CHAPTER 4

STOICHIOMETRY

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.

Balanced chemical equation

(i) The balanced chemical equation has the same number of atoms of each type on both sides of equation.

(ii) 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.

(iii) Stoichiometric calculations obey law of conservation of mass and law of definite proportions.

Law of conservation of mass

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 chemical equation.

Law of definite proportions

According to the law of definite proportions, a pure compound always contains the same element combined in the same ratio by mass.

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.

SHORT QUESTION

Define mole with examples.

Avogadro number

"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 **Amedeo 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^{-1}. In calculations we use the rounded off value 6.02×10^{23}.

Examples

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 x 10^{23} formula units of NaCl.
- 1 mole of Na^+ contains 6.02 x 10^{23} ions of Na^+.

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

Molar mass

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.

Definition

"*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.

Example

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 × 4 = 12.0 × 1 + 35.5 × 4
Molar mass of CCl_4 = 12.0 + 142.0 = 154.0 g

Other examples

1 mole of carbon atoms is 12.0 g
1 mole of CO_2 molecules 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

Calculation of number of moles (by using mass)

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

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

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

Sample Problem 4.2

Calculate the mass of 0.5 moles of $HC\ell$.

Solution

$$\text{Mass of } HC\ell = \text{Number of moles} \times \text{Molar mass}$$
$$= 0.5 \times 36.5$$
$$= 18.3 \text{ mol}$$

Sample Problem 4.3

Calculate the mass of 10^{-3} mole 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 $MgSO_4$	$= 10^{-3}\ mol \times 120\ g\ mol^{-1}$
	$= 120 \times 10^{-3}$
	$= 0.12\ g$

QUICK CHECK 4.1

(a) Calculate the molar mass of $KMnO_4$.

Ans. Molar mass of $KMnO_4$ = (A_r of K) + (A_r of (Mn) + 4(A_r of O)
= 39 +55 + 4(16)
= 158 $gmol^{-1}$

(b) Calculate the number of moles in 0.23 g of sodium.

Ans. $$\text{Number of moles} = \frac{\text{mass}}{\text{molar mass}}$$
$$= \frac{0.23}{23 gmol^{-1}}$$
$$= 0.01 \text{ mol}$$

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

Ans. Mass of substance = n × M
Molar mass of $Ca(OH)_2$= 40 + 32 + 2
= 74 g/mol^{-1}
Mass of $Ca(OH)_2$ = 1.5 × 74 g/mol
= 111 g

(d) The given mass of $KC\ell O_3$ is 24.5 g. Calculate the number of moles of potassium chlorate.

Ans. Mass of $KC\ell O_3$ = 24.5 g
Molar mass of $KC\ell O_3$ = 39 + 35.5 +(16) × 3
= 122.5 g/mol
$$n = \frac{m}{M}$$
= 24.5/122.5 = 0.2 mol

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

Ans. Molar mass of H_2O_2 = 34 $gmol^{-1}$
number of molecules $= n \times N_A$
$$= \frac{m}{M} \times N_A$$
$$= \frac{1.75}{34} \times 6.02 \times 10^{23}$$
$= 3.01 \times 10^{22}$ molecules

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

Ans. Molar mass of gold =197 $gmol^{-1}$
number of particles $= n \times N_A$
$$= \frac{m}{M} \times N_A$$
$$= \frac{15}{197} \times 6.02 \times 10^{23}$$
$= 4.57 \times 10^{22}$ atoms

(i) **Stoichiometric calculations obey**

(A) Law of conservation of mass (B) Law of definite proportions
(C) Distribution law **(D) Both A and B**

(ii) **Number of moles present in 20 g of NaOH**

(A) 1 **(B) 0.5** (C) 0.25 (D) 1.5

EXTENSIVE QUESTION

Define mole and Avogadro's number. Explain with two examples of each.

- 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., 6.02×10^{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 H = 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
- 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-}

Calculation of number of moles (by using number particles)

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

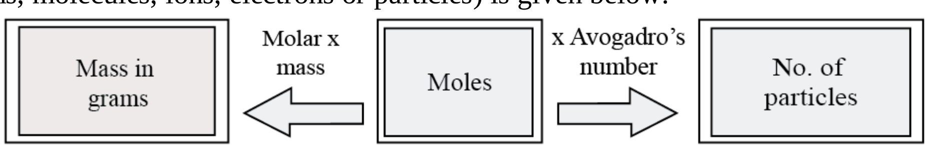
$$\text{Number of moles} = \frac{\text{Number of particles of substance}}{\text{Avogadro's number}}$$

SHORT QUESTION

How many atoms are present in 2.3 g of Na metal?

Relationship between moles and number of particles

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:



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

$$\text{No. of moles of glucose} = \frac{3.76 \times 10^{24}\,\text{molecules}}{6.02 \times 10^{23}\,\text{molecules mol}^{-1}}$$
$$= 6.25 \text{ moles}$$

Sample Problem 4.5

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

Solution

$$\text{No. of moles of sodium} = \frac{2.3}{23.0} = 0.1\,\text{mol}$$

Number of atoms of sodium = Number of moles of sodium × N_A
$= 0.1 \text{ x } 6.02 \times 10^{23}$
$= 0.602 \times 10^{23}$ atoms

Interesting Information

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

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.

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

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.

(i) **Juglone is a dye and produce from the hacks of black walnuts, its formula is**
(A) $C_6H_{12}O_6$ (B) $C_{12}H_{22}O_{11}$ **(C) $C_{10}H_6O_3$** (D) $C_{10}H_{12}O_5$

(ii) **12 g of carbon contains same number of atoms as**

(A) 4g He (B) 1 g of H_2 (C) 16 g of O_2 **(D) All of these**

EXTENSIVE QUESTION

Explain relationship between molar mass, moles and Avogadro's number with an example.

QUICK CHECK 4.2

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

Ans. Number of moles of $Cu = \frac{\text{No. of particles}}{\text{Avogadro's number}}$

$$= \frac{27.10\times10^{25}}{6.02\times10^{23}}$$

$= 4.50 \times 10^2$ 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?

Ans. Number of molecules of isopentyl acetate

Molar mass of isopentyl acetate ($C_7H_{14}O_2$) = (12)7 + (1)14 + (16)2

= 130 gmol

$$\text{Number of molecules} = \frac{m}{M}\times N_A = \frac{1\times10^{-6}\,g}{130\,g/mol}\times6.02\times10^{23}$$

Number of molecules $= 4.6 \times 10^{15}$ molecules

1 molecule of $C_7H_{14}O_2$ contains carbon atoms = 7

4.6×10^{15} molecules of $C_7H_{14}O_2$ contains carbon atom $= 7 \times 4.6 \times 10^{15}$

$= 3.22 \times 10^{16}$

1 molecule of $C_7H_{14}O_2$ contains hydrogen atoms = 14

4.6×10^{15} molecules of $C_7H_{14}O_2$ contains carbon atom $= 14 \times 4.6 \times 10^{15}$

$= 6.44 \times 10^{16}$

1 molecule of $C_7H_{14}O_2$ contains oxygen atoms = 2

4.6×10^{15} molecules of $C_7H_{14}O_2$ contains carbon atom $= 2 \times 4.6 \times 10^{15}$

$= 9.2 \times 10^{15}$

Molar volume

"The volume occupied by one mole of an ideal gas at standard temperature (0°C or 273K) and pressure (1 atm) is called molar volume."

OR

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

SHORT QUESTION

Justify 44g of CO_2 and 16g of CH_4 have equal volume at STP.

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.

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}.

Examples

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

Calculation of volume (by using number of moles)

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.

Volume of a gas = Number of moles × 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 $= 22.4\ dm^3 \times 2.5$

$= 56.0\ dm^3$

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 $= 4.75 \times 22.4$

$= 106.4\ dm^3$

(i) **2.24 dm^3 volume occupied by ______of CO_2**

(A) 44 (B) 22 (C) 11 **(D) 4.4**

(ii) **Equal volume of all the gases at same temperature and pressure contain equal number of molecule is statement of**

(A) Avogadro's law (B) Boyle's law (C) Charles' law (D) Dalton's law

EXTENSIVE QUESTION

3.5 g of an ideal gas occupies volume of 2.8 dm^3 at STP. Calculate the molar mass of that gas.

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

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

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

SHORT QUESTION

Calculate the molar mass of a gas having density of 1.43 g/dm^3 at STP.

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.

Sample Problem 4.9

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

Solution

$$m = d \times V$$

Mass of gas at STP = 1.97 × 22.4 = 44.1g mol^{-1}

QUICK CHECK 4.3

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

Density of gas $= \frac{\text{mass of gas}(m)}{\text{Volume of gas}(V)}$

Density of gas = 1.34 g/dm^3
1dm^3 of gas has mass = 1.34 g
22.4 dm^3 of gas has mass = 30.016 g
If $V = V_m$ then $m = M$
So, molar mass of the gas = 30.016 g/mol

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

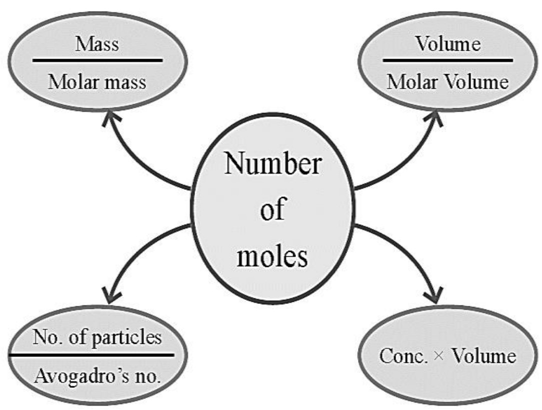
$$n = C \times V$$

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

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

$$\text{Mass of } K_2CO_3 = 27.64\,g$$

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

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

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

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

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

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

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

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 158 $gmol^{-1}$

Ans. Molar concentration $C = \frac{n}{V}$

Number of moles of $KMnO_4$ $= \frac{m}{M}$

Number of moles of $KMnO_4$ $= \frac{7.9\text{ g}}{158\text{ g/mol}}$

$= 0.05$ mol

Molar concentration $= \frac{0.05\,mol}{1dm^3}$

Molar concentration $= 0.05\ moldm^{-3}$

(i) **Which have highest density**

(A) H_2 (B) O_2 (C) N_2 **(D) Cl_2**

(ii) **Which is true for molar concentration**

(A) $C = \frac{n}{V}$ (B) $C = \frac{m}{V}$ (C) $c = \sqrt{\frac{3RT}{M}}$ (D) $C = \frac{M}{V}$

EXTENSIVE QUESTION

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

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.

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

Ans. Moles of CO_2 produced = ?
According to balanced chemical equation
1 mole of glucose produces CO_2 = 6 moles
2.25 mole of glucose produces CO_2 = 6 × 2.25
= 13.50 moles
So, 13.50 moles of CO_2 are produces when 2.25 moles of glucose are used by a person.

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

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

= 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 × 0.02 = 0.04 mol

Mass of Aℓ = n x M

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

= 1.08 g

QUICK CHECK 4.6

(a) **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

Mass of Fe produced from 160 g of Fe_2O_3 = ?

$$\text{mass of } Fe_2O_3 = 160\,g$$

Mass of Fe produced from 160 g of Fe_2O_3 = ?

Solution

To find moles of Fe_2O_3

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

Molar mass (M) of Fe_2O_3 = 2(56) + 3(16) =160g/mol

$$n = \frac{160\,g}{160\,g/mol} = 1\text{ mol}$$

To find moles of Fe produced from Fe_2O_3
According to balanced chemical equation
1 mol of Fe_2O_3 produces Fe = 2 mol
To find mass of Fe produced
Mass of Fe produced = n × M
= 2 mol × 56 g/mol

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:

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

Ans. Volume of methane (CH_4) = 4.5 dm^3
Volume of CO_2 produced = ?
According to volume-volume relationship
at STP 22.4 dm^3 of CH_4 produce = 22.4 dm^3 of CO_2
1 dm^3 of CH_4 produces = 1 dm^3 of CO_2
4.5 dm^3 of CH_4 produce = 4.5 dm^3 of CO_2
So,
4.5 dm^3 of CO_2 will be produced when 4.5 dm^3 of methane is burnt at STP.

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.

$$2LiOH_{(s)} + CO_{2(g)} \longrightarrow Li_2CO_{3(s)} + H_2O(l)$$

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	= No. of mole × 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 moles of LiOH.

Sample Problem 4.15 (Mass-mole Calculations)

Baking soda ($NaHCO_3$) acts as an acid. 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}

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

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.

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

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

Ans. Moles of Cl_2 = 2.25 mol

Mass of $NaOC\ell$

According to balanced chemical equation

1 mol of $C\ell_2$ produces = 1 mol of $NaOC\ell$

2.25 mol of $C\ell_2$ produces = 2.25 mol of $NaOC\ell$

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

= 74.5 g/mol

Mass produced = m × M

= 2.25 mol × 74.5 g/mol

= 167.625 g

Sample Problem 4.16 (Mass-Volume Calculations)

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

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

= 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

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

$= 2.8\text{ dm}^3$

(i) Which of the following stoichiometric relationship is correct?

(A) Mole-Mole (B) Mole-Mass (C) Mole-Volume **(D) All of these**

(ii) Which is not related to stoichiometric calculations?

(A) Reaction is reversible

(B) All the reactants completely converted into products

(C) No side reaction occurs
(D) Law of conservation of mass

EXTENSIVE QUESTION

Define stoichiometry. Write the assumptions and laws for stoichiometric calculations.

In many chemical processes, the quantities of the reactants are usually not present in the proportions indicated by the balanced chemical equation.

Reason to use excess reactant

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

O_2 as an excess reactant in combustion

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.

SHORT QUESTION

Define limiting and excess reactant. Also write down the steps for the determination of limiting reactant.

Limiting reactant

Definition

"This reactant which is consumed earlier is called the limiting reactant."

In this way, the amount of product that forms are 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.

Following numerical problem will make the idea clear.

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 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.

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

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

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

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

$$NH_3 : N_2$$

$$2 : 1$$

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

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

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.

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

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.

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

n = C×V

Moles of Pb^{+2} ions = 0.05 mol dm^{-3} × 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} × 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}):

Mass of $PbSO_4$ = 0.05 mol × 303.3 g mol^{-1}
= 15.2 g

4.7.2 Amount of Product and Unreacted Excess Reagent

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).

Or

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

Example

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. 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. 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, ½ mole of O_2 is required. The remaining ½ 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:

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 2: 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 = 1×1.875 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 produce CO_2 = 1 mol
1.563 moles of oxygen produce CO_2 = 0.5 × 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 3: 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 moles × 44 g mol^{-1}
= 34.39 g

Step 4: 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 = 1.563mol × 0.5
= 0.7815 mol

Step 5: 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 = Quantity taken – Quantity used

= 1.875 mol - 0.7815 mol
=1.0935 mol 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.

Ans. According to balanced chemical equation
1 mol of N_2 = 2 moles of NH_3
When reacts with H_2 produces

a. 1 mol of N_2 can give = 2 moles of NH_3
3 mol of H_2 can give = 2 mole of NH_3
Maximum mol of product = 2 mol (As, there is no L.R)

b. 2 mol of N_2 can give = 4 mol of NH_3
3 mol of H_2 can give = 2 mol of NH_3 (H_2 is L.R)
Maximum moles of product = 2 moles

c. 2 mol of N_2 can give = 2 mol of NH_3
3 mol of H_2 can give = 2 mol of NH_3
1 mol of H_2 can give = $\frac{2}{3}$ mol of NH_3
5 mol of H_2 can give $= \frac{2}{3} \times 5$ = 3.33 mol of NH_3
Maximum mol of product from this mixture = 2 mol

d. 3 mol of N_2 can give = 6 mol of NH_3
3 mol of H_2 can give = 2 mol of NH_3 (L.R)
Maximum moles of product from = 2 mol this mixture

e. Each produces the same amount. (Yes)

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."

Theoretical is larger than actual 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.

SHORT QUESTION

Why actual yield is always less than theoretical 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

Percentage yield

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

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

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}).

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

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 $(CH_3CO)_2O$. 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 mole
138.12g	102.09 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 $= \dfrac{3.00\,g}{138.12\ g\,mol^{-1}} = 0.022$ mol

Mass of salicylic acid = 6.00 g

Number of moles of salicylic acid $= \dfrac{6.00\,g}{102.09\ g\,mol^{-1}} = 0.059$ 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

Actual yield = 3.15 g

$$\%\text{age yield} = \frac{3.15}{3.96} \times 100 = 79.54\ \%$$

(i) Which of the following is true about limiting reactant?

(A) Consumes earlier
(B) Taken in small amount
(C) Control the amount of product
(D) All of these

(ii) Efficiency of a chemical reaction can be expressed in terms of

(A) Actual yield
(B) Theoretical yield
(C) Percentage yield
(D) Unpredictable

EXTENSIVE QUESTION

Define limiting reactant. Also explain the significance of excess reactant in a chemical reaction.

QUICK CHECK 4.10

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

$$CaCO_3 \longrightarrow CaO + CO_2$$

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 lime stone roasted	= 4.5 kg = 4500 g
Mass of quick lime produced (actual yield)	= 2.5 kg = 2500 g
Molar mass of $CaCO_3$	= 100 g mol^{-1}
Molar mass of CaO	= 56 g mol^{-1}

Solution

According to the balanced chemical equation

100 g of $CaCO_3$ should give CaO	= 56 g
1 g of $CaCO_3$ should give CaO	$= \frac{56}{100}$
4500 g of $CaCO_3$ should give CaO	$= \frac{56}{100} \times 4500$
	= 2520 g
Theoretical yield of CaO	= 2520 g
Actual yield of CaO	= 2500 g
$\%\text{yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$	$= \frac{2500\,g}{2520\,g} \times 100 = 99.2\%$

- 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.

SHORT QUESTION

Write any two importance of stoichiometry.

4.9.1 Significance of Stoichiometry in Medicine

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

1. Prepare the antibiotics

In the preparation of antibiotics, the stoichiometry ensures that each dose matches the active ingredient and target bacteria.

2. Determine the cholesterol

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. Determine the glucose level

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. Determine the steroid and other stimulants

To determine the steroid and other stimulants in the urine of athletes. Athletes use steroids and other stimulants to enhance performance and increase strength.

SHORT QUESTION

Elaborate the significance of stoichiometry in medicine industry.

5. Determine the concentration of viral antigens

To determine the concentration of viral antigens in the preparation of vaccine for effective results.

6. Determine the amount and number of drugs

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.

(i) **In which of the following stoichiometry can be uses significantly?**

(A) For preparation of antibiotics
(B) To determine the cholesterol level in blood
(C) to determine the glucose level in blood
(D) All of these

(ii) **An order dose of paracetamol may result**

(A) Blood thickness
(B) Blood thinning
(C) Sever liver damage
(D) Both B and C

EXTENSIVE QUESTION

Write the significance of stoichiometric calculations in medicine.

TEXT BOOK EXERCISE

Q.1 Choose the correct one from the given options.

(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 (At. mass of Fe = 56)
(B) 48 g of magnesium (At. mass of Mg = 24)
(C) 32 g of S_8 molecules (At. mass S = 32)
(D) 44 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 Na_2SO_4 (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}

ANSWER KEY

i	C	ii	D	iii	B	iv	D	v	B
vi	D	vii	A	viii	C	ix	D	x	B
xi	C	xii	C						

SHORT QUESTIONS

Q.2 Answer the following short questions:

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

Ans. Concept of mole

Mole is the amount of substance that contains Avogadro's number of particles:
For example,

$18g \text{ of } H_2O = 6.02\times10^{23}$ molecules of H_2O
$= 1mol \text{ of } H_2O$

$100g \text{ of } CaCO_3 = 6.02\times10^{23}$ formula units of $CaCO_3$ $= 1mol \text{ of } CaCO_3$

So, collection of 6.02×10^{23} particles of any substance (element, molecular compound, ionic compound or ionic specie) is equal to 1 mole of that substance

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

i. **Molar mass**
ii. **Molar volume**
iii. **Molar concentration**

Ans.

i. Molar mass

Mass of 1 mole of any substance is called molar mass
1 mole of H = 1.008 g
1 mole of O_2 = 32 g
1.008 g of H, 32 g of O_2 are Molar masses.

ii. Molar Volume

Volume of one mole of an ideal gas at STP is called molar volume.
It is denoted by V_m.
1 mole of H_2 = 2 g = 22.414 dm^3
1 mole of CO_2 = 44 g = 22.414 dm^3

iii. Molar concentration

It is the number of moles of a substance (reactants or products) dissolved per volume of a solution in dm^3. It is given as mol dm^{-3}. The relationship between number of moles and molar concentration is given by

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

C = molar concentration, n = number of moles, V = volume in dm^3

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

Ans. Molar Volume

Volume of one mole of an ideal gas at STP is called molar volume.
It is denoted by V_m.
1 mole of H_2 = 2 g = 22.414 dm^3

Relationship between Avogadro's number and molar volume

It can be explained on basis of Avogadro's law, which states that, "equal volume of all the ideal gases have equal number of molecules at STP".

$22.414\ dm^3$ of $H_2 = V_m = 1$mole
$= 6.02\times10^{23}$ molecules

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

Ans. Justification

39 g of K and 56 g of Fe have equal no. of atoms because both are equal to one mole, and one mole of any substance contains 6.02×10^{23} particles.
39 g of K = 1 mole of K = 6.02×10^{23} atoms
56 g of Fe = 1 mole of Fe = 6.02×10^{23} atoms

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 the three gases are very different from each other.

Ans. Reason

4 g of He, 17 g of NH_3 and 64 g of SO_2 are 1mole of He, NH_3 and SO_2 respectively. One mole of an ideal gas occupies 22.414 dm^3 of volume
4 g of He = 1 mole of He = 22.414 dm^3
17 g of NH_3 =1 mole of NH_3 = 22.414 dm^3
64 g of SO_2 = 1 mole of SO_2 = 22.414 dm^3
1mole of He, NH_3 and SO_2 have equal volume because masses and sizes of molecules does not affect the volume of the gases.

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?

Ans. Justification

Yes, 1 mole of H_2 and 1 mole of NH_3 at $0°C$ 1 atm (STP) have Avogadro's number of particles (molecules) as 1 mole of any substance contain N_a particles.
1 mole of H_2 = 6.02×10^{23} molecules
1 mole of NH_3 = 6.02×10^{23} 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.

Basic assumptions

(i) All the reactants are completely converted into products.
(ii) No side reaction occurs.
(iii) Law of conservation of mass and law of definite proportions are obeyed.

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

Ans. Limiting reactant

The reactant which controls the amount of product formed in a chemical reaction and is consumed earlier is called limiting reagent or reactant

Reason

Limiting reactant controls the amount of product due to its complete and earlier consumption.

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

Ans.

Actual Yield	Theoretical yield
Amount of product obtained by performing chemical reaction	Amount of product calculated from balanced chemical equation.
It is always lesser than theoretical yield.	It always greater than actual yield.

Percentage yield of reaction

It can be calculated from actual yield and theoretical yield as follows:

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

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

Ans. Factors

Following factors are responsible for low yield of products

(i) Mechanical loss of products due to inexperienced worker as the processes like filtration, separation by distillation, separation by separating funnel, washing, drying and crystallization is not properly carried out reduces the yield.
(ii) Side reaction may occur.
(iii) Reversibility of reaction.
(iv) Presence of impurities.
(v) Due to faulty apparatus.

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.

Limiting reactant	Non-limiting reactant
The reactant which is consumed earlier in the reaction.	The reactant which is not completely consumed in reaction.
It controls the amount of products (yield)	It does not control the amount of products (yield).

Determination of limiting reactant

(For answer to this question, see article 4.7).

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

Ans. (For answer to this question, see article 4.8).

NUMERICALS

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.

Ans.

Molar mass of NaOH=23+16+1=40 g mol^{-1}

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

Solution

$$\text{NaOH} = \frac{\text{mass}}{\text{Molar mass}}$$

Molar mass of NaOH=23+16+1=40 g mol^{-1}
So,

$$n=\frac{2.00\ g}{40\ gmol^{-1}}=0.05\ mol$$

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

Given data

Mass of NaOH = 2.00 g
Volume of solution = 250 cm^3

$$=\frac{250}{1000}\ dm^3$$
$$=0.250\,dm^3$$

Solution

$$\text{Concentration}=\frac{n(mol)}{V(dm^3)}$$
$$=\frac{0.05\ mol}{0.250\ dm^3}=0.2\ moldm^{-3}$$

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?

Solution.

Concentration of solution if volume becomes 500 cm^3 = C_2 = ?

$$V_2=\frac{500}{1000}dm^3=0.5dm^3$$
$$C_2=\frac{n(mol)}{V_2(dm^3)}$$
$$C_2=\frac{0.05}{0.5}=0.1\ moldm^{-3}$$

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

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

In an experiment, 34.0 g of ammonia is reacted with 96.0 g of oxygen.

a) Determine the limiting reactant.

Given data

$4NH_{3(g)}+5O_{2(g)}\longrightarrow 4NO_{(g)}+6H_2O_{(g)}$
mass of NH_3 = 34.0 g
mass of oxygen = 80 g
Limiting reactant = ?

Solution

Number of moles of $NH_3=\frac{34\ g}{17\ gmol^{-1}}$
= 2 moles

Number of moles of $O_2=\frac{96\ g}{32\ gmol^{-1}}$
= 3 moles

To find moles of NO that can be formed from NH_3
Compare the moles of NH_3 and NO

NH_3	:	NO
4	:	4
1	:	1
2	:	2

To find moles of NO that can be formed from O_2
Compare the moles of O_2 and NO

O_2	:	NO
5	:	4
1	:	$\frac{4}{5}$
3	:	$\frac{4}{5}\times 3$
3	:	2.4

As, NH_3 produces less number of moles of NO (2 mole). So, NH_3 is limiting reactant and oxygen is excess reactant.

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

Solution

Maximum mass of NO

mass = Moles × Molar mass

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

mass of NO = 2 mol×30 g mol^{-1}

Mass of NO = 60 g

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)

Solution

Since oxygen is limiting reactant so 2.5 mole of O_2 is used.

O_2	:	NO
5	:	4
2.5	:	2

Out of 5 moles of oxygen 2.5 moles are used and 2.5 moles are left as excess reagent

Mass of excess of oxygen = 2.5 × 32 = 80 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:

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

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)

Given data

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

Mass of $Fe_2O_3 = 1\text{kg}$

$= 1\times1000\text{g} = 1000\text{g}$

Mass of $Fe_{(s)}\left(\text{Actual yield}\right) = 650\text{g}$

Required

%age yield =?

Solution

$$\%\text{ age yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}\times100$$

To find moles of Fe_2O_3 and Fe

Number of moles of $Fe_2O_3 = \dfrac{\text{mass}}{\text{Molar mass}}$

Molar mass of $Fe_2O_3 = 2(56)+3(16)$

$= 112 + 48$

$= 160\text{ gmol}^{-1}$

No. of moles of $Fe_2O_3 = \dfrac{1000}{160} = 6.25\text{ mol}$

To find theoretical yield

Compare the moles of Fe_2O_3 and Fe

Fe_2O_3	:	Fe
1	:	2
6.25	:	2×6.25
6.25	:	12.50

So the theoretical yield of Fe is 12.50 mol.

Mass of $Fe = 12.50\text{mol}\times56\text{ g mol}^{-1} = 700\text{ g}$

$$\%\text{age yield} = \frac{650}{700}\times100$$

$= 92.85\%$

Q.8 1.5 g 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.

$$\mathbf{2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O}$$

Also calculate

a) No. of molecules of 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

$C_2H_6 = 1.5$ g

Volume of CO_2 = ?

Solution

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

Molar mass of C_2H_6 = 2(Ar of C) +6 (Ar of H)

= 2(12) +6 (1)

= 24 + 6

= 30 g mol^{-1}

No. of moles of C_2H_6 = $\begin{array}{c} \text{H}\quad\text{H} \\ |\quad\;| \\ \text{H}-\text{C}-\text{C}-\text{H} \\ |\quad\;| \\ \text{H}\quad\text{H} \end{array}$

$= \dfrac{1.5}{30}$

= 0.05 moles

C_2H_6	:	CO_2
2	:	4
0.05	:	1

Since 1 mole of CO_2 is produced so its volume produced will be 22.414 dm^3

a)

Given data

$C_2H = 0.5$ mole

No. of molecule of CO_2 produced = **?**

Solution

2 moles of C_2H_6 produce $CO_2 = 4$ moles

0.05 moles of C_2H_6 produce $CO_2 = \frac{4}{2} \times 0.05$

$= 0.1$ moles

No. of CO_2 molecule $=$ moles $\times N_A$

No. of CO_2 molecules $= 0.1 \times 6.02 \times 10^{23}$

No. of CO_2 molecules $= 6.02 \times 10^{22}$ molecules

b)

Given data

$C_2H_6 = 0.5$ mole

No. of O_2 molecules reacted = ?

Solution

2 moles of C_2H_6 with O_2 with = 1 moles

No. of O_2 molecule $=$ moles $\times N_A$

$= 7 \times 6.01 \times 10^{23}$

No. of O_2 molecules $= 4.2 \times 10^{24}$ molecules

c)

Given data

$C_2H_6 = 0.5$ mole

No. of CH bond broken = ?

Solution.

$$
\begin{array}{ccccccc}
 & & H & & H & & \\
 & & | & & | & & \\
H & - & C & - & C & - & H \\
 & & | & & | & & \\
 & & H & & H & &
\end{array}
$$

Total CH bonds in $C_2H_6 = 6$

As this reaction uses 2 moles of C_2H_6 so

No. of CH bonds broken

$= 6 \times 2$

$= 12$ moles

$= 12 \times 6.02 \times 10^{23}$

Total CH bonds broken $= 7.2 \times 10^{24}$

SELF IMPROVEMENT TEST

Time: 35 Minutes | **OBJECTIVE TYPE** | **Total Marks: 25**

Q.No.1 Choose the Correct Answer by Filling the Circle. **(7 × 1 = 7)**

1. **Which one of the following has maximum mass?**

Ⓐ **0.1 mol of Cl_2** Ⓑ **0.5 mol of NH_3** Ⓒ **0.25 mole of O_2** Ⓓ **0.5 mol of CO**

2. **The percentage yield of a reaction can be determined by**

Ⓐ Amount of reactants Ⓑ Amount of product

Ⓒ Theoretical yield Ⓓ Amount of excess reactant

3. **Which is not true for concept of limiting reactant?**

Ⓐ Only applicable to irreversible reactions

Ⓑ Present in less amount than stoichiometric amount

Ⓒ Consumed earlier

Ⓓ None

4. **Mass of 2 mole of chlorine is**

Ⓐ 71 g Ⓑ 71 amu Ⓒ 142 g Ⓓ 35.5 g

5. **The number of moles of CO which contain 8.0 g of oxygen**

Ⓐ 0.25 Ⓑ 0.50 Ⓒ 1.0 Ⓓ 1.50

6. **If the density of a gas is 1.43 g/dm^3, what is the molar mass?**

Ⓐ 32g/mol Ⓑ 12g/mol Ⓒ 22g/mol Ⓓ 44.g/mol

7. **Mass of ammonia produced, when 4 g of hydrogen reacts with 28 g of nitrogen.**

Ⓐ 17 g Ⓑ 22.5 g Ⓒ 25.5 g Ⓓ 30 g

Q.No.2 Write Short Answer. **(5 × 2 = 10)**

(i) Give the steps involved in identification of limiting reactant.

(ii) What is molar volume? Give its importance.

(iii) Find the number of moles of oxygen atoms in 0.98 grams of sulphuric acid.

(iv) One mole of different compounds has different masses but the same number of molecules. Justify the statement.

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

SECTION II

Q.No.3 Extensive Questions. **(2 × 4 = 8)**

(a) When 30 gram of silica reacts with 12 gram of carbon, silicon carbide is produced, an important ceramic material. **(4)**

$$SiO_2 + C \longrightarrow SiC + 2CO$$

Calculate which reactant is limiting reactant and how much silicon carbide is produced?

(b) What is stoichiometry? Give the basic assumptions of stoichiometric calculations. **(4)**