WHAT IS CHEMISTRY?

- Chemistry is the science that deals with the ***materials*** of the universe, and the ***changes*** they undergo.

- ***Materials*** of the universe can be of several forms:

  Gas: air, oxygen

  Liquid: water, gasoline, vinegar, orange juice,

  Solid: rocks, charcoal, table salt, sugar, wood, baking soda

- Some examples of **changes**:

Burning of charcoal

charcoal + oxygen $\longrightarrow$ carbon dioxide

Burning of gasoline

gasoline + oxygen $\longrightarrow$ carbon dioxide+ water vapor

Fermentation of grape juice

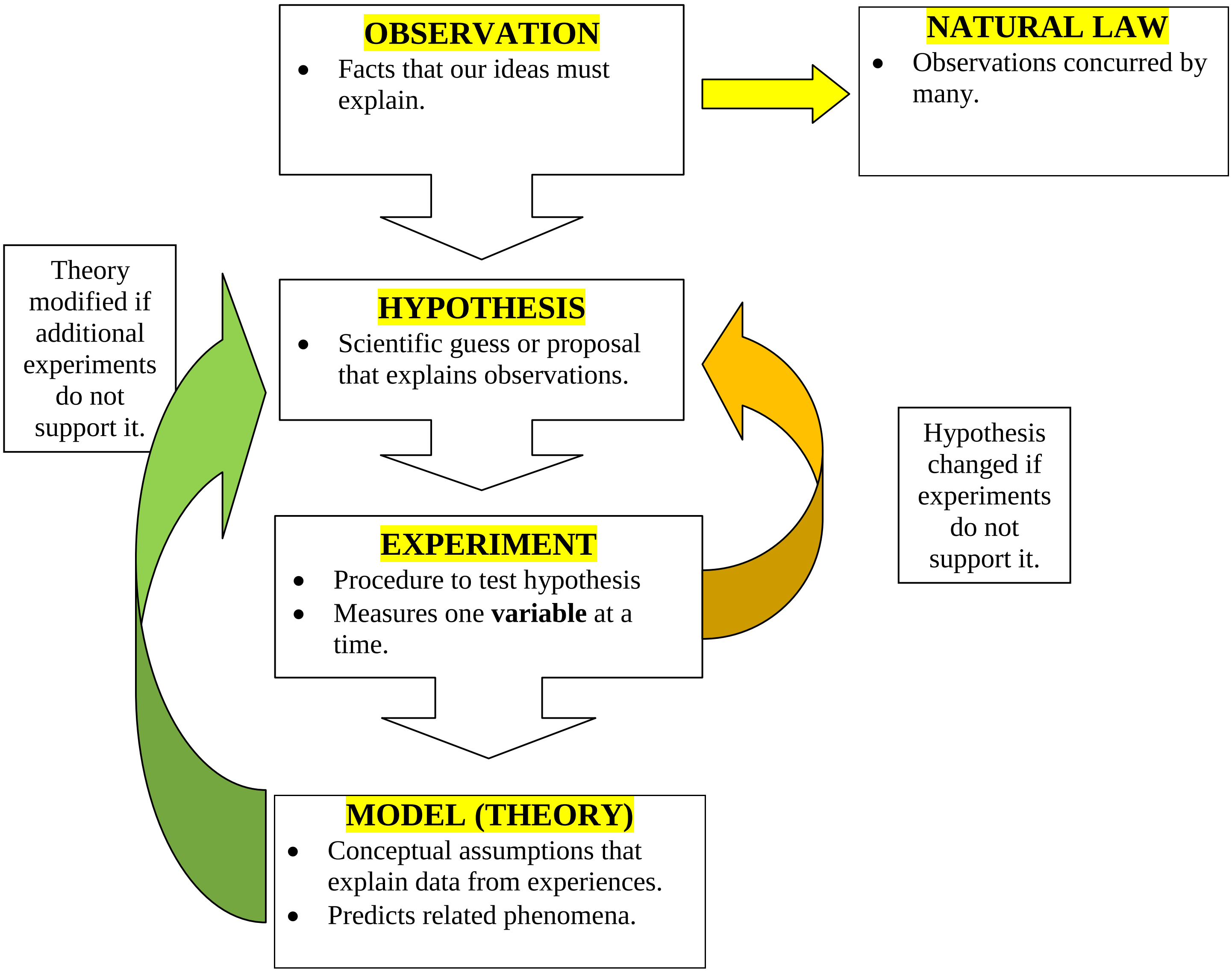
glucose $\longrightarrow$ ethyl alcohol + carbon dioxide
(in water) (in water)

Souring of wine

ethyl alcohol + oxygen $\longrightarrow$ acetic acid
(in water) (from air) (in water)

THE SCIENTIFIC METHOD

- The scientific method is a **process** of creative thinking and testing aimed at **objective** and **verifiable** discoveries. It is generally composed of the following steps:



MEASUREMENTS / SI (METRIC) UNITS

- **Measurements** are made by scientists to determine size, length and other **properties** of matter.
- For measurements to be useful, a measurement **standard** must be used.
- A **standard** is an exact quantity that people agree to use for **comparison**.
- **SI (Metric)** system is the **standard** system of measurement used **worldwide** by scientists.

SI BASE UNITS:

- There are 7 base SI units. Five of these are of importance in the study of chemistry.

Quantity Measured	***Metric Units***	***Symbol***	***English Units***
Length	Meter	m	yd
Mass	Kilogram	kg	lb
Time	Seconds	s	s
Temperature	Kelvin	K	°F
Amount of substance	Mole	mol	mol

DERIVED UNITS:

- In addition to the **fundamental** units above, several useful **derived** units are commonly used in SI system.

Quantity Measured	***Units***	***Symbol***
Volume	Liter	L
Density	grams/cc	g/cm^3

SCIENTIFIC NOTATION

- Scientific Notation is a convenient way to express **very large** or **very small** quantities.

General form:

$$A \times 10^{n} \qquad 1 \leq A < 10 \qquad n = \text{integer}$$

- Converting between decimal and scientific notation:
  1. Move the decimal point in the original number so that it is located after the first nonzero digit.
  2. Follow the new number by a multiplication sign and 10 with an exponent (***power***).
  3. The exponent is equal to the number of places that the decimal point was shifted.
  4. For numbers smaller than 1, the decimal moves to the left and the power becomes negative.

75,000,000 changes to 7.5×10^{7} (7 to the left)
0.00642 changes to 6.42×10^{-3} (3 to the right)

Examples:

1. Write 6419 in scientific notation:
2. Write 0.000654 in scientific notation:

- Addition and subtraction (NOT COVERED)
- Multiplication and division :
  1. Change numbers to exponential form.
  2. Multiply or divide coefficients.
  3. **Add** exponents if **multiplying**, or **subtract** exponents if **dividing**.
  4. If needed, reconstruct answer in **standard** exponential notation.

Examples:

1. Multiply 30,000 x 600,000

$$(3\times10^{4})\ (6\times10^{5}) = 18 \times 10^{(4+5)} = 18 \times 10^{9} = 1.8 \times 10^{10}$$

2. Divide 30,000 by 0.006

$$\frac{3 \times 10^{4}}{6 \times 10^{-3}} = \frac{3}{6} \times 10^{[4-(-3)]} = 0.5 \times 10^{7} = 5 \times 10^{6}$$

Follow-up Problems:

1. $(5.5 \times 10^{3})(3.1 \times 10^{5}) =$

2. $(9.7 \times 10^{14})(4.3 \times 10^{-20}) =$

3. $\frac{2.6\text{x}10^{6}}{5.8\text{x}10^{2}} =$

4. $\frac{1.7\text{x}10^{-5}}{8.2\text{x}10^{-8}} =$

5. $(3.7 \times 10^{-6}) \times (4.0 \times 10^{8}) =$

6. $(8.75\text{x}10^{14})(3.6\text{x}10^{8}) =$

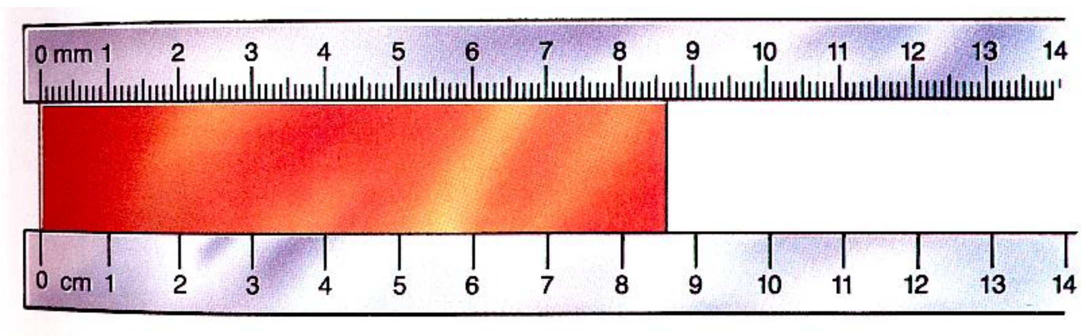
7. $\frac{1.48\text{x}10^{-28}}{7.25\text{x}10^{13}} =$

ERROR IN MEASUREMENTS

Two kinds of numbers are used in science:

- **Counted or Defined**: exact numbers; **no uncertainty**
- **Measured**: are subject to error; have **uncertainty**

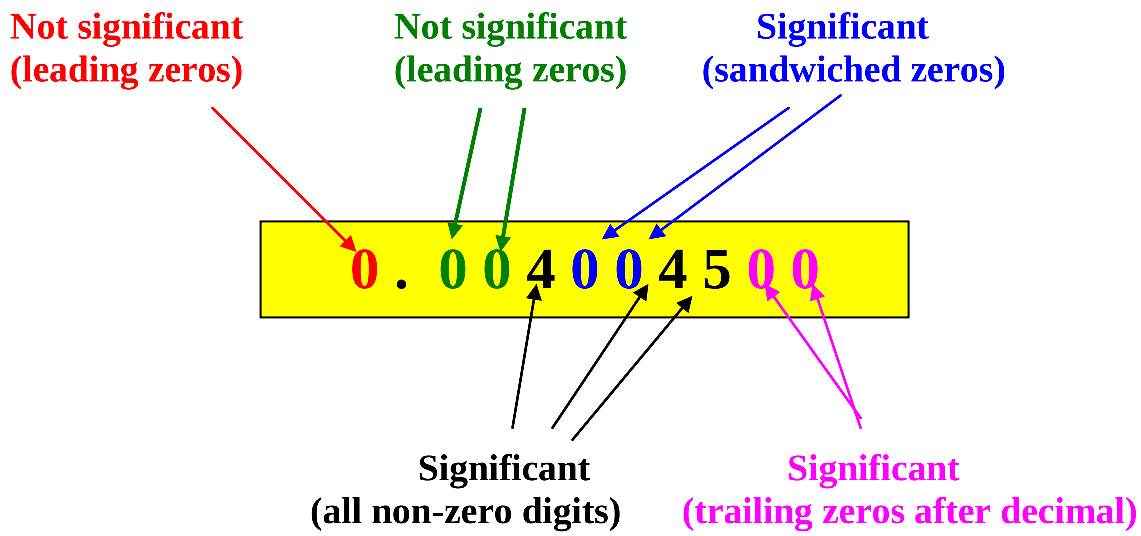
8.65 cm



8.6 cm

- Every measurement has **uncertainty** because of instrument limitations and human error.
- The **last** digit is the **estimated** one.
- **Significant numbers** are the **certain** and **uncertain** digits.

SIGNIFICANT FIGURE RULES



Examples:
Determine the number of significant figures in each of the following measurements:

461 cm	93.500 g
1025 g	0.006 m
0.705 mL	5500 km

Rounding Off Rules

- If the rounded digit is **<5**, the digit is simply dropped.
- If the rounded digit is **≥5**, the digit is increased.

Examples:

	3 sig. figs	**2 sig. figs.**
8.4234 rounds off to	8.42	8.4
14.780 rounds off to	14.8	15
3256 rounds off to	3260 (3.26×10^3)	3300 (3.3×10^3)

SIGNIFICANT FIGURES AND CALCULATIONS

- The results of a calculation cannot be more precies than the least precise measurement.

- For **multiplication and division,** the **answer** must contain the **same number of significant figures** as there are in the measurement with the **fewest significant figures**.

9.2 x 6.80 x 0.3744 = **23**.4225 (calculator answer)

= **23** (rounded answer)

- For **addition and subtraction**, the **answer** must have the **same number of decimal places** as there are in the measurement with the **fewest decimal places**.

$$\begin{array}{r} 83.\mathbf{5} \\ +\ \underline{23.28} \\ 106.78 \end{array}$$ (calculator answer)

106.**8** (rounded answer)

Examples:

1) 5.008 + 16.2 + 13.48 =

2) $\dfrac{3.15 \text{ x } 1.53}{0.78} =$

3) 104.45 mL – 0.838 mL + 46 mL =

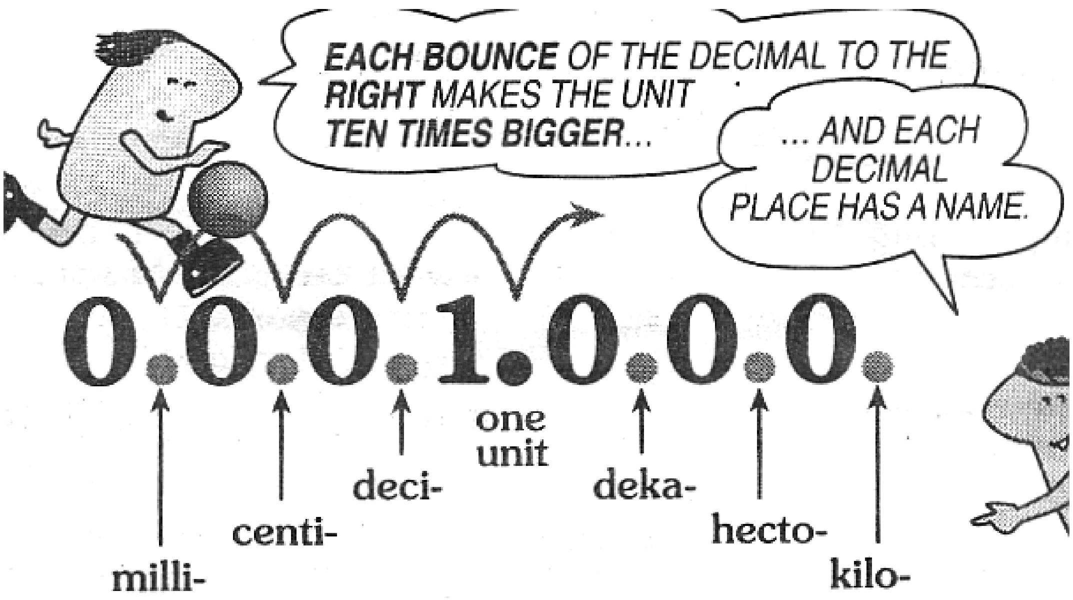
4) $\dfrac{4.0 \text{ x } 8.00}{16} =$

SI PREFIXES

- The **SI** system of units is easy to use because it is based on **multiples of ten**.
- Common **prefixes** are used with the base units to indicate the multiple of ten that the unit represents.

SI PREFIXES

Prefixes	**Symbol**	**Multiplying factor**	
mega-	M	1,000,000	10^6
kilo-	k	1000	10^3
centi-	c	0.01	10^{-2}
milli-	m	0.001	10^{-3}
micro-	μ	0.000,001	10^{-6}



CONVERSION FACTORS

- Many problems in chemistry and related fields require a change of units.
- Any unit can be converted into another by use of the appropriate **conversion factor**.
- Any equality in units can be written in the form of a fraction called a **conversion factor**. For example:

Equality: 1 m = 100 cm

Conversion factors: $$\frac{1\ \text{m}}{100\ \text{cm}} \quad \text{or} \quad \frac{100\ \text{cm}}{1\ \text{m}}$$

Metric-Metric

Equality: 1 kg = 2.20 lb

Conversion factors: $$\frac{1\ \text{kg}}{2.20\ \text{lb}} \quad \text{or} \quad \frac{2.20\ \text{lb}}{1\ \text{kg}}$$

Metric-English

- Sometimes a conversion factor is given as a percentage. For example:

Percent quantity: 18% body fat by mass

Conversion factors: $$\frac{18\ \text{kg body fat}}{100\ \text{kg body mass}} \quad \text{or} \quad \frac{100\ \text{kg body mass}}{18\ \text{kg body fat}}$$

Percentage

CONVERSION OF UNITS

- Problems involving conversion of units and other chemistry problems can be solved using the following step-wise method:

    1. Determine the intial unit given and the final unit needed.
    2. Plan a sequence of steps to convert the intial unit to the final unit.
    3. Write the conversion factor for each units change in your plan.
    4. Set up the problem by arranging cancelling units in the numerator and denominator of the steps involved.

$$\cancel{\textbf{beginning unit}} \times \frac{\textbf{final unit}}{\cancel{\textbf{beginning unit}}} = \textbf{final unit}$$

↑

conversion factor

Examples:

1. Convert 164 lb to kg (1 kg = 2.20 lb)

Step 1: Given 164 lb Need kg

Step 2: lb → Metric – English factor → kg

Step 3: $\frac{1\ \text{kg}}{2.20\ \text{lb}}$ or $\frac{2.20\ \text{lb}}{1\ \text{kg}}$

Step 4: $164\ \text{lb} \times \frac{1\ \text{kg}}{2.20\ \text{lb}} = 74.5\ \text{kg}$

2. The thickness of a book is 2.5 cm. What is this measurement in mm?

Step 1: Given Need

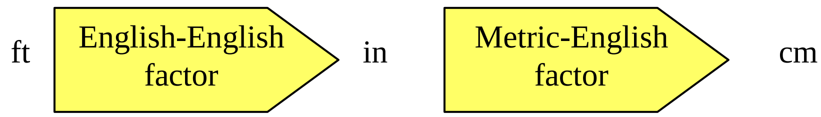
Step 2 & 3:

Step 4: cm x ________ = m

3. How many centimeters are in 2.0 ft? (1 in=2.54 cm)

Step 1: Given 2.0 ft Need cm

Step 2:



Step 3: $\frac{1\text{ ft}}{12\text{ in}}$ and $\frac{1\text{ in}}{2.54\text{ cm}}$

Step 4: 2.0 ft x ——— x ———— = cm

4. Bronze is 80.0% by mass copper and 20.0% by mass tin. A sculptor is preparing to cast a figue that requires 1.75 lb of bronze. How many grams of copper are needed for the brass figure?

Step 1: Given Need

Step 2:

Step 3:

Step 4: 1.75 lb bronze x ——— x ———— = g copper

VOLUME & DENSITY

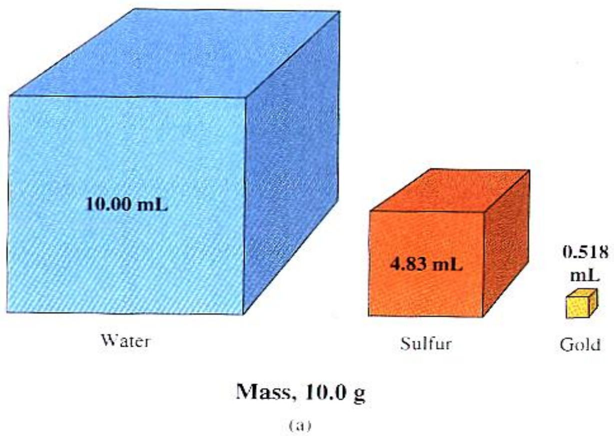
- **Volume** is the amount of **space** an object occupies. Common units are **cm^3** or **liter** (L) and **milliliter** (mL).

1 L = 1000 mL **1 mL=1 cm^3**

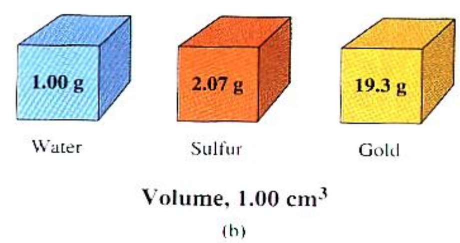
- **Density** is **mass per unit volume** of a material. Common units are **g/cm^3** or **g/mL**.

$$\text{Density}=\frac{\text{mass}}{\text{volume}} \qquad d=\frac{m}{v}$$

- Density **is** <u>directly</u> related to mass **of an object, and** <u>indirectly</u> related to the volume **of an object.**



Comparison of the *volume* of equal masses of 3 materials with different *densities*



Comparison of the *masses* of equal volumes of 3 materials with different *densities*

VOLUME & DENSITY

Examples:

1. A copper sample has a mass of 44.65 g and a volume of 5.0 mL. What is the density of copper?

   m = 44.65 g $\qquad$ $\mathbf{d = \dfrac{m}{v} =}$

   v = 5.0 mL

   d = ???

2. A silver bar with a volume of 28.0 cm^3 has a mass of 294 g. What is the density of this bar?

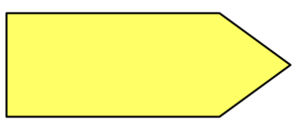
   m =

   v =

   d =

3. If the density of gold is 19.3 g/cm^3, how many grams does a 5.00 cm^3 nugget weigh?

   Step 1: Given Need

   Step 2: 

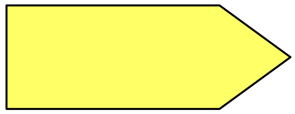
   Step 3:

   Step 4: x ———————— =

4. If the density of milk is 1.04 g/mL, what is the mass of 0.50 qt of milk? (1L = 1.06 qt)

   Step 1: Given Need

   Step 2:

   Step 3:

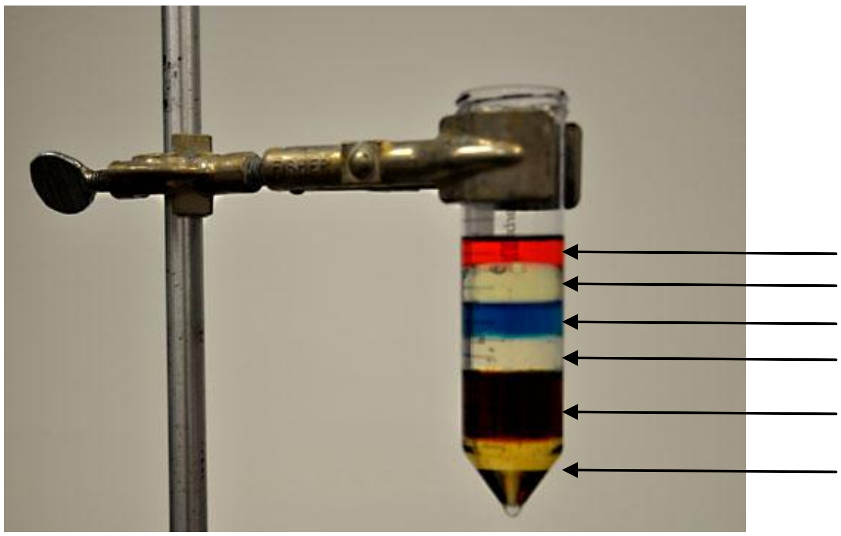
   Step 4:

   x ________ =

5. What volume of mercury has a mass of 60.0 g if its density is 13.6 g/ml?

DENSITY & FLOATING

- Objects float in liquids when their density is lower relative to the density of the liquid.
- The density column shown below was prepared by layering liquids of various densitites.



- The greater the density of the liquid, the lower it layers itself.