IB Chemistry Lesson 1.2

Units of Measurement

- Quantity – Something that has magnitude or size.
- Unit –
  - The standard used to measure a quantity.
  - Examples?

- SI –
  - International System of units (metric)

SI units

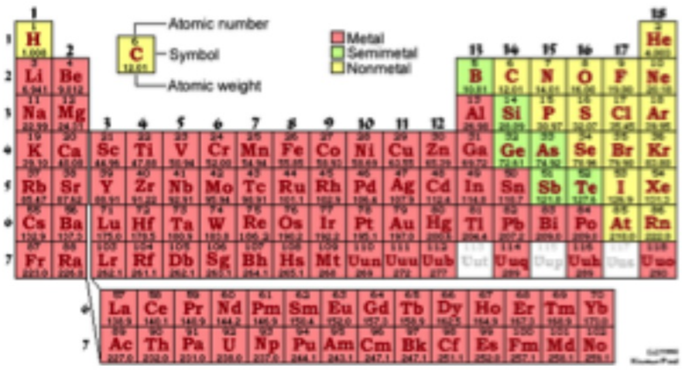
Quantity	Symbol	Unit	Abbreviation
Length			
Mass			
Time			
Temperature			
Amount of Substance			
Electric Current			
Luminous Intensity			
	I Iv T t n l m	mole meter kilogram candela kelvin ampere second	cd m kg K A s mol

Relative Atomic Masses (A_r)

- Relative atomic mass: average masses of isotopes:
  - Naturally occurring C: 98.892 % ^{12}C + 1.108 % ^{13}C.
- Average mass (A_r) of C:

  (0.98892)(12 amu) + (0.0108)(13.00335) = 12.011 amu.

- Relative atomic masses are listed on the periodic table.



Relative Molecular/Formula Mass (M_r)

- The sum of the atomic masses of the atoms that make up a compound (no units).

- Examples:
  1. Determine the relative Molecular Mass of H_2SO_4.

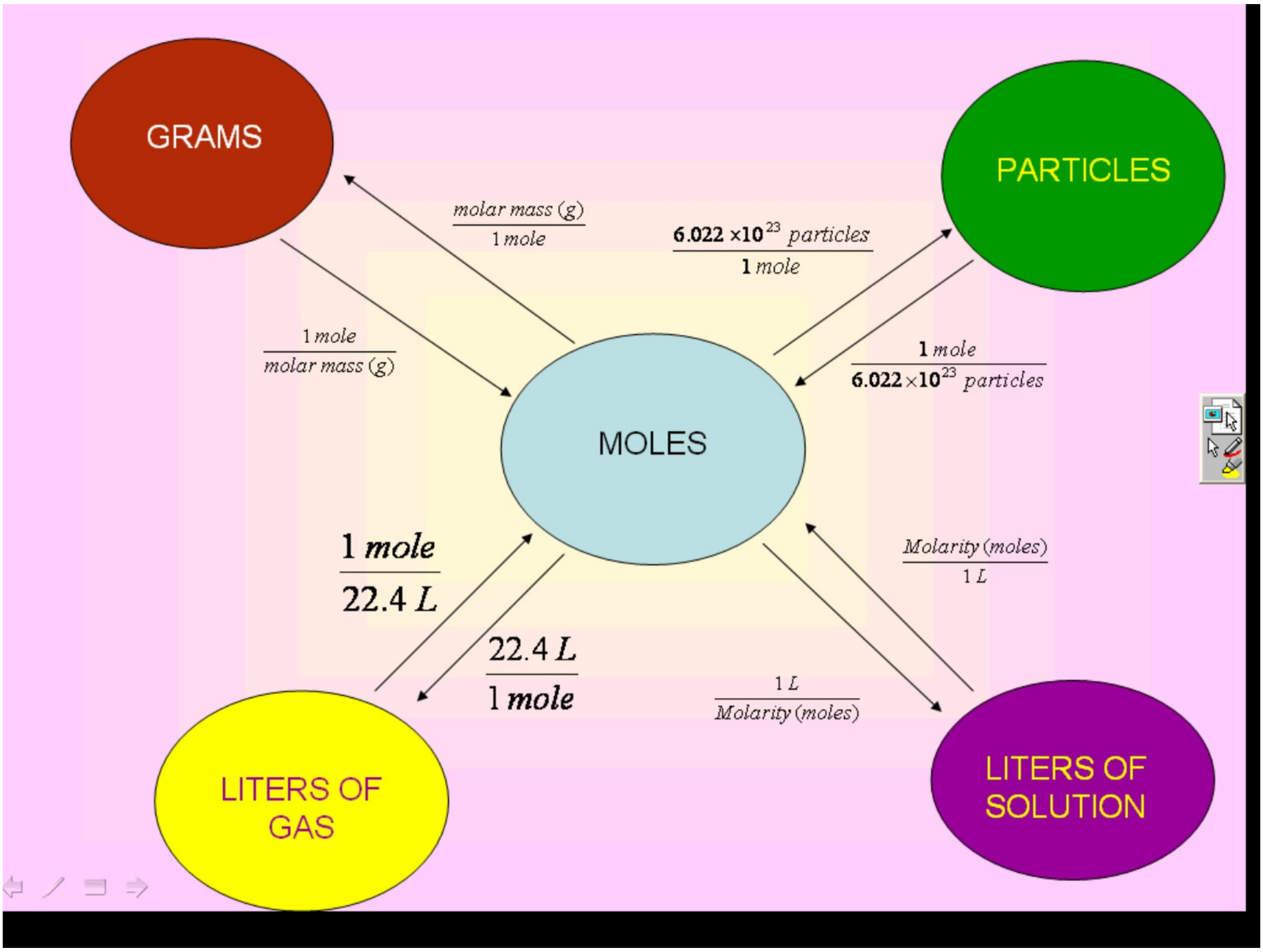
  2. Determine the relative formula mass of $Mg_3(PO_4)_2$.

  3. Determine the relative Molecular Mass of $Ca(NO_3)_2 \cdot 4H_2O$

Molar Mass

-The mass of one mole of a substance (element or compound).

-Numerically equal to A_r or M_r, but units are g mol^{-1}


GRAMS
PARTICLES
MOLES
LITERS OF GAS
LITERS OF SOLUTION
$\frac{molar\ mass\ (g)}{1\ mole}$
$\frac{1\ mole}{molar\ mass\ (g)}$
$\frac{6.022 \times 10^{23}\ particles}{1\ mole}$
$\frac{1\ mole}{6.022 \times 10^{23}\ particles}$
$\frac{1\ mole}{22.4\ L}$
$\frac{22.4\ L}{1\ mole}$
$\frac{Molarity\ (moles)}{1\ L}$
$\frac{1\ L}{Molarity\ (moles)}$

Examples - Make the following conversions:

1. Convert 21.98 g of CO to moles.
2. Convert 3.60 x 10^{24} molecules of N_2 to moles.
3. Convert 38.4 g of $Mg(BrO_3)_2$ to molecules.
4. Convert 5.03 x 10^{21} molecules of H_2SO_4 to grams.
5. Convert 33.9 grams of CO_2 to liters (at STP).

Percentage Composition from Formulas

•Percent composition is the atomic weight for each element divided by the formula weight of the compound multiplied by 100:

$$\%\,\text{Element} = \frac{(\text{Atoms of Element})(\text{molar mass of element})}{\text{molar mass of Compound}} \times 100$$

Examples: Calculate Percentage Composition:

1. CO_2
2. $Mg(ClO_2)_2$
3. $(NH_4)_3PO_4$

Empirical Formula: The simplest whole number ratio of atoms of each element in a compound.

Example: The empirical formula of glucose ($C_6H_{12}O_6$) is CH_2O.

What about...

1. H_2O_2

2. C_4H_{10}

3. $C_3H_8O_2$

To determine the empirical formula of a compound from experimental data:

- Assume any percentages are masses, in grams.
- Convert all masses to moles (divide by molar mass)
- Divide by the smallest number of moles – if all are whole numbers (within 0.1), the numbers are the subscripts.
- If not, make whole numbers by multiplying all moles by given integer (.5 x 2, .33 or .67 by 3, .25 or .75 by 4, etc.). Now, the numbers are the subscripts.

* Hydrates are compounds with water molecules incorporated into their structure. Their formulas are determined using % (compound) and % (water) to determine the ratio of water molecules per formula unit of the compound (example: $CuSO_4 \cdot 5H_2O$)

Examples: Determine the Empirical formulas:
1. 63.6 % N, 36.4 % O

2. 54.5 % C, 9.20 % H, 36.3 % O

3. 58.9 % $NiSO_4$, 41.1 % H_2O

To Determine Molecular Formula:

- First determine empirical formula
- Divide:

$$\frac{molar\ mass\ (given)}{empirical\ \ formula\ mass}$$

- You will get a whole number. Multiply all subscripts in the empirical formula by the resulting integer.

Examples: Determine Molecular Formula

1. A compound is 30.4 % N and 69.6 % O, with a molar mass of 92.14.

2. A compound contains 55.8 % C, 7.02 % H, and 37.2 % O, with a molar mass of 129.14.

STOICHIOMETRY

* Coefficients can be used to convert moles of one substance to moles of another substance.

* All other conversions are still used.

EXAMPLES

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

what mass of oxygen would be required to react completely with 275 g of propane?

What mass of carbon dioxide would be produced?

What mass of water would be produced?

EXAMPLES: $\_\_C_4H_{10} + \_\_O_2 \rightarrow \_\_CO_2 + \_\_H_2O$

1. What mass of water is produced from 4.39 dm^3 of oxygen gas at STP?

2. How many molecules of butane must react to produce 3.88 dm^3 of carbon dioxide?

LIMITING REACTANT (1.4.2)

- A chemical reaction proceeds until one (or more) of the reactants runs out.
- The reactant that runs out is called the limiting reactant or limiting reagent.
- To determine the limiting reagent, you can:
  - Use the amount of one reactant to determine how much of other reactant(s) is(are) needed, and compare that to how much you have, or
  - Use the amounts of each reactant to determine how much product will be produced. The reactant that produces the least product is the limiting reagent.

$Al\ (s) + Fe_2O_3\ (s) \longrightarrow Fe(s) + Al_2O_3\ (s)$

- If 35.9 g of aluminum and 63.7 g of iron (III) oxide are mixed, what is the limiting reagent?

- What mass of iron will be produced by this reaction?

YIELDS

- Theoretical Yield: The amount of product (usually in grams) calculated using stoichiometry. (1.4.1)
- Actual Yield: The amount of product actually produced by a reaction (measured in the laboratory)
- Percent Yield: The amount of product produced expressed as a percentage of the amount of product expected. (1.4.3)

$$\%\ Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100\%$$

$$Al\ (s) + Fe_2O_3\ (s) \longrightarrow Fe\ (s) + Al_2O_3\ (s)$$

- If 35.9 g of aluminum and 63.7 g of iron (III) oxide are mixed, what is the theoretical yield of iron metal?

- If the actual yield of iron metal is 18.7 g, what is the percent yield?