4 STOICHIOMETRY

Student Learning Outcomes [C-11-A-01 to C-11-A-25]

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)**

Stoichiometry is derived from Greek words *stoicheion* means element and *metron* means measure. 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 element 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.

Stoichiometric calculations obey law of conservation of mass and law of definite proportions. 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*. According to the law of definite proportions, a pure compound always contains the same element combined in the same ratio by mass.

4.1 CONCEPT OF MOLE

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?

Avogadro's number is a physical constant representing the molar number of entities. The exact value of it is $6.02214179 \times 10^{23}$ mol^{-}. 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^+.

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.** The mass of one mole of a substance expressed in grams is called molar mass. The unit of molar mass is g/mol. 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.

$$\text{Molar mass of } CCl_4 = \text{Molar mass of one C} + \text{Molar mass of Cl} \times 4$$
$$= 12.0 \times 1 + 35.5 \times 4$$
$$\text{Molar mass of } CCl_4 = 12.0 + 142.0 = 154.0 \text{ g}$$

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.

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:

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

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

Sample Problem 4.1

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

Solution:

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

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

Sample Problem 4.2

Calculate the mass of 0.5 moles of HCl.

Solution:

Mass of HCl = Number of moles x Molar mass

= 0.5 x 36.5 = 18.3 mol

Sample Problem 4.3

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

Solution:

Molar mass of $MgSO_4 = 24 + 96 = 120 \text{ g mol}^{-1}$

Number of moles of $MgSO_4 = 10^{-3}$

Mass of $MgSO_4 = 10^{-3} \text{ mol} \times 120 \text{ g mol}^{-1} = 120 \times 10^{-3} = 0.12 \text{ g}$

Quick Check 4.1

a) Calculate the molar mass of $KMnO_4$.
b) Calculate the number of moles in 0.23 g of sodium.
c) Calculate the mass of 1.5 moles of $Ca(OH)_2$.
d) The given mass of $KClO_3$ is 24.5 g. Calculate its number of moles.
e) How many molecules are present in 1.75 g of H_2O_2?
f) How many atoms are present in 15 g of a gold ring?

4.2 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., 6.02 x 10^{23}.
It is interesting to know that different masses of elements have the same number of atoms.

1.0 g of hydrogen = 1 mol of hydrogen = 6.02 x 10^{23} atoms of H

23.0 g of sodium = 1 mol of Na = 6.02 x 10^{23} atoms of Na

238.0 g of uranium = 1 mol of U = 6.02 x 10^{23} atoms of U

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 x 10^{23} molecules of water

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

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

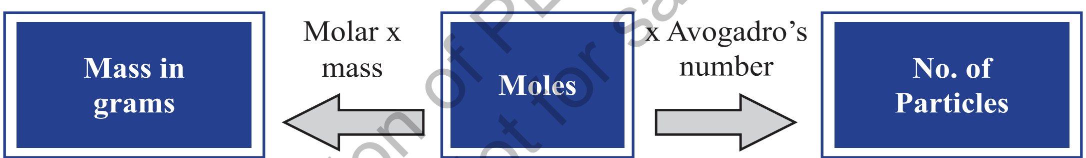
Similarly, the number of ions in one mole of different ionic species is always the same, i.e Avogadro's number.

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

62.0 g of NO_3^{1-} = 1 mole of NO_3^{1-} = 6.02 x 10^{23} ions of NO_3^{1-}

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

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



Sample Problem 4.4

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

Solution: No. of moles of glucose = $\dfrac{3.76 \text{ x } 10^{24} \text{ molecules}}{6.02 \text{ x } 10^{23} \text{ molecules mol}^{-1}}$

= 6.25 moles

Sample Problem 4.5

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

Solution:

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

Number of atoms of sodium = Number of moles of sodium x N_A

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

$= 0.602 \times 10^{23}$ atoms

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) $C_{10}H_6O_3$

$10 \times A_r\,(C) + 6 \times 1.0\, A_r\,(H) + 3 \times A_r\,(O)$

$(10 \times 12.0) + (6 \times 1.0) + (3 \times 16.0)$

$120 + 6 + 48 = 174$ g/mol

Mass of 1 mol of $C_{10}H_6O_3$ = 174 g/mol

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

Quick Check 4.2

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

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?

Interesting Information!

Isopentyl acetate ($C_7H_{14}O_2$) is the compound responsible for the scent of bananas. Interestingly, bees release about 1 μg (1×10^{-6} g) of this compound when they sting. The resulting scent attracts other bees to join the attack.

4.3 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 $\mathbf{V_m}$. By using molar volume relationship, mass or mole of a gas at STP can be converted into volume, and vice versa.

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

- 22.4 dm^3 of CO_2 at STP = 44.0 g of CO_2 = 6.02×10^{23} molecules of CO_2 =1mole of CO_2
- 22.4 dm^3 of any gas at STP = molar mass in grams = 6.02×10^{23} molecules = 1mole
- 22.4 dm^3 of H_2 gas at STP = 2 g = 6.02×10^{23} molecules = 1mole
- 22.4 dm^3 of NH_3 gas at STP = 17 g = 6.02×10^{23} molecules= 1mole

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

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

$$\mathbf{V = n \times V_m}$$

Sample Problem 4.7

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

Solution:

The formula for volume determination at STP is,

$$V = n \times V_m$$

Volume of 2.5 mole of Cl_2 = 22.4 dm^3 x 2.5 = 56.0 dm^3

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 is,

$$V = n \times V_m$$

Volume of methane in dm^3 at STP = 4.75 x 22.4 = 106.4 dm^3

4.4 MOLAR MASS AND DENSITY OF GASES

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

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

$$\mathbf{d = \frac{m}{V}}$$

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 the gas at STP is known, its molar mass can be calculated.

Sample Problem 4.9

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

Solution

$$m = d \times V$$

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

Quick Check 4.3

Calculate the molar mass of a gas which has density of 1.34 g/dm^3 at STP.

4.5 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

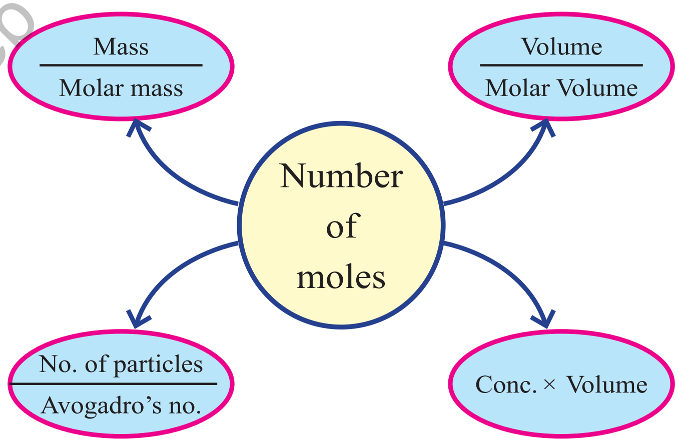
$$\mathbf{n = C \times V}$$

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

where C is the molar concentration and V is the volume of the solution.
For example concentration of 0.2 mol of a substance per dm^3 is.

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

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



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}

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

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

$$n = \frac{27.64}{138.2}$$

Volume of solution = 1 dm^3

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

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

Quick Check 4.4

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

4.6 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

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 different ways;

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

ii. 1 molecule of N_2 reacts with 3 molecules of H_2 to form 2 molecules of NH_3.

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

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

Keep in Mind!

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.
3) 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 form the mass 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

↓

Calculate the number of moles of the unknown reactant or product using the relevant formula

↓

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

Sample Problem 4.11 (Mole-Mole Conversion)

When 3.3 mol 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

3.6 moles of N_2 needs H_2 to produce NH_3 = 3 × 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 combustion of glucose is:

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

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:

$$2Al_{(s)} + Fe_2O_{3(s)} \longrightarrow Al_2O_{3(s)} + 2Fe_{(s)}.$$

Solution:

Molar mass of Fe_2O_3, M = 159.6 g mol^{-1}

Number of moles of Fe_2O_3 (n) $= \dfrac{m}{M}$

$$= \frac{32.0 \text{ g}}{159.6 \text{ g mol}^{-1}}$$

$= 0.02$ mol

From the balanced equation, 1 mol of Fe_2O_3 reacts with 2 moles of Al, therefore, number of moles of Al that reacts with 0.02 mole of Fe_2O_3= 2 x 0.02 = 0.04 mol

Mass of $A\ell$ = n x M

= 0.04 mol x 27 g mol^{-1}

= 1.08 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?

Sample Problem 4.13 (Volume-Volume)

Calculate volume of ammonia that can be produced by the reaction of 100 dm^3 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:

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

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 mol of LiOH produces = 1 mol Li_2CO_3

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

Mass of Li_2CO_3 produced = No. of mol × Molar mass

Mass of Li_2CO_3 produced = 10.0 mol × 73.9 g mol^{-1} = 739.0 g

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

Quick Check 4.8

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

$$2\ NaOH_{(aq)} + Cl_{2(g)} \longrightarrow NaOCl_{(aq)} + NaCl_{(aq)} + H_2O_{(\ell)}.$$

Sample Problem 4.15 (Mass-mole calculations)

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

$$NaHCO_{3\,(g)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(aq)}$$

How many moles of HCl will be neutralized by 2.1 g of baking soda

Solution:

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

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

Stoichiometrically, the mole ratio of HCl and $NaHCO_3$ is 1 : 1.

Hence moles of HCl used = 0.025 mol

Thus 2.1 g of $NaHCO_3$ will neutralize 0.025 moles of HCl.

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) $= \dfrac{m}{M}$

$= \dfrac{7.0\ g}{55.8\ gmol^{-1}}$

= 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

$= 2.8\ dm^3$

4.7 LIMITING AND EXCESS REACTANT

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. 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. 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.

4.7.1 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.

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.

Following numerical problem will make the idea clear.

Sample Problem 4.17 (Limiting Reactant)

Calculate the mass of N_2 produced from 1.81 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:

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

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

In balanced equation, $CuO : N_2$

$3 : 1$

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

$NH_3 : N_2$

$$2 : 1$$

$$1 : \frac{1}{2}$$

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

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 x M

= 0.38 mol x 17.0 g mol^{-1}

= 9.0 g

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.

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

Solution:

The net ionic equation is

$$Pb^{+2}_{(aq)} + SO_4^{-2}{}_{(aq)} \longrightarrow PbSO_{4(s)}$$

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

No. of moles = Concentration (mol dm^{-3}) x Volume (dm^3)

$$n = CV$$

moles of Pb^{2+} ions = 0.05 mol dm^{-3} × 1.25 dm^3 = 0.0625 mol

moles of SO_4^{-2} ions = 0.025 mol dm^{-3} × 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}):

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

4.7.2 Amount of Product and Unreacted Excess Reagent

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).

Consider the reaction between hydrogen and oxygen to form water.

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

- 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. Actually, we will get 2 moles (36 g) of water because 2 moles (4 g) of hydrogen react with 1 mole (32 g) of oxygen according to the balanced equation.

- 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 will be excess.

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.

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

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

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 mol of methane produces CO_2 = 1mol

1.875 mol of methane produces CO_2 = 1×1.875 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 mol of oxygen produces CO_2 = 1 mol

1.563 mol of oxygen produce CO_2 = 0.5 × 1.563 mol = 0.7815 mol 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.

Mass of CO_2 in grams = Moles of CO_2 × Molar mass of CO_2

$= 0.7815\ mol \times 44\ g\ mol^{-1} = 34.39\ g$

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

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

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

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

Number of moles of CH_4 in excess = Quantity taken – Quantity used

= 1.875 mol - 0.7815 mol =1.0935 mol

Excess mass of CH_4 (excess reagent) = 1.0935 × 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

b) 2 moles of N_2 and 3 moles of H_2

c) 1 mole of N_2 and 5 moles of H_2

d) 3 moles of N_2 and 3 moles of H_2

e) Each produces the same amount of product.

4.8 THEORETICAL YIELD AND ACTUAL YIELD

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

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

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

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

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 g mol^{-1}) and acetic anhydride, $C_4H_6O_3$ (molar mass 163.93 g mol^{-1}).

$$C_7H_6O_{3(\ell)} + C_4H_6O_{3(\ell)} \longrightarrow C_9H_8O_{4(s)} + CH_3COOH_{(aq)}$$

Calculate the theoretical yield of aspirin, (molar mass 180.16 g mol^{-1}) when 3.00 g of salicylic acid is heated with 6.00 g of $C_4H_6O_3$. What is % yield when actual yield is 3.15 g?

Solution:

$C_7H_6O_3$	+	$C_4H_6O_3$	$\longrightarrow$	$C_9H_8O_4$	+	CH_3COOH
Salicylic acid		Acetic anhydride		Aspirin		
1 mol		1 mol		1 mol		
138.12 g		163.93 g		180.16 g		

1 mol of salicylic acid produces aspirin = 1 mol

Mass of salicylic acid = 3.00 g

Number of moles of salicylic acid $= \dfrac{300\ \text{g}}{138.12\ \text{g mol}^{-1}} = 0.022$ mol

Mass of acetic anhydride = 6.00 g

Number of moles of acetic anhydride $= \dfrac{6.00\ \text{g}}{163.93\ \text{g mol}^{-1}} = 0.037$ 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 of salicylic acid produces aspirin = 0.022 mol

0.022 mol Salicylic acid produces Aspirin =0.022 mol

Mass of Aspirin = 0.021 mol x 180.16 g mol^{-1} = 3.96 g

Theoretical yield = 3.96

Actual yield = 2.85

% age yield $= \dfrac{2.85}{3.96 \times 100} = 71.97$ %

Quick Check 4.10

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

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

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?

4.9 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.

4.9.1 Significance of Stoichiometry in Medicine

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

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

EXERCISE

MULTIPLE CHOICE QUESTIONS

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 mol 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 dioxide (molar mass of CO_2 = 44)

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 Na2SO4 (formula mass 142) d) 8.5 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 g dm^{-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 g dm^{-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}

SHORT-ANSWER QUESTIONS

Q.2 Attempt the following short-answer questions:

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

b) Define the following terms with one example in each case.

(a) Molar mass (b) Molar volume (c) Molar concentration

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

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

e) 4g 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 molecular masses of molecules of the three gases are very different from each other. Explain.

f) Do you think that 1 mole of H_2 and 1 mole of NH_3 at 0 ^oC and 1 atm will have Avogadro's number of particles?

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

g) What is a limiting reactant? How does it control the quantity of the product formed?

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

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

DESCRIPTIVE QUESTIONS

Q3. Differentiate limiting and non-limiting reactants. How a limiting reactant is determined from a balanced chemical equation and given data?

Q4. Differentiate actual and theoretical yields. Why the theoretical yield is always greater than actual yield?

(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) Determine the molar mass of sodium hydroxide.

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

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 500 cm^3, what would be the concentration now?

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

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

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)

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:

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

If 1.00 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)

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. $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

Also calculate:

a) No. of molecules of solid CO_2 produced. b) No. of O_2 molecules reacted

c) No. of CH bond of C2H6 are broken in this reaction.