Basic Concepts of Chemistry and Chemical Calculations

Relative Atomic mass:

The relative atomic mass is defined as the ratio of the average atomic mass factor to the unified atomic mass unit. (Or)

The ratio of the average mass factor to one twelfth of the mass of an atom of carbon-12

$$\textbf{Relative atomic mass (Ar)} = \frac{\textbf{Average mass of the atom}}{\textbf{Unified atomic mass}}$$

Example: Relative atomic mass of hydrogen $(Ar)_H = \frac{\text{Average mass of H atom (in Kg)}}{\text{Unified atomic mass}}$

$$= \frac{1.6736 \times 10^{-27}\ kg}{1.6605 \times 10^{-27}\ kg} = 1.0078 \approx 1.008 \text{ u.}$$

Average atomic mass:

Average atomic mass is defined as the average of the atomic masses of all atoms in their naturally occurring isotopes.

Example: Chlorine consists of two naturally occurring isotopes $_{17}Cl^{35}$ and $_{17}Cl^{37}$ in the ratio 77 : 23, the average relative atomic mass of chlorine is

$$\textbf{Average atomic mass } \overline{A} = \frac{A_1 X_1 + A_2 X_2}{X_1 + X_2}$$

$$= \frac{(35 \times 77) + (37 \times 23)}{100} = 35.46 \text{ u}$$

(or) Average atomic mass $\overline{A} = \frac{\sum A_i X_i}{\sum X_i}$

A-atomic mass, X- percentage

- An element X has the following isotopic composition ^{200}X = 90 %, ^{199}X = 8 % and ^{202}X = 2 %. The weighted average atomic mass of the element X is closest to **(Q. No. 2)**

$$\text{Average atomic mass } \overline{A} = \frac{A_1 X_1 + A_2 X_2}{X_1 + X_2} = \frac{(200 \times 90) + (199 \times 8) + (202 \times 2)}{100} = 199.96 \text{ u}$$

$$= 200 \text{ u}$$

Molecular Mass:

Relative molecular mass is defined as the ratio of the mass of a molecule to the unified atomic mass unit.

The relative molecular mass of any compound can be calculated by adding the relative atomic masses of its constituent atoms.

Relative molecular mass of H_2 molecule **(H_2)** = 2 × (relative atomic mass of hydrogen atom)

= 2 × 1.008 u = 2.016 u.

Calculate the molecular mass (or) Formula mass of the following.

i. Ethanol(C_2H_5OH)

= (2x12) + (1x5) + 16 + 1

= 24 + 5 + 16 + 1 = 46 u

ii. Potassium permanganate ($KMnO_4$)

=39 + 55 + (4x16) = 158 u

iii. Potassium dichromate ($K_2Cr_2O_7$)

= (2x39) + (2x52) + (7x16)

=78+104+112 = 294 u

iv. Sucrose ($C_{12}H_{22}O_{11}$)

= 144 + 22 + 176 = 342 u

Mole Concept:

One mole is the amount of substance of a system, which contains as many elementary particles as there are atoms in 12 g of carbon-12 isotope. The elementary particles can be molecules, atoms, ions, electrons or any other specified particles.

Avogadro Number:

The total number of entities present in one mole of any substance is equal to 6.022 x 10^{23}. This number is called Avogadro number.

Avogadro number does not have any unit.

∴ ***The number of atom (or) Molecule = Mole x Avogadro Number***

Example: How many molecules present in two moles of CO_2

= 2 x 6.022 x 10^{23} = 12.044 x 10^{23}

- Which contains the greatest number of moles of oxygen atoms? **(Q. No. 33)**

  i) 1 mol of ethanol ii) 1 mol of formic acid iii) 1 mol of H_2O

  Answer:

Compound	Given no. of moles	No. of oxygen atoms
Ethanol - C_2H_5OH	1	$1 \times 6.022 \times 10^{23}$
Formic acid - HCOOH	1	$2 \times 6.022 \times 10^{23}$
Water - H_2O	1	$1 \times 6.022 \times 10^{23}$
Answer : Formic acid		

Molar Mass:

Molar mass is defined as the mass of one mole of a substance.

$$\text{Molar mass} = \frac{\text{Mass}}{\text{Mole}} \qquad \therefore \ \text{Mole} = \frac{\text{Mass}}{\text{Molar mass}}$$

The molar mass of a compound is equal to the sum of the relative atomic masses of its constituents expressed in g mol^{-1}

Calculate the molar mass of the following compounds. (Q. No. 31)

i. Urea [$CO(NH_2)_2$]
= 12 + 16 + 2[14 + (2x1)]
= 12 + 16 + 2[16] = 60 g/mol

ii. Acetone [CH_3COCH_3]
=12+3+12+16+12+3 = 58 g/mol

iii. Boric acid [H_3BO_3]
=3+10+48 = 61 g mol^{-1}

iv. Sulphuric acid [H_2SO_4]
=2+32+64
= 98 g mol^{-1}

Molar Volume:

The volume occupied by one mole of any substance in the gaseous state at a given temperature and pressure is called molar volume.

$$\text{Molar Volume} = \frac{\text{Volume}}{\text{Volume at one Mole}} \ (or) \ \frac{\text{Volume}}{22.4\ L}$$

At 273 K and 1 bar pressure (STP)	*- 22.71 L (Standard volume per litre)*
At 273 K and 1 atm pressure (SATP)	*- 22.4 L (Standard volume per litre)*
At 298 K and 1 atm pressure	*- 24.5 L (Room Temperature & pressure i.e 25°C)*

Calculate the number of molecules of oxygen gas that occupies a volume of 224 ml at 273 K and 3 atm pressure.

Ans: Volume of Oxygen = 224 ml = 0.224 L

∴ No. of moles of Oxygen (n) = Volume / 22.4 L

$$= \frac{0.224\ L\ X\ 3\ atm}{22.4\ L} = 0.03 \text{ mole}$$

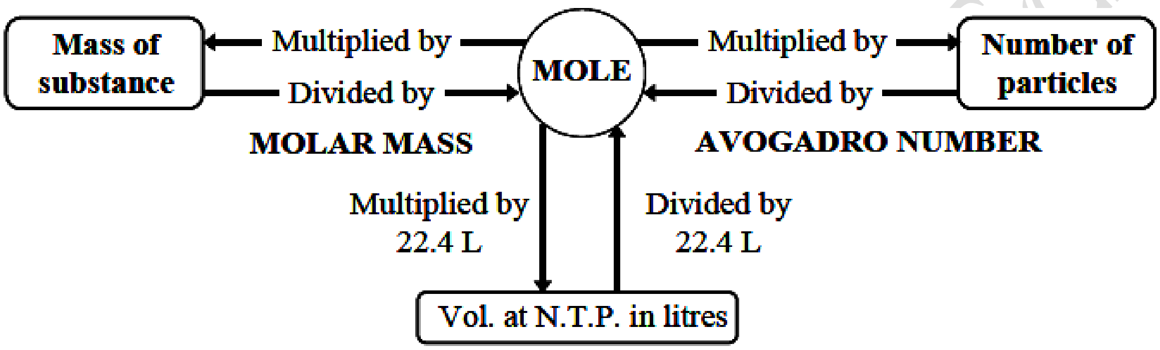
The mass of a gas that occupies a volume of 612.5 ml at room temperature and pressure (25°C and 1 atm pressure) is 1.1g. The molar mass of the gas is (Q. No. 21)

Molar mass = Mass / no. of moles

∴ Moles = Volume / 24.5 L = 0.6125 L / 24.5 L = 0.025 moles

∴ Molar mass = 1.1g / 0.025 moles =44 g mol^{-1}

<u>Interconverting Moles, Mass, Volumes and Number of Chemical Entities:</u>



Calculate the number of moles present in 9 grams of ethane.

Ans: Molar mass of ethane, C_2H_6 = (2 x 12) + (6 x 1) = 30 g mol^{-1}

No. of moles (n) = mass / molar mass = 9 g / 30 g mol^{-1} = 0.3 mole

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$$No.of\ atoms\ or\ molecule = \frac{Mass\ of\ substance\ (molar\ mass)}{Mass\ of\ single\ atom\ or\ molecule}$$

$$\therefore \boldsymbol{Mass\ of\ substance\ (molar\ mass)}$$
$$= \boldsymbol{Mass\ of\ single\ atom\ or\ molecule\ X\ Avogadro\ Number}$$

Mass of one atom of an element is 6.645 x 10^{-23} g. How many moles of element are there in 0.320 kg. (Q. No. 36)

Ans: Given: mass of one atom = 6.645 × 10^{-23} g

∴ Mass of 1 mole of atom (molar mass) = 6.645 × 10^{-23} g × 6.022 × 10^{23} = 40 g

∴ Number of moles of element in 0.320 kg = mass / malar mass

= 320 g / 40 g mol^{-1}

= 8 mol

The number of water molecules in a drop of water weighing 0.018 g is (Q. No. 6)

Ans: Weight of the water drop = 0.018 g

No. of moles of water in the drop = Mass of water / molar mass

= 0.018 / 18 = 10^{-3} mole

∴ No. of water molecules in one drop of water (10^{-3} mole) = mole X Avogadro number

= 10^{-3} x 6.022 x 10^{23}

= 6.022 x 10^{20}

Which one of the following represents 180g of water? (Q. No. 13)

No. of moles of water present in 180 g = Mass of water / Molar mass of water

= 180 g / 18 g mol^{-1} = 10 moles

One mole of water contains = 6.022 x 10^{23} water molecules

10 mole of water contains = 10 x 6.022 x 10^{23} = 6.022 x 10^{24} water molecules

Total number of electrons present in 1.7 g of ammonia is (Q. No. 15)

No. of moles of 1.7 g of ammonia = 1.7 g / 17 g mol^{-1} = 0.1 moles

∴ No. of NH_3 molecules in 0.1 mole = mole X Avogadro number

= 0.1 x 6.022 x 10^{23} = 6.022 x 10^{22} molecules

No. of electrons present in one ammonia (NH_3) molecule = 7 + 3 = 10 electrons

∴ No. of electrons present in 6.022 x 10^{22} molecules (1.7 g) of ammonia is

= 6.022 x 10^{22} x 10 = 6.022 x 10^{23} electrons.

Two 22.4 litre containers A and B contains 8 g of O_2 and 8 g of SO_2 respectively at 273 K and 1 atm pressure, then (Q. No. 19)

No, of moles of oxygen = mass / malar mass = 8 g/32 g = 0.25 moles of oxygen

No. of moles of sulphur dioxide = 8 g / 64 g = 0.125 moles of sulphur dioxide

∴ Ratio between the no. of molecules = 0.25: 0.125 = 2:1

Equivalent mass:

Gram equivalent mass of an element or compound or ion is the mass that combines or displaces 1.008 g hydrogen or 8 g oxygen or 35.5 g chlorine.

$$Gram\ equivalent\ \ \text{mass} = \frac{Molar\ mass\ (g\ mol^{-1})}{Equivalence\ factor\ (eq\ mol^{-1})}$$

i. **Equivalent mass of the Acid** = $\frac{\textit{Molar mass of the Acid}}{\textit{Basicity of the Acid}}$

` (Basicity is no. of moles of ionisable H^+ ions present in 1 mole of the acid)

Example: **Equivalent mass of H_2SO_4**

Molar mass of the H_2SO_4 = 98 g mol^{-1}

∴ Equivalent mass = 98 / 2 = 49 g eq^{-1}

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ii. **Equivalent mass of the Base** = $\frac{\textit{Molar mass of the Base}}{\textit{Acidity of the Base}}$

(Acidity is no. of moles of ionisable OH^- ion present in 1 mole of the base)

Example: **Equivalent mass of KOH**

Molar mass of the KOH = 39 + 16 + 1 = 56 g mol^{-1}

∴ Equivalent mass = 56 / 1 = 56 g mol^{-1}

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iii. **Equivalent mass of Oxidising agent** = $\frac{\textit{Molar mass of the Oxidising agent}}{\textit{no. of electrons gained by one molecule}}$

Example: **The equivalent mass of potassium permanganate in alkaline medium is**

$$MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$$ **(Q. No. 12)**

Ans: The reduction reaction of the oxidising agent ($MnO4^-$) involves gain of 3 electrons. Hence the equivalent mass = (Molar mass of $KMnO_4$) / 3

= 158.1 / 3 = 52.7 g eq^{-1}

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iv. **Equivalent mass of Reducing agent** = $\frac{\textit{Molar mass of the Reducing agent}}{\textit{no. of electrons lost by one molecule}}$

Example: **The equivalent mass of potassium permanganate in alkaline medium is**

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

Ans: The oxidation reaction of the redusing agent (MnO_4^-) involves gain of 5 electrons. Hence the equivalent mass = (Molar mass of $KMnO_4$) / 5

= 158.1 / 5 = 31.6 g eq^{-1}

Calculate the equivalent mass of potassium dichromate. The reduction half-reaction in acid medium is, $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ (Evaluate Q, 4 .b)

Ans: Equivalent mass of a oxidising agent = molar mass / number of moles of electrons gained

$= 292.2\ g\ mol^{-1} / 6\ eq\ mol^{-1} = 48.7\ eq^{-}$

The equivalent mass of a trivalent metal element is 9 g eq^{-1} the molar mass of its anhydrous oxide is (Q. No. 5)

Ans: Equivalent mass = mass of the metal / valance factor

∴ Mass of the metal = Equivalent mass x valance factor = 9 g eq^{-1} x 3 eq = 27 g.

∴ Mass of the Oxide formed (M_2O_3) = (2 x 27) + (3 x 16) = 102 g.

Determination of equivalent masses of elements

Equivalent masses can be determined by the following methods:

1. Hydrogen displacement method 2. Oxide method 3. Chloride method

1. Hydrogen displacement method:

$$Equivalent\ \ mass\ of\ the\ Metal = \frac{Mass\ of\ the\ Metal}{Combines\ or\ displaces\ mass\ of\ Hydrogen}\ X\ 1.008$$

2. Oxide method:

$$Equivalent\ \ mass\ of\ the\ Metal = \frac{Mass\ of\ the\ Metal}{Combines\ or\ displaces\ mass\ of\ Oxygen}\ X\ 8$$

3. Chloride method:

$$Equivalent\ \ mass\ of\ the\ Metal = \frac{Mass\ of\ the\ Metal}{Combines\ or\ displaces\ mass\ of\ Chlorine}\ X\ 35.5$$

0.456 g of a metal gives 0.606 g of its chloride. Calculate the equivalent mass of the metal. (Evaluate Q, 4 .a)

Ans: Mass of the metal = 0.456 g

Mass of the metal chloride = 0.606 g

∴ Mass of chlorine = 0.15 g

$$Equivalent\ \ mass\ of\ the\ Metal = \frac{Mass\ of\ the\ Metal}{Combines\ or\ displaces\ mass\ of\ Chlorine}\ X\ 35.5$$

$= (0.456\ g / 0.15\ g) \times 35.5 = 107.92\ g\ eq^{-1}$

Carbon forms two oxides, namely carbon monoxide and carbon dioxide. The equivalent mass of which element remains constant? (Q. No. 4)

Ans: Reaction 1 : $2\,C + O_2 \rightarrow 2\,CO$

2 × 12 g carbon combines with 32 g of oxygen. Hence,

$$Equivalent\ \ mass\ of\ the\ Carbon = \frac{Mass\ of\ the\ Carbon}{Combines\ or\ displaces\ mass\ of\ Oxygen}\ X\ 8$$

$$= \frac{24}{32}\ \text{x}\ 8 = 6\ \text{g eq}^{-1}$$

Reaction 2 : $C + O_2 \rightarrow CO_2$ here 12 g carbon combines with 32 g of oxygen. Hence,

Equivalent mass of Carbon $= \frac{12}{32}\ \text{x}\ 8 = 3\ \text{g eq}^{-1}$

Carbon equivalent mass is change but Oxygen is same.

***The equivalent mass of the salt is equal to its molar mass. E.g KCℓ* = 39.1+ 35.5 = 74.6 g eq^{-1}**

STOICHIOMETRIC CALCULATIONS:

In Greek, stoicheion means *element* and metron means *measure* that is, stoichiometry gives the numerical relationship between chemical quantities in a balanced chemical equation.

$$\textbf{\textit{Mole / Mass / Volume of the reactant / product}} = \frac{\textbf{\textit{Which one required}}}{\textbf{\textit{Which one compared}}}\ \textbf{x}\ \textbf{\textit{How much}}$$

(Based on Stoichiometric Equation)

Types of Stoichiometric relationships:

On the basis of unit, stoichiometric problems may be classified into following relationships.

I. Mole – Mole relationship

Mole of the reactant / product is to be calculated, if mole of the other one is given.

Example: **How many moles of hydrogen is required to produce 10 moles of ammonia?**

Ans: $N_{2(g)} + 3\,H_{2\ (g)} \rightarrow 2\,NH_{3\ (g)}$

2 moles of ammonia produced by 3 moles of Hydrogen

∴ 10 moles of ammonia will be produced by $= \frac{3}{2}$ x 10 = 15 moles of Hydrogen.

II. Mass – Mass relationship

Mass of the reactant / product is to be calculated, if Mass of the other one is given.

Example: **Calculate the amount of water produced by the combustion of 32 g of methane.**

Ans:

$$\underset{16\ g}{CH_{4\,(g)}} + 2O_{2(g)} \rightarrow CO_{2(g)} + \underset{36\ g}{2H_2O_{(g)}}$$

16 g of methane to produce 36 g of water

$\therefore$ 32 g of methane will produce water is = $\frac{36}{16}$ x 32 = 72 g of water.

III. Volume – Volume relationship

Volume of the reactant / product is to be calculated, if Volume of the other one is given.

Example: **How much volume of chlorine is required to form 11.2 L of HCl at 273 K and 1 atm pressure?**

Ans:

$$H_{2\,(g)} + \underset{22.4\ L}{Cl_{2\,(g)}} \rightarrow \underset{44.8\ L}{2\ HCl_{(g)}}$$

44.8 L of HCl produced by 22.4 L of Chlorine

$\therefore$ 11.2 L of HCl will produced by the required Chlorine is = $\frac{22.4}{44.8}$ x 11.2

= 5.6 litre of Chlorine

IV. Mass – Mole relationship

Mass / Mole of the reactant / product is to be calculated, if we have the mole or mass of other.

Example: **How many moles of ethane is required to produce 44 g of CO_2 (g) after combustion. (Q. No. 40)**

Method 1

$$2\ C_2H_{6\,(g)} + 7\ O_{2\,(g)} \rightarrow 4\ CO_{2\,(g)} + 6\ H_2O_{(g)}$$

No. of moles of 44 g CO_2 = mass / mole

= 44 / 44 = 1 mole

4 moles of CO_2 produce by 2 moles of ethane

$\therefore$ 1 mole of CO_2 will produce by = $\frac{2}{4}$ x 1

= 0.5 moles of ethane is required

Method 2

$$\underset{2\ mole}{2\ C_2H_{6\,(g)}} + 7\ O_{2\,(g)} \rightarrow \underset{176\ g}{4CO_{2\,(g)}} + 6H_2O_{(g}$$

176 g of CO_2 produced by 2 moles of ethane

$\therefore$ 44 g of CO_2 will produced by = $\frac{2}{176}$ x 44

= 0.5 moles of ethane is required

V. Mass – Volume relationship

Mass / Volume of the reactant / product is to be calculated, if we have the volume or mass of other.

Example: **How much volume of carbon dioxide is produced when 50 g of calcium carbonate is heated completely under standard conditions?**

Ans: $\underset{100\ g}{CaCO_{3\,(s)}} \xrightarrow{\Delta} CaO_{(s)} + \underset{22.7\ L\ (at\ STP)}{CO_{2\,(g)}}$

1 mole (100g) $CaCO_3$ on heating at std condition to produces 22.7 L CO_2

$\therefore$ 50 g $CaCO_3$ will produce CO_2 is = $\frac{22.7}{100}$ x 50 = 11.35 L of CO_2

METHOD 2 (all values converts to moles or litre)

Ans: $\underset{100\ g\ (1\ mol)}{CaCO_{3\,(s)}} \xrightarrow{\Delta} CaO_{(s)} + \underset{1\ mol}{CO_{2\,(g)}}$

No. of moles of 50 g $CaCO_3$ = mass / molar mass = 50 / 100 = 0.5 mole.

1 mole of $CaCO_3$ produces 1 mole of CO_2

$\therefore$ 0.5 mole of $CaCO_3$ will produce = $\frac{1}{1}$ x 0.5 = 0.5mol of CO_2

$\therefore$ At STP the volume of 0.5 mole CO_2 is = mole x molar volume = 0.5 x 22.7 = 11.35 L

Limiting Reagents:

When a reaction is carried out using non-stoichiometric quantities of the reactants, the product yield will be determined by the reactant that is completely consumed.

Limiting Reagents: The reactant that is completely used up or reacted .

- Is used up first
- Stops the reaction

Excess Reagents: The reactant that are not use up when reaction is finished.

"Products formation is based only **Limiting Reagents**"

Sodium Chloride can be prepared by the reaction of Sodium metal with Chlorine gas

$2\,Na_{(g)} + Cl_{2(g)} \longrightarrow 2\,NaCl_{(s)}$ .Suppose that 6.70 mol Na reacts with 3.20 mol Cl_2.

(i) What is limiting reagent

(ii) How many moles of NaC*l* are produced?

Ans: i) $2\,Na_{(g)} + Cl_{2(g)} \longrightarrow 2\,NaCl_{(s)}$

1 mole of Cl_2 reacted with 1 mol of Na

∴ The moles Cl_2 required to react with 6.7 moles Na is = $\frac{1}{2}$ x 6.7

= 3.35 moles of Cl_2 required

So, **Chlorine** is limiting reagents.

ii) Products formation is based only **Limiting Reagents**

So, 1 mole of Cl_2 produce 2 mole NaC*l*

∴ 3.35 moles of Cl_2 will produce NaC*l* is = $\frac{2}{1}$ x 3.35

= 6.7 moles of NaC*l*

The balanced equation for a reaction is given below: $2x+3y \rightarrow 4\,l + m$

When 8 moles of x react with 15 moles of y, then

i) Which is the limiting reagent?

ii) Calculate the amount of products formed.

iii) Calculate the amount of excess reactant left at the end of the reaction.

Ans:

Content	Reactant		Products	
	x	y	*l*	m
Stoichiometric coefficient	2	3	4	1
No. of moles allowed to react	8	15	–	–
No. of moles of reactant reacted and product formed	8	12	16	4
No. of moles of un-reacted reactants and the product formed	–	3	16	4

Limiting reagent : x

Product formed : 16 moles of *l* & 4 moles of m

Amount of excess reactant : 3 moles of y

Urea is prepared by the reaction between ammonia and carbon dioxide as follows.

$2\ NH_{3\ (g)} + CO_{2\ (g)} \rightarrow H_2N–CO–NH_{2\ (aq)} + H_2O_{(l)}$. In a process, 646 g of ammonia is allowed to react with 1.144 kg of CO_2 to form urea.

i) Which is the limiting reagent?

ii) Calculate the quantity of urea formed and unreacted quantity of the excess reagent.

Ans:

Content	Reactant		Products	
	NH_3	CO_2	Urea	H_2O
Stoichiometric coefficient	2	1	1	1
Amount of reactant allowed to react	646 g	1144 g	1	1
No. of moles allowed to react Moles = Mass / Molar mass	646 / 17 = 38	1144 / 44 = 26	–	–
No. of moles of reactant reacted and product formed	38	19	19	19
No. of moles of un-reacted reactants and the product formed	–	7	19	19

i) So, ammonia is the limiting reagent. Some quantity of CO_2 remains unreacted.

ii) Quantity of urea formed

= number of moles of urea formed × molar mass of urea

= 19 moles × 60 g mol^{-1}

= 1140 g = 1.14 kg

iii) Excess reagent leftover at the end of the reaction is carbon dioxide.

= 7 moles × 44 g mol^{-1} = 308 g

The reaction between A*l* and ferric oxide can generate temperatures up to 3273 K and is used in welding metals. (Atomic mass of A*l* = 27 u , Atomic mass of O = 16 u)

$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$. If in this process, 324 g of aluminium is allowed to react with 1.12 kg of ferric oxide. i) Calculate the mass of Al_2O_3 formed.

ii) How much of the excess reagent is left at the end of the reaction? (Q. No. 39)

Ans:

Content	Reactant		Products	
	Al	Fe_2O_3	Al_2O_3	Fe
Stoichiometric coefficient	2	1	1	2
Amount of reactant allowed to react	324 g	1120 g	1	1
No. of moles allowed to react Moles = Mass / Molar mass	324 / 27 = 12	1120 /160 = 7	–	–
No. of moles of reactant reacted and product formed	12	6	6	12
No. of moles of un-reacted reactants and the product formed	-	1	6	12

i) Mass of Al_2O_3 formed = moles x molar mass = 6 molx 102 g mol^{-1} = 612 g

ii) Excess Reagent = Fe_2O_3

∴ Amount of excess reagent left at the end of the reaction = 1 mol × 160 g mol^{-1}

= 160 g

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