SLO 2: Stoichiometry

2.1 Avogadro's Number and Mole

2.1.1 Define The Terms:

Gram Atomic Mass (GAM)
The atomic mass of an element expressed in grams is called gram atomic mass.

Gram Molecular Mass (GMM)
The molecular mass of an element or a compound expressed in grams is called gram molecular mass.

Gram Formula Mass (GFM)
The formula mass of an ionic compound expressed in grams is called gram formula mass.

Formula Unit
The simplest ratio of ions in an ionic compound. Ionic compounds don't form molecules; they form crystal lattices. NaCl (1 Na^+ : 1 Cl^-)

Mole
The atomic mass, molecular mass, and formula mass of a substance expressed in grams is known as mole. A mole is defined as "amount of substance containing 6.02 x 10^{23} particles.

$$Number\ of\ Moles\ (n) = \frac{Given\ Mass\ (g)}{Molar\ Mass\ (g/mol)}$$

Avogadro's Number
6.022×10^{23} is the number of atoms, molecules, or ions in 1 mole of a substance. Symbol N_A

1 mole = Avogadro's number of particles

2.1.2 Relate Gram Atomic Mass, Gram Molecular Mass And Gram Formula Mass To Mole And Avogadro's Number

- 1 mole of any substance contains 6.022×10^{23} particles.
- 1 mole of atoms = gram atomic mass
- 1 mole of molecules = gram molecular mass
- 1 mole of ionic compounds = gram formula mass

2.1.3 Calculate The Number Of Moles, Atoms And Molecules Of Substances

Number of Moles

$$Moles\ (n) = \frac{Mass\ (g)}{GAM\ , GMM\ , GFM\ (g/mol)}$$

Number of Particles (Atoms, Molecules, Ions)

Number of Particles = [Moles] × [6.022×10^{23}]

Molar Mass

$$Molar\ mass = \frac{Mass\ (g)}{Number\ of\ moles}$$

2.2 Formulae and percentage composition

2.2.1 Differentiate Between Empirical And Molecular Formula

Empirical Formula	Molecular Formula
The formula showing minimum relative numbers of each type of atoms in a molecule is called Empirical Formula.	The Molecular formula is the formula which shows actual number of atoms of each element present in a molecule.
Empirical Formula shows simplest ratio of each atom present in a molecule.	Molecular formula is derived from empirical formula.
Empirical Formula does not show the actual number of atoms in the molecule.	Molecular formula Mass calculated by adding atomic masses of its atoms.
Empirical Formula tells us the type of element present in it.	Molecular formula of a compound may be same or multiple of empirical formula.
Empirical formula of Benzene is CH.	Benzene has molecular formula C6H6.
Empirical formula of glucose is CH2O	Glucose has molecular formula C6H2O6.

2.2.2 Calculate Percentage Composition By Mass Of An Element In A Compound

Percentage composition by mass tells you what portion of a compound's mass comes from each element.

$$Percentage\ of\ an\ element = \left(\frac{Total\ mass\ of\ the\ element\ in\ 1\ mole\ of\ compound}{Molar\ mass\ of\ the\ compound}\right) \times 100$$

2.2.3 Calculate Empirical Formula Using The Percentages Of Elements

1. Write down the percentage of each element (assume 100 g of the compound).
2. Convert percentages to mass (in grams).
3. Convert mass to moles.

$$Moles = \frac{mass\ (g)}{atomic\ mass\ (Ar)}$$

4. Divide all moles by the smallest mole value.
5. Multiply all ratios to get whole numbers (if needed).
6. Write the empirical formula.

Example:

A compound contains 40% carbon, 6.7% hydrogen, and 53.3% oxygen. Find the empirical formula.

Element	% Composition	Ar	Moles
C	40.0	12	40 ÷ 12 = 3.33
H	6.7	1	6.7 ÷ 1 = 6.7
O	53.3	16	53.3 ÷ 16 = 3.33

Divide by smallest (3.33)

C = 3.33 ÷ 3.33 = 1 H = 6.7 ÷ 3.33 ≈ 2 O = 3.33 ÷ 3.33 = 1

Empirical formula = CH_2O

2.2.4 molecular formula using molecular mass and empirical formula

1. Calculate empirical formula mass (EFM).
2. Use the given molecular mass (MM).

$$n = \frac{molecular\ mass}{Empirical\ formula\ Mass}$$

3. Multiply the subscripts in the empirical formula by n.

Example

- Empirical formula = CH_2O
- Empirical formula mass = 12 + 2 + 16 = 30 g/mol
- Given molecular mass = 180 g/mol

$$\text{n} = \frac{180}{30} = 6$$

Molecular formula = $(CH_2O)_6$ = $C_6H_{12}O_6$

2.3 Chemical Reactions and Calculations

2.3.1 Define The Terms ‘Chemical Reaction’ And ‘Chemical Equation’

A chemical reaction is a process in which one or more substances (called reactants) are converted into new substances (called products) with different chemical properties.

A chemical equation is a symbolic representation of a chemical reaction using chemical formulas and symbols to show the reactants, products, and their quantities.

2.3.2 Illustrate The Following Types Of Chemical Reactions:

TYPE	EXPLANATION	GENERAL REACTION
Single displacement reaction	One element takes place with another element in the compound.	A + BC → AC + B
Double displacement reaction	Elements take place with another element in the compounds.	AB + CD → AD + CB
Decomposition reaction	The opposite of a combination reaction – a complex molecule breaks down to make simpler ones.	AB → A + B
Combination reaction	Two or more compounds combine to form one compound.	A + B → AB
Combustion reaction	Oxygen combines with a compound to form carbon dioxide and water. These reactions are exothermic, meaning they give off heat.	A + O_2 → H_2O + CO_2
Neutralization reaction	An acid and a base react with each other. Generally, the product of this reaction is salt and water.	Acid + Base → Salt + Water
Hydrolysis	A compound reacts with water and breaks down into two or more products.	AB + H_2O → AH + BOH

2.3.3 Evaluate The Effectiveness Of Chemical Reactions Based On

A. Change Of State

- What it shows: A new phase (solid, liquid, gas) is formed.
- Effectiveness: Strong indicator that a new substance has formed.
- Example: Ice melting (physical) vs. gas bubbling in vinegar + baking soda (chemical).

B. Change In Color

- What it shows: Formation of new compounds with different optical properties.
- Effectiveness: Clear evidence, especially when the change is sudden and unexpected.
- Example: Iron rusting (silver to reddish-brown).

C. Evolution Of Gas

- What it shows: A gas is released during the reaction.
- Effectiveness: Highly reliable for identifying a chemical change.
- Example: Hydrochloric acid + zinc → hydrogen gas (seen as bubbling).

D. Change In Temperature

- What it shows: Reaction absorbs or releases energy as heat.
- Effectiveness: Confirms energy transfer; common in both endothermic and exothermic reactions.
- Example: Neutralization of acid and base releases heat (exothermic).

E. Formation Of A Precipitate

- What it shows: A solid forms from a solution, indicating insoluble product.
- Effectiveness: Strong visual clue for reactions in aqueous solutions.
- Example: Mixing lead(II) nitrate and potassium iodide → yellow precipitate.

F. Occurrence Of Sound

- What it shows: Often associated with rapid reactions, especially gas formation or combustion.
- Effectiveness: Less common, but confirms violent or fast chemical change.
- Example: Fireworks, explosive gas reactions.

2.3.4 Construct Balanced Chemical Equations For Chemical Reactions

2.3.5 Balance The Chemical Equations By Inspection Or Trial And Error Method

Balance this equation: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

Step 1: List the elements that constitute the chemical equation

ELEMENTS	REACTANT		PRODUCT	
	C_3H_8	O_2	CO_2	H_2O
Carbon [C]	3	0	1	0
Hydrogen [H]	8	0	0	2
Oxygen [O]	0	2	2	1

Step 2: Equalize the first element by multiplying the reactant/product with required coefficient

ELEMENTS	REACTANT		PRODUCT	
	C_3H_8	O_2	$3CO_2$	H_2O
Carbon [C]	3	0	1 x 3	0
Hydrogen [H]	8	0	0	2
Oxygen [O]	0	2	2 x 3	1

Step 3: Equalize second element by multiplying the reactant/product with required coefficient

ELEMENTS	REACTANT		PRODUCT	
	C_3H_8	O_2	$3CO_2$	$4H_2O$
Carbon [C]	3	0	1 x 3	0
Hydrogen [H]	8	0	0	2 x 4
Oxygen [O]	0	2	2 x 3	1 x 4

Step 4: Equalize third element by multiplying the reactant/product with required coefficient

ELEMENTS	REACTANT		PRODUCT	
	C_3H_8	$5O_2$	$3CO_2$	$4H_2O$
Carbon [C]	3	0	1 x 3 = 3	0
Hydrogen [H]	8	0	0	2 x 4 = 8
Oxygen [O]	0	2 x 5 = 10	2 x 3 = 6	1 x 4 = 4

The equation is now balanced: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

2.3.6 Solve Problems Based On Stoichiometric Relationships Of Substances

To calculate mass, number of moles, number of molecules, mole ratio, volume of gases at Room Temperature and Pressure (RTP) (24 L or dm^3 and 24000 mL or cm^3) based on a balanced chemical equation of combustion, let's use the combustion of methane (CH_4) as our balanced chemical equation:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Molar masses:

CH_4 = 16 g/mol O_2 = 32 g/mol CO_2 = 44 g/mol H_2O = 18 g/mol

Mass:
To calculate the mass of each substance, use:

Mass = Moles × Molar Mass

- Mass of CH_4 = 1 mol × 16 g/mol = 16 g
- Mass of O_2 = 2 mol × 32 g/mol = 64 g
- Mass of CO_2 = 1 mol × 44 g/mol = 44 g
- Mass of H_2O = 2 mol × 18 g/mol = 36 g

Number of Moles:
From the balanced equation:

1 mol CH_4 reacts with 2 mol O_2 and Produces 1 mol CO_2 and 2 mol H_2O

This gives us the mole values directly from the coefficients.

Number of Molecules:
To find the number of molecules, use Avogadro's number:

1 mole = 6.022×10^{23} molecules

- Molecules of CH_4 / CO_2 = 1 mol × 6.022 × 10^{23} = 6.022 × 10^{23}
- Molecules of O_2 / H_2O = 2 mol × 6.022 × 10^{23} = 1.204 × 10^{24}

Mole Ratio

The mole ratio comes from the balanced equation:

1:2:1:2

This means that 1 mole of methane reacts with 2 moles of oxygen Produces 1 mole of carbon dioxide and 2 moles of water.

Volume of gases at Room Temperature and Pressure (RTP):

$$\text{Volume (dm}^3\text{)} = \text{Moles of gas} \times 24$$

$$\text{Volume (cm}^3\text{)} = \text{Moles of gas} \times 24{,}000$$

Example: What volume of oxygen gas is needed to completely burn 4 moles of methane (CH_4) at RTP?

1 mole CH_4 reacts with 2 moles O_2. So, 4 moles CH_4 will need:

$$4 \times 2 = 8 \text{ moles of } O_2$$

$$\text{Volume} = 8 \times 24 = 192 \text{ dm}^3$$

Example: How many cm^3 of hydrogen are produced when 5g of zinc reacts with excess HCL acid at RTP?

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2Zn + 2HCl$$

$$Ar(Zn) = 65$$

$$\text{Moles} = 5/65 \approx 0.077 \text{ mol}$$

From the equation: 1 mole Zn → 1 mole H_2, so:

$$0.077 \text{ mol Zn} \rightarrow 0.077 \text{ mol } H_2$$

$$\text{Volume} = 0.077 \times 24{,}000 = 1{,}848 \text{ cm}^3$$