SELF-ASSESSMENT Chapter # 03

Total Mark: 30

Q.1 Encircle the correct option. **(1 × 6 = 6)**

(i) The shape of ammonia (NH_3) molecule according to VSEPR theory is:
A. Linear B. Tetrahedral C. Trigonal pyramidal D. Bent

(ii) The numbers of σ and π bonds in the N_2 molecule are:
A. one σ and one π bonds B. one σ and two π bonds
C. three σ bonds only D. two σ and one π

(iii) Which of the following molecules has zero dipole moment?
A. NH_3 B. $CHCl_3$ C. H_2O D. BF_3

(iv) Which of the following species contains a dative bond?
A. CH_4 B. NaCl C. NH_4^+ D. O_2

(v) Which of the following molecules shows hydrogen bonding?
A. CH_4 B. NH_3 C. CO_2 D. CF_4

(vi) Which type of hybridization of C occurs in CO_2?
A. sp^3 B. sp^2 C. sp D. dsp^2

Q.2 Write short answers of the following questions. **(2 × 8 = 16)**

(i) Explain the difference between the formation of σ and π bonds.
(ii) What are necessary conditions for the hydrogen bond to form?
(iii) Why does bond polarity lead to the development of dipole moments?
(iv) State the shape and bond angle of BF_3 using VSEPR theory.
(v) Draw the orbital structures of the CO_2 molecule in terms of VBT.
(vi) Write the hybridization of carbon in ethyne (C_2H_2).
(vii) Calculate the bond order of O_2 using molecular orbital theory.
(viii) SO_2 is a polar molecule but SO_3 not. Justify.

Q.3 Extensive Questions. **(2 × 4 = 8)**

(a) Draw the molecular orbital diagrams of the following molecules. Calculate their bond orders? (i) N_2 (ii) O_2
(b) What are the postulates of VSEPR model? Discuss the structures of the following species with reference to this theory.
(i) H_3O^+ (ii) PCl_5 (iii) SO_3 (iv) SF_6

Chapter 04 STOICHIOMETRY

Student Learning Outcomes

After studying this chapter, students will be able to:

- Derive measurements of mass, volume, and number of particles using moles. **(Application)**
- State the volume of one mole of a gas at STP. **(Knowledge)**
- Use the volume of one mole of gas at STP to solve mole-volume problems. **(Knowledge)**
- Calculate the gram molecular mass of a gas from density measurements at STP. **(Application)**
- Express balanced chemical equations in terms of moles, representative particles, masses, and volumes of gases at STP. **(Application)**
- Explain the concept of limiting reagents. **(Understanding)**
- Calculate the maximum amount of product and amount of any unreacted excess reagent. **(Application)**
- Calculate theoretical yield, actual yield, and percentage yield when given appropriate information. **(Application)**
- Calculate the quantities of reactants and products involved in a chemical reaction using stoichiometric principles. (Some examples include calculations involving reacting masses, volumes of gases, volumes, and concentrations of solutions, limiting reagent and excess reagent, percentage yield calculations). **(Application)**
- Explain with examples, the importance of stoichiometry in the production and dosage of medicine. **(Understanding)**

Introduction

Stoichiometry is derived from Greek words *stoicheion* means element and *metron* means measure. Collectively, stoichiometry means quantitative measure of reactants and products.

Stoichiometry

Stoichiometry (pronounced as stoy-key-om.eh-tree) is the branch of chemistry in which the relationship between the amounts of reactants and products in a balanced chemical equation is studied.

- The balanced chemical equation has the same number of atoms of each type on both sides of equation.
- It has definite ratios of reactants and products just as compounds have definite ratios of elements. Such ratios are used to calculate the mass or mole of other substances.

Laws in Stoichiometry:

Stoichiometric calculations obey law of conservation of mass and law of definite proportions.

1. **According to the law of conservation of mass, "matter (mass) can neither be created nor destroyed".** It states in terms of stoichiometry that the total mass of reactants is equal to the total mass of products in a balanced equation.
2. **According to the law of definite proportions, a pure compound always contains the same element combined in the same ratio by mass.**

CONCEPT OF MOLE

➢ **Definition**

The mole is the amount of a substance which contains as many elementary entities as there are atoms in 0.012 kg(12 g) of carbon-12.

- The elementary entities may be atoms, molecules, ions, electrons, and other particles.
- It is represented by 'n'.
- The number of entities present in one mole of a substance is a constant number named Avogadro's Number, i.e. 6.02×10^{23}. It is represented by N_A.
- This value is attributed to an Italian scientist Amedo Avogadro (1776-1856).

Did You Know?

The word 'mole' was introduced in 1896 by Wilhelm Ostwald who derived the term from Latin word moles meaning a 'heap' or 'pile'.

Did You Know?

Avogadro's number is a physical constant representing the molar number of entities. The exact value of it is 6.02214179×10^{23} mol^{-1}. In calculations we use the rounded off value 6.02×10^{23}.

➢ **Examples are Given Below:**

- 1 mole of ^{12}C contains 6.02×10^{23} atoms of ^{12}C.
- 1 mole of H_2O contains 6.02×10^{23} molecules of H_2O.
- 1 mole of $NaCl$ contains 6.02×10^{23} formula units of NaCl.
- 1 mole of Na^+ contains 6.02×10^{23} ions of Na^+.

➲ Molar Mass

- **Definition:** "The mass of one mole of a substance expressed in grams is called molar mass".
- **Unit:** The unit of molar mass is g/mol.
- **Example:** The molar mass is the sum of the masses of the component atoms.

The mass of one mole of CCl_4 can be found by adding the masses of carbon and chlorine present.

Molar mass of CCl_4 = Molar mass of one C + Molar mass of $Cl\times4$
Molar mass of $CCl_4 = 12.0\times1+35.5\times4$
Molar mass of $CCl_4 = 12.0+142.0 = 154.0$ g

Rack Your Mind!

1. Which of the following quantities represents one mole of oxygen molecules (O_2)?
A) 16.0 g of O_2
B) 6.02×10^{23} atoms of oxygen
C) 32.0 g of O_2
D) 22.4 cm^3 of O_2 at STP

The chemists use the mole as the SI unit to weigh and count atoms, molecules, formula units or ions.

The mass of one mole of a substance (element, compound or ionic species) is equal to the atomic mass, molecular mass, formula mass or ionic mass of a substance when expressed in grams and is known as molar mass, represented by M.

➢ **Other Examples**

- 1 mole of carbon atoms is 12.0 g.
- 1 mole of CO_2 molecule is 44.0 g.
- 1 mole of CaO formula units is 56.1 g.
- 1 mole of CO_3^{2-} ions is 60.0 g.

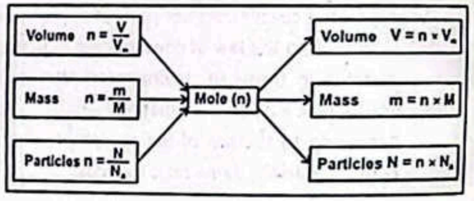
Rack Your Mind!

2. One mole of H_2O has two moles of bonds, three moles of atoms, ten moles of electrons and 28 moles of total fundamental particles. Explain.

The number of moles of a substance can be calculated by dividing mass in grams by molar mass. The formula for number of moles is:

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$n = \frac{m}{M}$$



Sample Problem 4.1

Calculate the number of moles present in 20 g of NaOH.

Solution:

Given mass of NaOH = 20 g
Molar mass of NaOH = 23 + 16 + 1 = 40 g/mol

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$n = \frac{20}{40} = 0.5 \text{ mol}$$

Sample Problem 4.2

Calculate the mass of 0.5 moles of HCl.

Solution:

No. of moles of HCl = 0.5 mol
Molar mass of HCl = 1 + 35.5 = 36.5 g/mol
Mass of HCl = Number of moles × Molar mass
= 0.5 × 36.5 = 18.25 g

Sample Problem 4.3

Calculate the mass of 10^{-3} mol of $MgSO_4$.

Solution:

Molar mass of $MgSO_4$ = 24 + 96 = 120 g mol^{-1}
Number of moles of $MgSO_4$ = 10^{-3}
Mass of $MgSO_4$ = 10^{-3} mol × 120 g mol^{-1} = 120×10^{-3} = 0.12 g

☑ QUICK CHECK 4.1

a. **Calculate the molar mass of $KMnO_4$.**

Ans. $KMnO_4$ = 39 + 55 + 16(4)
= 39 + 55 + 64
= 158 g/mol

b. **Calculate the number of moles in 0.23 g of sodium.**

Given: mass of sodium = m = 0.23 g
Molar mass of sodium = M = 23 g

To Find: Number of moles = n = ?

Solution:

$$n = \frac{m}{M} = \frac{0.23g}{23g}$$

n = 0.01 moles

c. **Calculate the mass of 1.5 moles of $Ca(OH)_2$.**

Ans. Molecular mass of 1 mole of $Ca(OH)_2$ = 40 + 2(16 + 1) = 74 g
Molecular mass of 1.5 mole of $Ca(OH)_2$ = 74 × 1.5 = 111 g

d. **The given mass of $KClO_3$ is 24.5 g. Calculate the number of moles of potassium chlorate.**

Given: Mass of $KClO_3$ = 24.5 g

To Find:
Number of moles of $KClO_3$ = ?
Number of formula units of $KClO_3$ = ?

Solution:

$$\text{Number of moles of } KClO_3 = \frac{\text{Mass of } KClO_3}{\text{Molar mass of } KClO_3} \qquad \boxed{n = \frac{m}{M}}$$

Molar Mass of $KClO_3$ = 122.5 g mol^{-1}

$$\text{Number of moles of } KClO_3 = \frac{24.5 \text{ g}}{122.5 \text{ g/mol}} = 0.2 \text{ moles}$$

e. How many molecules are present in 1.75 g of H_2O_2?

Given: Mass of H_2O_2 = m = 1.75 g

Molar mass of H_2O_2 = M = 2 + (16×2) = 34 g

To Find: Number of molecules N = ?

Solution:

$$\text{Number of molecules } H_2O_2 = N = \frac{m}{M}\times N_A$$

$$= \frac{1.75g}{34\,g/mol}\times 6.02\times10^{23} = 0.051\times6.02\times10^{23}$$

$$= 3.1\times10^{22}$$

f. How many atoms are present in 15 g of gold ring?

Given: mass of Gold (Au) = m = 15.0 g

Molar mass of Gold = M = 196.97 g/mol

To Find: Number of atoms = N = ?

$$N = \frac{m}{M}\times N_A$$

$$= \frac{15g}{196.97\,g/mol}\times6.02\times10^{23}$$

$$= 4.58\times10^{22}\text{ atoms}$$

RELATIONSHIP BETWEEN MOLE, MOLAR MASS AND AVOGADRO'S NUMBER

A sample of 12.0 grams of natural carbon contains the same number of atoms as 4.0 grams of natural helium. Both samples contain 1 mole of atoms i.e., *Note that different masses of elements have the same number of atoms.*

- 1.0 g of hydrogen = 1 mol of hydrogen = 6.02×10^{23} atoms of H
- 23.0 g of sodium = 1 mol of Na = 6.02×10^{23} atoms of Na
- 238.0 g of uranium = 1 mol of U = 6.02×10^{23} atoms of U

Q. Justify the following statements:

(i) 23 g of sodium and 238 g of Uranium have equal number of atoms in them.

(ii) 180 g of glucose and 342 g of sucrose have same number of molecules but different number of atoms present in them.

An atom of sodium is 23 times heavier than an atom of hydrogen. In order to have equal number of atoms, sodium should be taken 23 times greater in mass than hydrogen.

- 18.0 g of H_2O = 1 mol of water = 6.02×10^{23} molecules of water

180.0 g of glucose = 1 mol of glucose = 6.02×10^{23} molecules of glucose

Hence, one mole of different compounds has different masses but the same number of molecules.

Similarly, ***the number of ions in one mole of different ionic species is always the same, i.e., Avogadro's number.***

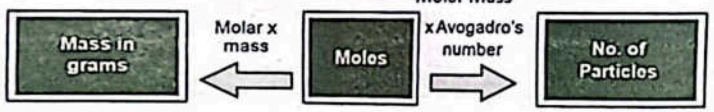
- 96.1 g of SO_4^{2-} = 1 mole of SO_4^{2-} = 6.02×10^{23} ions of SO_4^{2-}
- 62.0 g of NO_3^{1-} = 1 mole of NO_3^{1-} = 6.02×10^{23} ions of NO_3^{1-}

One can calculate the number of moles by dividing the number of particles by Avogadro's number.

$$\text{Number of moles} = \frac{\text{No. of particles of a substance}}{\text{Avogadro's Number}}$$

The relationship between amounts of substances in terms of their moles and the number of particles (atoms, molecules, ions, electrons or particles) is given below:

$$\text{Number of particles} = \frac{\text{Mass of subtance}\times N_A}{\text{Molar Mass}}$$



Sample Problem 4.4

A sample of glucose, contains 3.76×10^{24} molecules of glucose. What is the number of moles in this quantity?

Solution:

$$\text{No. of moles of glucose} = \frac{3.76\times10^{24}\text{ molecules}}{6.02\times10^{23}\text{ molecules mol}^{-1}} = 6.24\text{ moles}$$

Sample Problem 4.5

How many atoms are there in a sodium metal that contains 2.3 g?

Solution:

$$\text{Number of moles of sodium} = \frac{2.3}{23.0} = 0.1\text{ mol}$$

$$\text{Number of atoms of sodium} = \text{Number of moles of sodium}\times N_A$$

$$= 0.1\times6.02\times10^{23}$$

$$= 0.602\times10^{23}\text{ atoms}$$

$$= 6.022\times10^{22}\text{ g}$$

Sample Problem 4.6

Juglone, is a dye and is produced from the husks of black walnuts. The formula for juglone is $C_{10}H_6O_3$.

a) Calculate the molar mass of juglone.

b) Calculate number of moles in 0.87 g of a sample of juglone extracted from black walnut husks.

Interesting Information!

Juglone, is a natural herbicide (weed killer). It kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants.

Solution:

a) $C_{10}H_6O_3$

$$-10\times A_r(C)+6\times1.0\,A_r(H)+3\times A_r(O)(10\times12.0)+(6\times1.0)+(3\times16.0)120+6+48=174\text{ g/mol}$$

Mass of 1 mol of $C_{10}H_6O_3 = 174\,g\,mol^{-1}$

b) $$\text{Moles of juglone} = \frac{\text{Mass}}{\text{Molar Mass}} = \frac{0.87\text{ g}}{174\text{ g mol}^{-1}} = 0.005\text{ mol}$$

QUICK CHECK 4.2

a) A copper wire contains 27.10×10^{25} atoms of copper. Calculate the number of moles of copper.

Solution: $$\text{Number of moles of Copper} = \frac{27.10\times10^{25}\text{ atoms}}{6.02\times10^{23}\text{ atoms/mol}} = 4.50\times10^{2}\text{mol}$$

b) Calculate the molecules of 1×10^{-6} g of isopentyl acetate, $C_7H_{14}O_2$ which are released in a typical bee sting. How many atoms of carbon, hydrogen and oxygen are present in it?

Solution: Mass of isopentyl acetate = m = 1×10^{-6} g

Molecular mass of isopentyl acetate = M = $C_7H_{14}O_2$

= (12 × 7) + (14 × 1) + (16 × 2)

= 84 + 14 + 32

= 130 g/mol

$$\text{No. of molecules} = N = \frac{m}{M}\times N_A = \frac{1\times10^{-6}}{130}\times6.02\times10^{23}$$

$$= 46.29\times10^{14} = 4.629\times10^{15}\text{ molecules}$$

Number of C-atoms = $4.629\times10^{15}\times7 = 3.2382\times10^{16}$

Number of H-atoms = $4.629\times10^{15}\times14 = 6.4764\times10^{16}$

Number of O-atoms = $4.629\times10^{15}\times2 = 9.252\times10^{15}$

> **Interesting Information!**
>
> Isopentyl acetate $(C_7H_{14}O_2)$ is the compound responsible for the scent of bananas. Interestingly, bees release about $1\mu g (1\times10^{-6} g)$ of this compound when they sting. The resulting scent attracts other bees to join the attack.

MOLAR VOLUME (V_m)

"The volume of one mole of an ideal gas at STP (Standard temperature and pressure) is called molar volume".

- Its value is equal to $22.414 dm^3$.
- The value of molar volume is commonly rounded to $22.4 dm^3$.
- It is denoted by V_m. By using molar volume relationship, mass or mole of a gas at STP can be converted into volume, and vice versa.

> **Rack Your Mind!**
>
> 3. The volume of one mole of an ideal gas at STP is equal to:
>   A) $1 dm^3$ B) $22.4 dm^3$
>   C) $273 dm^3$ D) $6.02\times10^{23} dm^3$

According to Avogadro's law,

"Equal volumes of all ideal gases at the same temperature and pressure contain equal numbers of molecules".

This statement is indirectly the same when we say that one mole of an ideal gas at 273.16 K and one atm pressure has a volume of $22.414 dm^3$. Since one mole of a gas has Avogadro's number of particles, so $22.414 dm^3$ of various ideal gases at STP will have Avogadro's number of molecules i.e., 6.02×10^{23}. $22.4 dm^3$ of a gas at STP = Molar mass of a gas = 6.02×10^{23} particles of a gas = 1 mole of a gas

> **Smart Thinking**
>
> **Interesting Facts:**
> The molar volume of gas is $24 dm^3$ at RTP (Room temperature and pressure)

- $22.4 dm^3$ of CO_2 at STP = 44.0 g of $CO_2 = 6.02\times10^{23}$ molecules of CO_2 = 1 mole of CO_2
- $22.4 dm^3$ of any gas at STP = molar mass in grams $= 6.02\times10^{23}$ molecules = 1 mole
- $22.4 dm^3$ of H_2 gas at STP = 2 g $= 6.02\times10^{23}$ molecules = 1 mole
- $22.4 dm^3$ of NH_3 gas at STP = 17 g $= 6.02\times10^{23}$ molecules = 1 mole

If the number of moles of a gas is known, one can calculate the volume of a gas by multiplying number of moles of the gas by molar volume.

$$\text{Volume of a gas} = \text{Number of moles} \times \text{Molar volume}$$
$$V = n \times V_m$$

Sample Problem 4.7

Determine the volume of 2.5 moles of chlorine molecules at STP.

Solution:

The formula for volume determination at STP

$$V = n \times V_m$$

2.5 mole of Cl_2 occupy a volume $= 22.4 dm^3 \times 2.5 = 56.0 dm^3$

> **Rack Your Mind!**
>
> 4. 2 g of H_2, 16 g of CH_4 and 44 g of CO_2 occupy separately the volumes of $22.414 dm^3$, although the sizes and masses of three gases are very different from each other.

Sample Problem 4.8

What is the volume in dm^3 of 4.75 mol of methane (CH_4) gas at STP?

Solution:

The formula for volume determination at STP

$$V = n \times V_m$$

Volume of methane in dm^3 at STP $= 4.75 \times 22.4 = 106.4 dm^3$

MOLAR MASS AND DENSITY OF GASES

Density is defined as the mass per unit volume of a substance.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$
$$d = \frac{m}{V}$$

As molar mass of all the gases occupies same volume at STP, therefore, density of a gas depends on its molar mass. A gas having higher molar mass will have higher density and vice versa. If the density of gas at STP is determined, its molar mass can be calculated.

> **Rack Your Mind!**
>
> 5. How is the density of a gas related to its molar mass at STP?

Sample Problem 4.9

Calculate the molar mass of a gas which has density of $1.97 g/dm^3$ at STP.

Solution

$$m = d \times V$$
$$\text{Mass of gas at STP} = 1.97 \times 22.4 = 44.1 g mol^{-1}$$

> **QUICK CHECK 4.3**
>
> **Q.** Calculate the molar mass of a gas which has density of $1.34 g/dm^3$ at STP.
>
> **Ans.** $m = d \times V$
> $m = 1.34 \times 22.4 = 30.016 g mol^{-1}$

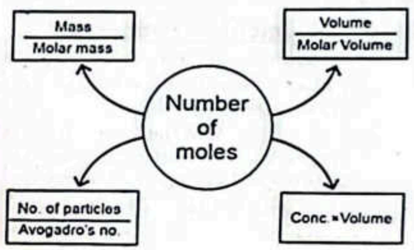
MOLAR CONCENTRATION

Molar concentration of solutions is given as mol/dm^3, which is the number of moles of a substance (reactant or product) dissolved per volume of a solution in dm^3.

The relationship between number of moles and molar concentration is given by

$$n = C \times V$$
$$C = \frac{n}{V}$$

where C is the molar concentration and V is the volume of the solution.



$$\text{Molar Concentration (C)} = \frac{\text{Number of moles (n)}}{\text{Volume in } dm^3 \text{ (V)}}$$

Example: If n = 0.2 mol and volume of solution = V = $1 dm^3$ then,

$$C = \frac{0.2 \text{ mol}}{1 dm^3} = 0.2 \text{ mol } dm^{-3}$$

Sample Problem 4.10

Calculate the molar concentration of a substance containing 27.64 g of K_2CO_3 dissolved in $1 dm^3$ of the given solution.

Solution:

Mass of K_2CO_3 = 27.64 g

Molar mass of K_2CO_3 = $138.2 g mol^{-1}$

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{Molar mass}}$$
$$n = \frac{m}{M}$$

$$n = \frac{27.64}{138.2} = 0.2\ \text{mol}$$

Volume of Solution $= 1\text{dm}^3$

$$\text{Molar Concentration} = \frac{\text{Number of moles}}{\text{Volume in dm}^3}$$

$$C = \frac{0.2\text{mol}}{1\text{dm}^3} = 0.2\ \text{moldm}^{-3}$$

QUICK CHECK 4.4

Calculate the molar concentration of a solution containing 7.9 g of $KMnO_4$ dissolved in 1dm^3 of the given solution. The molar mass of $KMnO_4$ is 158gmol^{-1}.

Solution:

No. of moles of $KMnO_4$ $= \dfrac{7.9\text{g}}{158\text{gmol}^{-1}} = \dfrac{1}{20}\text{mol}$

$n = 0.05\ \text{mol}$

Now, $C = \dfrac{n}{V} = \dfrac{0.05\text{mol}}{1\text{dm}^3}$

$C = 0.05\ \text{mol dm}^{-3}$

Alternate Method:

$$C = \frac{m}{M_r} \times \frac{1}{V(\text{dm}^3)}$$

$$= \frac{7.9}{158} \times \frac{1}{1}$$

$$= 0.05\text{moldm}^{-3}$$

STOICHIOMETRIC RELATIONSHIPS

Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.

With the knowledge of mole, Avogadro's number, molar mass, molar volume, and molar concentration, we can establish quantitative relationships between reactants and products using the balanced chemical equations.

Stoichion (element) and metron (measure). He defined it as 'the art of chemical measurements aming to express chemical changes in mathematical from. He is referred to as father of stoichiometry.

➲ Stoichiometric Relationships

The following types of relationship can be studied with the help of a balanced chemical equation involving quantities of reactant(s) and product(s).

I. Mole-Mole Relationship
II. Mass-mass relationship
III. Volume-Volume Relationship
IV. Mole-Mass Relationship
V. Mole-Volume Relationship
VI. Mass-Volume Relationship

Explanation:

Formation of Ammonia

To understand these relationships, we need to interpret information hidden in a balanced chemical equation which is used to make stoichiometric calculations. For example:

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

This equation can be described in three different ways;

I. 1 mole of N_2 reacts with 3 moles of H_2 to form 2 moles of NH_3.

II. 22.4dm^3 of N_2 reacts with 67.2dm^3 of H_2 to form 44.8dm^3 of NH_3.

III. 28.0 g of N_2 react with 6 g of H_2 to form 34.0 g of NH_3.

Rack Your Mind!

6. **Stoichiometry is fundamentally concerned with:**
   A) The speed at which chemical reaction occur.
   B) The energy changes involved in chemical reactions.
   C) The quantitative relationship between substances participating in a balanced chemical equation.
   D) The structure and bonding of chemical compound.

Rack Your Mind!

7. **All of the following are common types of stoichiometric calculations except for:**
   A) Determining the number of moles from given mass of reactant.
   B) Calculating the volume of gaseous reactant needed to completely react with another gaseous reactant.
   C) Predicting the time required for a reaction to reach equilibrium.
   D) Converting the mass of one reactant to the mass of product.

KEEP IN MIND

Important Assumptions of Stoichiometry

The following assumptions must be made while performing stoichiometric calculations:

1. All the reactants are completely converted into the products.
2. Law of conservation of mass and law of definite proportions are obeyed.
   No side reaction occurs.

➲ Approach to do Stoichiometric Calculations

Mass of known solid or volume of a known gas, or molar concentration of known solution

↓

Calculate number of moles of known solid or volume of a known gas, or molar concentration of known solution using the relevant formula

↓

Calculate number of moles of known solid or volume of a known gas, or molar concentration of known solution using the relevant formula

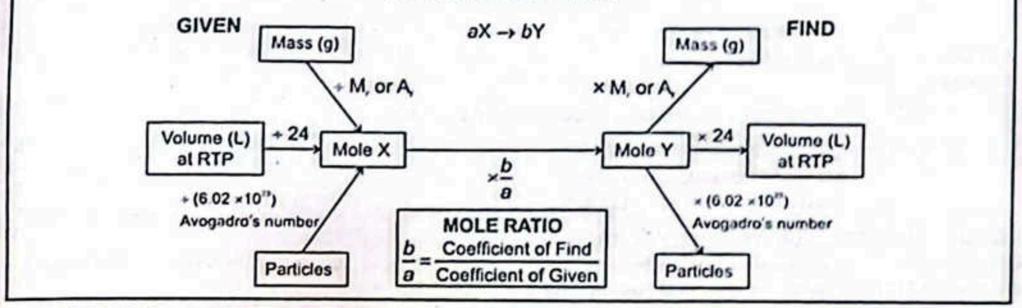
↓

Find the ratio of the known and the unknown reactant or product from the balanced chemical equation

↓

Convert the number of moles of the unknown to mass, volume, or concentration of the substance

STOICHIOMETRY MAP



Sample Problem 4.11 (Mole-Mole Conversion)

When 3.3 moles of nitrogen reacts with hydrogen to form ammonia, how many moles of hydrogen are consumed in the process? The equation for this reaction is

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

Solution:

Number of moles of $N_2 = 3.3$ mol

Number of moles of $H_2 = ?$

1 mole of N_2 needs H_2 to produce $NH_3 = 3$ mol

3.3 moles of N_2 needs H_2 to produce $NH_3 = 3 \times 3.3 = 9.9$ mol

QUICK CHECK 4.5

How many moles of carbon dioxide are produced when 2.25 moles of glucose are used by a person? The oxygen is in excess. The equation for the reaction is:

$$C_6H_{12}O_{6(s)} + 6O_{2(g)} \longrightarrow 6CO_{2(g)} + 6H_2O_{(\ell)}$$

Solution:

1 mole of $C_6H_{12}O_6$ produces CO_2 = 6 moles

2.25 moles of $C_6H_{12}O_6$ produces CO_2 = $2.25 \times 6 = 13.5$ moles

Sample Problem 4.12 (Mass-Mass Conversion)

Calculate the mass of Al needed to react completely with 32.0 g of iron (III) oxide according to the equation given below:

$$2A\ell_{(s)} + Fe_2O_{3(s)} \longrightarrow A\ell_2O_{3(s)} + 2Fe_{(s)}$$

Solution:

Molar mass of Fe_2O_3, M = $160\ g\ mol^{-1}$

Number of moles of Fe_2O_3 (n) = $\frac{m}{M} = \frac{32.0\ g}{160\ g\ mol^{-1}} = 0.2$ mole

From the balanced equation, 1 mol of Fe_2O_3 reacts with 2 moles of $A\ell$, therefore, number of moles of $A\ell$ that reacts with 0.02 mole of $Fe_2O_3 = 2 \times 0.2 = 0.4$ mol

Mass of $A\ell$ = n × M

= $0.4\ mol \times 27\ g\ mol^{-1} = 10.8\ g$

QUICK CHECK 4.6

Fe_2O_3, an ore of iron is called Hematite. CO can reduce it to get free Fe as below:

$$Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

How much Fe can be produced from 160 g of Fe_2O_3?

Given Data: Molar mass of Fe_2O_3 = $(56 \times 2) + (16 \times 3)$

= 160 g/mol

mass of Fe_2O_3 = 160 g

To Find: Mass of Fe produced = ?

Equation:

$$Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

Mole of Fe_2O_3 produces moles of Fe = 2

Mass of Fe produced = 2×56

Mass of Fe produced = 112 g

Sample Problem 4.13 (Volume-Volume)

Calculate volume of ammonia that can be produced by the reaction of 100 dm³ of hydrogen with excess of nitrogen at STP? The balanced chemical equation for the reaction is:

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

Solution: Volume of hydrogen = $100\ dm^3$

Volume of ammonia = ?

$67.2 dm^3$ (3 mol) of H_2 produce ammonia = $44.8\ dm^3$ (2 mol)

$1\ dm^3$ of H_2 produce ammonia = $\frac{44.8}{67.2} = \frac{2}{3}$

$100\ dm^3$ of H_2 produce ammonia = $\frac{2}{3} \times 100 = 66.7\ dm^3$

So, the volume of ammonia produced by the reaction of $100 dm^3$ of H_2 with excess nitrogen is $66.7 dm^3$.

QUICK CHECK 4.7

Calculate the volume of carbon dioxide produced at STP when $4.5 dm^3$ of methane is burnt by a person. The oxygen is in excess. The equation for the reaction is:

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(\ell)}$$

Solution:

$22.44\ dm^3$ of CH_4 produces = $22.414\ dm^3$ of CO_2

For gaseous reaction mole ratio is equal to volume. Ratio.

One mole CH_4 produce one mole CO_2. SO, $4.5\ dm^3$ will produce $4.5\ dm^3$ of CO_2.

Sample Problem 4.14 Mole-Mass Calculations

Solid lithium hydroxide LiOH is used in space vehicles. It is employed to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. Calculate the mass of Li_2CO_3 that can be produced by 20.0 mol of LiOH.

$$2LiOH_{(s)} + CO_{2(g)} \longrightarrow Li_2CO_{3(s)} + H_2O_{(\ell)}$$

Solution:

According to the given balanced chemical equation,

2 moles of LiOH produces = 1 mole Li_2CO_3

20.0 moles of LiOH produces = $\frac{1}{2} \times 20.0 = 10.0$ mol Li_2CO_3

Mass of Li_2CO_3 produced = Number of mole × Molar mass

Mass of Li_2CO_3 produced = $10.0\ mol \times 73.9\ g\ mol^{-1} = 739.0\ g$

Thus, 739.0 g Li_2CO_3 will be produced from 20.0 moles of LiOH.

QUICK CHECK 4.8

Calculate the mass of sodium hypochlorite ($NaOC\ell$), a household bleach, produced by the reaction of 2.25 moles of chlorine with excess sodium hydroxide. The balanced equation is

$$2NaOH_{(aq)} + C\ell_{2(g)} \longrightarrow NaOC\ell_{(aq)} + NaC\ell_{(aq)} + H_2O_{(\ell)}$$

Solution:

Because NaOH is in excess so $C\ell_2$ is limiting reactant. One mole $C\ell_2$ produce one mole of $NaOC\ell$. So, 2.25 mole of $C\ell_2$ will produce 2.25 mole of $NaOC\ell$.

Molar mass of $NaOC\ell$ = $23 + 16 + 35.5 = 74.5$ g

Mass of $NaOC\ell$ = $2.25 \times 74.5 = 167.625$ g

Sample Problem 4.15 (Mass-mole calculations)

Baking soda ($NaHCO_3$) acts as an antacid. It can neutralize excess hydrochloric acid ($HC\ell$) secreted by the stomach according to equation.

$$NaHCO_{3(s)} + HC\ell_{(aq)} \longrightarrow NaC\ell_{(aq)} + H_2O_{(\ell)} + CO_{2(aq)}$$

How many moles of $HC\ell$ will be neutralized by 2.1 g of baking soda?

Solution:

Molar mass of $NaHCO_3$ = $84.0\ g\ mol^{-1}$

Moles of $NaHCO_3$ = $\frac{2.1\ g}{84.0\ g\ mol^{-1}} = 0.025$ mol

Stoichiometrically, the mole ratio of $HC\ell$ and $NaHCO_3$ is 1:1.

Hence moles of $HC\ell$ used = 0.025 mol

Thus 2.1 g of $NaHCO_3$ will neutralize 0.025 moles of $HC\ell$.

Rack Your Mind!

8. Why is it important to convert between mass and moles in stoichiometric calculations?

Sample Problem 4.16 (Mass-Volume Conversion)

What volume of hydrogen at STP will be produced when 7.0 g of iron are reacted with an excess of sulphuric acid?

$$Fe_{(s)} + H_2SO_{4(aq)} \longrightarrow FeSO_{4(aq)} + H_{2(g)}$$

Solution:

Molar mass of Fe (M) = 55.8 g/mol

Number of moles of iron (n) $= \frac{m}{M}$

$= \frac{7.0\,g}{55.8\,g\,mol^{-1}}$

$= 0.125$ mol

From the balanced equation, 1 mol of iron produces 1 mole of hydrogen.

So, number of moles of $H_2 = 0.125$ mol

Volume of H_2 in dm^3 = molar volume x moles of H_2

$= 22.4\,dm^3\,mol^{-1} \times 0.125\,mol$

$= 2.8\,dm^3$

LIMITING AND EXCESS REACTANT

(Exercise LQ)

Q. Differentiate limiting and non-limiting reactants. How a limiting reactant is determined from a balanced chemical equation and given data?

In many chemical processes, the quantities of the reactants are usually not present in the proportions indicated by the balanced chemical equation. Frequently, a large amount of inexpensive reactant is supplied because of the following reasons:

a. ***To ensure that whole of the mass of the expensive reactant is completely converted to the desired product.***

b. ***To produce maximum amount of product.***

c. ***To increase the rate of reaction.***

We know that a large quantity of oxygen in a chemical reaction makes things burn more rapidly. In this way, excess of oxygen is left behind at the end of reaction and the other reactant, i.e. fuel, is consumed earlier.

Definition 1:

"This reactant which is consumed earlier is called the limiting reactant". In this way, the amount of product that forms is limited by the reactant that is completely used. Once this reactant is consumed, the reaction stops and no additional product is formed.

Definition 2:

"The reactant which controls the amounts of products formed in a chemical reaction and is consumed earlier is called the limiting reactant or reagent".

The maximum amount of the product formed depends upon the amount of limiting reactant in the reaction mixture.

Properties of Limiting Reactant:

i. Limiting reactant controls the amount of product formed.
ii. Limiting reactant is taken in lesser amount.
iii. Limiting reactant is consumed earlier.
iv. Limiting reactant produces least amount of product.

Rack Your Mind!

9. Why does a chemical reaction stop even when one reactant is still present?

Did You Know?

Fire is a combustion reaction in which fuel and oxygen, O_2, combine, usually at high temperatures, to form water and carbon dioxide. Once the fire has started, it is self-supporting. An effective way to quench a fire is smothering, which reduces the amount of available oxygen below the level needed to support combustion. In other words, smothering decreases the amount of the excess reactant. Foams, inert gas, and CO_2 are effective substances for smothering.

➲ Strategy for the Identification of Limiting Reactant

To identify a limiting reactant, the following three steps are performed.

i. Calculate the number of moles from the given amounts of reactants.
ii. Find out the number of moles of product with the help of a balanced chemical equation.
iii. Identify the reactant which produces the least amount of product as limiting reactant and the other as an excess reactant.

Limiting Reactant	Non-Limiting (Excess) Reactant
1. The reactant used up first in a chemical reaction is called limiting reactant. 2. It controls the amount of product. 3. It is not present at the end of reaction. 4. It limits the amount of product formed from chemical reaction.	1. The reactant which left unreacted after completion of a chemical reaction is called excess reactant. 2. It does not control the amount of product. 3. Some amount is present at the end of reaction. 4. It has no effect on product formed from reaction.

Following numerical problem will make the idea clear.

Sample Problem 17 (Limiting Reactant)

Calculate the mass of N_2 produced from 18.1 g of NH_3 (molar mass = $17.0\,g\,mol^{-1}$) and 90.4 g of CuO (molar mass = $79.5\,g\,mol^{-1}$) according to following balanced equation:

$$2NH_{3(g)} + 3CuO_{(s)} \longrightarrow N_{2(g)} + 3Cu_{(s)} + 3H_2O_{(g)}$$

Solution:

Moles of $NH_3 = 18.1$ g of $NH_3 / 17.0\,g\,mol^{-1} = 1.06$ mol

Moles of CuO = 90.4 g of CuO / $79.5\,g\,mol^{-1} = 1.14$ mol

In balanced equation,

CuO : N_2
3 : 1
$1.14 : \frac{1}{3} \times 1.14 = 0.38$ mol

NH_3 : N_2
2 : 1
$1 : \frac{1}{2}$
$1.06 : \frac{1}{2} \times 1.06 = 0.53$ mol

Thus, CuO is the limiting reactant and the number of moles of N_2 produced will be 0.38 mol.

Hence, mass of N_2 produced = n × M
= 0.38 × 28
mass of N_2 produced = 10.64 g

Rack Your Mind!

10. For the reaction: $N_2 + 3H_2 \rightarrow 2NH_3$, if 2 moles of N_2 react with 4 moles of H_2, how many moles of NH_3 are produced?
A) 2 moles B) 4 moles
C) 2.67 moles D) 1.33 moles

Sample Problem 4.18 (Limiting Reactant)

When aqueous solutions of Na_2SO_4 and $Pb(NO_3)_2$ are mixed, $PbSO_4$ precipitates down. Calculate the mass of $PbSO_4$ formed when $1.25\,dm^3$ of $0.05\,mol\,dm^{-3}$ $Pb(NO_3)_2$ and $2.00\,dm^3$ of $0.025\,mol\,dm^{-3}$ Na_2SO_4 are mixed.

$$Na_2SO_{4(aq)} + Pb(NO_3)_{2(aq)} \longrightarrow PbSO_{4(s)} + NaNO_{3(aq)}$$

Solution:

The net ionic equation is

$$Pb^{+2}_{(aq)} + SO^{-2}_{4(aq)} \longrightarrow PbSO_{4(s)}$$

Since $0.05\,mol\,dm^{-3}$ $Pb(NO_3)_2$ contains $0.05\,mol\,dm^{-3}$ Pb^{2+} ions.

The formula: No. of moles = Concentration ($mol\,dm^{-3}$)×(Volume dm^3)

n = CV

moles of Pb^{2+} ions $= 0.05\,mol\,dm^{-3} \times 1.25\,dm^3 = 0.0625$ mol

Thus, $0.025\,mol\,dm^{-3}$ Na_2SO_4 solution contains $0.025\,mol\,dm^{-3}$ SO_4^{-2} ions.

moles of SO_4^{-2} ions $= 0.025\,mol\,dm^{-3} \times 2.00\,dm^3 = 0.05$ mol

As Pb^{2+} and SO_4^{-2} react in a 1:1 ratio, here, SO_4^{-2} (0.05 mol) will be consumed earlier than Pb^{2+} (0.0625 mol). The amount of SO_4^{-2} will be limiting. The reason is that 0.05 mole of SO_4^{-2} is less than 0.0625 mole of Pb^{2+}. Since the Pb^{2+} ions are present in excess, only 0.05 mole of solid $PbSO_4$ will be formed. The mass of $PbSO_4$ formed can be calculated using the molar mass of $PbSO_4$ (303.3 g mol^{-1}):

$$\text{Mass of } PbSO_4 = 0.05\ mol \times 303.3\ g\ mol^{-1} = 15.2\ g$$

Maximum Amount of Product and Amount of Any Unreacted Excess Reagent

Excess Reactant:

"The reactants which are in larger amounts (according to stoichiometry of reaction) and remain unreacted at the end of the reaction are called "excess reagents" (or excess reactants)".

OR

"The reactant that is left unreacted after the completion of reaction is called non-limiting or excess reactant".

Consider the reaction between hydrogen and oxygen to form water.

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

- When we take 2 moles of hydrogen (4 g) and allow it to react with 2 moles of oxygen (64 g), then we will get only 2 moles (36 g) of water.
- **Reason:** It is because 2 moles (4 g) of hydrogen react with 1 mole (32 g) of oxygen according to the balanced equation. Since less hydrogen is present as compared to oxygen, so hydrogen is a limiting reactant.
- When 1 mole of O_2 and 1 mole of H_2 are mixed, all the H_2 will react completely and O_2 will be left unreacted because for 1 mole of H_2, $\frac{1}{2}$ mole of O_2 is required. The remaining $\frac{1}{2}$ mole of O_2 will be excess and H_2 is a limiting reactant.

Sample Problem 4.19 (Excess Reactant)

Natural gas consists primarily of methane (CH_4). The complete combustion of methane (CH_4) gives carbon dioxide (CO_2) and water.

$$CH_{4(g)} + 2O_{2(g)} \xrightarrow{\text{Ignition}} CO_{2(g)} + 2H_2O_{(g)}$$

a) How many grams of CO_2 can be produced when 30 g of CH_4 and 50 g of O_2 are allowed to combine?

b) How many grams of excess reagent are left unreacted after the completion of reaction?

Solution (a):

Step 1: Write balanced chemical equation.

Step 2: Convert the given mass of both the reactants into their moles.

$$\text{Moles of } CH_4 = \frac{\text{given mass of } CH_4}{\text{molar mass of } CH_4} = \frac{30\ g}{16 g mol^{-1}} = 1.875\ mol$$

$$\text{Moles of } O_2 = \frac{\text{given mass of } O_2}{\text{molar mass of } O_2} = \frac{50\ g}{32 g mol^{-1}} = 1.563\ mol$$

Step 3: Calculate the number of moles of product from each reactant.

Compare the number of moles of CH_4 with those of CO_2. From the balanced chemical equation.

1 mole of methane produces $CO_2 = 1$ mol

1.875 mole of methane produces $CO_2 = 1 \times 1.875\ mol = 1.875$ mol of CO_2

Compare the number of moles of O_2 with those of CO_2. From the balanced chemical equation, we know:

2 moles of oxygen produces $CO_2 = 1$ mol

1.563 moles of oxygen produce $CO_2 = 0.5 \times 1.563\ mol = 0.7815$ moles of CO_2

From the above calculation, it is clear that the limiting reactant is O_2 because it produces lesser amount (moles) of product CO_2 than CH_4.

Step 4: Convert the moles of the product into mass.

$$\text{Mass of } CO_2 \text{ in grams} = \text{Moles of } CO_2 \times \text{Molar mass of } CO_2$$
$$= 0.7815 moles \times 44\ g\ mol^{-1} = 34.39\ g$$

Step 5: The quantity of limiting reactant can also be used to calculate the quantity of excess reactant used.

2 moles of O_2 reacts with moles of $CH_4 = 1$ mol

1.563 moles of O_2 reacts with moles of $CH_4 = \frac{1}{2} \times 1.563\ mol = 0.7815$ mol

Step 6: The mass of methane (excess reagent) is equal to the starting quantity minus the amount used during the reaction.

$$\text{Number of moles of } CH_4 = \text{Quantity taken} - \text{Quantity used}$$
$$= 1.875\ mol - 0.7815\ mol = 1.0935\ mol$$

$$\textbf{Mass of } CH_4 \textbf{ (excess reagent)} = 1.0935 \times 16.0 = 17.5\ g$$

☑ QUICK CHECK 4.9

Which of the following reaction mixtures could produce the greatest amount of product when they combine according to the reaction given below?

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

a) 1 mole of N_2 and 3 moles of H_2
Ans. 2 moles of NH_3

b) 2 moles of N_2 and 3 moles of H_2
Ans. 2 moles of NH_3

c) 1 mole of N_2 and 5 moles of H_2
Ans. 2 moles of NH_3

d) 3 moles of N_2 and 3 moles of H_2
Ans. 2 moles of NH_3

e) Each produces the same amount of product.
Ans. Yes, This is the correct answer.

THEORETICAL YIELD AND ACTUAL YIELD

(Exercise L.Q.4)

Q. Differentiate actual and theoretical yields. Why the theoretical yield is always greater than actual yield?

Actual Yield:

"The amount of the products obtained in a chemical reaction is called the actual yield of that reaction".

Theoretical Yield:

"The amount of the products calculated from the balanced chemical equation represents the theoretical yield". The theoretical yield is the maximum amount of the product that can be produced by a given amount of a reactant, according to balanced chemical equation.

NOTE: In most chemical reactions the ***amount of the product obtained is less than the theoretical yield.***

Theoretical Yield is always Greater than Actual Yield

Reasons:

There are following reasons for that:

- The processes like filtration, separation by distillation, separation by a separating funnel, washing, drying and crystallization, if not properly carried out, decrease the actual yield.
- Some of the reactants might take part in a competing side reaction and reduce the amount of the desired product. So, in most of the reactions the actual yield is less than the theoretical yield.
- A reaction may be reversible. Therefore, the amount of the product will be reduced by the backward reaction.

A chemist is usually interested in the efficiency of a reaction. The efficiency of a reaction is expressed by comparing the actual and theoretical yields in the form of percentage (%) yield.

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

Percentage Yield & Efficiency of the Reaction:

Greater the % age yield, higher will be the efficiency of reaction and vice versa.

Differentiate Between Theoretical and Actual Yield

Theoretical Yield	Actual Yield
1. The amount of product calculated from a balanced chemical equation is called **theoretical yield.**	1. The amount of product obtained while performing a chemical reaction is called **actual yield.**
2. It is also called expected yield or calculated yield.	2. It is also called experimental yield.
3. Theoretical yield is always greater than actual yield.	3. Actual yield is always lesser than theoretical yield.
4. There is no need to perform experiment. It is easily calculated from balanced chemical equation.	4. In order to get actual yield experiment has to be performed.

Sample Problem 4.20 (% age Yield)

Aspirin ($C_9H_8O_4$) is prepared by heating salicylic acid, $C_7H_6O_3$ (molar mass 138.12 $gmol^{-1}$) and acetic anhydride, $C_4H_6O_3$ (molar mass 102.09 $gmol^{-1}$).

$$C_7H_6O_3 + C_4H_6O_3 \longrightarrow C_9H_8O_4 + CH_3COOH$$

Calculate the theoretical yield of aspirin, (molar mass = 180.16 $gmol^{-1}$) when 3.00 g of salicylic acid is heated with 6.00 g of $(CH_3CO)_2O$. What is % yield when actual yield is 3.15 g?

Solution:

$$\underset{\substack{\text{Salicylic acid}\\ 1\text{ mol}\\ 138.12\text{g}}}{C_7H_6O_3} + \underset{\substack{\text{Acetic anhydride}\\ 1\text{ mol}\\ 102.09\text{g}}}{C_4H_6O_3} \longrightarrow \underset{\substack{\text{Aspirin}\\ 1\text{ mol}\\ 180.16\text{g}}}{C_9H_8O_4} + CH_3COOH$$

1 mole of Salicylic acid produces aspirin = 1 mol

Mass of Salicylic acid = 3.00 g

$$\text{Number of moles of Salicylic acid} = \frac{3.00\text{ g}}{138.12\text{ g mol}^{-1}} = 0.022\text{ mol}$$

Mass of Salicylic acid = 6.00 g

$$\text{Number of moles of Salicylic acid} = \frac{6.00\text{ g}}{102.09\text{g mol}^{-1}} = 0.058\text{ mol}$$

Here, salicylic acid is limiting reactant while acetic anhydride is an excess reactant. The amount of salicylic controls the yield of product i.e., aspirin.

0.022 mol Salicylic acid produces Aspirin = 0.022 mol

Mass of Aspirin = 0.022 mol × 180.16 g mol^{-1} = 3.96 g

Theoretical yield = 3.96 g

Actual yield = 3.15 g

$$\% \text{ age yield} = \frac{3.15}{3.96} \times 100$$

$$= \boxed{79.55\%}$$

QUICK CHECK 4.10

When limestone ($CaCO_3$) is roasted, quicklime (CaO) is produced according to the following equation.

$$CaCO_3 \longrightarrow CaO_{(s)} + CO_{2(g)}$$

The actual yield of CaO is 2.5 kg, when 4.5 kg of limestone is roasted. What is the percentage yield of this reaction?

Given Data:

Mass of limestone roasted = 4.5 kg = 4.5 × 1000 = 4500 g

Mass of quick lime (actual yield) = 2.5 kg = 2.5 × 1000 = 2500 g

Molar Mass of $CaCO_3$ = 40 + 12 + 16 × 3

= 40 + 12 + 48 = 100 g mol^{-1}

Molar mass of CaO = 40 + 16 = 56 g mol^{-1}

Required: Percentage yield of CaO = ?

Solution:

Equation: $CaCO_{3(s)} \longrightarrow CaO_{(s)} + CO_{2(g)}$

According to balanced chemical equation

100g of $CaCO_3$ give CaO = 56g

$$1\text{g of } CaCO_3 \text{ gives CaO} = \frac{56}{100}\text{g}$$

$$4500\text{g of } CaCO_3 \text{ give CaO} = \frac{56}{100} \times 4500$$

Theoretical yield of CaO = 2520g

Actual yield of CaO = 2500g

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{2500}{2520} \times 100$$

% yield of CaO = $\boxed{99.2\%}$

IMPORTANCE OF STOICHIOMETRY IN PRODUCTION AND DOSAGE OF MEDICINE

While preparing required dose of a medicine, the optimum amount of the active ingredient in a medicine is essential to produce desired effects in the patient. Stoichiometry ensures the accuracy of drug synthesis. Any deviation can result in incomplete reaction or contamination with un-reacted reactants or by-products. Stoichiometry allows chemists to precisely control chemical reactions to produce drugs, to ensure its efficiency, effectiveness and safe use.

➲ Significance of Stoichiometry in Medicine

Stoichiometry is very important in the field of medicine and is used to:

1. In the preparation of antibiotics, the stoichiometry ensures that each dose matches the active ingredient and target bacteria.
2. To determine the cholesterol level in the blood of patients. Cholesterol is a form of fat that is not all bad. However, cholesterol can have harmful effects.
3. To determine the glucose level in the blood of diabetic patient. Use of insulin relies on the stoichiometry to precise control of blood sugar levels.
4. To determine the steroid and other stimulants in the urine of athletes. Athletes use steroids and other stimulants to enhance performance and increase strength.
5. To determine the concentration of viral antigens in the preparation of vaccine for effective results.
6. To determine the amount and number of drugs to give a dosage to a patient. The medicine has no effect when given in small amounts and can cause toxic state or death when given in large amounts.

Example: Paracetamol is used as a pain killer and to decrease fever. An overdose may result a blood thinning, organ damage and severe liver damage.

Solution File — Rack Your Brain!

Sr. #	Option	Explanation
1.	C	O_2 gas = 32 grams = 1 mole of O_2 gas = 6.02×10^{23} molecule of O_2
2.	S.Q	Water (H–O–H) molecule has two bonds. So two moles of bonds are present in water. H_2O contains two moles of H-atoms and one mole of O-atoms so total 3 moles of atoms. H_2 = 2e, O = 8e. So H_2O contains 10 moles of electrons. H_2 = 2p, 2e = 4 O = 8p, 8e, 8n = 24 Total = 28 So, 1 mole of H_2O has 28 moles of fundamental particles.
3.	B	Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.
4.	S.Q	2g of H_2 = 1 mol H_2 = 6.02×10^{23} molecule = 22.414 dm^3 16 g of CH_4 = 1 mol CH_4 = 6.02×10^{23} molecules = 22.414 dm^3 44 g of CO_2 = 1 mol CO_2 = 6.02×10^{23} molecules = 22.414 dm^3 According to Avogadro's law, equal moles of gases at same temperature and pressure (STP) occupy same volume. One mole of any gas can occupy 22.414 dm^3 at STP. So 2g of H_2, 16g of CH_4 and 44g of CO_2 occupy 22.414 dm^3. Volume occupied by a gas does not depend on size and mass of gas molecule. It only depends on number of molecules. Reason is that at STP, distance between gas molecules is 300 times greater than their own diameters.
5.	S.Q	At STP, gases with higher molar masses have higher densities because all gases occupy the same volume per mole (22.4 dm^3). NOTE: Formula to calculate molar mass from density at STP: Molar Mass = Density × 22.4 dm^3
6.	C	The quantitative relationship between substances participating in a balanced chemical equation.
7.	C	Predicting the time required for a reaction to reach equilibrium.
8.	S.Q	The purpose is to convert between mass and moles allows us to use the mole ratio from the balanced equation to calculate quantities.
9.	S.Q	The reaction stops when the limiting reactant is completely used up. Even if another reactant remains, no more product can form without the limiting reactant.
10.	C	4 moles of H_2 will produce 8/3 (2.66) moles of NH_3 according to the stoichiometry of $N_2 + 3H_2 \rightarrow 2NH_3$, where H_2 is the limiting reagent.

Exercise

MULTIPLE CHOICE QUESTIONS (MCQs)

Q.1 Four choices are given for each question. Select the correct choice.

I. Which one of the following statements is incorrect?
a) One mole of nitrogen gas contains Avogadro's number of molecules
b) One mole of ozone gas contains Avogadro's number of molecules
c) One mole of ozone contains Avogadro's number of O atoms
d) One mole of hydrogen gas contains Avogadro's number of molecules

II. Which one of the following has greatest mass?
a) 0.5 mol of N_2 b) 0.5 mol of NH_3 c) 0.5 mole of He d) 0.5 mol of CO_2

III. Which one of the following gases will have greatest volume at STP?
a) 22 g of CO_2 b) 88 g of N_2O c) 28 g of CO d) 28 g of N_2

IV. Which of the following contains same number of particles as present in 12 g of carbon?
a) 28 g of iron (Atomic mass of Fe = 56) b) 48 g of magnesium (Atomic mass of Mg = 24)
c) 32 g of S_8 molecules (Atomic mass S = 32) d) 40 g of carbon monoxide (molar mass of CO = 28)

V. Volume at S.T.P. of 22 g of CO_2 is same as that of:
a) 2 g of hydrogen b) 8.5 g of NH_3 c) 64 g of gaseous SO_2 d) 7 g of CO

VI. 4.0 g of NaOH (molar mass 40 g mol^{-1}) contains same number of sodium ions as are present in:
a) 10.6 g of Na_2CO_3, (molar mass 106) b) 58.5 g of NaCl (molar mass 58.5)
c) 76 g Na_2SO_4 (formula mass 142) d) 85 g of $NaNO_3$ (molar mass 85)

VII. A container holds 0.5 moles of an ideal gas at STP. What is the volume of the gas in dm^3?
a) 11.2 dm^3 b) 22.4 dm^3 c) 44.8 dm^3 d) 12.2 dm^3

VIII. A solution contains 4.0 g of sodium hydroxide (NaOH, molar mass = 40.0 g mol^{-1}) in 250 cm^3 of solution. What is the molar concentration of this solution?
a) 0.10 mol dm^{-3} b) 0.20 mol dm^{-3} c) 0.40 mol dm^{-3} d) 0.80 mol dm^{-3}

IX. A solution contains 10.0 g of an unknown solute in 250 cm^3 of solution. If the molar concentration of the solution is 0.20 mol dm^{-3}, what is the molar mass of the solute?
a) 50 g mol^{-1} b) 100 g mol^{-1} c) 200 g mol^{-1} d) 400 g mol^{-1}

X. A sample of nitrogen gas (N_2, molar mass = 28.0 g mol^{-1}) has a mass of 14.0 g. How many nitrogen atoms are present in this sample?
a) 3.01×10^{23} atoms b) 6.02×10^{23} atoms c) 1.20×10^{24} atoms d) 2.40×10^{24} atoms

XI. A gas has a density of 1.43 gdm^{-3} at STP. What is the molar mass of the gas? (Molar volume at STP = 22.4 dm^3 mol^{-1})
a) 14.3 g mol^{-1} (b) 22.4 g mol^{-1} c) 32.0 g mol^{-1} (d) 64.0 g mol^{-1}

XII. A gas has a density of 1.96 gdm^{-3} at STP (0°C and 1.00 atm). What is its molecular mass? (Molar volume at STP = 22.4 dm^3 mol^{-1})
a) 11.2 g mol^{-1} b) 22.4 g mol^{-1} c) 44.0 g mol^{-1} d) 88.0 g mol^{-1}

Answer Key with Explanations

Sr.No.	Option	Answer	Explanation
I.	c	One mole of ozone contains Avogadro's number of O atoms	• 1 mole of O_3 contains Avogadro's number of O_3 molecules, not O atoms (which would be 3 × Avogadro's number).
II.	d	0.5 mol of CO_2	• Mass = moles × molar mass: ○ N_2: 0.5 × 28 = 14 g ○ NH_3: 0.5 × 17 = 8.5 g, He: 0.5 × 4 = 2 g, CO_2: 0.5 × 44 = 22 g (Highest).
III.	d	28 g of N_2	• At STP, volume ∝ moles: ○ 28 g N_2 = 1 mol (22.4 dm^3). ○ 22 g CO_2 = 0.5 mol (11.2 dm^3), 88 g N_2O = 2 mol (44.8 dm^3), 28 g CO = 1 mol (22.4 dm^3). • Correction: (b) 88 g of N_2O has maximum volume (2 mol = 44.8 dm^3).
IV.	d	28 g of carbon monoxide	• 12 g C = 1 mol C = 6.02×10^{23} atoms. • 28 g CO = 1 mol CO = 6.02×10^{23} molecules (each with 2 atoms ⇒ 1.2×10^{24} atoms). • Correction: (b) 48 g Mg = 2 mol Mg = 1.2×10^{24} atoms (matches C).
V.	b	8.5 g of NH_3	• 22 g CO_2 = 0.5 mol ⇒ 11.2 dm^3 at STP. • 8.5 g NH_3 = 0.5 mol ⇒ same volume.
VI.	a	10.6 g of Na_2CO_3	• 4 g NaOH = 0.1 mol ⇒ 0.1 mol Na^+ ions. • 10.6 g Na_2CO_3 = 0.1 mol ⇒ 0.2 mol Na^+ ions. • Correction: (b) 58.5 g NaCℓ = 1 mol ⇒ 1 mol Na^+ ions. • Actual answer: (a) 10.6 g Na_2CO_3 has 0.2 mol Na^+, but none match exactly.
VII.	a	11.2 dm^3	• At STP, 1 mol = 22.4 dm^3 ⇒ 0.5 mol = 11.2 dm^3.
VIII.	c	0.40 mol dm^{-3}	• Molarity = moles/volume = (4 g / 40 g/mol) / 0.25 dm^3 Molarity = 0.40 mol dm^{-3}
IX.	c	200 g mol^{-1}	• Moles = M × V = 0.2 × 0.25 = 0.05 mol. • Molar mass = mass/moles = 10 g / 0.05 mol = 200 g mol^{-1}.
X.	b	6.02×10^{23} atoms	• 14 g N_2 = 0.5 mol ⇒ 0.5 × 6.02 × 10^{23} molecules = 6.02×10^{23} N atoms (each N_2 has 2 atoms).
XI.	c	32.0 g mol^{-1}	• Molar mass = density × molar volume Molar mass = 1.43 g/dm^3 × 22.4 dm^3/mol = 32.0 g mol^{-1}.
XII.	c	44.0 g mol^{-1}	• Molar mass = 1.96 g/dm^3 × 22.4 dm^3/mol Molar mass = 44.0 g mol^{-1} (matches CO_2)

SHORT ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

a. How is the concept of mole derived from Avogadro's number?

Ans. Amedeo Avogadro proposed that:

"Equal volumes of gases, at the same temperature and pressure, contain equal numbers of molecules."

$$N_A = 6.02214076 \times 10^{23} \text{ particles per mole}$$

The mole (mol) was then defined as:

"The amount of substance containing exactly N_A elementary entities (atoms, molecules, ions, etc.)."

This means:

$$1 \text{ mol} = N_A \text{ particles} = 6.022 \times 10^{23} \text{ particles}$$

Derivation of Molar Mass

- The atomic mass unit (u) is defined based on ^{12}C (12 u = mass of one ^{12}C atom).
- Since N_A atoms of ^{12}C weigh 12 grams, the molar mass (mass per mole) is numerically equal to the atomic/molecular mass in u:

$$\text{Molar mass of } ^{12}C = 12 \text{ g/mol}$$

We can calculate the number of moles by dividing the number of particles by Avogadro's number.

b. Define the following terms with one example in each case.
(a) Molar mass (b) Molar volume (c) Molar concentration

Ans. (a) Molar Mass

The mass of one mole of a substance expressed in grams is called molar mass.

Molar mass of $CC\ell_4$ = Molar mass of one C + Molar mass of $C\ell \times 4$
Molar mass of $CC\ell_4$ = 12.0×1 + 35.5×4
Molar mass of $CC\ell_4$ = 12.0 + 142.0 = 154.0 g

(b) Molar VOLUME

The volume of one mole of an ideal gas at STP (Standard temperature and pressure) is called molar volume. Its value is equal to 22.414 dm^3. The value of molar volume is commonly rounded to 22.4 dm^3. It is denoted by V_m. By using molar volume relationship, mass or mole of a gas at STP can be converted into volume, and vice versa.

(c) Molar Concentration

Molar concentration of solutions is given as mol/dm^3, which is the number of moles of a substance (reactant or product) dissolved per volume of a solution in dm^3. The relationship between number of moles and molar concentration is given by

$$n = C \times V$$

$$C = \frac{n}{V}$$

where C is the molar concentration and V is the volume of the solution.

$$\text{Molar Concentration (C)} = \frac{\text{Number of moles (n)}}{\text{Volume in } dm^3 \text{ (V)}}$$

c. What do you mean by molar volume of a gas? How Avogadro's number is related with molar volume?

Ans. Molar Volume of a Gas:

The (molar volume) of a gas is the volume occupied by one mole of any ideal gas at standard temperature and pressure (STP).

- At STP (0°C or 273.15 K and 1 atm pressure),
- 1 mole of any ideal gas occupies 22.414 dm^3 (or liters).

Relation with Avogadro's Number:

Avogadro's number (6.022×10^{23} particles/mol) tells us that 1 mole of gas contains this many molecules. Therefore, 22.414 dm^3 of any ideal gas at STP contains exactly 6.022×10^{23} molecules.

Conclusion:

- Molar volume links the physical volume of a gas to the number of particles (via Avogadro's number) under standard conditions.

d. 39 g of potassium and 56 g of iron have equal number of atoms in them. Justify.

Ans. Molar mass of K (potassium) = 39 g/mol
Molar mass of Fe (iron) = 56 g/mol
It means Moles of K = 39 g/mol; 39 g = 1 mol
Moles of Fe = 56 g/mol 56; g = 1 mol
We know that 1 mole of any element contains Avogadro's number of atoms = 6.022×10^{23} atoms
Thus, 39 g of K = 1 mol = 6.022×10^{23}
56 g of Fe = 1 mol = 6.022×10^{23}

In short, 39 g of potassium and 56 g of iron contain an equal number of atoms, because both represent 1 mole of their respective elements.

e. 4 g of He, 17 g of NH_3 and 64 g of SO_2 occupy separately the volumes of 22.414 dm^3 although the sizes and masses of molecules of three gases are very different from each other. Explain.

Ans. **Molar Masses:** He (helium) 4 g/mol, NH_3 (ammonia) 17 g/mol, SO_2 (sulfur dioxide) 64 g/mol

At STP (Standard Temperature and Pressure), 1 mole of any ideal gas occupies:

$$22.414\ dm^3$$

Even though the **masses and sizes** of individual He, NH_3, and SO_2 molecules differ:

The **number of molecules** in each case is the same:

$$1\ mol = 6.022 \times 10^{23}\ molecules$$

The volume a gas occupies depends mainly on:

- Number of particles (moles)
- Temperature
- Pressure

It does not depend on **mass or size** of individual molecules (according to ideal gas behaviour).

Although He, NH_3, and SO_2 differ in molecular mass:

If you take **1 mole of each**, they will all occupy **22.414 dm^3** at STP.

Because the **volume of a gas** is determined by **number of moles**, not the mass of molecules.

f. Do you think that 1 mole of H_2 and 1 mole of NH_3 at 0 °C and 1 atm will have Avogadro's number of particles?

Ans. According to Avogadro's Law:

Equal volumes of all gases at the same temperature and pressure contain an equal number of molecules.

1 mole of H_2 (hydrogen gas) and 1 mole of NH_3 (ammonia gas) at 0°C and 1 atm will each contain Avogadro's number of particles, i.e.: 6.022×10^{23} molecules

$$1\ mole\ of\ H_2\ gas = 6.022 \times 10^{23}\ H_2\ molecules,$$

$$1\ mole\ of\ NH_3\ gas = 6.022 \times 10^{23}\ NH_3\ molecules$$

g. What is stoichiometry? Give the basic assumptions of stoichiometric calculations.

Ans. Stoichiometry

Stoichiometry is a branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.

Assumptions of Stoichiometry:

The following assumptions must be made while performing stoichiometric calculations:

- All the reactants are completely converted into the products.
- Law of conservation of mass and law of definite proportions are obeyed.
- No side reaction occurs.

h. What is a limiting reactant? How does it control the quantity of the product formed?

Ans. The limiting reactant (or limiting reagent) is the reactant that is completely consumed first during a chemical reaction. Once it is used up, the reaction stops, and no more product can form, even if other reactants are still available.

Chemical reactions require specific mole ratios between reactants.

If one reactant runs out, the remaining reactants can't react further — they become excess.

The amount of product is directly determined by how much of the limiting reactant was present at the start.

Example:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

Suppose we have:

1 mole of N_2 and 2 moles of H_2. According to the equation, we need 3 moles of H_2 for every 1 mole of N_2. But we only have 2 moles of H_2. Thus H_2 is the limiting reactant.

i. Differentiate theoretical and actual yields. How is the percentage yield of a reaction calculated?

Ans. Differentiate Between Theoretical and Actual Yield

Theoretical Yield	Actual Yield
1. The amount of product calculated from a balanced chemical equation is called **theoretical yield**.	1. The amount of product obtained while performing a chemical reaction is called **actual yield**.
2. It is also called expected yield or calculated yield.	2. It is also called experimental yield.
3. Theoretical yield is always greater than actual yield.	3. Actual yield is always lesser than theoretical yield.
4. There is no need to perform experiment. It is easily calculated from balanced chemical equation.	4. In order to get actual yield experiment has to be performed.

We can calculate the percentage yield of a chemical reaction with the help of actual yield and theoretical yield of the reaction. The efficiency of a reaction is also expressed in the form of percentage yield.

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Greater the percentage yield, more amount of reactants will convert into product and high will be the efficiency of the reaction i.e. % age yield ∝ Efficiency of the reaction.

j. What are the factors which are mostly responsible for the low yield of the products in chemical reactions?

Ans. Actual yield of a chemical reaction is usually lesser than theoretical yield due to many reasons. They are:

i. Practically inexperienced worker has many shortcomings and cannot get expected yield.
ii. Mechanical loss during experimentation e.g., filtration, separation by distillation or by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the yield.
iii. Some of the reactants might take part in a competing side reaction.
iv. The reaction might be reversible.
v. The reactants might be impure.
vi. Sometimes the reaction conditions are not suitably maintained.

DESCRIPTIVE QUESTIONS

Q.3 Differentiate limiting and non-limiting reactants. How a limiting reactant is determined from a balanced chemical equation and given data?

Ans. See Page No. (118)

Q.4 Differentiate actual and theoretical yields. Why the theoretical yield is always greater than actual yield?

Ans. See Page No. (121)

NUMERICAL PROBLEMS

Q.5 A solution of sodium hydroxide (NaOH) is prepared by dissolving 2.00 g of solid sodium hydroxide in water to make a final volume of 250 cm^3.

a) Calculate the number of moles of sodium hydroxide used.

b) Determine the molar mass of sodium hydroxide.

c) Calculate the concentration of the sodium hydroxide solution in mol dm^{-3}.

d) If more water is added to the above solution to raise the volume of solution to cm^3, what would be the concentration now?

(a) Number of moles of sodium hydroxide used

Formula:

$$\text{Number of moles (n)} = \frac{\text{Mass of solute (g)}}{\text{Molar mass (g/mol)}}$$

$$n = \frac{2g}{40gmol^{-1}} = 0.05mol$$

(b) Molar mass of sodium hydroxide (NaOH)

- Atomic mass of Na (Sodium) = 23 g/mol
- Atomic mass of O (Oxygen) = 16 g/mol
- Atomic mass of H (Hydrogen) = 1 g/mol

Molar mass of NaOH = 23 + 16 + 1 = 40 g/mol

(c) Concentration of NaOH solution in mol dm^{-3}

Formula:

$$\text{Concentration (C)} = \frac{\text{Number of moles}(n)}{\text{Volume of solution in } dm^3}$$

Convert cm^3 to dm^3

$$1\ dm^3 = 1000\ cm^3 \Rightarrow 250\ cm^3 = \frac{250}{1000} = 0.250\ dm^3$$

$$\text{Concentration} = \frac{0.050}{0.250} = 0.20\ mol/dm^3$$

(d) New concentration if volume is increased to 500 cm^3

No. of moles of NaOH = n = 0.05 mol

Volume of solution = v = 500 cm^3 = 0.50 dm^3

$$\text{Volume of solution} = v = 500\ cm^3 = \frac{500}{1000} = 0.500\ dm^3$$

$$\text{Molar concentration} = c = \frac{n}{v}$$

$$\text{New Concentration} = c = \frac{0.050}{0.500\ dm^3} = 0.10\ mol/dm^3$$

Q.6 Ammonia gas (NH_3) reacts with oxygen gas (O_2) according to the following balanced equation:

$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$

In an experiment, 34.0 g of ammonia is reacted with 96 g of oxygen.

a) Determine the limiting reactant.

b) Calculate the maximum mass of nitrogen monoxide (NO) that can be formed.

c) Calculate the mass of the excess reactant remaining after the reaction is complete. (Relative atomic masses: H = 1.0, N = 14.0, O = 16.0)

Ans. (a) Reaction Given: $4NH_{3(g)} + 5O_{2(g)} \longrightarrow 4NO_{(g)} + 6H_2O_{(g)}$

Given Data:

Given mass of NH_3 = 34.0 g

Given mass of O_2 = 96 g

Molar mass of NH_3 = 14 + (3 × 1) = 17 g/mol

Molar mass of O_2 = 2 × 16 = 32 g/mol

$$\text{No. of moles of } NH_3 = n = \frac{\text{mass}}{\text{molar mass}} = \frac{34.0}{17} = 2.00\ mol$$

$$\text{No. of moles of } O_2 = n = \frac{96g}{32} = 3\ mol$$

Comparing the mole ratio, from the balanced chemical equation

NH_3 : NO

4 : 4

2 : 2

Now O_2 : NO

5 : 4

$1 : \frac{4}{5}$

$3 : \frac{4}{5} \times 3 = 2.4\ mol$

As NH_3 gives least amount of NO (i.e 2 moles) so it is the Limiting reactant.

(b) No. of moles of NO produced = 2 mol

Molar mass of NO = 14 + 16 = 30 g mol^{-1}

Mass of NO produced = m = n × M

= 2 mol × 30 g mol^{-1}

= 60 g

(c) Mass of excess reactant:

NH_3 is the Limiting reactant, so

NH_3 : O_2

4 : 5

$1 : \frac{5}{4}$

$2 : \frac{5}{4} \times 2 = 2.5\ mol$

No. of moles of O_2 left (unreacted) = No. of moles of O_2 initially taken − No. of moles of O_2 consumed

= 3 − 2.5 = 0.5 mol

Mass of O_2 left (unreacted) = n × M

= 0.5 mol × 32 gmol^{-1}

= 16 g

Q.7 When iron (III) oxide (Fe_2O_3) reacts with carbon monoxide (CO) in a blast furnace, iron metal (Fe) is produced according to the following equation:

$$Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

If 1000 kg of iron (III) oxide is reacted with excess carbon monoxide, and 650 g of iron is obtained, what is the percentage yield of iron? (A_r of O = 16.0, Fe = 55.8)

Ans. Chemical Equation:

$$Fe_2O_{3(s)} + 3CO_{(g)} \longrightarrow 2Fe_{(s)} + 3CO_{2(g)}$$

Given Data:

Mass of Fe_2O_3 used = 1.00 kg = 1000 g

Mass of Fe actually obtained = 650 g

Atomic Masses: Fe = 55.8 O = 16.0

Molar mass of Fe_2O_3 = (2×55.8)+(3×16.0)=111.6+48.0=159.6 g/mol

$$\text{No. of moles of } Fe_2O_3 = n = \frac{\text{mass}}{\text{molar mass}} = \frac{1000}{159.6} \approx 6.265\ mol$$

To Find Theoretical Mass of Fe

From the equation:

$$1\ mol\ Fe_2O_3 \longrightarrow 2\ mol\ Fe$$

So,

$$6.265\ mol\ Fe_2O_3 \longrightarrow 6.265 \times 2 = 12.53\ mol\ Fe$$

Mass of 12.53 mol Fe:

$$\text{Mass} = n \times M = 12.53 \times 55.8 \approx 699.2 \text{ gFe (theoretical)}$$

Step 4: Use Percentage Yield Formula

$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

$$= \frac{650}{699.2} \times 100 \approx 92.96\%$$

Final Answer:

Percentage Yield of Fe = 92.96%

Q.8 1.5g of C_2H_6 is burnt in excess of O_2 to produce CO_2 and H_2O. What volume of CO_2 is produced at STP.

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

Also calculate:

a) No. of molecules of solid CO_2 produced.

b) No. of O_2 molecules reacted.

c) No. of CH bond of C_2H_6 are broken in this reaction.

Given Data:

Mas of C_2H_6 burnt = 1.5g

Molar Mass of C_2H_6 = 30g mol^{-1}

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

$$\text{No. of moles of } C_2H_6 = \frac{1.5 \text{ g}}{30 \text{ g/mol}} = 0.05 \text{ mol}$$

C_2H_6 : CO_2

2 : 4

0.05 : x

Moles of CO_2 = 0.1 moles

Volume of CO_2 = $0.1 \times 22.4 = 2.24 \text{ dm}^3$

(a) No. of molecules of Solid CO_2 produced

$= n \times N_A$

$= 0.1 \times 6.02 \times 10^{23} = 6.02 \times 10^{22}$ molecules

(b) No. of molecules reacted

C_2H_6 : O_2

2 : 7

0.05 : x

$$\text{Moles of } O_2 \text{ reacted} = \frac{7}{2} \times 0.05 = 0.175 \text{ mole}$$

Molecules of Oxygen = $0.175 \times 6.02 \times 10^{23} = 1.0535 \times 10^{23}$ or 1.05×10^{23}

(c) No. of C – H bonds broken

1 mole of C_2H_6 has C – H bonds = 6 moles

0.05 moles of C_2H_6 has C – H bonds = $6 \times 0.05 = 0.3$ moles

No. of bonds = $n \times N_A$

$= 0.3 \times 6.02 \times 10^{23} = 1.8 \times 10^{23}$ C – H bonds

ADDITIONAL SLOs BASED MCQs

1. How many molecules are there in one mole of H_2O?
   6.023 × 10^{19} B. 6.023 × 10^{23} C. 1.084 × 10^{18} D. None of these
2. A flask contains 500 cm^3 of SO_2 at STP. The flask contains
   A. 40 g B. 100 g C. 50 g D. 1.427g
3. A necklace has 6 g of diamond in it. What are the numbers of atoms in it?
   A. 6.02 × 10^{23} B. 12.04 × 10^{23} C. 1.003 × 10^{23} D. 3.01 × 10^{23}
4. What is the mass of aluminium in 204 g of the aluminium oxide, Al_2O_3?
   A. 26 g B. 27 g C. 54 g D. 108 g
5. The reactant which is consumed earlier and gives least quantity of product is called.
   A. Reactant in excess B. Stoichiometry C. Limiting reactant D. Stoichiometric amount

6. Which one of the following compounds contains the highest percentage by mass of nitrogen?
   A. NH_3 B. N_2H_4 C. NO D. NH_4OH
7. Vitamin-A has a molecular formula $C_{20}H_{30}O$. The number of vitamin-A molecules in 500 mg of its capsule will be:
   A. 6.02 × 10^{23} B. 1.05 × 10^{21} C. 3.01 × 10^{22} D. 3.01 × 10^{23}
8. When one mole of each of the following is completely burnt in oxygen, which will give the largest mass of CO_2?
   A. Carbon Monoxide B. Diamond C. Ethane D. Methane
9. One mole of ethanol and one mole of ethane have an equal:
   A. Mass B. Number of Atoms C. Number of electrons D. Number of molecules
10. Methane reacts with steam to form H_2 and CO as shown below.
    $$CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{(g)} + 3H_{2(g)}$$
    What volume of H_2 can be obtained from 100 cm^3 of methane at the standard temperature and pressure?
    A. 300 cm^3 B. 200 cm^3 C. 150 cm^3 D. 100 cm^3
11. The Avogadro's constant is the number of:
    A. Atoms in 1g of Helium B. Molecules in 35.5g of Chlorine
    C. Electrons needed to deposit 24g of Mg D. Atoms in 24g of Mg
12. How many moles of oxygen are needed for the complete combustion of two moles of butane?
    A. 2 B. 8 C. 10 D. 13
13. If four moles of SO_2 are oxidized to SO_3, how many moles of oxygen molecules are required?
    A. 0.5 B. 1.0 C. 1.5 D. 2.0
14. Which of the following statements is incorrect?
    A. 12 g of Carbon gas contains one mole of atoms
    B. 28 g of Nitrogen gas contains one mole of molecules of N_2
    C. 1 dm^3 of 1.0 mole dm^{-3} solution of NaCl contains one mole of Chloride ions
    D. None of above
15. What volume of SO_2 at room temperature and pressure is produced on heating 9.7g of Zinc Sulphide (ZnS) if reaction takes place as follows;
    $$2ZnS_{(s)} + 3O_{2(g)} \longrightarrow 2ZnO_{(s)} + 2SO_{2(g)}$$
    A. 1.2 dm^3 B. 2.4 dm^3 C. 3.6 dm^3 D. 4.8 dm^3

Answer Key with Explanations

Sr.No.	Option	Explanations
1.	B	Water is a molecular substance. One mole of a molecular substance contains Avogadro's number (6.023 × 10^{23}) molecules.
2.	D	1 mol of a gas at STP = 22.414 dm^3 = 22414 cm^3 1 mol of SO_2 = 64g 22414 cm^3 of SO_2 at STP has mass = 64g 500 cm^3 of SO_2 at STP has mass = $\frac{64}{22414} \times 500 = 1.427$g
3.	D	Diamond is an allotrope of carbon. 12g carbon (Diamond) = 1 mol = 6.023 × 10^{23} atoms 6g carbon (Diamond) = 0.5 mol = 3.011 × 10^{23} atoms

4.	D	Molar mass of $Al_2O_3 = 27(2) + 16(3) = 102g\ mol^{-1}$ $102g\ Al_2O_3$ contains $Al = 54g$ $204g\ Al_2O_3$ contains $Al = \frac{54}{102} \times 204 = 108g$
5.	C	A reactant which is used or consumed earlier due to its lesser amount and controls the amount of product formed in a chemical reaction is called limiting reactant.
6.	B	NH_3: % of N = $\frac{14}{17} \times 100 = 82.3\%$ N_2H_4: % of N = $\frac{28}{32} \times 100 = 87.5\%$ NO: % of N = $\frac{14}{30} \times 100 = 46.6\%$ NH_4OH: % of N = $\frac{14}{35} \times 100 = 40\%$
7.	B	Number of molecules of a compound = $\frac{\text{Mass of compound}}{\text{Molecular mass}} \times N_A$ Molecular mass of vitamin-A ($C_{20}H_{30}O$) = 276g Number of molecules of vitamin-A = $\frac{0.5}{276} \times 6.023 \times 10^{23} = 1.05 \times 10^{21}$
8.	C	$CO + ½O_2 \longrightarrow CO_2$ $C + O_2 \longrightarrow CO_2$ $C_2H_6 + 3½O_2 \longrightarrow 2CO_2 + 3H_2O$ $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$
9.	D	Both ethanol and ethane are molecular substances. So 1 mole of each ethanol and ethane contains Avogadro's number (6.023×10^{23}) molecules.
10.	A	$CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{(g)} + 3H_{2(g)}$ The above equation shows that 1 mole of methane (CH_4) produces 3 moles of hydrogen (H_2). So 100 cm^3 of CH_4 will produce 300 cm^3 of H_2.
11.	D	24g of Mg = 1 mol of Mg = 6.023×10^{23} atoms (Avogadro's constant)
12.	D	According to balance chemical equation; $2C_4H_{10} + 13O_2 \longrightarrow 8CO_2 + 10H_2O$
13.	D	$2SO_2 + O_2 \longrightarrow 2SO_3$ $4SO_2 + 2O_2 \longrightarrow 4SO_3$
14.	D	(a), (b) and (c) are correct statements because in each statement the number of stated particles are Avogadro's number.
15.	B	$2ZnS_{(s)} + 3O_{2(g)} \longrightarrow 2ZnO_{(s)} + 2SO_{2(g)}$ $n_{ZnS} = \frac{9.7}{97} = 0.1$ mol ZnS : SO_2 2mol : 2mol 0.1 mol : 0.1 mol 1 mole of SO_2 = 22.414 dm^3 at STP 0.1 mole of SO_2 = 2.4 dm^3 at STP

ADDITIONAL SHORT ANSWER QUESTIONS

Q.1 Define Avogadro's number and molar volume.

Ans. Avogadro's Number:

"The number of particles (atoms, molecules, formula units or ions) which are present in one mole of a chemical substance is called Avogadro's number."

Equal volumes of all the ideal gases at same temperature and pressure contain equal number of molecules.

- It is represented by N_A
- Its unit is mol^{-1}
- Its value is 6.022×10^{23}

Example:

16g CH_4 = 1 mole of CH_4 = 22.414 dm^3 of CH_4 = 6.02×10^{23} molecules of CH_4.

Molar Volume:

"The volume occupied by one mole of an ideal gas at standard temperature and pressure (STP) i.e. 0°C and 1 atm is called molar volume."

Value: Its value is 22.414 dm^3 (0.022414 m^3 or 22414 cm^3)

Examples:

(i) 2g of H_2 = 1 mole of H_2 = 22.414 dm^3
(ii) 16g of CH_4 = 1 mole of CH_4 = 22.414 dm^3
(iii) 44g CO_2 = 1 mole of CO_2 = 22.414 dm^3

From above examples it is clear that 1 mole of different gases e.g., 2g H_2, 16g CH_4, 44g CO_2 have same number of molecules and same volume (22.414 dm^3) at STP but different masses.

Q.2 Define mole with two examples.

Ans. Mole: "The amount of a substance which contains Avogadro's number of particles (atoms, molecules, formula units, or ions) is called mole."

Examples:

1 mole of Na = 23 g
1 mole of H_2O = 18 g
1 mole of NaCl = 58.5 g
1 mole of HCO_3^{-1} = 61 g

Formula: Mole = $\frac{\text{Given mass of the substance}}{\text{Molar mass}}$

Q.3 Why 180 g of glucose and 342 g of sucrose have same number of molecules but different number of atoms?

Ans. Molar mass of glucose ($C_6H_{12}O_6$) = 180 g mol^{-1}
Molar mass of sucrose ($C_{12}H_{22}O_{11}$) = 342 g mol^{-1}

Both glucose and sucrose are molecular species. Therefore one mole contains same number of molecules i.e., Avogadro's number (6.02×10^{23}).

180 g of glucose = 1 mol = 6.02×102^{23} Molecules of glucose
342 g of sucrose = 1 mol = 6.02×102^{23} Molecules of sucrose

One molecule of glucose has different number of atoms than one molecule of sucrose so one mole of each molecule contains different number of atoms in them.

As 1 molecule of glucose (contains) = 24 atoms
So, 1 mole glucose molecules (contains) = $24 \times 6.02 \times 10^{23}$ atoms
Similarly,
As 1 molecule of sucrose (contains) = 45 atoms
So, 1 mole of sucrose molecules (contains) = $45 \times 6.02 \times 10^{23}$ atoms

Q.4 Calculate the number of oxygen atoms in 10.037g of $CuSO_4.5H_2O$ (Atomic mass of Cu = 63, H = 1, O = 16).

Given Data: Mass of $CuSO_4.5H_2O$ = 10.037 g
Required: Number of O atoms = ?

Solution:

$$\text{Number of moles of } CuSO_4.5H_2O = \frac{\text{Mass of } CuSO_4.5H_2O}{\text{Molar mass of } CuSO_4.5H_2O}$$

$$\text{Molar Mass of } CuSO_4.5H_2O = 63.5 + 32 + (4 \times 16) + (5 \times 18) = 249.5 \text{ g mol}^{-1}$$

$$\text{Number of moles of } CuSO_4.5H_2O = \frac{10.037}{249.5} = 0.04 \text{ moles}$$

1 mole of $CuSO_4.5H_2O$ have moles of O = 9 moles

0.04 moles of $CuSO_4.5H_2O$ have moles of O = $9 \times 0.04 = 0.36$ moles

$$\text{Number of Oxygen atoms} = \text{Number of moles of oxygen atoms} \times N_A$$

$$= 0.36 \times 6.02 \times 10^{23}$$

$$\text{Number of Oxygen atoms} = 2.167 \times 10^{23} \text{ atoms}$$

Q.5 Calculate mass in Kg of 2.6×10^{20} molecules of SO_2.

Given Data: Number of SO_2 molecules = 2.6×10^{20} molecules

Required: Mass in kg of SO_2 molecules = ?

Solution:

$$\text{Number of } SO_2 \text{ molecules} = \frac{\text{Mass of } SO_2 \text{ molecules}}{\text{Molar mass of } SO_2} \times N_A$$

$$\text{Molar Mass of } SO_2 = 32 + (16 \times 2) = 64 \text{ g mol}^{-1}$$

$$2.6 \times 10^{20} = \frac{\text{Mass of } SO_2 \text{ molecules}}{64} \times 6.02 \times 10^{23}$$

$$\frac{2.6 \times 10^{20} \times 64}{6.02 \times 10^{23}} = \text{Mass of } SO_2 \text{ molecules}$$

$$27.641 \times 10^{-3} g = \frac{27.641 \times 10^{-3}}{1000}$$

$$= 2.764 \times 10^{-5} \text{ kg} = \text{Mass of } SO_2 \text{ molecules}$$

Q.6 Calculate the number of water molecules in 10 g of ice.

Given Data: Mass of ice (water) = 10.0 gm

Molar mass of water = 18 gmol^{-1}

To Find: Number of molecules of water = ?

Solution:

$$\text{Number of molecules of water} = \frac{\text{Mass of water in gm}}{\text{Molar mass of water in gmol}^{-1}} \times \text{Avogadro's number}$$

$$= \frac{10}{18\text{gmol}^{-1}} \times 6.02 \times 10^{23}$$

$$= 0.55 \times 6.02 \times 10^{23} = 3.31 \times 10^{23}$$

Q.7 Write down only steps to determine Limiting Reactant.

Ans. Limiting Reactant:

"A reactant which is used or consumed earlier due to its lesser amount and controls the amount of product formed in a chemical reaction is called limiting reactant."

A limiting reactant is that which

(i) controls the amount of product formed (ii) is taken in lesser amount

(iii) is consumed earlier (iv) produces least amount of product.

Identification of Limiting Reactant:

The following three steps should be followed to find out the limiting reactant:

(i) Calculate the number of moles from the given amount of reactant.

(ii) Find out the number of moles of product with the help of a balanced chemical equation.

(iii) Identify the reactant which produces the least amount of product as limiting reactant.

Q.8 Many reactions taking place in our surrounding involve limiting reactant. Justify with examples.

Ans. "The reactant which is consumed earlier and controls the amount of product in a chemical reaction is called a limiting reactant."

Mostly combustion reactions taking place in our surroundings involve the limiting reactant.

For example:

- $CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$ • $C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)}$

In above examples, O_2 gas is present in excess while CH_4 and C are the reactants which are consumed earlier so they are limiting reactants.

Q.9 Law of conservation of mass has to be obeyed during stoichiometric calculations. Justify it.

Ans. According to law of conservation of mass.

"Mass can neither be created nor destroyed during a chemical reaction but it changes its form from one form to another."

In stoichiometric calculations, we use balanced chemical equation where total mass of reactants is equal to the total mass of products.

For Example:
- $2H_2 + O_2 \longrightarrow 2H_2O$
- $4g + 32g \longrightarrow 36g$

So while doing stoichiometric calculations law of conservation of mass has to be obeyed.

Q.10 What is stoichiometry? Which laws are obeyed during stoichiometric calculation?

Ans. Stoichiometry:

"The branch of chemistry which deals with the study of quantitative relationship between reactants and products in a balanced chemical equation is called stoichiometry."

Assumptions:

When stoichiometric calculations are performed, we have to assume the following Conditions:

(i) All the reactants are completely converted into products.

(ii) No side reaction occurs.

During stoichiometric calculations, law of conservation of mass and law of definite proportions are obeyed.

Q.11 Why theoretical yield of a chemical reaction is greater than the actual yield?

Ans. Actual yield of a chemical reaction is usually lesser than theoretical yield due to many reasons. They are:

(i) Practically inexperienced worker has many shortcomings and cannot get expected yield.

(ii) Mechanical loss during experimentation e.g., filtration, separation by distillation or by a separating funnel, washing, drying and crystallization if not properly carried out, decrease the yield.

(iii) Some of the reactants might take part in a competing side reaction.

(iv) The reaction might be reversible.

(v) The reactants might be impure.

(vi) Sometimes the reaction conditions are not suitably maintained.

Q.12 Differentiate between Theoretical and Experimental yield.

Ans.

Theoretical Yield	Actual Yield
1) The amount of product calculated from a balanced chemical equation is called theoretical yield.	1) The amount of product obtained while performing a chemical reaction is called actual yield.
2) It is also called expected yield or calculated yield.	2) It is also called experimental yield.
3) Theoretical yield is always greater than actual yield.	3) Actual yield is always lesser than theoretical yield.
4) No need to perform experiment. Just to calculate from balanced chemical equation.	4) In order to get actual yield experiment has to be performed.