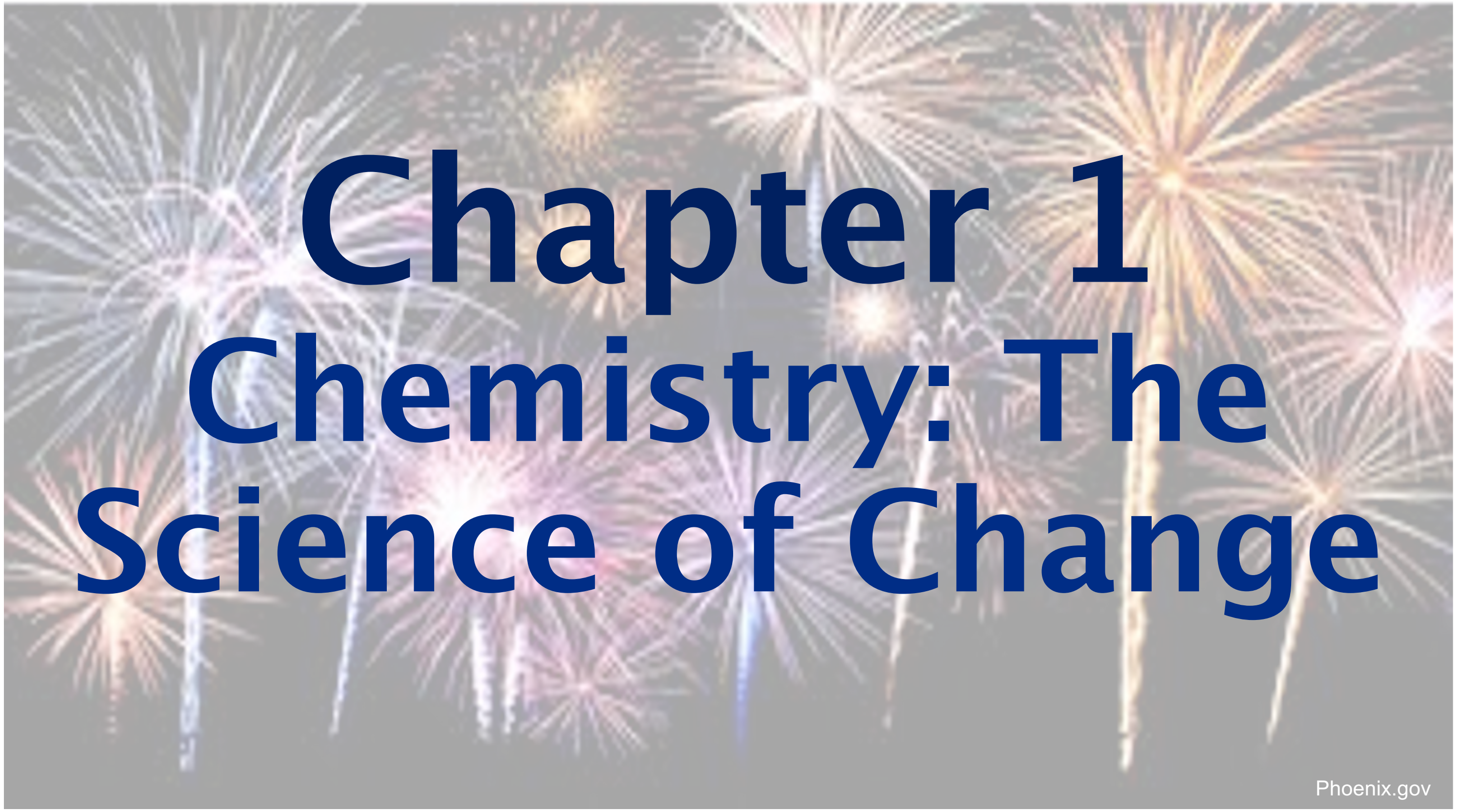
Chapter 1

Chemistry: The Science of Change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The science that studies the properties of substances & how substances react with one another.

How stuff works on a molecular/atomic/subatomic level

Chemistry!

MATTER

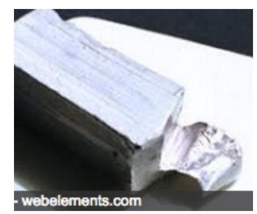
Has mass & takes up space

ENERGY

The capacity to do work or cause change

REACTIONS

How materials interact & change



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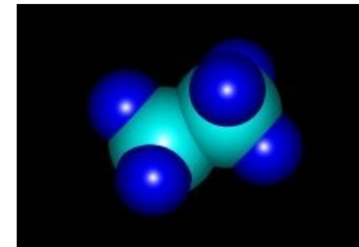
Learning the Language

Chemistry describes materials and predicts behavior using three basic concepts



Composition: What is in a material

- Mass percent of elements/compounds
- Atomic/molecular ratios within material
- Stoichiometry

Structure

- Molecular/ionic/atomic arrangement
- Phase (solid, liquid, gas)

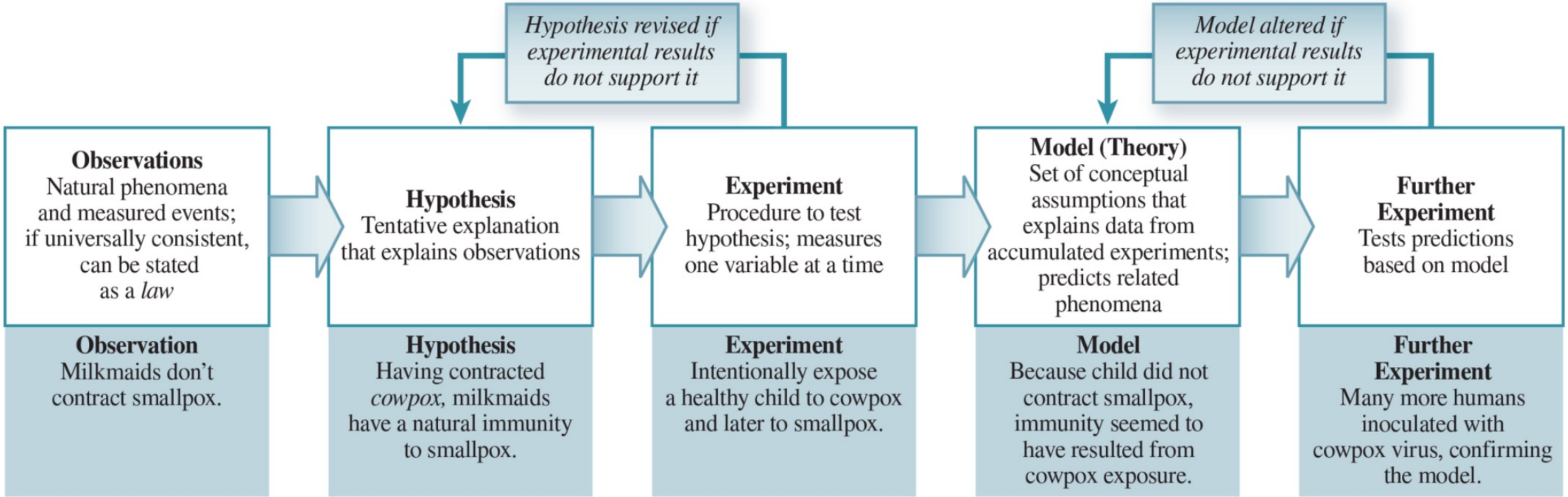
Properties – chemical & physical

- Specific to a particular material
- ex: boiling point, color, odor, reactivity
- Used for identification

Often looking at materials at the submicroscopic level – too small to see with the human eye

The Scientific Method

Series of steps that explain an observation

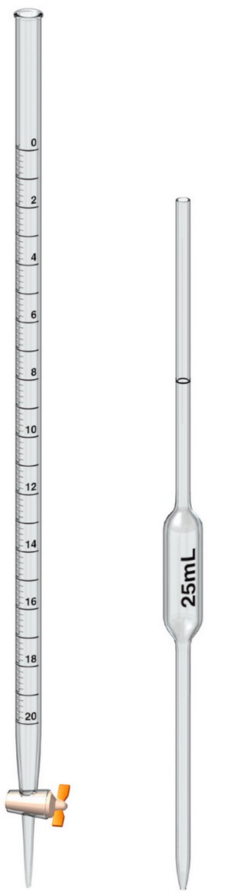


Exposure to a virus can enable humans to build an immunity to that virus – enabled the development of vaccines

Most vaccines today use inactivated viruses - safer

Measurements

Determining how much matter is present



Burette
(a)

Volumetric pipette
(b)

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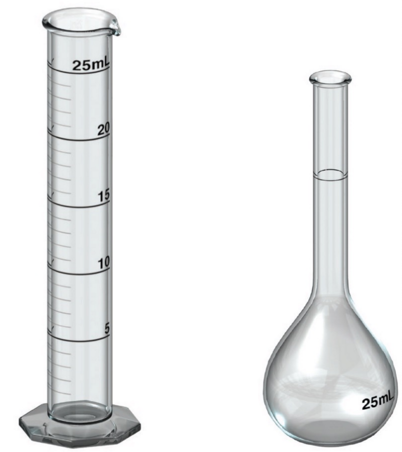


Graduated cylinder
(c)

Volumetric flask
(d)

Base Units of Measurement
International System of Units (SI)

TABLE 1.1 Base SI Units

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

$K = °C + 273.15$

Will be used frequently in CHM 101; you are expected to know them! (Depending on other classes, will likely need to know ampere in the future.)

SI Prefixes

Yes, you need to know these too

TABLE 1.2 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
Tera-	T	1×10^{12} (1,000,000,000,000)	1 teragram (Tg) = 1×10^{12} g
Giga-	G	1×10^{9} (1,000,000,000)	1 gigawatt (GW) = 1×10^{9}
Mega-	M	1×10^{6} (1,000,000)	1 megahertz (MHz) = 1×10^{6}
Kilo-	k	1×10^{3} (1,000)	1 kilometer (km) = 1×10^{3} m
Deci-	d	1×10^{-1} (0.1)	1 deciliter (dL) = 1×10^{-1} L
Centi-	c	1×10^{-2} (0.01)	1 centimeter (cm) = 1×10^{-2} m
Milli-	m	1×10^{-3} (0.001)	1 millimeter (mm) = 1×10^{-3} m
Micro-	μ	1×10^{-6} (0.000001)	1 microliter (μL) = 1×10^{-6} L
Nano-	n	1×10^{-9} (0.000000001)	1 nanosecond (ns) = 1×10^{-9} s
Pico-	p	1×10^{-12} (0.000000000001)	1 picogram (pg) = 1×10^{-12} g

hecto (10^2)
deca (10^1)
Base

The Great Majestic King Henry Died By Drinking Chocolate Milk at Mad Nick's Palace

The Great Majestic King Henry Died By Drinking Chocolate Milk at Mad Nick's Palace

Metric System is Base 10 – essentially just moving the decimal point

tera	giga	Mega	kilo	hecto	deca	base	deci	centi	milli	micro	nano	pico
(10^{12})	(10^{9})	(10^{6})	(10^{3})	(10^{2})	(10)	(1)	(10^{-1})	(10^{-2})	(10^{-3})	(10^{-6})	(10^{-9})	(10^{-12})

1 Tm = 1×10^{12} m

1 Gm = 1×10^{9} m

1 Mm = 1×10^{6} m

1 km = 1,000 m

1 hm = 100 m

1 dam = 10 m

1 m = 1 m

10 dm = 1 m

100 cm = 1 m

1000 mm = 1 m

1×10^{6} μm = 1 m

1×10^{9} nm = 1 m

1×10^{12} pm = 1 m

Metric Conversion Examples

1.) Convert 256.74g to kg

(0.25674 kg)

2.) How many milliliters are in 3.78 L?

(3780 mL)

3.) Convert 18000000 cm into Mm

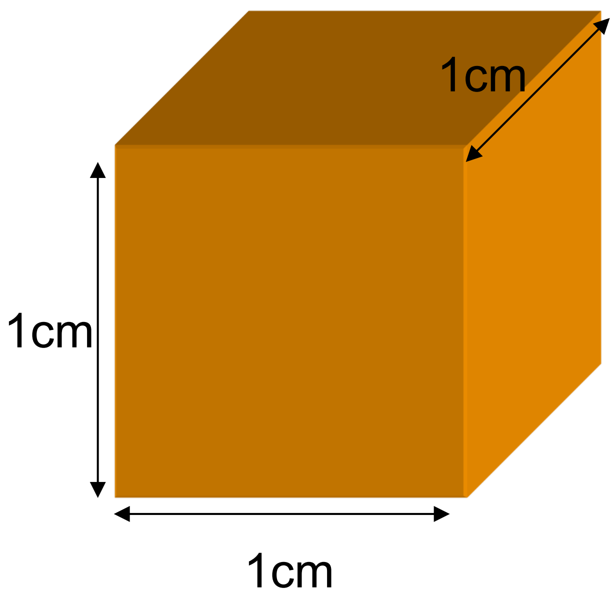
(0.18 Mm)

Derived Units: Volume

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

Common unit is a "**Liter (L)**"

$$1L = 1000cm^3 = \frac{1000cm}{1} x \frac{1cm}{1} x \frac{1cm}{1} x \frac{1m}{100cm} x \frac{1m}{100cm} x \frac{1m}{100cm} = 1x10^{-3} m^3$$



$1\ L \neq 1\ m^3$

$1L = 1x10^{-3}m^3$

$1\ mL = 1\ cm^3$

Metric Conversions with Units that are squared (s^2), cubed (cm^3), etc. can be tricky:

ex.) Convert 87856 cm^3 to m^3

Note: 1 m = 100 cm but $1 m^3 \neq 100\ cm^3$

Need to do the conversion 3x for cubed numbers

(2x for squared, etc.)

87856 cm^3 = 0.087856 m^3

Derived Units: Density

Density: Ratio of mass to volume of a material

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

SI derived unit for density is kg/m^3

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

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

Intensive property

- Can be used to identify a material
- Units of mass and volume may vary

Handling Numbers

Math Review

Significant Figures:

Number of Digits to Report in Final Answer

1. All non-zero digits are significant
2. Use decimal point to decide if zeros are significant

Between 2 numbers	significant	50.002	5 sig figs
Before decimal point	not significant	0.502	3 sig figs
Before the first digit	not significant	0.0052	2 sig figs
End of # after decimal	significant	0.0200	3 sig figs
No decimal point:	not significant	500	1 sig fig

3. **Exact** numbers have unlimited number of sig. figs.

Inherently an integer:	e.g. 4 sides to a square
Inherently a fraction:	e.g. ½ of a pie
Obtained by counting:	e.g. 47 people in a class
Defined quantity:	e.g. 12 eggs in a dozen

Determining the correct number of significant figures (sigfigs) in math problems:

Answer is based on the LEAST significant value

Addition/subtraction – Sig figs based on decimal

$$\begin{array}{r} 1500 \\ +\ 2976 \\ \hline 4476 \end{array} \longrightarrow 4500 \qquad \begin{array}{r} 12.45XX \\ -\ \ 9.2680 \\ \hline 3.1820 \end{array} \longrightarrow 3.18$$

Multiplication/Division – Sig figs based on all sig digits

4 sig figs

3 < 4 so 3 sig figs

$$3.182 \times 3.57 = 11.35974 \longrightarrow 11.4$$

3 sig figs

Rounding is based on number after last sigfig:

≥ 5 round up

≤ 5 round down

Multiple math functions – follow order of ops

(12.45 - 9.2680) x 3.575 = 11.37565

Step one: Subtraction → Sigfigs based on decimal

(12.<u>45</u> - 9.<u>2680</u>) = 3.182

2 sigfigs after decimal

3 sigfigs overall in final answer

$$\begin{array}{r} 12.45\text{XX} \\ -\ \ 9.2680 \\ \hline \underline{\mathbf{3.18}}20 \end{array}$$

Step two: Multiplication → Sigfigs based on all sig digits

<u>**3.18**</u>2 x 3.575 = <u>**11.3**</u>7565

3 sigfigs in 1st number, 4 in 2nd → 3 in final answer

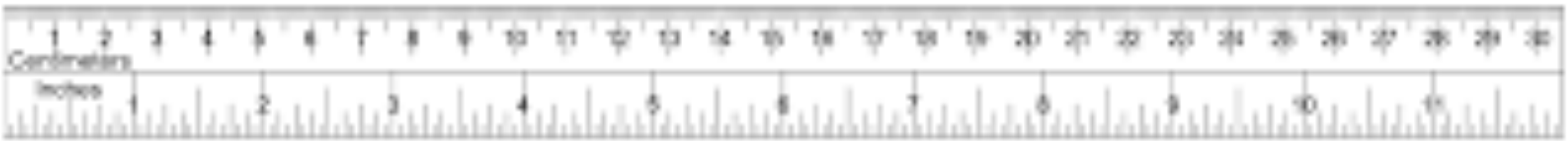
Here addition limits sigfigs

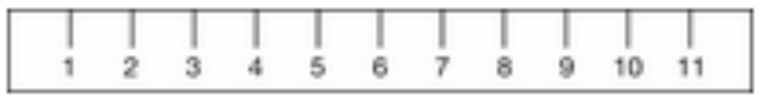
Round up because the next number is >5

<u>11.3</u>7565 → $\boxed{11.4}$

Why do significant figures matter?

123.52 cm





121 cm

Width of room: 244.6 cm
Will the two desks fit?

$$
\begin{array}{r}
123.52 \text{ cm} \\
+\ 121.?? \text{ cm} \\
\hline
244.52 \text{ cm}
\end{array}
\longrightarrow 245 \text{ cm}
$$

What if this is actually 121.1?!?

Fitting desks in a room may not seem all that important – but the same concept is true for the design of buildings & bridges!

Scientific Notation

For very large or very small numbers

Significant digits → **1.7×10^{6}** ← Size of number (multiplier)

$1700000 \rightarrow 1.7 \times 10^{6}$ ← Positive exp = large number (>1)

$0.0000017 \rightarrow 1.7 \times 10^{-6}$ ← Negative exp = small number (<1)

Rules:

- Keep all significant numbers
- Place decimal after 1st significant number (1.7)
- To get exponent:
  - Count number of places decimal moved to get to correct location (after 1st significant number). This value is your exponent.
  - If the number is >1, exp is positive $1700000 \rightarrow 1.7 \times 10^{6}$
  - If the number is <1 exp is negative $0.0000017 \rightarrow 1.7 \times 10^{-6}$

Scientific Notation Examples

Write the Following in Scientific Notation:

1.) 280000000

2.) 280.0

3.) 0.000000004577

4.) 0.00000060

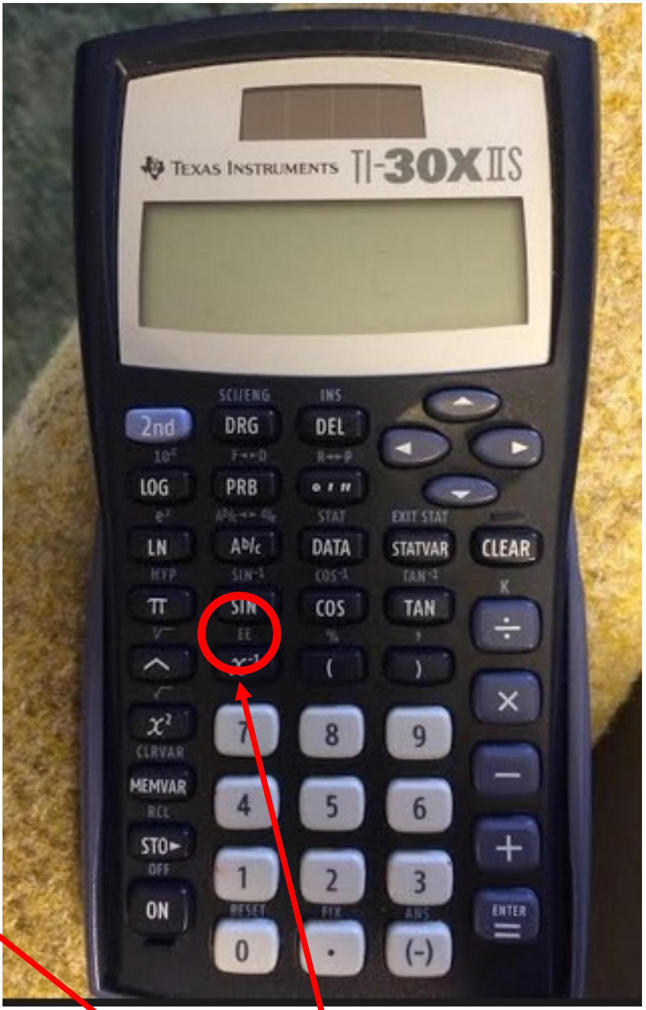
Write the Following in Standard Format:

1.) 2.45×10^{2}

2.) 3.98×10^{6}

3.) 4.29×10^{-3}

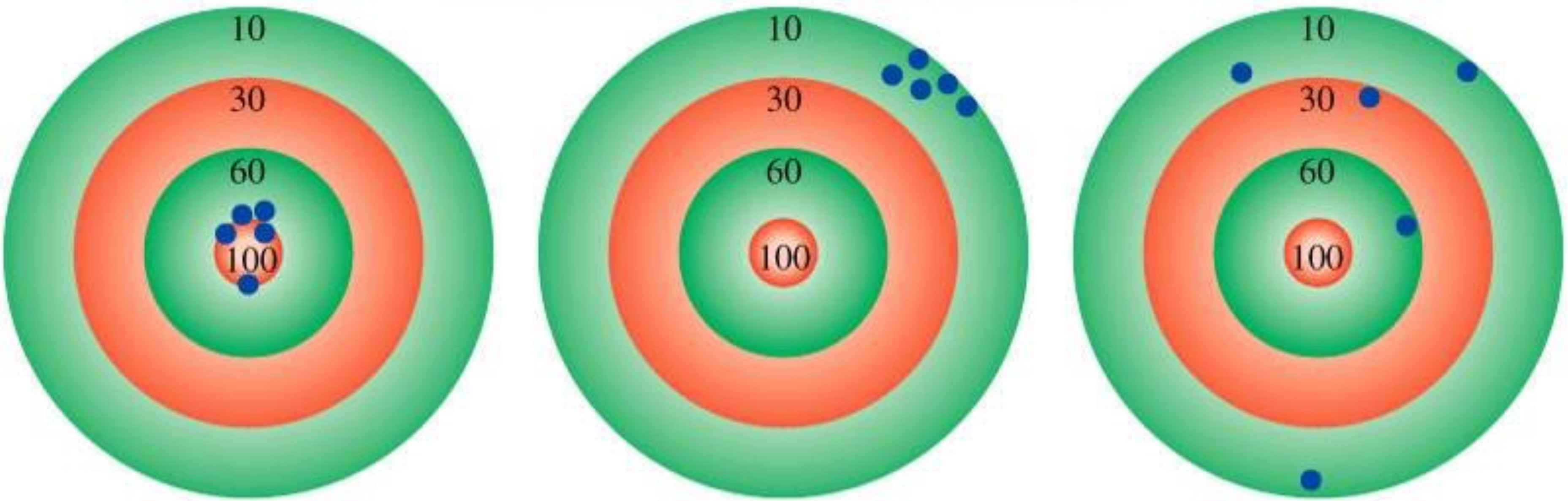
4.) 8.0×10^{-6}



Use EXP, SCI, EE or x10^x keys on calculator

Precision and Accuracy

Accuracy – how close a measurement is to the true value
Precision – how close measurements are to each other



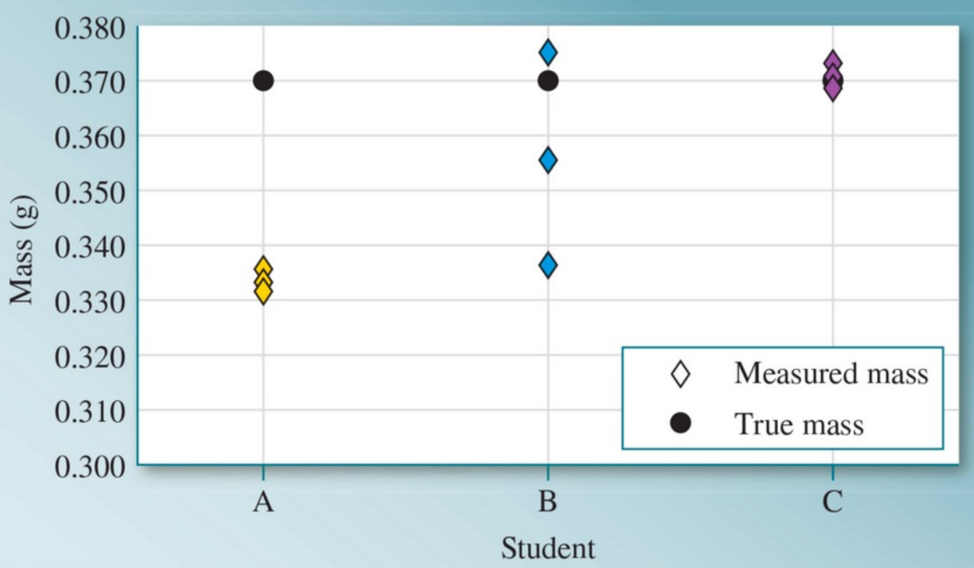
accurate & precise

precise but not accurate

not accurate & not precise

Precision and Accuracy

	Student A	Student B	Student C
Measurement 1	0.335 g	0.357 g	0.369 g
Measurement 2	0.331 g	0.375 g	0.373 g
Measurement 3	0.333 g	0.338 g	0.371 g



Percent Error

Comparison of experimental results to expected or real values

- Usually reported without a + or - sign

$$\% \text{ error} = \left(\frac{|\text{Experimental value} - \text{Real value}|}{\text{Real value}} \right) \times 100$$

Experimental value - Real value = **Deviation**

- Often reported with a + or – sign

Real value:

- Widely accepted, often an industry standard value
- Average of several experiments can sometimes be used if real value is unknown

Dimensional Analysis
Problem Solving & Canceling Units

Look at question:

How many kilograms of methanol will fill a 15.5 gallon fuel tank of a car modified to run on methanol? (Density of methanol = 0.791 g/mL)

What unit do you want to solve for? kilograms (kg)

What information do you need?

Data in problem: Volume = 15.5 gallons
Density of methanol= 0.791 g / mL

Data to look up: Gallon to Liter conversion: 1gal= 3.785L

Data to know: 1000mL = 1L & 1000g = 1kg

$$\frac{kg}{1} = \frac{0.791g}{1ml} x \frac{1kg}{1000g} x \frac{1000mL}{L} x \frac{3.785L}{1gal} x \frac{15.5gal}{1} = 46.4kg$$

Dimensional Analysis Problems

1) How many kilograms of methanol will fill a 15.5 gallon fuel tank of a car modified to run on methanol? (Density of methanol = 0.791 g/mL; 1 gal = 3.785 L) A: 46.4 kg

2) How many liters are equal to 500. cm^3? A: 0.500 L

3) A cube with sides measuring 7.50 m has a mass of 0.04567 mg. What is the density of the cube in μg/mL? A: 1.08×10^{-7} μg/mL

4.) A solution of NaCl has a concentration of 0.579 mol/L. How many moles are present in 250.0mL of solution? A: 0.145mol

Temperature Units: Celsius & Kelvin

- Kelvin is the official SI unit but degrees Celsius are often used.
- 0K is absolute zero – lowest possible temp.
  - Never actually reached – will not have 0K
- Temp in Kelvin = °C + 273.15
- Temp in °C = Kelvin – 273.15
- Fahrenheit – rarely used in science today

Classifications of Matter

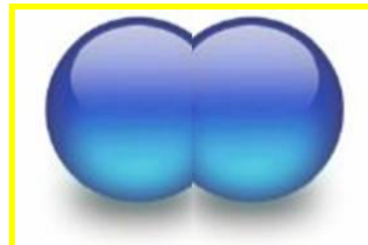
What is in the material you are investigating?

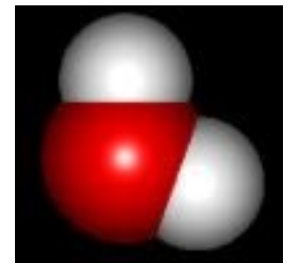
Pure materials

Atom:

- Smallest distinctive unit w/ properties of element
- Ions are charged atoms

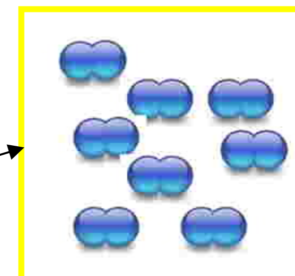
Molecule:



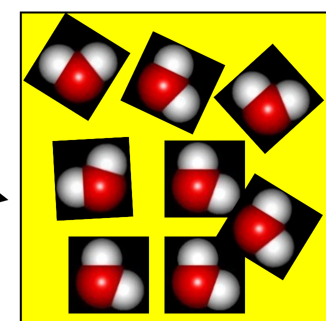


- 2 or more atoms together

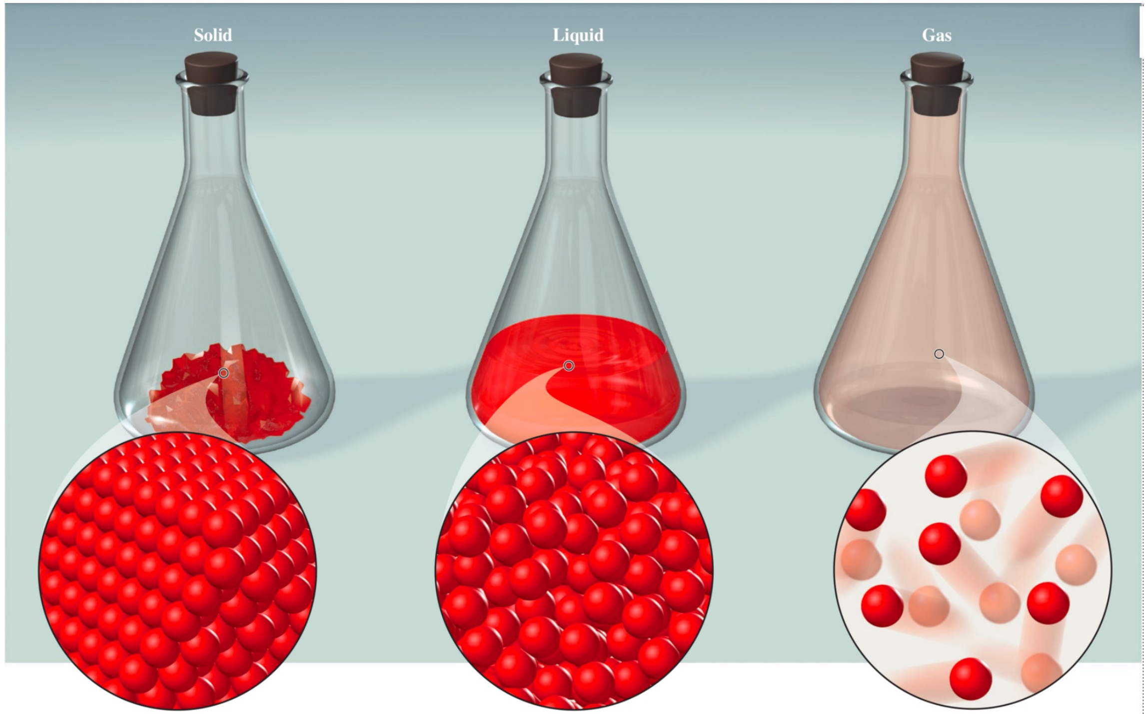
Pure Substance:



- specific composition & distinct properties
- **TWO** types of pure substances:
  - **Element** → one type of atom
  - **Compound**→ more than one type of atom chemically bonded
    - Compounds contain more than one element – still a pure substance!!!



States (Phases) of Matter



Solid:

- Particles close together
- Orderly arrangement
- Little freedom of motion
- Specific shape & volume

Liquid:

- Particles free to move around each other
- Specific volume
- No specific shape

Gas:

- Particles very far apart
- Particles free to move around
- No specific shape or volume

Liquids & gases are fluids – they can "flow"

Mixtures

Mixture: Combination of 2 or more pure substances

- Can be separated by physical means

Homogeneous Mixture

- Substances stay mixed
- No distinct layers
- Uniform properties
- **Also called a "solution"**

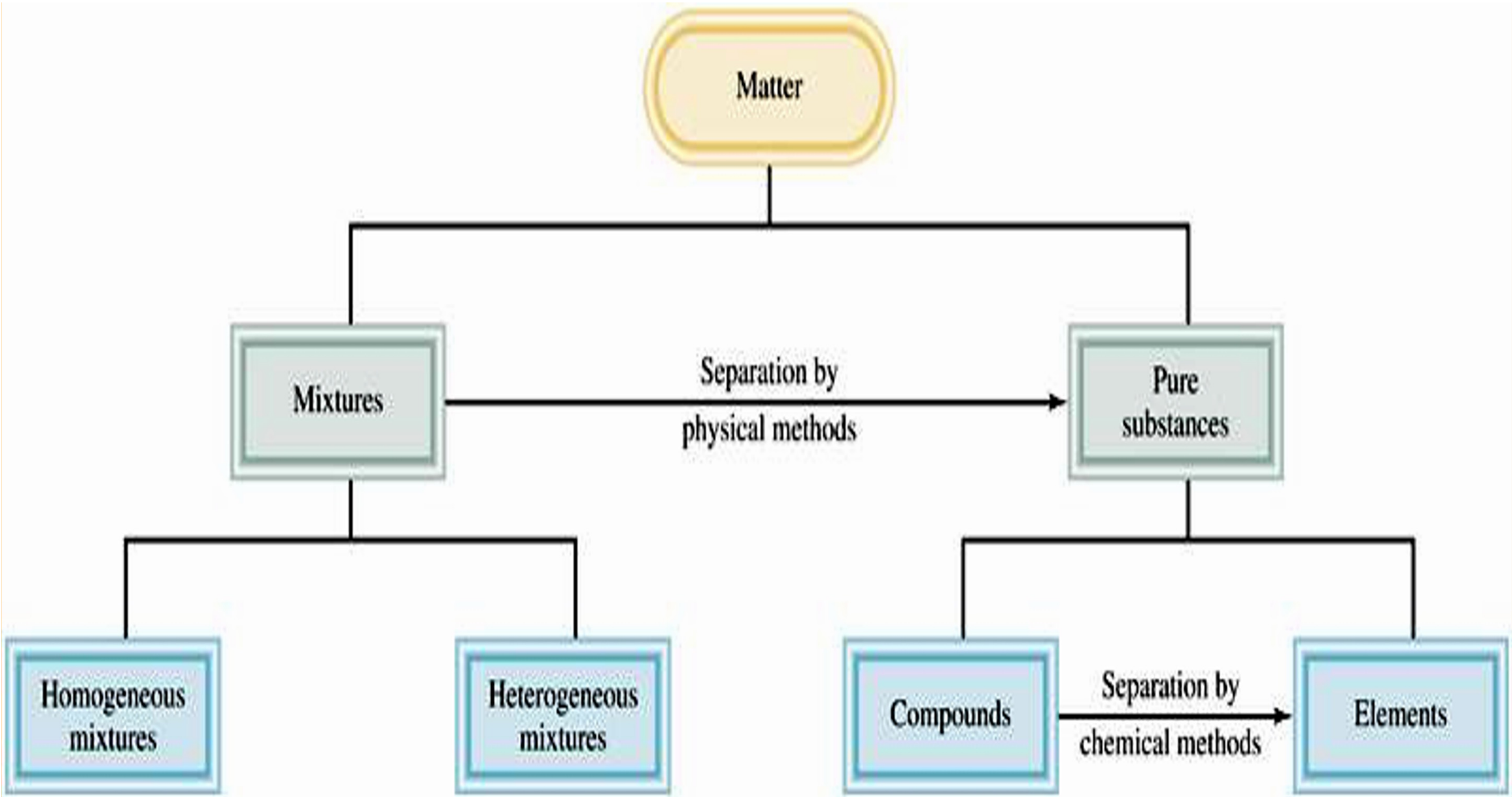
14 karat gold
Mixture of gold and silver

Heterogeneous Mixture

- Substances separate easily
- Distinct layers often seen
- Properties may not be uniform

Iron filings and sand

Matter Summary



	Heterogeneous mixture	Homogeneous mixture	Pure Substance
			
			
			
			
			
			

Physical and Chemical Properties of Matter

Can be used to identify & separate substances

Physical Properties of Matter

Can be changed without changing molecular composition

Chemical identity is NOT CHANGED
eg: smashing a window – still glass
melting ice – still water

Phase changes are physical changes
(solid to liquid to gas etc.)
Melting, freezing, boiling, etc.

CHEMICAL BONDS ARE NOT BROKEN DURING PHASE CHANGES!

Can be used to ID a substance without damage
Color, odor, solubility, conductivity, density
molecular mass, boiling/melting points
Original compound can be recovered

Chemical Properties of Matter

Describe how chemicals react with each other

What will they react with?

How will they react?

- Generate heat or light?
- Burn? Explode?
- Decompose slowly? (Rusting, rotting)

Compositional changes to molecules

- Often called a chemical change
- Original material changed on an atomic level

Original compound no longer present

- Compound cannot be restored to its original form without another chemical change

Extensive and Intensive Properties

Extensive Property: Depends on amount of matter present

ex: mass, length, volume, heat, intensity of color or odor

Intensive Property: Independent of amount of matter present

ex: Temperature, boiling point, color, odor

Often a calculated ratio

ex: Density (mass/vol ratio)

Molar mass (grams/mol)

Specific heat (J/g)

Intensive properties can be used to identify a material, extensive properties cannot. Why?