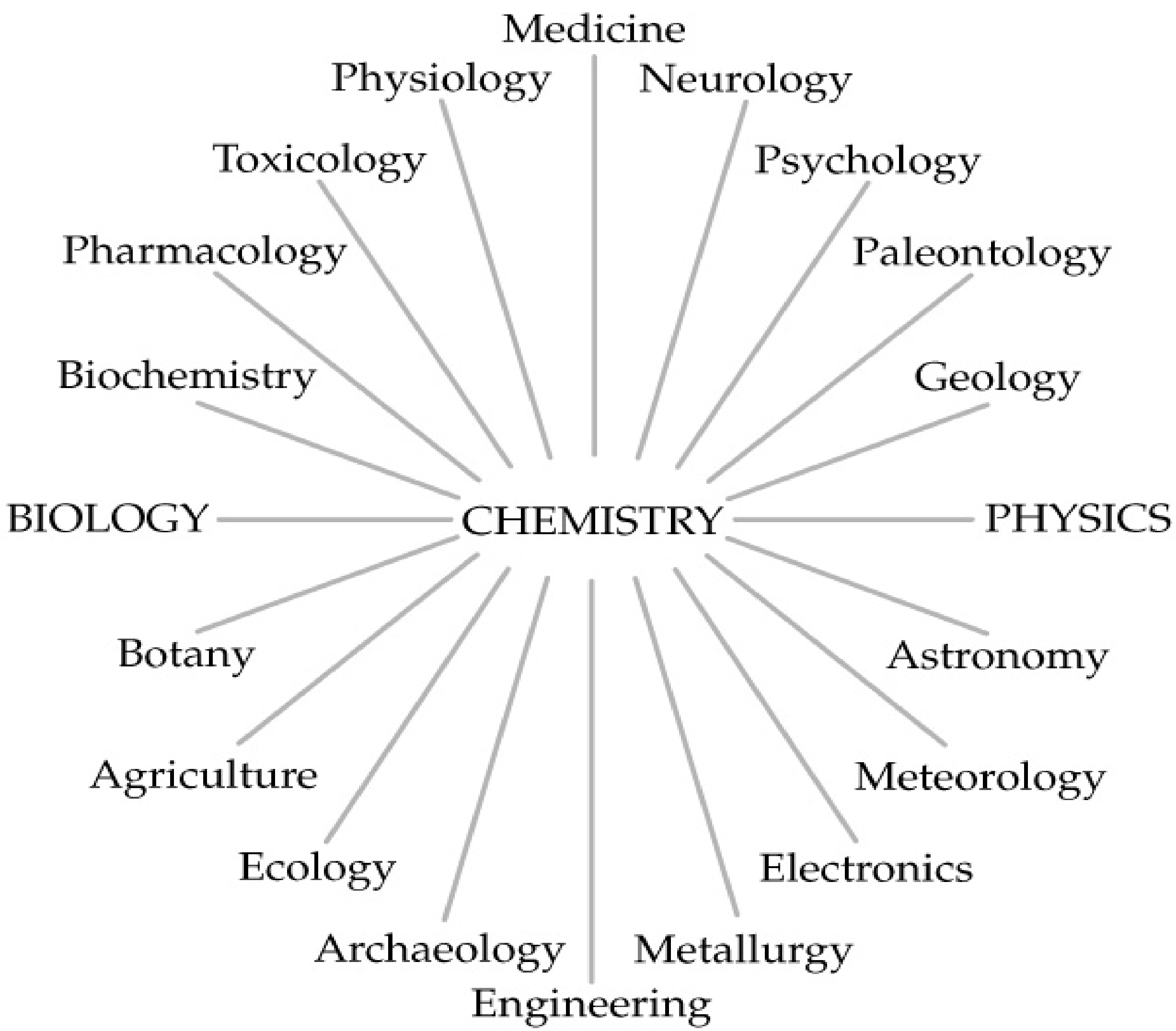
Enjoy!

Good luck !!!

Lecture Plan

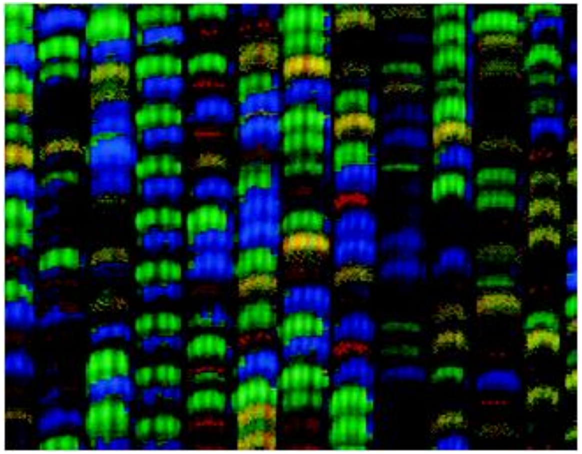
- Brief Introduction
- Scientific method of research
- Matter and its classification
- Three states of matter
- Measurement-SI Units
- Temperature conversions of matter
- Scientific notation
- Scientific figures

Chemistry: the Central Science



Chemistry: A Science for the 21st Century

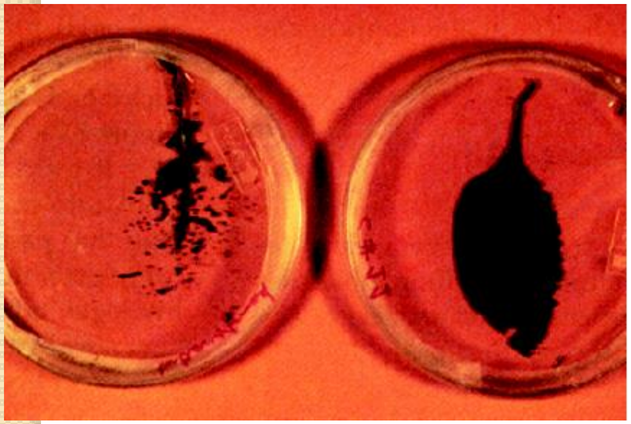
- Health and Medicine
  - Sanitation systems
  - Surgery with anesthesia
  - Vaccines and antibiotics
  - Gene therapy



- 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?

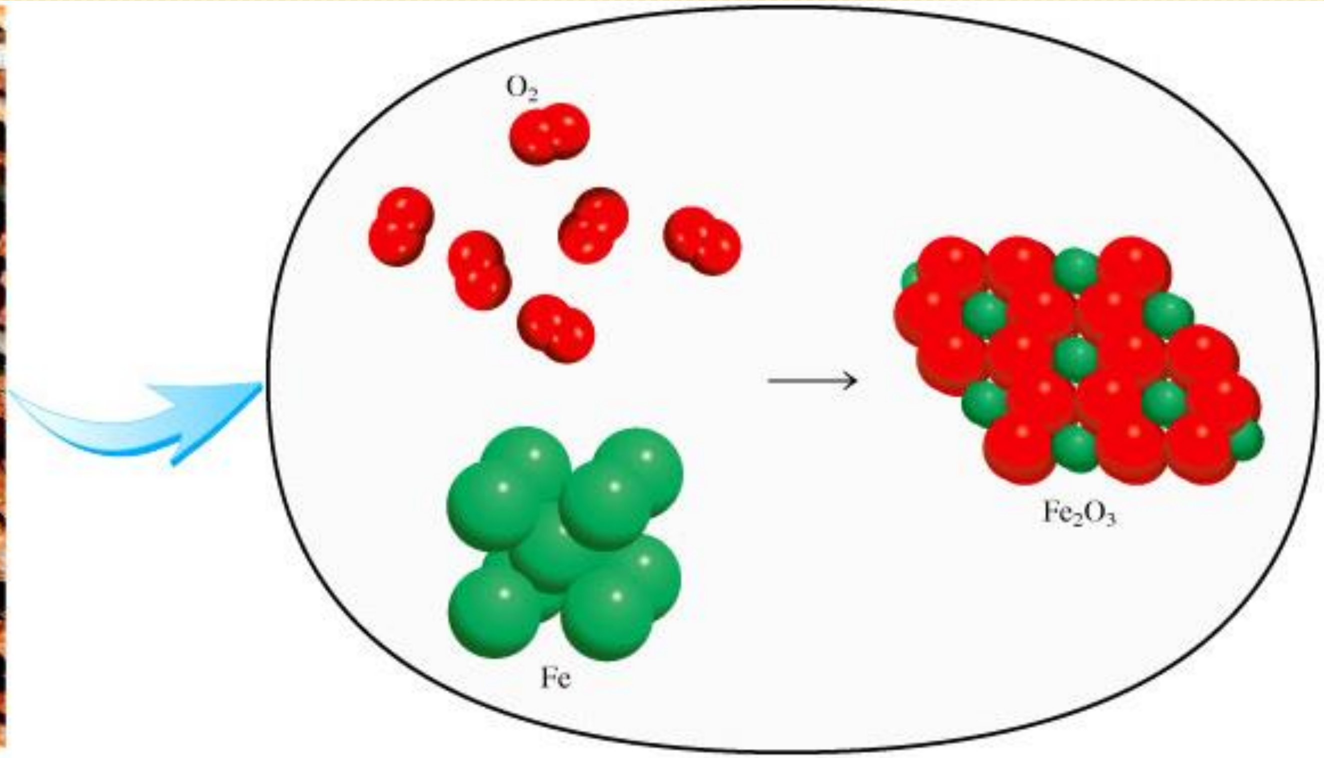


- Food and Agriculture
  - Genetically modified crops
  - “Natural” pesticides
  - Specialized fertilizers

The Study of Chemistry

Macroscopic

Microscopic



The **overall chemical equation** for the formation of rust is

Iron + water + oxygen $\longrightarrow$ rust

$4\ Fe(s) + 6\ H_2O(l) + 3\ O_2(g) \longrightarrow 4\ Fe(OH)_3(s)$

Iron(III) hydroxide, $Fe(OH)_3$ then dehydrates to produce $Fe_2O_3.nH_2O(s)$ or rust

The Scientific Method

- All science has a systemic approach to research

TYPES OF DATA

- Quantitative data

  measurements use scale with equal intervals

  examples include mass (g), length (cm), volume (mL), temperature (^{o}C or K)

- Qualitative data

  non-standard scales with unequal intervals or discrete categories

  examples include gender, choice, color scales

Data Types

- Data obtained in a research can be-
  - Qualitative
  - Quantitative
- Examples:

(a) *The sun is approximately 93 million mi from Earth.*

(b) *Ice is less dense than water.*

(c) *Butter tastes better than margarine.*

(d) *A stitch in time saves nine.*

The Scientific Method

- **Hypothesis:**
  - Educated guess, based on observation.
  - Further experiments are required to test the validity.
- **Theory:**
  - Summarization of hypothesis or group of hypothesises that have been supported with repeated testing.
  - When a Hypothesis is accepted its called the theory
- **Law:**
  - Scientific laws explain things, but they do not describe them.
  - A law generalizes a body of observations.

The Scientific Method

Hypothesis: A proposed explanation for a phenomenon made as a starting point for further investigation. Example: "It's bright outside because the sun is probably out."

Theory: A well-substantiated explanation acquired through the scientific method and repeatedly tested and confirmed through observation and experimentation. Example: "When the sun is out, it tends to make it bright outside."

Law: A statement based on repeated experimental observations that describes some phenomenon of nature. Proof that something happens and how it happens, but not why it happens. Example: Newton's Law of Universal Gravitation.

Classify each of the following statements as a hypothesis, a law or a theory:

a) If you get at least 6 hours of sleep, you will do better on tests than if you get less sleep.

b) The total energy in the universe is fixed.

c) Earth's magnetic field is generated by a conducting fluid in its core.

Matter

- Matter is anything that occupies space and has mass.
- Things that we can see and touch as well as we cannot (air)!
- Chemistry is the study of matter and the <u>changes</u> it undergoes.

Types of Changes

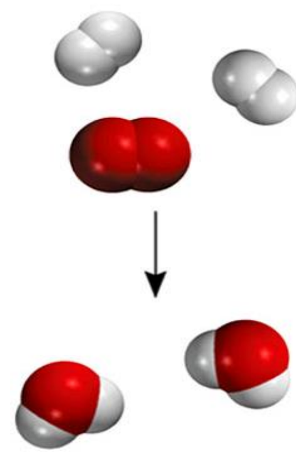
A ***physical change*** does not alter the composition or identity of a substance.

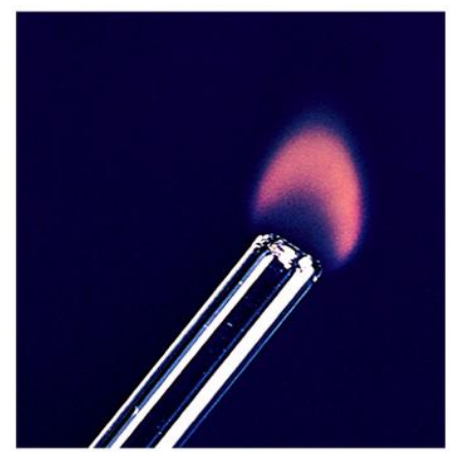
ice melting

sugar dissolving in water

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

hydrogen burns in air to form water

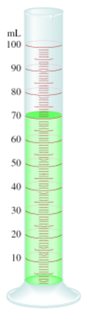




Extensive and Intensive Properties

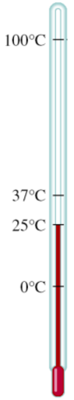
An ***extensive property*** of a material depends upon how much matter is being considered.

- mass
- length
- volume

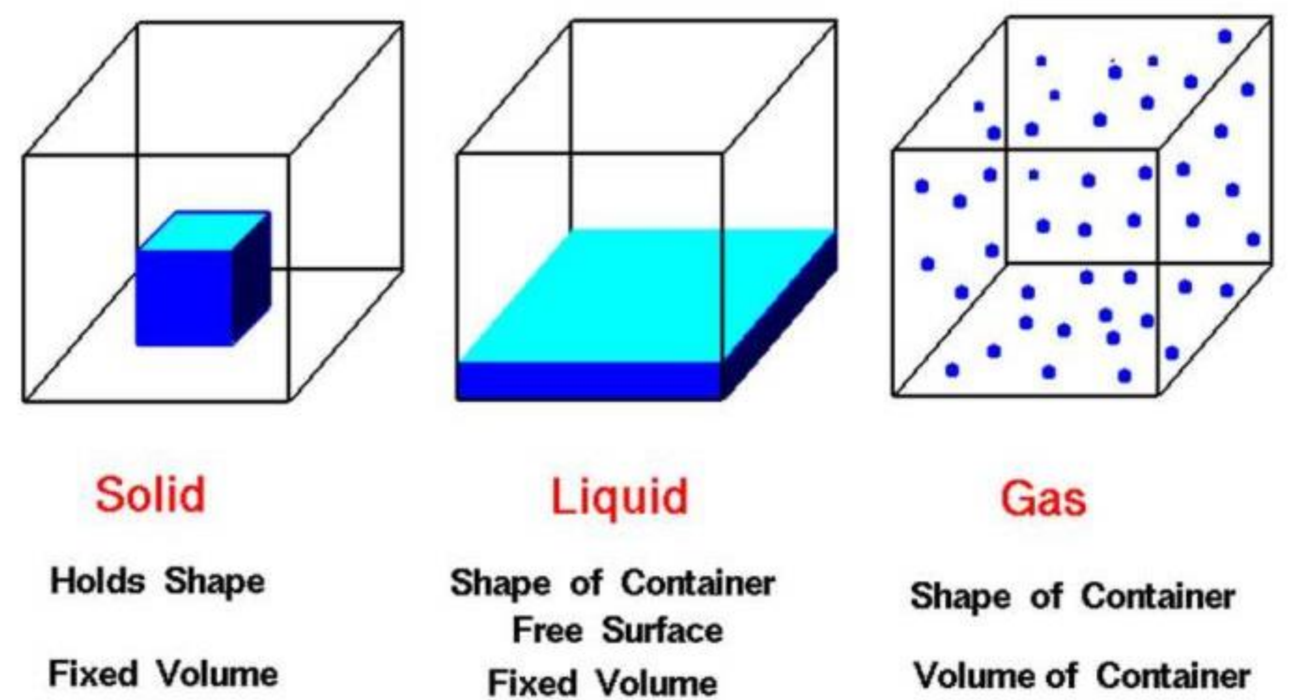


An ***intensive property*** of a material does not depend upon how much matter is being considered.

- density
- temperature
- color



The States/Phases of Matter



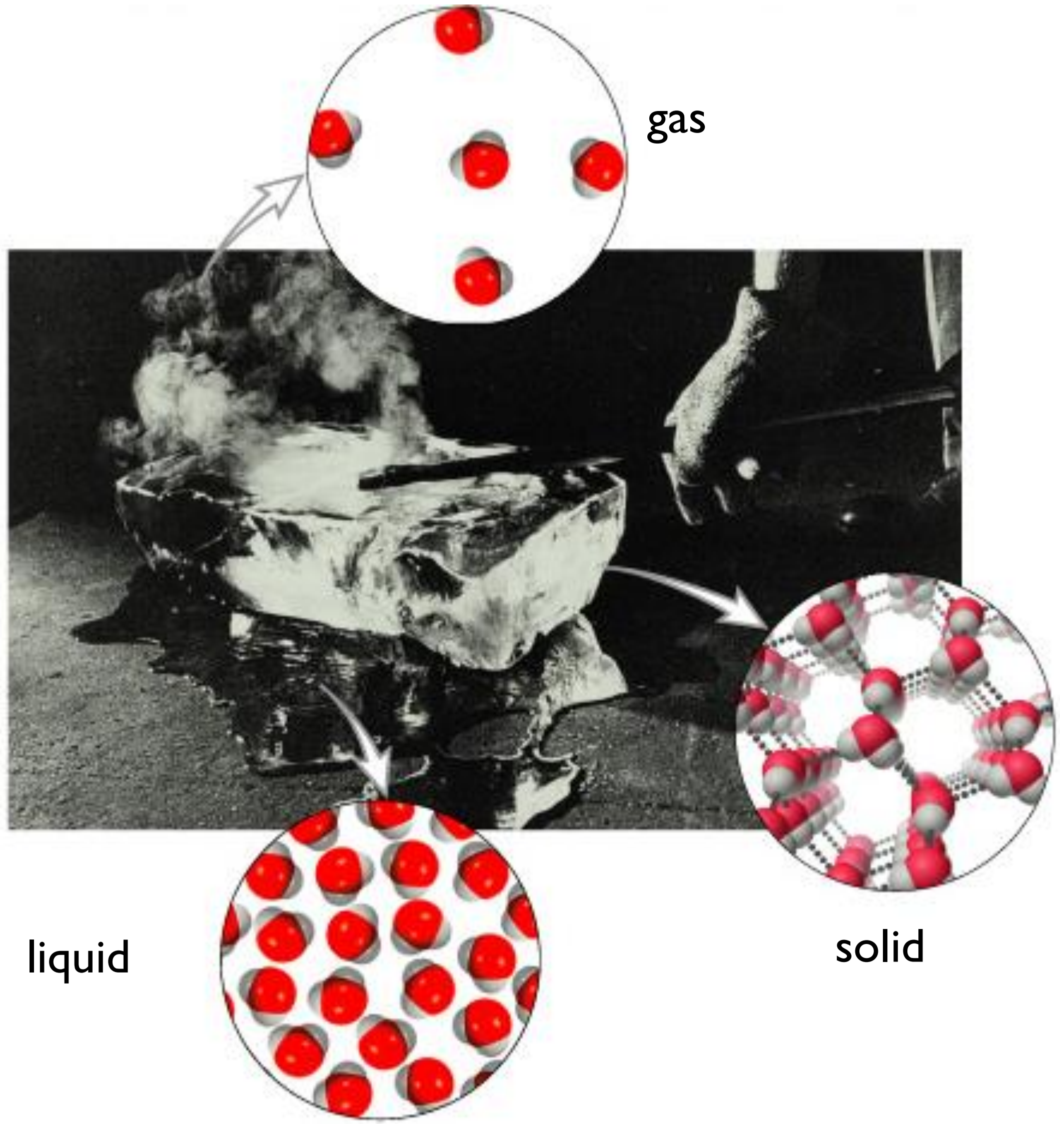
Difference in the Three States of Matter

Difference in the characteristics of states of matter		
Solid	Liquid	Gas
Definite shape	Indefinite shape	Indefinite shape
Definite volume	Definite volume	Indefinite volume
Maximum force of attraction between particles	Less forces of attraction between particles compare to solid	Negligible force of attraction between particles
Particles are closely packed	Particles are loosely packed compared to solid	Particles are loosely packed
Cannot be compressed	Cannot be compressed	Can be compressed
Kinetic energy of particles is minimum	Kinetic energy of particles is more than solid	Kinetic energy of particles is maximum
Particles cannot move rather they vibrate only at their fixed position	Particles can slide over one another	Particles can move freely
Highest density	Density is lower than solid	Lowest density
Cannot flow	Flow	Flow

Conversion among States of Matter

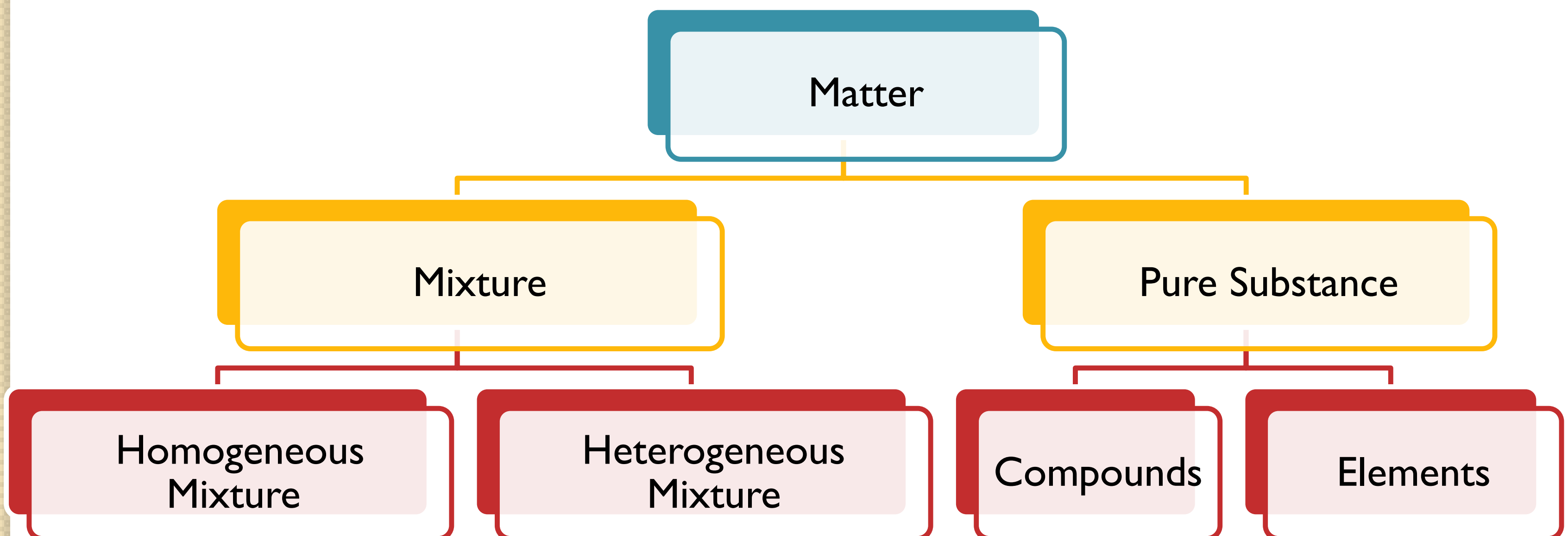
- The three states of matter can be inter-converted without changing the composition of the substance.
- Upon heating or by removing the heat states of matter can be changeable .
- Melting Point: *The temperature at which the transition of a matter from solid state to liquid state occurs.*
- Boiling Point: *The temperature at which the transition of a matter from liquid state to gaseous state occurs.*

The States of Matter



Physical change

Classification of Matter



Classification of Matter

- Substance:
  - A substance is a form of matter that has a definite (constant) composition and distinct properties.
  - Example: Water, $NH_{3,}$ Sugar, Gold, Oxygen
  - Differ form one another by their composition.
  - Identified by their appearance- smell, taste etc.
- Substances can be two types-
  - Elements
  - Compounds

Classification of Matter

- Elements:
  - An element is a substance that cannot be separated into simpler substances by chemical means.
  - To date 118 elements have been identified
  - Most of them occur naturally, and a few are made by scientists via nuclear processes.
  - Elements are written with symbols of 1 or2 letters. If it is a two letter symbol, then the 1st letter must be in capital and the 2nd is small.
  - Example: C(carbon), O (oxygen), Co(cobalt)

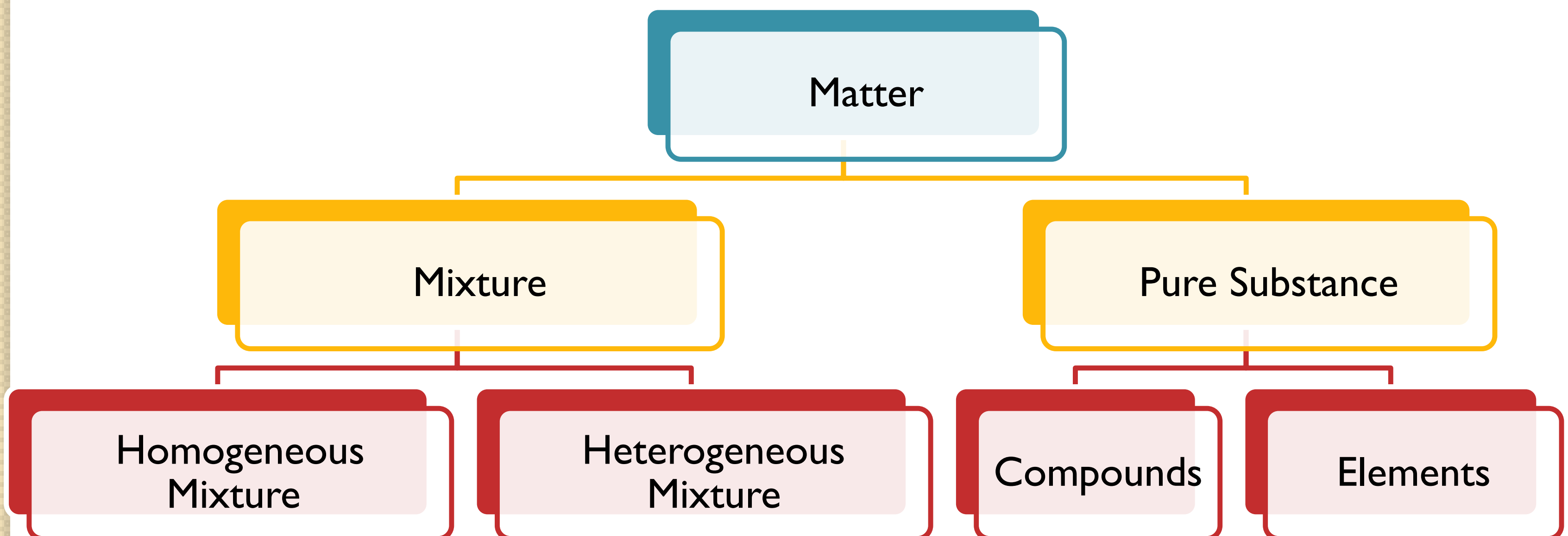
TABLE 1.1 Some Common Elements and Their Symbols

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

Classification of Matter

- Compound:
  - Compound is a substance composed of atoms of 2 or more elements, chemically united in fixed proportions.
  - They can be separated by chemical means into their pure components.
  - They are written with the symbols of elements they contain in exact proportions.
  - Example: CO(Carbon monooxide), CO_2 (Carbon di oxide)

Classification of Matter



Classification of Matter

- Mixture:
  - Combination of two or more substances, in which substances will retain their distinct identities.
  - Example: Air, Coke, milk, cement etc.
  - Do not have any constant composition.

magnet

Classification of Matter

- Mixtures are mainly two types:
  - Homogenous Mixture
  - Heterogeneous Mixture
- Homogenous Mixture:
  - Composition of mixture same throughout. i.e. sugar in water.
  - Usually solutions are homogenous mixture.
- Heterogeneous Mixture:
  - Composition of the mixture is not same. i.e. sand and iron particles.
  - Usually components are visible while mixing as made up of very diverse substances.

Classify each of the following as an element, a compound a homogeneous mixture, or a heterogeneous mixture:

(a) Seawater,

(b) Helium gas,

(c) Sodium chloride (table salt),

(d) A bottle of soft drink, ?

(e) A milkshake, ?

(f) Air in a bottle,

(g) Concrete.

Measurement

- As a chemist, things can be viewed at the macroscopic and microscopic level.

- <u>Macroscopic Property Measurement:</u>

The macroscopic level includes anything seen with the naked eye and they can be determined directly

1. <u>Length/Scale</u>: meter stick
2. <u>Volume</u>: Burette, pipette, graduated cylinder, volumetric flask
3. <u>Mass</u> :balance measures
4. <u>Temperature</u>: thermometer

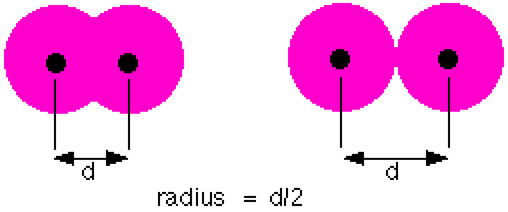
Measurement

❑<u>Microscopic Property Measurement</u>:

• The microscopic level includes atoms and molecules, things not seen with the naked eye.

• They are determined by an indirect method.

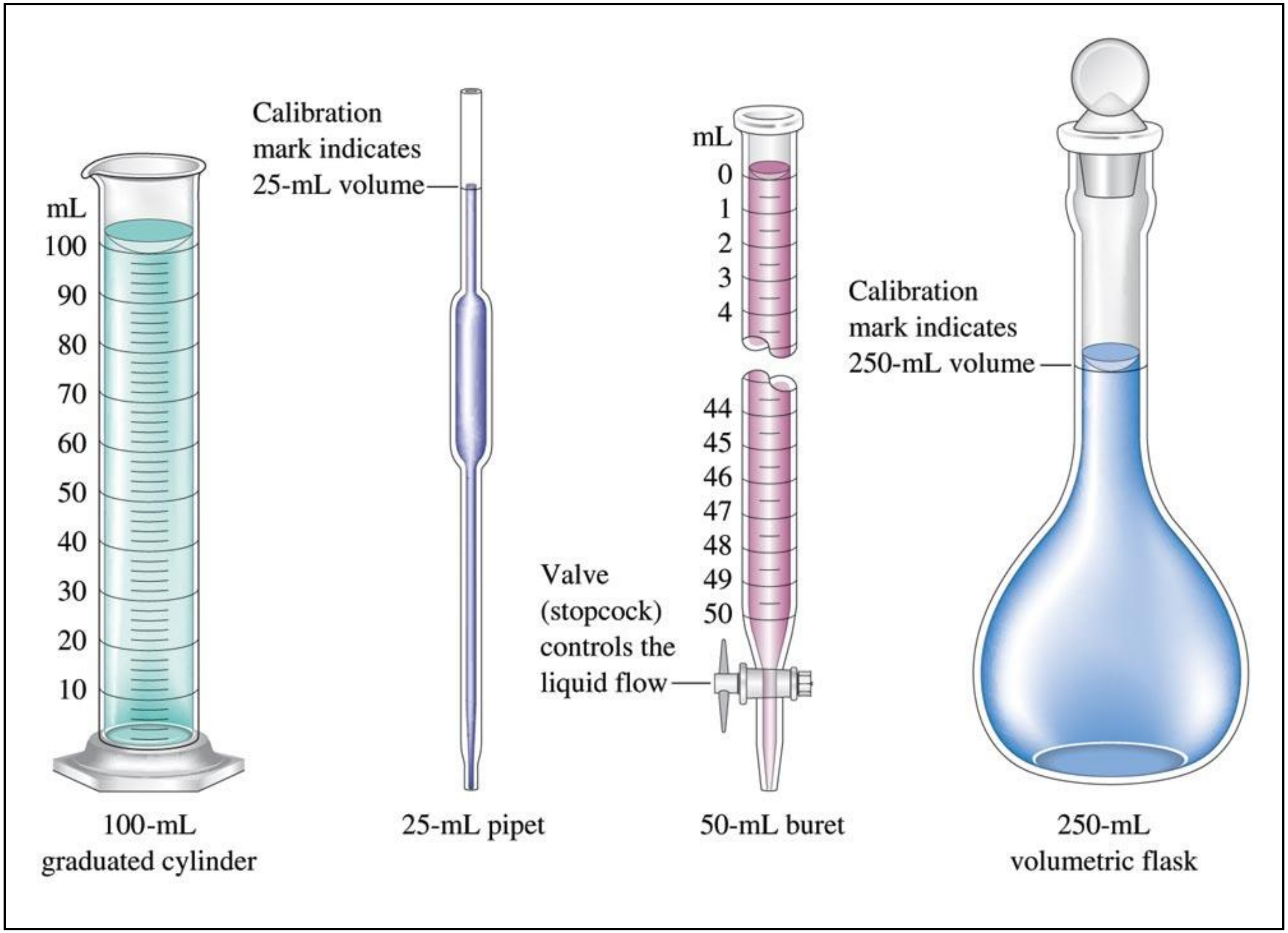
-<u>Atomic radius</u>: The radius of an atom is found by measuring the distance between the nuclei of two touching atoms, and then halving that distance.



- <u>Electronegativity</u>: Electronegativity of an element is measured only in relation to the electronegativity of other elements.

Common Types of Laboratory Equipment Used to Measure Mass & Length

Common Types of Laboratory Equipment Used to Measure Liquid Volume



SI Units: International System of Units

❖ The SI (Système International d'Unités) is a globally agreed system of units, with seven base units

❖ The SI system specifies a particular and modified choice of metric units for use in scientific measurements.

❖ SI units and their relative values were adopted by the 11th General Conference on Weights and Measures (CGPM) in Paris in 1960.

International System of Units (SI)

❖ The SI system has seven base units from which all other units are derived. The table below lists these base units and their symbols. In this lecture series we will mostly consider the base units for length, mass, and temperature. The seven SI base units are:

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

International System of Units (SI)

❖ Prefixes are used to indicate multiples or decimal fractions of various units.

❖ The table presents prefixes commonly encountered in chemistry:

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

SI Units: International System of Units

Mass:

- Mass is the amount of matter in an object.
- Mass is always fixed for a specific object.
- The usual symbol for mass is 'm' and its SI unit is kilogram (kg).

Weight:

- The weight of an object is the force of gravity on the object and may be defined as the mass times the acceleration of gravity, $w = mg$.
- Weight is a force and its SI unit is Newton (N).

SI Units: International System of Units

Volume:

- Volume is the amount of three-dimensional space an object occupies.
- SI unit: $\mathbf{m^3}$; as unit for length is Meter(m).
- For smaller objects chemists use cm^3 (cubic centimeter) and dm^3(cubic decimeter)
- Liter is also an unit for volume of water/ liquid but not SI.

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

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

$$\mathbf{1\ L = 1\ dm^3} \quad / \quad \mathbf{1\,mL = 1\ cm^3}$$

International System of Units (SI)

Density:

- It is the amount of mass in a unit volume of that substance:
  - Mass in per unit volume.

$$d = m/v$$

  - Unit:

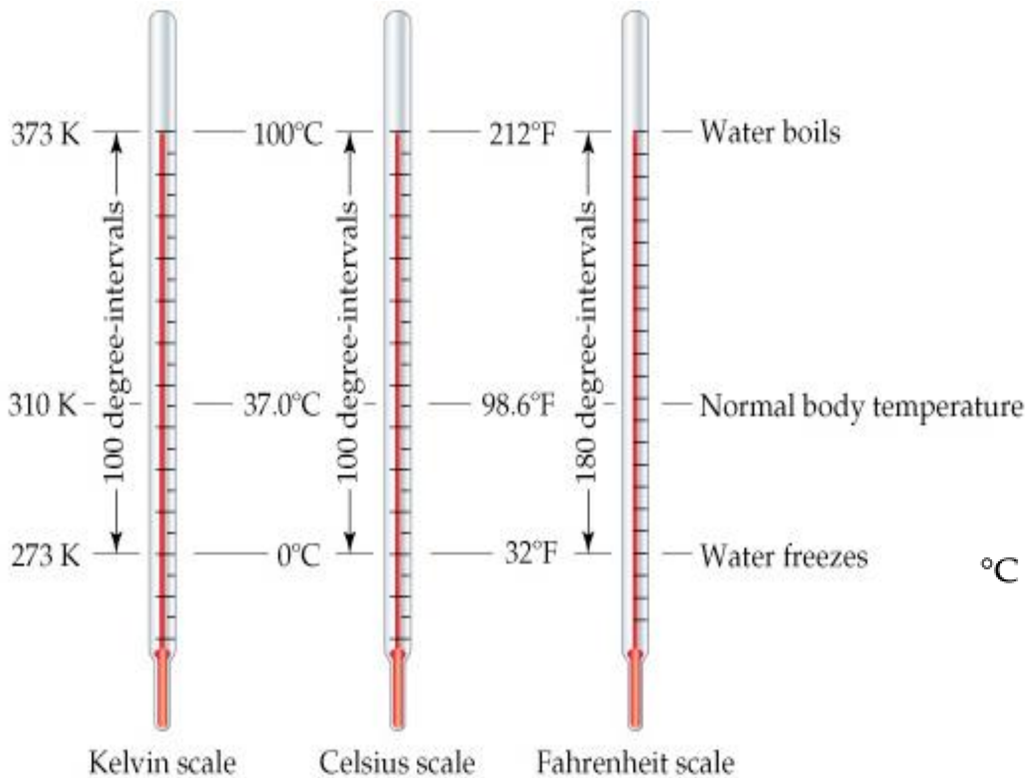
g/cm^3 or, g/mL

Practice Problem:

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.

Temperature Scales

We sense temperature as a measure of the hotness or coldness of an object. Indeed, temperature determines the direction of heat flow.



$$K = °C + 273.15$$

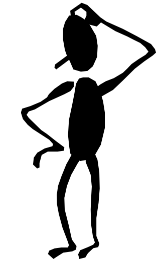
273 K = 0 ^{0}C
373 K = 100 ^{0}C

$$°C = \frac{5}{9}(°F - 32) \quad \text{or} \quad °F = \frac{9}{5}(°C) + 32$$

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

Comparison of the Kelvin, Celsius and Fahrenheit scale

Convert 172.9 ^{0}F to degrees Celsius:



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

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

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

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

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

Practice Exercises:

Convert the following temperature to °Celsius/ °Fahrenheit/Kelvin –

- 95 ° F→ C

- 77K → F

Ans: 35 °C, -321.07 °F

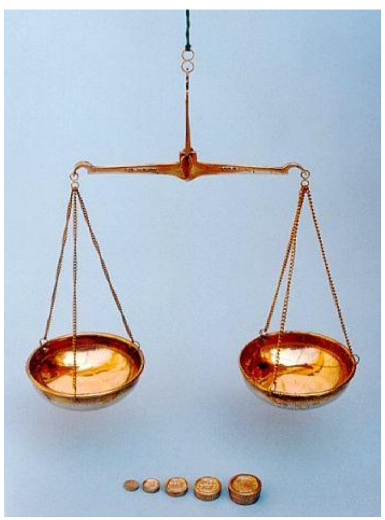
Significant Figures

10.5583 g

$\pm$0.0001 g

10.55 g ?

Last digit is uncertain.



1.55 kg

$\pm$ 0.01 kg

1.5583 kg ?

The **significant figures** of a number are those digits that carry meaning contributing to its precision.

Who are 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 equals/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 are in the middle of nonzero digits 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	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.

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

+1.1 ← one significant figure after decimal point

90.432 ← round off to 90.4

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

3.70 ← two significant figures after decimal point

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

Scientific Notations

- Chemists often work with numbers which are either extremely large or extremely small
- 1 g Hydrogen has

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

  Mass of Hydrogen atom

  0.00000000000000000000000166 g

- To work with very large and very small numbers, we use a system called *Scientific Notation*

Scientific Notations

- We can think of 5.6 x 10ˆ-9 as the product of two numbers: 5.6 (the digit term) and 10-9 (the exponential term)
- Some examples:

$10000 = 1 \times 10^4$	$24327 = 2.4327 \times 10^4$
$1000 = 1 \times 10^3$	$7354 = 7.354 \times 10^3$
$100 = 1 \times 10^2$	$482 = 4.82 \times 10^2$
$10 = 1 \times 10^1$	$89 = 8.9 \times 10^1$ (not usually done)
$1 = 10^0$	
$1/10 = 0.1 = 1 \times 10^{-1}$	$0.32 = 3.2 \times 10^{-1}$ (not usually done)
$1/100 = 0.01 = 1 \times 10^{-2}$	$0.053 = 5.3 \times 10^{-2}$
$1/1000 = 0.001 = 1 \times 10^{-3}$	$0.0078 = 7.8 \times 10^{-3}$
$1/10000 = 0.0001 = 1 \times 10^{-4}$	$0.00044 = 4.4 \times 10^{-4}$

- The exponent of 10 is the number of places the decimal point must be shifted to give the number in long form.

Scientific Notations

- A positive exponent shows that the decimal point is shifted that number of places to the right. A negative exponent shows that the decimal point is shifted that number of places to the left.

- In scientific notation, the digit term indicates the number of significant figures in the number. The exponential term only places the decimal point.

As an example,
46600000 = 4.66 x 10^7
This number only has 3 significant figures. The zeros are not significant; they are only holding a place.

As another example,
0.00053 = 5.3 x 10^-4
This number has 2 significant figures. The zeros are only place holders.

Scientific Notations- Addition & Subtraction

- $(7.4\times10^3) + (2.1\times 10^3) = 9.5\times10^3$
- $(4.31\times10^4)+(3.9\times10^3)=(4.31\times10^4)+(0.39\times10^4)$
  $=4.70\times10^4$

- $(2.22\times10^{-2})-(4.10\times10^{-3}) = (2.22\times10^{-2})-(0.41\times10^{-2})$
  $=1.81\times10^{-2}$

Scientific Notations- Multiplication & division

- $(8.0\times10^4) \times (5.0\times10^2) = (8.0\times5.0)\ (10^{4+2})$

  $= 40 \times 10^6$

  $= 4.0 \times 10^7$

  $(4.0\times10^{-5}) \times (7.0\times10^3) = (4.0\times7.0)\ (10^{-5+3})$

  $= 28 \times 10^{-2}$

  $= 2.8 \times 10^{-1}$

- $6.9\times10^7 / 3.0\times10^{-5} = 6.9/3.0 \times 10^{7-(-5)}$

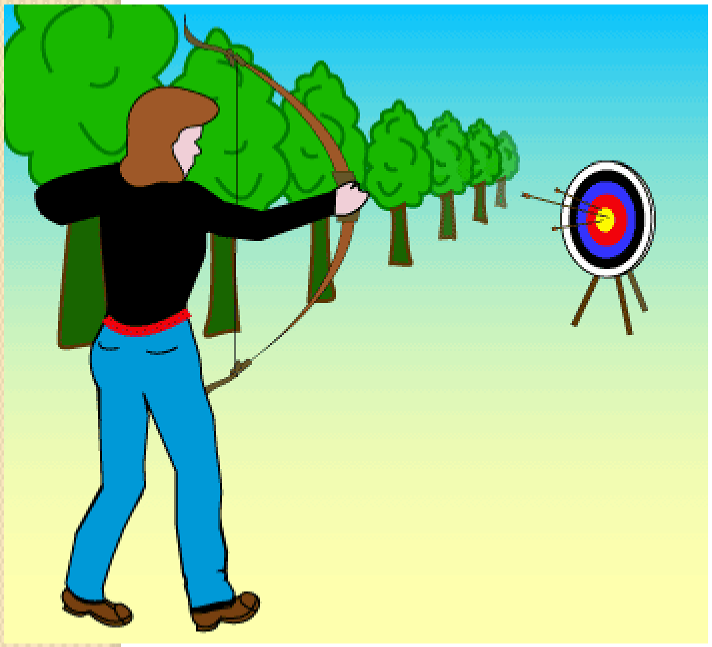
  $= 2.3 \times 10^{12}$

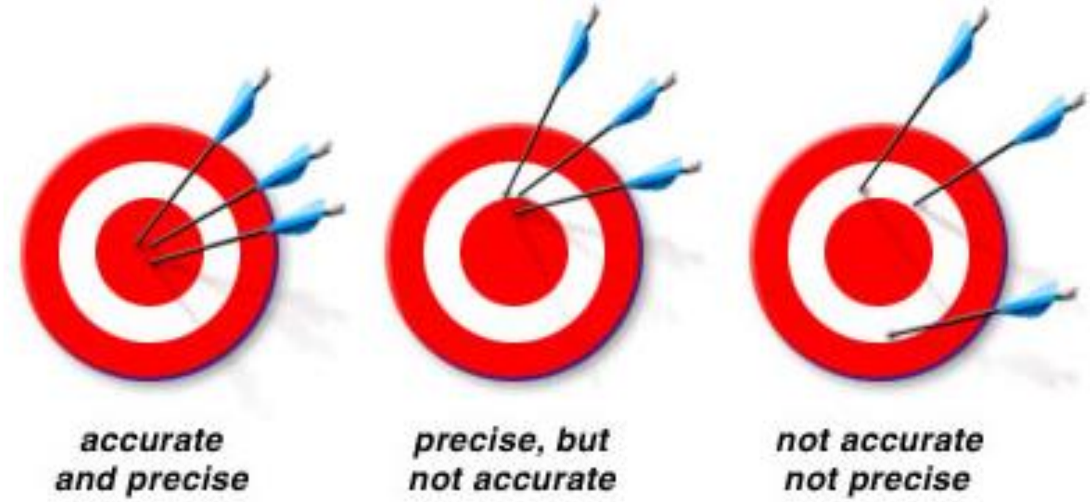
Accuracy and Precision

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

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

• If you shoot a quiver of arrows at a target, several outcomes are possible with different accuracy and precision.





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

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

How many mL are in 1.63 L?

Conversion Unit 1 L = 1000 mL

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

Conversion factor = 1000ml/1L

EXAMPLE 1.5

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams (1 mg = 1×10^{-3} g). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversion is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors:

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}$$

(Continued)

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = \boxed{3.78 \times 10^4 \text{ mg}}$$

Practice Exercise:

The density of the lightest metal, Lithium (Li) is 5.34 X 10^2 Kg/m3. Convert the density to g/cm3.

Key Equations

$d = \frac{m}{V}$ (1.1)	Equation for density
$? \,°C = (°F - 32°F) \times \frac{5°C}{9°F}$ (1.2)	Converting °F to °C
$? \,°F = \frac{9°F}{5°C} \times (°C) + 32°F$ (1.3)	Converting °C to °F
$? \,K = (°C + 273.15°C)\frac{1\ K}{1°C}$ (1.4)	Converting °C to K

THANK YOU

Practice related mathematical problems from Raymond Chang, 9th Ed.