UNIT 15

STOICHIOMETRY

Class 10 — Chemistry Notes

Complete Chapter Notes — Printable Format

Contents

MOLAR VOLUME

Q.1 Define molar volume. Why is it necessary to measure volume of a gas at standard conditions? What is its importance?

Molar Volume

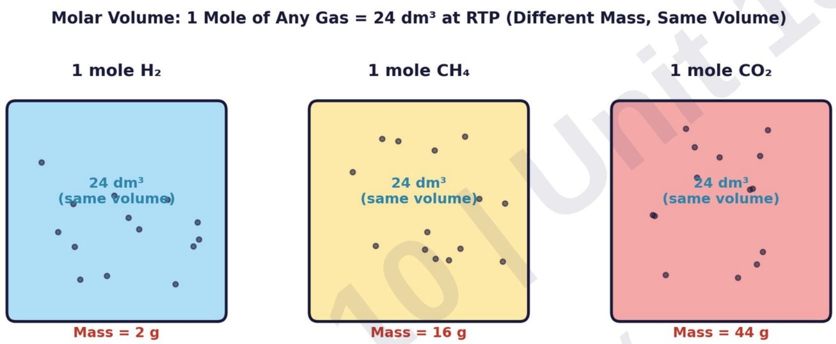
It has been found that one mole of all gases occupies a volume of 24 dm³ at room temperature and pressure (RTP). This is called molar volume.

Examples:

24 dm³ of any gas at RTP = 1 mole

24 dm³ of H_2 gas at RTP = 1 mole

24 dm³ of CH_4 gas at RTP = 1 mole



Molar Volume: Different Gases, Same Volume, Different Mass

Formula

Number of moles of gas = Volume of the gas at RTP / Molar volume of the gas

Measurement of Volume at Room Temperature and Pressure (RTP)

The volume of a gas varies with changes in pressure and temperature. In order to compare the volumes occupied by different gases, we must adopt a set of standard conditions of temperature and pressure called room temperature and pressure (RTP).

Standard Conditions

- Temperature: 25°C (298.15 K)
- Pressure: 1 atmosphere (760 torr)

These are used as standard conditions for measurement of volumes of gases at room temperature.

Explanation

It is worth noting that:

- One mole of each gas contains the same number of molecules (Avogadro's number = 6.02×10^{23})
- One mole of each gas occupies the same volume (24 dm³ at RTP)
- One mole of each gas has a different mass

Important Point: The masses and the sizes of the molecules of gases do not change the volume occupied by them.

Importance of Molar Volume

The concept of molar volume is useful because it enables us to calculate:

- The number of moles of gases
- The masses of the gases

...if we know their volumes.

SOLVED EXAMPLES

Example 15.1

Calculate the number of moles of nitrogen gas present when it occupies a volume of 2.5 dm^3 at RTP.

Solution:

24 dm^3 is occupied by how many moles of nitrogen = 1 mole

1 dm^3 will be occupied by how many moles of nitrogen = 1/24 mole

2.5 dm^3 will be occupied by how many moles of nitrogen = $(1/24) \times 2.5 = 0.10$ mole

Answer: 0.10 mole

Example 15.2

Calculate the volume of 0.5 mol of hydrogen at RTP.

Solution:

Number of moles of gas = Volume of the gas at RTP / Molar volume of the gas

0.5 = Volume of hydrogen / 24

Volume of hydrogen = $0.5 \times 24 = 12$ dm^3

Answer: 12 dm^3

Example 15.3

Calculate the volume occupied by 200 g of carbon dioxide at RTP.

Solution:

Number of moles = Volume / Molar Volume

Volume = Number of moles × Molar volume

Since Number of moles = Mass / Molar Mass

Volume = (Mass / Molar Mass) × Molar volume

Volume = (200 g / 44 g/mol) × 24 dm^3 = 4.545 × 24 dm^3

Volume = 109 dm^3

Answer: 109 dm^3

Quick Check — Calculate the amount of CO_2 in 100 dm^3 at RTP
Given data: Volume of CO_2 at RTP = 100 dm^3; Molar volume at RTP = 24 dm^3; Molar mass of CO_2 = 12 + 16(2) = 44 g/mol.
Required: Amount (mass) of CO_2 = ?
Solution: Number of moles of CO_2 = Volume / Molar volume = 100 / 24 = 4.166 moles
Amount (mass) of CO_2 = Number of moles × Molar mass = 4.166 × 44 = 183.33 g
Answer: 183.33 g

IMPORTANT FORMULAS — MOLAR VOLUME

1. Number of Moles from Volume

Number of moles = Volume of gas at RTP / 24 dm^3

2. Volume from Number of Moles

Volume = Number of moles × 24 dm^3

3. Mass from Volume

Mass = (Volume / 24) × Molar mass

4. Volume from Mass

Volume = (Mass / Molar mass) × 24 dm^3

SUMMARY TABLE: MOLAR VOLUME RELATIONSHIPS

Given Quantity	Formula to Find	Calculation
Volume (dm^3)	Number of moles	n = V/24
Number of moles	Volume (dm^3)	V = n × 24
Mass (g)	Volume (dm^3)	V = (m/M) × 24
Volume (dm^3)	Mass (g)	m = (V/24) × M

Where: n = number of moles; V = volume in dm^3; m = mass in grams; M = molar mass in g/mol

KEY CONCEPTS — MOLAR VOLUME

- Molar Volume: Volume occupied by 1 mole of any gas at RTP = 24 dm^3
- **RTP Conditions:** Temperature = 25°C (298.15 K); Pressure = 1 atm (760 torr)
- Avogadro's Law Connection: One mole of any gas contains the same number of molecules (6.02×10^{23}), therefore occupies the same volume at same temperature and pressure.
- Independent of Gas Type: The volume of 1 mole of gas at RTP is 24 dm^3 regardless of the type of gas (H_2, CH_4, CO_2, N_2, etc.)
- Mass Varies: While volume is same for all gases (24 dm^3), the mass of 1 mole varies depending on the molar mass of the gas.

QUICK REFERENCE: MOLAR MASSES

Gas	Formula	Molar Mass (g/mol)
Hydrogen	H_2	2
Methane	CH_4	16
Nitrogen	N_2	28
Oxygen	O_2	32
Carbon Dioxide	CO_2	44
Ammonia	NH_3	17
Chlorine	Cl_2	71

PRACTICE PROBLEMS

Problem 1: Calculate the number of moles of oxygen gas present in 48 dm^3 at RTP.

n = V/24 = 48/24 = 2 moles

Problem 2: Calculate the volume occupied by 3 moles of ammonia gas at RTP.

V = n × 24 = 3 × 24 = 72 dm^3

Problem 3: Calculate the mass of 60 dm^3 of methane (CH_4) gas at RTP.

n = V/24 = 60/24 = 2.5 moles; Mass = n × M = 2.5 × 16 = 40 g

Problem 4: Calculate the volume occupied by 88 g of CO_2 at RTP.

$$V = (m/M) \times 24 = (88/44) \times 24 = 2 \times 24 = 48 \text{ dm}^3$$

KEY POINTS TO REMEMBER

- 1 mole of any gas at RTP = 24 dm^3
- RTP = 25°C and 1 atm pressure
- Molar volume is independent of the type of gas
- Different gases have different masses but same volume at RTP
- Molar volume concept is derived from Avogadro's law
- Useful for converting between volume, moles, and mass of gases

DEFINITIONS

$$\text{1 mole of any gas at RTP} = 24 \text{ dm}^3$$

SECTION 1: CONCENTRATION OF SOLUTION

Q.2 What do you mean by concentration of a solution? How can you calculate concentration of a solution?

Concentration of a Solution

- It is a measure of the amount of solute that has been dissolved in a given amount of solution or solvent.
- **Amount of solute:** Expressed either in grams or in moles.
- **Quantity of solution:** Expressed in terms of its volume, dm^3 or cm^3.

Unit:

- Mass/volume concentration = g/dm^3
- Molar concentration = mol/dm^3

Calculation of Concentration

$$\textbf{Concentration of a solution} = \textbf{Mass of solute (g) / Volume of solution (dm}^3\textbf{)}$$

Example 15.4

Calculate concentration of a solution in g/dm^3 which contains 20 g of sodium chloride dissolved in 0.2 dm^3 of water.

Solution:

$$\text{Concentration} = 20 \text{ g} / 0.2 \text{ dm}^3 = 100 \text{ g/dm}^3$$

Example 15.5

500 cm^3 of hydrochloric acid contains 1.8 g of dissolved hydrogen chloride. Calculate the concentration of the acid in g/dm^3.

Solution:

$$\text{Volume in dm}^3 = 500 \text{ cm}^3 / 1000 = 0.5 \text{ dm}^3$$

$$\text{Concentration} = 1.8 / 0.5 = 3.6 \text{ g/dm}^3$$

Calculate the mass of copper sulphate in grams dissolved in 2 dm^3 of the solution to get a solution with a concentration 15 g/dm^3.

Solution:

$$\text{Mass of solute} = \text{Concentration} \times \text{Volume} = 15 \text{ g/dm}^3 \times 2 \text{ dm}^3 = 30 \text{ g}$$

Example 15.7

A solution is prepared by dissolving 10 g of NaOH in 1.2 dm^3 of water. Calculate the concentration of the solution in mol/dm^3.

Solution:

$$\text{Amount of NaOH} = 10\text{ g};\;\; \text{Number of moles of NaOH} = 10/40 = 0.25\text{ mol}$$

$$\text{Concentration} = 0.25/1.2 = 0.21\text{ mol/dm}^3$$

Quick Check 15.2 — What mass of $CuSO_4$ is present in 50 cm^3 of a 5×10^{-2} mol/dm^3 aqueous solution?
Given data: Volume of solution = 50 cm3; Concentration of solution = 5×10^{-2} mol/dm3; Molar mass of CuSO4 = 63.5 + 32 + 16×4 = 159.5 g/mol
Solution: Volume in dm3 = 50/1000 = 0.05 dm3
Number of moles = Concentration × Volume = $5\times10^{-2} \times 0.05 = 0.0025$
Mass of CuSO4 = 0.0025 × 159.5 = 0.3987 g
Answer: 0.3987 g

SECTION 2: TITRATION USING EMPIRICAL DATA

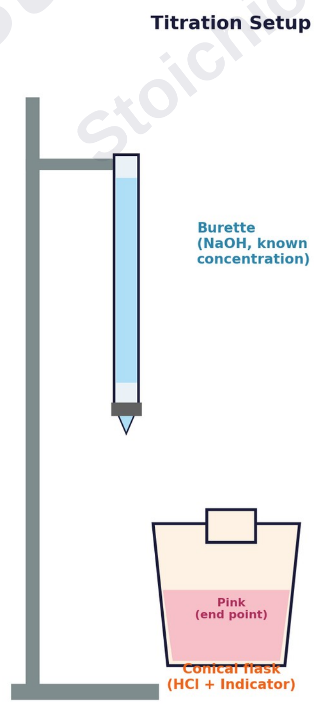
Q.3 How concentration of a solution in titration using empirical formula is calculated?

Calculation of Concentration in Titration

- The concentration of a solution can be determined by titrating it with a solution of known concentration till the end point is reached.

Requirements:

  - Balanced chemical equation for the reaction
  - Volumes of the two solutions used
  - Concentration of the other solution



Titration Setup — Burette and Conical Flask

Example 15.8

A titration is performed between NaOH and HCl till the end point is reached. 15 cm^3 of solution of sodium hydroxide having a molar concentration 0.5 mol/dm^3 is used to neutralize 23 cm^3 of HCl. Find out the concentration of HCl in mol/dm^3.

Solution:

1. Balanced chemical equation:

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

2. Convert volumes to dm^3:

$$\text{Volume of NaOH} = 15/1000 = 0.015\ dm^3; \quad \text{Volume of HCl} = 23/1000 = 0.023\ dm^3$$

3. Number of moles of NaOH:

$$\text{Moles of NaOH} = \text{Concentration} \times \text{Volume} = 0.5 \times 0.015 = 0.0075$$

4. From balanced equation:

$$\text{1 mol NaOH reacts with 1 mol HCl} \rightarrow \text{Moles of HCl} = 0.0075$$

5. Concentration of HCl:

$$\text{Concentration} = \text{Moles/Volume} = 0.0075/0.023 = 0.326\ mol/dm^3$$

Answer: 0.326 mol/dm^3

Quick Check 15.3 — Calculate the volume of 0.1 mol/dm^3 oxalic acid ($C_2O_4H_2$) solution to exactly neutralize 25 cm^3 of 0.1 mol/dm^3 KOH solution.
Given data: Concentration of oxalic acid = 0.1 mol/dm3; Concentration of KOH = 0.1 mol/dm3; Volume of KOH = 25 cm3
Solution: (COOH)2 + 2KOH → (COOK)2 + 2H2O
Mole ratio: KOH : Oxalic acid = 2 : 1
2 cm3 of KOH neutralizes 1 cm3 of oxalic acid; 1 cm3 of KOH neutralizes 1/2 cm3 of oxalic acid
25 cm3 of KOH neutralizes = 1/2 × 25 = 12.5 cm3
Answer: 12.5 cm^3

SECTION 3: PERCENTAGE COMPOSITION BY MASS

Q.4 Define percentage composition by mass. Write a formula to determine it.

Percentage Composition by Mass

- The percentage composition of a compound refers to the percentage mass of each element present in the compound.
- **Percentage of an element:** The number of grams of that element present in 100 grams of the compound.

Formula

Percentage of an element = (Mass of the element in the compound / Formula mass of the compound) × 100

Interesting Information
The composition of all the chemical products which are in our daily use, such as shampoos, cleaners, soaps, perfumes and fertilizers are formed using stoichiometric calculations.

Importance of Percentage Composition

- Useful concept to determine the purity of a compound
- Helps determine empirical and molecular formulae
- Can be determined both theoretically and by performing quantitative analysis

Example 15.9

Find out the percentage composition of $Ca(OH)_2$.

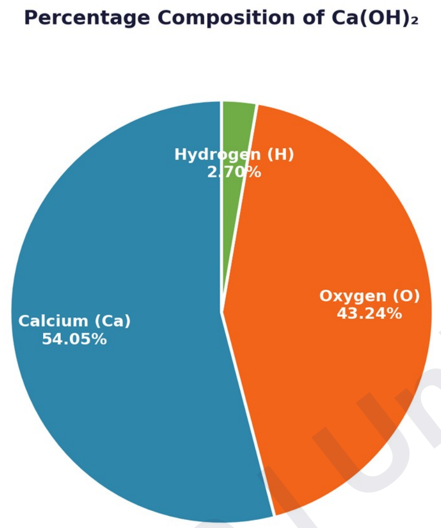
Solution:

Formula mass of $Ca(OH)_2$ = 40 + 32 + 2 = 74 g/mol

Percentage of Ca = (40/74) × 100 = 54.05%

Percentage of Oxygen = (32/74) × 100 = 43.24%

Percentage of Hydrogen = (2/74) × 100 = 2.70%



Percentage Composition of Ca(OH)2

Example 15.10

8.657 g of a compound were decomposed into its elements. It gave 5.217 g of carbon, 0.962 g of hydrogen and 2.478 g of oxygen. Find out the percentage composition of the compound.

Solution:

Percentage of carbon = (5.217/8.657) × 100 = 60.26%

Percentage of hydrogen = (0.962/8.657) × 100 = 11.11%

Percentage of oxygen = (2.478/8.657) × 100 = 28.62%

So, 100 g of the compound contains: Carbon: 60.26 g; Hydrogen: 11.11 g; Oxygen: 28.62 g

Calculate the percentage of sulphate ions SO_4^{2-} in $Al_2(SO_4)_3$

Given data: Formula of sulphate ion = SO_4^{2-}; Formula of compound (Aluminium Sulphate) = $Al_2(SO_4)_3$

Solution:

Molar mass of SO_4^{2-} ion = 96 × 3 = 288 (as there are 3 SO_4^{2-} ions)

Molar mass of $Al_2(SO_4)_3$ = 27×2 + 96×3 = 342

Percentage of SO_4^{2-} ions = (288/342) × 100 = 84.21%

Answer: 84.21%

SECTION 4: EMPIRICAL FORMULA

Q.5 Define empirical formula. Give examples. Write steps to write empirical formula of a compound.

Empirical Formula

- The formula which shows the simplest whole number ratio of atoms present in a compound is called empirical formula.

Examples:

- Hydrogen peroxide (H_2O_2) → HO
- Water (H_2O) → H_2O
- Benzene (C_6H_6) → CH

Empirical Formulae of Ionic Compounds

- All ionic compounds are represented by their empirical formulae.
- These formulae show the simplest ratio present between their ions.
- Example: NaCl shows simplest ratio between sodium and chloride ions.
- Example: $CaCl_2$ shows the ratio between calcium and chloride ions.

Similar Empirical Formulae of Different Compounds

- Since empirical formula does not tell about actual number of atoms, different compounds may have same empirical formula.

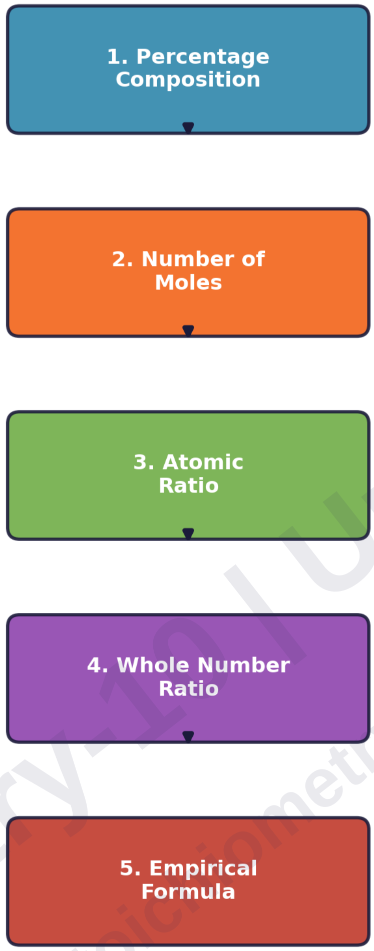
Examples:

- Acetylene (C_2H_2) and Benzene (C_6H_6) → same empirical formula CH
- Acetic acid (CH_3COOH) and Glucose ($C_6H_{12}O_6$) → same empirical formula CH_2O

Steps to Calculate Empirical Formula of a Compound

- **(i) Percentage composition:** Determine the percentage composition of the compound.
- **(ii) Number of moles:** Find out the number of moles of each element by dividing the mass of each element (percentage of element) by its atomic mass.
- **(iii) Atomic ratio:** Determine the atomic ratio of each element by dividing the number of moles of each element by the smallest number of moles.
- **(iv) Whole number atomic ratio:** If the atomic ratio obtained is simple whole number, it gives the empirical formula, otherwise multiply this ratio with suitable digit to get the whole number ratio.
- **(v) Subscripts for the elements:** Use this whole number ratio as a subscript for the elements to get the empirical formula.

Steps to Find Empirical Formula



Steps to Find Empirical Formula

Example 15.11

The percentage composition of a compound shows it to contain 25.26% Mg and 74.74% Chlorine. Find out its empirical formula.

Solution:

1. Change percentage to grams: Mg = 25.26 g, Cl = 74.74 g

2. Number of moles: Mg = 25.26/24 = 1.05, Cl = 74.74/35.5 = 2.10

3. Atomic ratio: Mg = 1.05/1.05 = 1, Cl = 2.10/1.05 = 2

Empirical formula = $MgCl_2$

Example 15.12

A compound contains 64.8% C, 13.62% H and 21.58% O. What is its empirical formula?

Solution:

1. Change percentage to grams: C = 64.8 g, H = 13.62 g, O = 21.58 g

2. Number of moles: C = 64.8/12 = 5.4, H = 13.62/1 = 13.62, O = 21.58/16 = 1.35

3. Atomic ratio: C = 5.4/1.35 = 4, H = 13.62/1.35 = 10, O = 1.35/1.35 = 1

Empirical formula = $C_4H_{10}O$

SECTION 5: MOLECULAR FORMULA

Q.6 Define molecular formula of a compound. Give examples. How molecular formula can be calculated by knowing its empirical formula.

Molecular Formula

- The formula which represents the actual number of atoms present in the molecule of a substance.

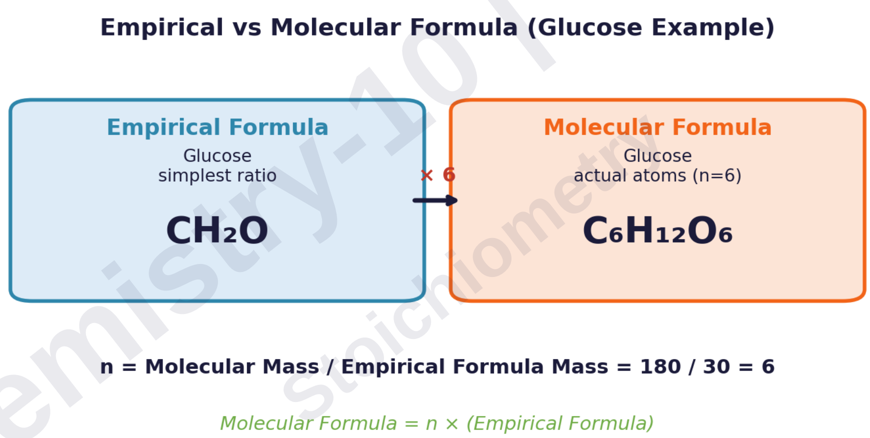
Examples:

- Water: H_2O
- Hydrogen peroxide: H_2O_2
- Ethylene: C_2H_4
- Sulphur: S_8

Calculation of Molecular Formula from Empirical Formula

Molecular formula = n (Empirical formula)

where n = Molecular Mass / Empirical Formula Mass



Empirical vs Molecular Formula (Glucose Example)

Example 15.13

Benzene is represented by its empirical formula CH while its molecular mass is 78. Find out its molecular formula.

Solution:

Empirical formula of benzene = CH; Empirical formula mass = 12 + 1 = 13

Molecular mass = 78; n = 78/13 = 6

Molecular formula = 6(CH) = C_6H_6

Example 15.14

Glucose has an empirical formula CH_2O and its molecular mass is 180 g mol^{-1}. What will be its molecular formula?

Solution:

Empirical formula of glucose = CH_2O; Empirical formula mass = 12 + 2 + 16 = 30

Molecular mass = 180; n = 180/30 = 6

Molecular formula = $6(CH_2O) = C_6H_{12}O_6$

SECTION 6: CHEMICAL EQUATIONS

Q.7 Define a chemical equation. Explain with an example how a balanced chemical equation can help us to determine masses of reactants and products.

Chemical Equation

- Symbolic representation of a chemical reaction in terms of symbols of elements and formulae of compounds.

Complete and Balanced Chemical Equation

- Tells us the mole ratio or mass ratio between the reactants and the products.

With this ratio we can find out:

 - Masses of products from given masses of reactants
 - Masses of reactants from given masses of products

Example

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(g)} + CO_{2(g)}$$

1 mole	2 moles	1 mole	1 mole	1 mole
100 g	73 g	111 g	18 g	44 g

This equation tells us that one mole (100 g) calcium carbonate reacts with two moles (73 g) of hydrochloric acid to produce:

- One mole (111 g) of $CaCl_2$
- One mole (18 g) of water
- One mole (44 g) of carbon dioxide

Quick Check 15.7 — Why is it necessary for a chemical equation to be balanced before it can be used in calculation?
Ans. It is necessary to balance the chemical equation before calculations because balanced chemical equations provide a standard for accurate calculations. Without balancing of equations, calculations cannot give right answer.

Example 15.15

25 g of lime stone ($CaCO_3$) react with an excess of hydrochloric acid. How much calcium chloride ($CaCl_2$) will be produced?

Solution:

According to the equation: 100 g of lime stone produce calcium chloride = 111 g

1 g of lime stone produce calcium chloride = 111/100 g

25 g of lime stone produce calcium chloride = (111/100) × 25 = 27.75 g

Answer: 27.75 g

SECTION 7: LIMITING REACTANT

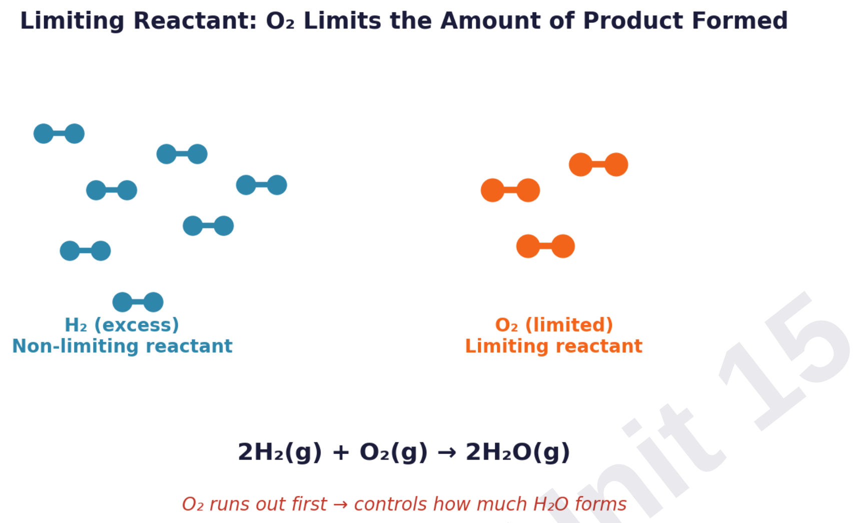
Q.8 What do you know about limiting reactant? Which steps are followed to find limiting reactant in a chemical reaction?

Limiting Reactant

- The reactant which is consumed earlier is called the limiting reactant.
- It can also be defined as the reactant that produces the lowest moles of products in a reaction.
- It is the limiting reactant which controls the amount of the product formed in a chemical reaction.

Non-limiting or Excess Reactant

- The reactant which is used in excess and left behind at the end of the chemical reaction.



Limiting Reactant: O2 Limits the Amount of Product Formed

Steps to Identify Limiting Reactant

- **(i) Number of moles of reactant:** Calculate the number of moles of the reactants from the amounts provided.
- **(ii) Number of moles of products:** Calculate the number of moles of the product which will form with the given moles of reactants.
- **(iii) Least amount of product:** Identify the reactant which gives the least amount of the product. This reactant will be the limiting reactant.

SECTION 8: YIELD

Q.9 What do you understand by yield? Differentiate between actual and theoretical yield.

Yield

- The amount of product obtained from a chemical reaction.

Actual Yield vs Theoretical Yield

Actual Yield	Theoretical Yield
The quantity of the product obtained when chemical reaction is performed	The amount of product which we calculate assuming that all the reactants react according to the balanced chemical equation
Mostly not fixed for a chemical reaction	Always fixed for a chemical reaction
Lesser than theoretical yield	Greater than actual yield

Percentage Yield

Percentage yield = (Actual yield / Theoretical yield) × 100

Example 15.16

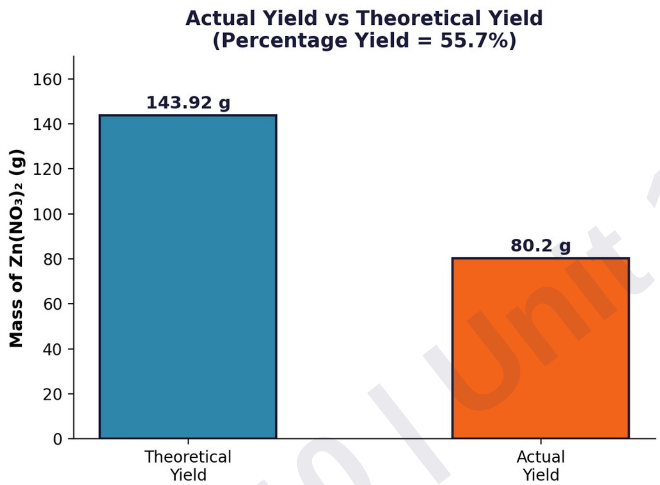
In a reaction, 0.76 mol of zinc gives 0.76 mol of $Zn(NO_3)_2$. If the actual amount of $Zn(NO_3)_2$ obtained is 80.2 g, what is the percentage yield?

Solution:

$$\text{Molar mass of } Zn(NO_3)_2 = 65.4 + 2(14 + 48) = 189.38 \text{ g/mol}$$

$$\text{Theoretical yield} = 0.76 \times 189.38 = 143.92 \text{ g}$$

$$\textbf{Percentage yield} = (80.2/143.92) \times 100 = 55.7\%$$



Actual Yield vs Theoretical Yield — Zn(NO3)2

Answer: 55.7%

Find out the theoretical yield and percent yield of oxygen generated by heating 40 g of $KClO_3$ (M = 122.5). The mass of oxygen gas produced is 14.9 g.

Given data:

$$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$$

Mass of $KClO_3$ = 40 g; Molar mass of $KClO_3$ = 122.5 g/mol; Mass of O_2 (Actual yield) = 14.9 g; Molar mass of O_2 = 32 g/mol

Solution:

$$\text{Number of moles of } KClO_3 = 40/122.5 = 0.3265 \text{ mol}$$

$$2 \text{ mol } KClO_3 \text{ gives} = 3 \text{ mol } O_2; \quad 1 \text{ mol } KClO_3 \text{ gives} = 3/2 \text{ mol } O_2$$

$$0.3265 \text{ mol } KClO_3 \text{ gives} = 3/2 \times 0.3265 = 0.490 \text{ mol } O_2$$

(i) Mass of O_2 (Theoretical yield) = 0.490 × 32 = 15.68 g

(ii) Percentage yield = (14.9/15.68) × 100 = 95.025%

SECTION 9: PERCENTAGE PURITY

Q.10 What do you understand by term percentage purity of a substance?

Percentage Purity

- In chemistry, purity refers to the degree to which a substance is free from impurities.
- Percentage purity of a compound indicates the amount of pure compound present in an impure sample.

Formula

Percentage purity = (Mass of the pure sample / Total mass of the impure sample) × 100

Example 15.18

12.0 g sample of a pharmaceutical product was found to contain 11.57 g of the active drug. Find out the percentage purity of the sample.

Solution:

$$\text{Percentage purity} = (11.57/12.0) \times 100 = 96.4\%$$

Answer: 96.4%

SECTION 10: MULTIPLE CHOICE QUESTIONS

Molar Volume

1. Volume occupied by 15 moles of ammonia gas at RTP: (a) 224 dm^3 (b) 160 dm^3 (c) 360 dm^3 (d) 265 dm^3

2. How many moles of natural gas CH_4 will be present in a cylinder if its volume is 18 dm^3 at RTP? (a) 0.75 mol (b) 0.67 mol (c) 0.85 mol (d) 0.56 mol

3. What will the molar concentration of potassium hydroxide solution if 2.0 g of it are dissolved in 100 cm^3 of solution? (a) 0.36 mol/dm^3 (b) 0.53 mol/dm^3 (c) 0.70 mol/dm^3 (d) 0.45 mol/dm^3

4. What will be the mass of NaCl needed to make up 15 dm^3 of its solution with a concentration of 0.7 g/dm^3? (a) 21 g (b) 10.5 g (c) 31.5 g (d) 5.4 g

5. Find out the concentration of H_2SO_4 in mol/dm^3 if its 10 cm^3 react with 20 cm^3 of NaOH solution having concentration 0.1 mol/dm^3. (a) 0.01 mol/dm^3 (b) 0.001 mol/dm^3 (c) 0.1 mol/dm^3 (d) 0.15 mol/dm^3

6. If the molar mass of a compound is 178 g/mol and its empirical formula is $C_3H_5O_3$. What is its molecular formula? (a) $C_3H_5O_3$ (b) $C_6H_{10}O_6$ (c) $C_9H_{15}O_9$ (d) $C_8H_{10}O_8$

7. When natural gas burns in air, which component is the limiting reactant? (a) Both components (b) Neither component is a limiting reactant (c) Air (d) Natural gas

8. If the actual yield of a compound is 0.198 g while the theoretical yield is 0.217 g. What will be the percentage yield? (a) 50% (b) 91.2% (c) 80% (d) 90%

SLO Based MCQs

9. The room temperature and pressure (RTP) is: (a) 25°C, 10 atm (b) 20°C, 1 atm (c) 20°C, 10 atm (d) 25°C, 1 atm

10. Calculate the number of moles of CH_4 gas, when its volume is 5 dm^3 at RTP: (a) 0.1 moles (b) 0.25 moles (c) 0.21 moles (d) 1.5 moles

11. 7.5 moles of the gas contain ____ dm^3 of volume at RTP. (a) 0.3125 dm^3 (b) 180 dm^3 (c) 1.8 dm^3 (d) 3.125 dm^3

12. Calculate the volume of CO_2 at RTP when its mass is 88 g. (a) 48 dm^3 (b) 22.414 dm^3 (c) 8.8 dm^3 (d) 24 dm^3

13. The measure of amount of solute that has been dissolved in a given amount of solution is: (a) Pressure of solution (b) Volume of solution (c) Condition of solution (d) Concentration of solution

14. Calculate the mass of sodium carbonate (Na_2CO_3) in a solution of volume 2 dm^3 having concentration of 0.5 g/dm^3. (a) 0.1 g (b) 0.01 g (c) 1 g (d) 0.001 g

15. Calculate concentration of a solution in g/dm^3 which contains 0.2 g of calcium acetate dissolved in 4.5 dm^3 of solution: (a) 0.044 g/dm^3 (b) 4.04 g/dm^3 (c) 40.2 g/dm^3 (d) 0.022 g/dm^3

16. 100 cm^3 of hydrochloric acid contains 0.4 g of dissolved hydrogen chloride. Calculate the concentration of the acid in g/dm^3. (a) 0.4 g/dm^3 (b) 0.004 g/dm^3 (c) 0.04 g/dm^3 (d) 4.0 g/dm^3

17. Molar concentration and ____ both refer to the concentration of a solute in a solution. (a) Molarity (b) Molality (c) Molar volume (d) Percentage composition

18. What is the mass of potassium fluoride in grams dissolved in 0.1 dm^3 of solution having concentration of 0.001 g/dm^3? (a) 0.0001 g (b) 0.01 g (c) 0.005 g (d) 0.1 g

Calculate the Concentration in Titration using Empirical Formula

19. The concentration of a solution can be determined by titrating it with a solution of ____ till the end point reached. (a) Unknown concentration (b) Normal concentration (c) Known concentration (d) Low concentration

20. In titration ____ will change its colour to show end point. (a) Indicator (b) Salt (c) Acid (d) Base

21. Stoichiometry is essential for ____ chemical reactions in many industries. (a) Securing (b) Minimizing (c) Maximizing (d) Optimizing

Percentage Composition by Mass, Empirical and Molecular Formula

22. Percentage composition of a compound is the percentage mass of each ____ present in the compound. (a) Specie (b) Element (c) Particle (d) Molecule

23. The percentage of H-atom in H_2O is: (a) 88.89% (b) 1.11% (c) 2% (d) 11.11%

24. To determine molecular and empirical formula, ____ is useful concept. (a) Percentage composition (b) Molar volume (c) Titration (d) Molarity

25. The percentage of calcium (Ca) in $CaCO_3$: (a) 30% (b) 60% (c) 40% (d) 74%

26. The percentage of an element is ____ type analysis: (a) Qualitative (b) Quantitative (c) Normal (d) All of these

27. What is the percentage of C in CH_4? (a) 25% (b) 65% (c) 50% (d) 75%

28. Calculate percentage of chloride (Cl^{-1}) ions in $CaCl_2$: (a) 40% (b) 24% (c) 71% (d) 64%

Empirical Formula

29. Which of the following pair is correct empirical formula and atomic ratio of H_2SO_4? (a) HSO_4 1:1:1 (b) H_2SO_4 2:1:2 (c) H_2SO_4 2:1:7 (d) HSO_2 1:1:2

30. All ____ compounds are represented only by their empirical formula: (a) Ionic (b) Covalent (c) Co-ordinate covalent (d) Unstable

31. The empirical formula of a gas ethyne (C_2H_2) used in cutting and welding of metals is: (a) CH_2 (b) C_2H (c) CH_3 (d) CH

32. Empirical formula shows the ____ number ratio of atoms in a compound: (a) Complex (b) Simplest (c) Normal (d) Abnormal

33. Number of steps used to calculate empirical formula of a compound are: (a) 5 (b) 4 (c) 7 (d) 3

34. The atomic ratio in empirical formula can be found by dividing the number of moles of each element by the: (a) Largest number of moles (b) Maximum number of moles (c) Vital number of moles (d) Smallest number of moles

35. The empirical formula of acetic acid $C_2H_4O_2$ is: (a) CH_2O (b) CHO (c) CHO_2 (d) CH_2O_2

36. Which one of the following is not an empirical formula? (a) $CaCl_2$ (b) $C_6H_{12}O_6$ (c) Na_2CO_3 (d) KCl

Molecular Formula

37. Which one of the following compound will have same molecular and empirical formula? (a) H_2O (b) $C_6H_{12}O_6$ (c) $C_2H_4O_2$ (d) C_6H_6

38. ____ is used to determine the optimal dosages of medicines for patients based on their weight and age etc. (a) Limiting reactant (b) Yield (c) Thermochemistry (d) Stoichiometry

39. "n" is a number found by dividing molecular mass with: (a) Ionic mass (b) Formula mass (c) Empirical formula mass (d) Atomic mass

40. Which one of the following is the molecular formula of ethylene? (a) C_2H_4 (b) CH_4 (c) C_2H_6 (d) C_2H_2

Calculations Based on Chemical Equation

41. A complete balanced chemical equation tells about ____ between reactants and products: (a) Nature (b) Density ratio (c) Reactivity (d) Mole ratio

42. $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$, when 64 g of O_2 reacts completely with hydrogen. How many grams of H_2O will be produced? (a) 72 g (b) 6.4 g (c) 64 g (d) 36 g

43. $NaOH + HCl \rightarrow NaCl + H_2O$, the mole ratio between reactants in this reaction is: (a) 1:2 (b) 2:1 (c) 1:1 (d) 1:4

44. In the decomposition of calcium carbonate, when 200 g of calcium carbonate $CaCO_3$ is heated, the mass of CaO will be: (a) 74 g (b) 112 g (c) 7.4 g (d) 56 g

Limiting Reactant

45. In a chemical reaction, the reactant which is consumed earlier is: (a) Excess reactant (b) Non-limiting reactant (c) Stable reactant (d) Limiting reactant

46. When 80 g of H_2 and 32 g of O_2 reacted, the limiting reactant will be: (a) H_2 (b) O_2 (c) H_2 and O_2 (d) None

47. Number of steps to identify limiting reactant is: (a) 1 (b) 2 (c) 4 (d) 3

48. The reactant that controls the amount of product formed in a chemical reaction is: (a) Limiting reactant (b) Non-limiting reactant (c) Reactive reactant (d) Excess reactant

Yield

49. The amount of product obtained when chemical reaction is performed is called: (a) Theoretical yield (b) Actual yield (c) Real yield (d) Supposed yield

50. To check efficiency of reaction, _____ is calculated: (a) Actual yield (b) Molarity (c) Volume (d) Percentage yield

Percentage Purity

51. The unit of percentage purity is: (a) g/mol (b) dm^3 (c) mol (d) no unit

52. Percentage purity of compound indicates the amount of _____ present in an impure sample: (a) Real compound (b) Pure compound (c) Stable compound (d) Unstable compound

Answer Key for MCQs

1	**2**	**3**	**4**	**5**	**6**	**7**	**8**
c	a	a	b	c	b	d	b
9	**10**	**11**	**12**	**13**	**14**	**15**	**16**
d	c	b	a	d	c	a	d
17	**18**	**19**	**20**	**21**	**22**	**23**	**24**
a	a	c	a	d	b	d	a
25	**26**	**27**	**28**	**29**	**30**	**31**	**32**
c	b	d	d	d	a	d	b
33	**34**	**35**	**36**	**37**	**38**	**39**	**40**
a	d	a	b	a	d	c	a
41	**42**	**43**	**44**	**45**	**46**	**47**	**48**
d	a	c	b	d	b	d	a
49	**50**	**51**	**52**				
b	d	d	b				

SECTION 11: SHORT ANSWER QUESTIONS (EXERCISE)

1. What is the molar volume of a gas at RTP?

Answer: Molar Volume: It is the volume occupied by one mole of a gas at RTP. It has been found that one mole of all gases occupies a volume of 24 dm3 at RTP. Examples: 1 mole of H2 gas at RTP = 2.016 g = 24 dm3; 1 mole of CH4 gas at RTP = 16 g = 24 dm3

2. How does the concept of molar volume useful for us?

Answer: The concept of molar volume is useful because it enables us to calculate the number of moles or mass of gas if we know their volume. The concept of molar volume gives a consistent reference point for comparing different gases at the same conditions.

3. How does the molar volume relate to the Avogadro's law?

Answer: Molar volume and Avogadro's law are directly connected to each other. Molar volume actually is the outcome of Avogadro's law. Because of Avogadro's law, one mole of any gas contains the same number of molecules that is 6.02 × 10^23. Therefore, one mole of any gas occupies the same volume of 24 dm3 at RTP. This fixed volume 24 dm3 is called molar volume.

4. Why two different compounds may show the same empirical formula?

Answer: Two different compounds can show the same empirical formula because the empirical formula only represents the simplest whole number ratio of atoms in a compound, not the actual number of atoms present in a compound. Example: Acetic acid (CH3COOH) and glucose (C6H12O6) have the same empirical formula CH2O.

5. Why is percentage yield important?

Answer: Importance of percentage yield: Percentage yield tells about the efficiency of a reaction. A higher percentage yield means the reaction produces more products with less waste. In chemical industries, it also helps to decide whether a process is economical and profitable or not.

6. What is the concentration of a solution in mol/dm³ if it contains 49 g of H_2SO_4 in one dm³ of solution?

Answer:

Mass of H_2SO_4 = 49 g

Number of moles of H_2SO_4 = 49/98 = 0.5 mol

Concentration = 0.5 mol/dm³

SLO BASED SHORT ANSWER QUESTIONS

7. Calculate the number of moles of CH_4 in 0.5 dm³ at RTP.

Answer: Volume of CH4 at RTP = 0.5 dm3; Molar volume at RTP = 24 dm3. Number of moles = Volume/Molar volume = 0.5/24 = 0.021 mol

8. What is RTP?

Answer: The standard condition of temperature 25°C (298.15 K) and pressure 1 atm (760 torr) is called room temperature and pressure (RTP). These standard conditions are used to measure the volume of gases.

9. Calculate the number of moles of O_2 gas, when its volume is 3.7 dm³ at RTP.

Answer: Volume of O2 at RTP = 3.7 dm3; Molar volume at RTP = 24 dm3. Number of moles = 3.7/24 = 0.154 moles

10. Define concentration of a solution. Give example.

Answer: Concentration of a solution: The measure of the amount of solute that has been dissolved in a given amount of solution or solvent is called concentration of a solution. Unit: The unit of molar concentration is mol/dm3.

11. What is mass of sodium chloride in grams dissolved in 5 dm³ of the solution having concentration of 0.4 g/dm³?

Answer: Mass of NaCl = Concentration × Volume = 0.4 g/dm3 × 5 dm3 = 2.0 g

12. What is percentage composition? Give example.

Answer: Percentage composition: The percentage composition of a compound refers to the percentage mass of each element present in the compound. Example: 11.11% H and 88.89% O is the percentage composition of H2O.

13. What is percentage of an element in a compound?

Answer: The percentage of an element in a compound is the number of grams of that element in 100 grams of the compound. Example: The percentage of carbon element (C) is 75% in CH4.

14. How percentage composition is measured?

Answer: The percentage composition is the measure of percentage of all the elements in a compound and it is calculated by following formula: Percentage of an element = (Mass of element in compound / Molar mass of compound) × 100

15. Give advantages of percentage composition.

Answer: Percentage composition is a useful concept to determine: The purity of a compound; Empirical and molecular formulae

16. What is empirical formula? Give example.

Answer: Empirical formula: The formula which shows the simplest whole number ratio of atoms present in a compound is called empirical formula. Example: Empirical formula of hydrogen peroxide (H2O2) is HO.

17. What are the steps to calculate empirical formula?

Answer:

- **(i) Percentage composition:** Determine the percentage composition of the compound.
- **(ii) Number of moles:** Find out the number of moles of each element by dividing the mass of each element (percentage of element) by its atomic mass.
- **(iii) Atomic ratio:** Determine the atomic ratio of each element by dividing the number of moles of each element by the smallest number of moles.
- **(iv) Whole number atomic ratio:** If the atomic ratio obtained is simple whole number, it gives the empirical formula, otherwise multiply this ratio with suitable digit to get the whole number ratio.
- **(v) Subscripts for the elements:** Use this whole number ratio as a subscript for the elements to get the empirical formula.

18. Define molecular formula. Give example.

Answer: Molecular formula: Molecular formula of an element or a compound represents the actual number of atoms present in the molecule of these substances. Examples: Water: H2O; Hydrogen peroxide: H2O2; Ethylene: C2H4; Sulphur: S8

19. How molecular formula can be calculated?

Answer: Molecular formula of a compound can be calculated by multiplying a whole number with the empirical formula of that compound: Molecular Formula = n(Empirical Formula), where n = Molecular Mass / Empirical Formula Mass

20. What is limiting reactant?

Answer: The reactant which is consumed earlier and controls the amount of product formed in a chemical reaction is called limiting reactant.

21. What is non-limiting or excess reactant?

Answer: The reactant which is used in excess and left behind at the end of the chemical reaction is called non-limiting or excess reactant.

22. How limiting reactant can be identified?

Answer: A limiting reactant present in a chemical reaction can be identified by carrying out the following steps: (i) Number of moles of reactant: Calculate the number of moles of the reactants from the amounts provided. (ii) Number of moles of products: Calculate the number of moles of the product which will form with the given moles of reactants. (iii) Least amount of product: Identify the reactant which gives the least amount of the product. This reactant will be the limiting reactant.

23. Differentiate between actual and theoretical yield.

Answer:

Actual Yield	Theoretical Yield
The quantity of the product obtained when chemical reaction is performed is called actual yield.	The amount of product which we calculate assuming that all the reactants react according to the balanced chemical equation is called theoretical yield.
It is mostly not fixed for a chemical reaction.	It is always fixed for a chemical reaction.
It is lesser than theoretical yield.	It is greater than actual yield.

24. How percentage yield is measured?

Answer: Percentage yield is measured by comparing the actual and theoretical yield. It is calculated by following formula: Percentage yield = (Actual yield / Theoretical yield) × 100

25. Define percentage purity. How is it measured?

Answer: Percentage purity: Percentage purity of a compound indicates the amount of pure compound present in the impure sample. Percentage purity = (Mass of the pure sample / Total mass of the impure sample) × 100

SECTION 12: CONSTRUCTED RESPONSE QUESTIONS

1. How would you identify the limiting reactant in a chemical reaction?

Answer: See Q.8 from theory (Steps to Identify Limiting Reactant).

2. Why is it important to find the purity of a compound which is used as a medicine?

Answer: Purity of a compound used in medicine: It is very important to find the purity of a compound used in medicine because: Impurities may be harmful or toxic; Impure medicines can give incorrect dosage; Impurities reduce effectiveness; Impurities cause side effects. The percentage purity of a medicine ensures that medicines are safe, effective and reliable for patient use.

3. Will there be a limiting reactant in a reversible reaction?

Answer: Limiting Reactant: "The reactant which is consumed earlier is called limiting reactant." It may also be defined as "the reactant that produces the lowest moles of products in a reaction." Limiting reactant in a reversible reaction: Reversible reactions are those reactions in which reactants are not completely converted into products. The limiting reactant controls the amount of product formed but in reversible reactions, as no product is formed or reactions are incomplete, there is no limiting reactant in reversible reactions.

4. How do you know if a formula is empirical or not?

Answer: We can know whether a formula is empirical or not by checking the ratio of atoms in the formula. Identification of empirical formula: If the simplest whole number ratio is formed between the atoms of an element, then the formula is empirical. Example: CH2O is empirical formula of glucose C6H12O6. If the subscripts in the formula can be reduced simpler further, then the formula is not empirical.

5. Write down one method to find out the molecular mass of a compound without knowing its molecular formula.

Answer: Determination of molecular mass without molecular formula: Normally a molecular formula is written by multiplying its empirical formula with a number "n" and "n" is the ratio of molecular mass to formula mass: Molecular formula = n(Empirical formula); n = Molecular mass / Formula mass. There are few ways to determine the molecular mass without knowing the molecular formula, e.g.: Mass spectrometry (MS); Freezing point depression; Boiling point elevation; Vapour density method

SECTION 13: NUMERICAL PROBLEMS (EXERCISE)

Problem 1: Calculate the number of moles, volume and number of molecules in 4.0 g of CH_4 at RTP.

Given data: Mass of CH_4 = m = 4.0 g; Molar mass of CH_4 = 12 + (1×4) = 16 g/mol

Solution:

$$\text{Number of moles (n)} = \text{Mass/Molar mass} = 4/16 = 0.25 \text{ mol}$$

$$\text{Volume of } CH_4 = n \times \text{Molar volume} = 0.25 \times 24 = 6 \text{ dm}^3$$

$$\text{Number of molecules} = n \times N_A = 0.25 \times 6.02 \times 10^{23} = 1.5 \times 10^{23}$$

Answers: Number of moles = 0.25 mol; Volume = 6 dm³; Number of molecules = 1.5 × 10^{23}

Problem 2: What mass of NaCl needs to be dissolved in 150 cm³ of a solution if the concentration of the solution is 0.4 mol/dm³?

Given data: Volume of solution = 150 cm³; Concentration = 0.4 mol/dm³; Molar mass of NaCl = 23 + 35.5 = 58.5 g/mol

Solution:

$$\text{Volume in dm}^3 = 150/1000 = 0.15 \text{ dm}^3$$

$$\text{Number of moles} = \text{Concentration} \times \text{Volume} = 0.4 \times 0.15 = 0.06 \text{ mol}$$

$$\text{Mass of NaCl} = \text{Number of moles} \times \text{Molar mass} = 0.06 \times 58.5 = 3.51 \text{ g}$$

Answer: 3.51 g

Problem 3: A compound contains by mass, 40.0% C, 6.71% H and 53.3% oxygen. A 0.320 mole of this compound weighs 28.8 g. What is the molecular formula of this compound?

Given data: Percentage of C = 40.0%; Percentage of H = 6.71%; Percentage of O = 53.3%; Number of moles of compound = 0.320 mol; Mass of compound = 28.8 g

Solution:

Step 1: Number of moles of each element

$$C = 40/12 = 3.33; \quad H = 6.71/1.008 = 6.65; \quad O = 53.3/16 = 3.33$$

Step 2: Atomic ratio

$$C = 3.33/3.33 = 1; \quad H = 6.65/3.33 = 2; \quad O = 3.33/3.33 = 1$$

$$\textbf{Empirical formula} = \mathbf{CH_2O}$$

Step 3: Molar mass of compound

$$\text{Number of moles} = \text{Mass/Molar mass}; \quad 0.320 = 28.8/\text{Molar mass}$$

$$\textbf{Molar mass = 90 g/mol}$$

Step 4: n = Molar mass / Empirical formula mass

$$\text{Empirical formula mass} = 12 + 2 + 16 = 30; \quad n = 90/30 = 3$$

$$\textbf{Molecular formula} = \mathbf{3(CH_2O) = C_3H_6O_3}$$

Answer: $C_3H_6O_3$

Problem 4: A sample of metal has a total mass of 3.66 kg and contains 2.45 kg of gold. What is the percentage purity of gold in this sample?

Given data: Total mass of sample = 3.66 kg; Mass of gold = 2.45 kg

Solution:

$$\text{Percentage purity} = (\text{Mass of pure gold / Total mass of sample}) \times 100$$

$$\textbf{Percentage purity} = \mathbf{(2.45/3.66) \times 100 = 66.94\%}$$

Answer: 66.94%

Problem 5: Calculate the percentage yield when 305 g of $AgNO_3$ reacts with $MgCl_2$ to give 23.7 g of $Mg(NO_3)_2$.

Given data:

$$2AgNO_{3(aq)} + MgCl_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2AgCl_{(aq)}$$

Molar mass of $AgNO_3$ = 108 + 14 + 48 = 170 g/mol; Mass of $AgNO_3$ = 305 g

Molar mass of $Mg(NO_3)_2$ = 24 + 14×2 + 16×6 = 148 g/mol; Actual yield of $Mg(NO_3)_2$ = 23.7 g

Solution:

$$\text{Moles of } AgNO_3 = 305/170 = 1.794 \text{ mol}$$

From equation: 2 mol $AgNO_3$ gives 1 mol $Mg(NO_3)_2$

$$1.794 \text{ mol } AgNO_3 \text{ gives} = 1/2 \times 1.794 = 0.897 \text{ mol}$$

$$\text{Mass of } Mg(NO_3)_2 \text{ (Theoretical yield)} = 0.897 \times 148 = 132.756 \text{ g}$$

$$\text{Percentage yield} = (\text{Actual yield / Theoretical yield}) \times 100$$

$$\textbf{Percentage yield} = \mathbf{(23.7/132.756) \times 100 = 17.85\%}$$

Answer: 17.85%

IMPORTANT FORMULAS SUMMARY

1. Concentration

Concentration = Mass of solute (g) / Volume of solution (dm^3)

Concentration (mol/dm^3) = Number of moles / Volume (dm^3)

2. Percentage Composition

Percentage of element = (Mass of element in compound / Formula mass) × 100

3. Empirical Formula

Steps: Percentage → Moles → Atomic Ratio → Whole Number Ratio → Empirical Formula

4. Molecular Formula

Molecular Formula = n(Empirical Formula); n = Molecular Mass / Empirical Formula Mass

5. Percentage Yield

Percentage yield = (Actual yield / Theoretical yield) × 100

6. Percentage Purity

Percentage purity = (Mass of pure sample / Total mass of impure sample) × 100

7. Molar Volume at RTP

1 mole of any gas at RTP = 24 dm^3; Number of moles = Volume (dm^3) / 24

KEY CHEMICAL EQUATIONS

1. Neutralization Reaction

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

2. Decomposition of Potassium Chlorate

$$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$$

3. Decomposition of Calcium Carbonate

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

4. Reaction of Calcium Carbonate with HCl

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(g)} + CO_{2(g)}$$

5. Reaction of Silver Nitrate with Magnesium Chloride

$$2AgNO_{3(aq)} + MgCl_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + 2AgCl_{(s)}$$

6. Combustion of Hydrogen

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

IMPORTANT MOLAR MASSES

Compound	Formula	Molar Mass (g/mol)
Water	H_2O	18
Sodium Hydroxide	NaOH	40
Hydrochloric Acid	HCl	36.5
Sodium Chloride	NaCl	58.5
Calcium Carbonate	$CaCO_3$	100

Calcium Oxide	CaO	56
Calcium Hydroxide	$Ca(OH)_2$	74
Magnesium Chloride	$MgCl_2$	95
Potassium Chlorate	$KClO_3$	122.5
Glucose	$C_6H_{12}O_6$	180
Methane	CH_4	16
Benzene	C_6H_6	78
Copper Sulphate	$CuSO_4$	159.5
Sulphuric Acid	H_2SO_4	98
Aluminium Sulphate	$Al_2(SO_4)_3$	342

KEY CONCEPTS SUMMARY

- **Concentration:** Measure of amount of solute in solution (g/dm^3 or mol/dm^3)
- **Titration:** Method to determine concentration using known solution
- **Percentage Composition:** Mass percentage of each element in compound
- **Empirical Formula:** Simplest whole number ratio of atoms
- **Molecular Formula:** Actual number of atoms in molecule
- **Limiting Reactant:** Reactant consumed first, controls product amount
- **Theoretical Yield:** Calculated product amount
- **Actual Yield:** Experimentally obtained product amount
- **Percentage Yield:** (Actual/Theoretical) × 100
- **Percentage Purity:** (Pure mass/Total mass) × 100
- **Molar Volume:** 24 dm^3 for 1 mole of gas at RTP

End of Chapter Notes