

UNIT – I – BASIC CONCEPTS OF CHEMISTRY & CHEMICAL CALCULATIONS

| Mr. S.JOHNSON., M.Sc., M.Sc., B.Ed.,

UNIT – 1 – BASIC CONCEPTS OF CHEMISTRY & CHEMICAL CALCULATIONS

II. WRITE BRIEF ANSWER TO THE FOLLOWING QUESTIONS.

26. Define relative atomic mass. [FMT-18]

The relative atomic mass of element is defined as the ratio of mass of one atom of the element to the mass of $\frac{1}{12}$ th mass of one atom of carbon-12

$$\text{Relative atomic mass (Ar)} = \frac{\text{Mass of one atom of the element}}{\text{Mass of } \frac{1}{12}\text{th mass of one atom of Carbon-12}}$$

$$= \frac{\text{Mass of one atom of an element}}{1.6605 \times 10^{-27} \text{ Kg}}$$

27. What do you understand by the term mole. [JUN-19, CRT-22]

The mole is defined as the amount of a substance which contains 6.023×10^{23} particles such as atoms, molecules or ions. It is denoted by the symbol "n".

28. Define equivalent mass. [GMQP-2018; QY-2018]

The equivalent mass of an element, compound or ion is the number of parts of mass of an element which combines with or displaces 1.008 parts of hydrogen or 8 parts of oxygen or 35.5 parts of chlorine.

29. What do you understand by the term oxidation number.

Oxidation number refers to the number of charges an atom would have in a molecule or an ionic compound, if electrons were transferred completely the oxidation numbers reflect the number of electron transferred.

30. Distinguish between oxidation and reduction. [HY-19, SEP-21]

	Oxidation	Reduction
(i)	Addition of oxygen and removal of hydrogen	Additional of hydrogen and removal of oxygen
(ii)	This process involves loss of electrons $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$	This process involves gain electrons. $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
(iii)	Oxidation number increases	Oxidation number decreases
(iv)	$\text{Ca} + \text{S} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$	$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$
(v)	Removal of Metal $2\text{KI} + \text{H}_2\text{O}_2 \rightarrow 2\text{KOH} + \text{I}_2$	Addition of metal $\text{HgCl}_2 + \text{Hg} \rightarrow \text{Hg}_2\text{Cl}_2$

31. Calculate the molar mass of the following compounds.

i) Urea [$\text{CO}(\text{NH}_2)_2$]

ii) Acetone [CH_3COCH_3]

iii) Boric acid [H_3BO_3]

iv) Sulphuric acid [H_2SO_4]

(i) urea [$\text{CO}(\text{NH}_2)_2$] :

$$\text{C} : 1 \times 12.01 = 12.01$$

$$\text{O} : 1 \times 16 = 16.00$$

$$\text{N} : 2 \times 14.01 = 28.02$$

$$\text{H} : 4 \times 1.01 = 4.04$$

$$\underline{\underline{60.07 \text{ g mol}^{-1}}}$$

(ii) acetone [CH_3COCH_3]

$$\text{C} : 3 \times 12.01 = 36.03$$

$$\text{H} : 6 \times 1.01 = 6.06$$

$$\text{O} : 1 \times 16 = 16.00$$

$$\underline{\underline{58.09 \text{ g mol}^{-1}}}$$

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(iii) boric acid [H_3BO_3] :

$$\begin{aligned} \text{H} : 3 \times 1.01 &= 3.03 \\ \text{B} : 1 \times 10 &= 10.00 \\ \text{O} : 3 \times 16 &= \underline{48.00} \end{aligned}$$

$$\underline{61.03 \text{ g mol}^{-1}}$$

(iv) sulphuric acid [H_2SO_4] :

$$\begin{aligned} \text{H} : 2 \times 1.01 &= 2.02 \\ \text{S} : 1 \times 32.06 &= 32.06 \\ \text{O} : 4 \times 16 &= \underline{64.00} \end{aligned}$$

$$\underline{98.08 \text{ g mol}^{-1}}$$

32. The density of carbon dioxide is equal to 1.965 kgm^{-3} at 273 K and 1 atm pressure. Calculate the molar mass of CO_2 .

Given : The density of CO_2 at 273 K and 1 atm pressure = 1.965 kgm^{-3}

Molar mass of CO_2 = ?

At 273 K and 1 atm pressure, 1 mole of CO_2 occupies a volume of 22.4 L

$$\begin{aligned} \text{Mass of 1 mole of } \text{CO}_2 &= \frac{1.965 \text{ Kg}}{1 \text{ m}^3} \times 22.4 \text{ L} \\ &= \frac{1.965 \times 10^3 \text{ g} \times 22.4 \times 10^{-3} \text{ m}^3}{1 \text{ m}^3} \\ &= 44.01 \text{ g} \end{aligned}$$

$$\text{Molar mass of } \text{CO}_2 = \underline{44 \text{ gmol}^{-1}}$$

33. Which contains the greatest number of moles of oxygen atoms

i) 1 mol of ethanol

ii) 1 mol of formic acid

iii) 1 mol of H_2O

(i) 1 mol of ethanol : $\text{C}_2\text{H}_5\text{OH}$ (ethanol) - Molar mass = $24 + 6 + 16 = 46$

46g of ethanol contains $1 \times 6.023 \times 10^{23}$ number of oxygen atoms.

(ii) 1 mol of formic acid : HCOOH (Formic acid) - Molar mass = $2 + 12 + 32 = 46$

46g of HCOOH contains $2 \times 6.023 \times 10^{23}$ number of oxygen atoms

(iii) 1 mol of H_2O : H_2O (Water) - Molar mass = $2 + 16 = 18$

18g of water contains $1 \times 6.023 \times 10^{23}$ number of oxygen atoms.

\therefore mol of formic acid contains the greatest number of oxygen atoms.

34. Calculate the average atomic mass of naturally occurring magnesium using the following data

Isotope	Isotopic atomic mass	Abundance (%)
Mg^{24}	23.99	78.99
Mg^{25}	24.99	10.00
Mg^{26}	25.98	11.01

Isotopes of Mg

$$\text{Atomic mass} = \text{Mg}^{24} = 23.99 \times \frac{78.99}{100} = 18.95$$

$$\text{Atomic mass} = \text{Mg}^{25} = 24.99 \times \frac{10}{100} = 2.499$$

$$\text{Atomic mass} = \text{Mg}^{26} = 25.98 \times \frac{11.01}{100} = 2.860$$

$$\text{Average atomic mass} = 24.309$$

$$\text{Average atomic mass of Mg} = 24.309$$

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35. In a reaction $x + y + z_2 \rightarrow xyz_2$ identify the Limiting reagent if any, in the following reaction mixtures.

(a) 200 atoms of x + 200 atoms of y + 50 molecules of z_2

(b) 1 mol of x + 1 mol of y + 3 mol of z_2

(c) 50 atoms of x + 25 atoms of y + 50 molecules of z_2

(d) 2.5 mol of x + 5 mol of y + 5 mol of z_2

Reaction : $x + y + z_2 \rightarrow xyz_2$

Question	Number of moles of reactants allowed to react			Number of moles of reactants consumed during reaction			Limiting reagent
	x	y	z_2	x	y	z_2	
(a)	200 atoms	200 atoms	50 molecules	50 atoms	50 atoms	50 molecules	z_2
(b)	1 mol	1 mol	3 mol	1 mol	1 mol	1 mol	X and y
(c)	50 atom	25 atom	50 molecules	25 atom	25 atom	25 molecules	y
(d)	2.5 mol	5 mol	5 mol	2.5 mol	2.5 mol	2.5 mol	x

36. Mass of one atom of an element is 6.645×10^{-23} g. How many moles of element are there in 0.320 kg.

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

\therefore mass of 1 mole of atom = 6.645×10^{-23} g $\times 6.022 \times 10^{23}$ = 40 g

\therefore number of moles of element in 0.320 kg = $\frac{1 \text{ mole}}{40 \text{ g}} \times 0.320 \text{ kg}$
 $= \frac{1 \text{ mol} \times 320 \text{ g}}{40 \text{ g}} = 8 \text{ mol.}$

37. What is the difference between molecular mass and molar mass? Calculate the molecular mass and molar mass for carbon monoxide.

	Molecular mass	Molar mass
(i)	Molecular mass is defined as the ratio of the mass of a molecule to the unified this is relative molecular mass atomic mass unit.	Molar mass is defined as the mass of one mole of a substance.
(ii)	The relative molecular mass of any compound is calculated by adding the relative atomic masses of its constituent atoms	The molar mass of a compound is equal to the sum of the relative atomic masses of its constituents.
(iii)	Its unit is u or amu	Its unit is g mol^{-1}
(iv)	Molecular mass of CO : $(1 \times \text{at.mass of C}) + (1 \times \text{at.mass of O})$ $1 \times 12.01 \text{ amu} + 1 \times 16 \text{ amu}$ $= 28.01 \text{ amu}$	Molar mass of CO : $1 \times 12.01 + 1 \times 16 = 28.01 \text{ g mol}^{-1}$

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38. What is the empirical formula of the following ? [FMT-18, QY-18, SEP-21]

i) Fructose (C₆H₁₂O₆) found in honey

ii) Caffeine (C₈H₁₀N₄O₂) a substance found in tea and coffee. [FIRST MID-2018; QY-2018]

Compound	Molecular formula	Empirical formula
Fructose	C ₆ H ₁₂ O ₆	CH ₂ O
Caffeine	C ₈ H ₁₀ N ₄ O ₂	C ₄ H ₅ N ₂ O

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

2Al + Fe₂O₃ → Al₂O₃ + 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 Al₂O₃ formed.

ii) How much of the excess reagent is left at the end of the reaction? [GMQP-2018]

i) 2Al + Fe₂O₃ → Al₂O₃ + 2Fe

54g 160g 102g 112g

As per balanced equation 54g Al is required for 112g of Iron and 102g of Al₂O₃.

∴ 324g of Al will give $\frac{102}{54} \times 324 = 612\text{g}$ of Al₂O₃

ii) 54g of Al required 160g of Fe₂O₃ for welding reaction

∴ 324g of Al will require $\frac{160}{54} \times 324 = 960\text{g}$ of Fe₂O₃

∴ Excess Fe₂O₃ – unreacted Fe₂O₃ = 1120 – 960 = 160g.

40. How many moles of ethane is required to produce 44 g of CO₂(g) after combustion. [FMT-18, QY-19]

Balanced equation for the combustion of ethane



To produce 4 moles of CO₂, 2 moles of ethane is required

∴ To produce 1 mole (44 g) of CO₂ required

$$\text{number of moles of ethane} = \frac{2 \text{ mol ethane}}{4 \text{ mol CO}_2} \times 1 \text{ mol CO}_2$$

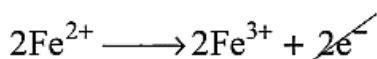
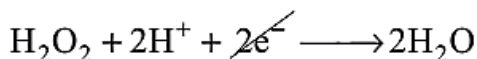
$$= \frac{1}{2} \text{ mole of ethane} = 0.5 \text{ mole of ethane}$$

41. Hydrogen peroxide is an oxidising agent. It oxidises ferrous ion to ferric ion and reduced itself to water. Write a balanced equation.

H₂O₂ – Oxidising agent

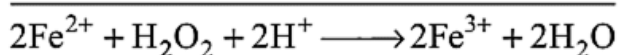
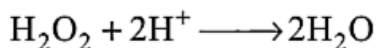


Ferrous ion is oxidized by H₂O₂ to Ferric ion. The balanced eq. is Fe²⁺ → Fe³⁺ + e⁻ × 2



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42. Calculate the empirical and molecular formula of a compound containing 76.6% carbon, 6.38 % hydrogen and rest oxygen its vapour density is 47. [MAR-19, SEP-20, AUG-21]

Element	Percentage	Atomic mass	Relative number of atoms	Simple ratio	Whole no
C	76.6	12	$\frac{76.6}{12} = 6.38$	$\frac{6.38}{1.06} = 6$	6
H	6.38	1	$\frac{6.38}{1} = 6.38$	$\frac{6.38}{1.06} = 6$	6
O	17.02	16	$\frac{17.02}{16} = 1.06$	$\frac{1.06}{1.06} = 1$	1

Empirical formula = $\text{C}_6\text{H}_6\text{O}$

$$n = \frac{\text{Molar mass}}{\text{Calculated empirical formula mass}}$$

$$= \frac{2 \times \text{vapour density}}{94} = \frac{2 \times 47}{94} = 1,$$

since Molar mass = 2 x Vapour density

molecular formula $n \times$ empirical formula

\therefore molecular formula $(\text{C}_6\text{H}_6\text{O}) \times 1 = \text{C}_6\text{H}_6\text{O}$

43. A Compound on analysis gave Na = 14.31% S = 9.97% H = 6.22% and O = 69.5% calculate the molecular formula of the compound, if all the hydrogen in the compound is present in combination with O_2 as H_2O of crystallization. (molecular mass of the compound is 322).

Element	%	Relative number of atoms	Simple ratio
Na	14.31	$\frac{14.31}{23} = 0.62$	$\frac{0.62}{0.31} = 2$
S	9.97	$\frac{9.97}{32} = 0.31$	$\frac{0.31}{0.31} = 1$
H	6.22	$\frac{6.22}{1} = 6.22$	$\frac{6.22}{0.31} = 20$
O	69.5	$\frac{69.5}{16} = 4.34$	$\frac{4.34}{0.31} = 14$

\therefore Empirical formula is = $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$

Empirical formula mass = $(23 \times 2) + (32 \times 1) + (16 + 4) + (10 \times 18)$
 $= 46 + 32 + 64 + 180 = 322$

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{322}{322} = 1$$

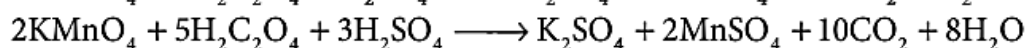
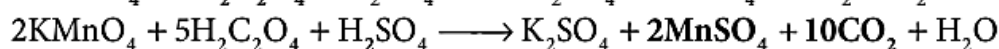
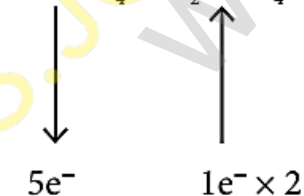
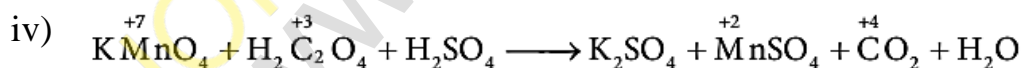
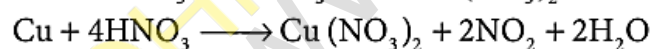
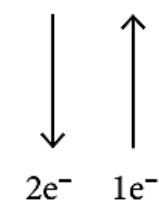
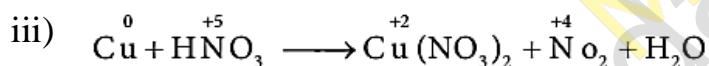
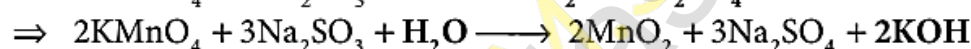
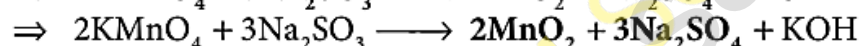
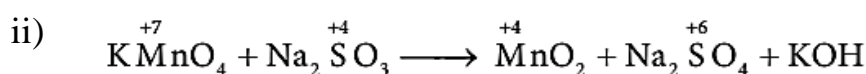
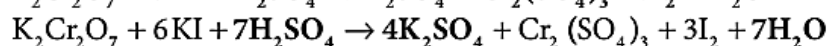
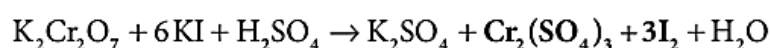
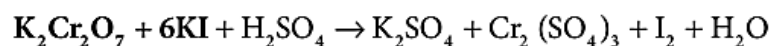
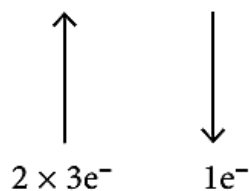
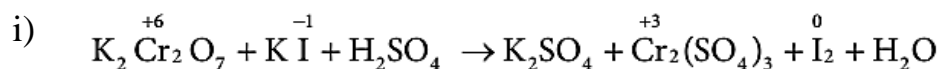
Molecular formula = $\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$

44. Balance the following equations by oxidation number method



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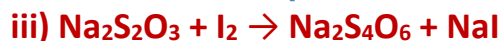


45. Balance the following equations by ion electron method.

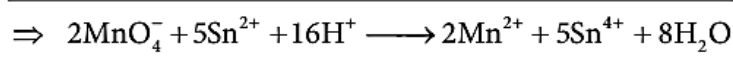
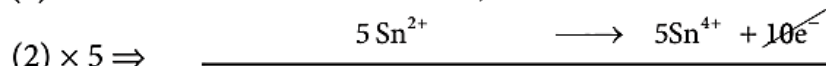
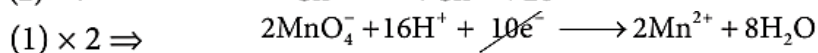
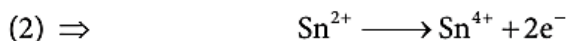
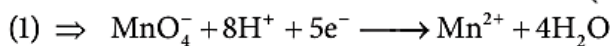
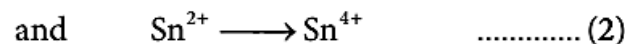
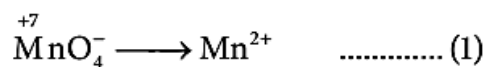


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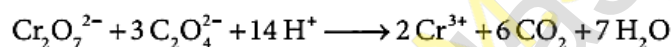
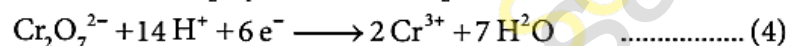
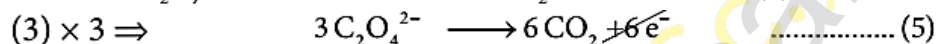
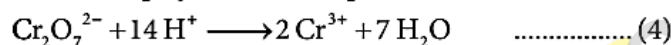
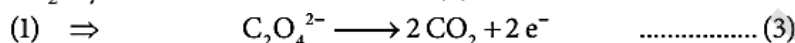
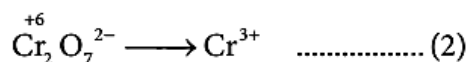
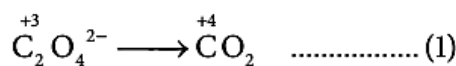
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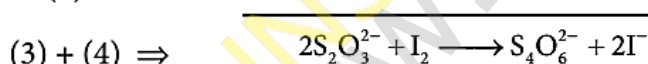
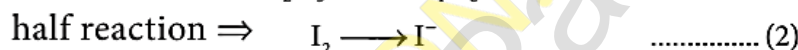
(i) Half reaction are



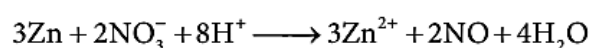
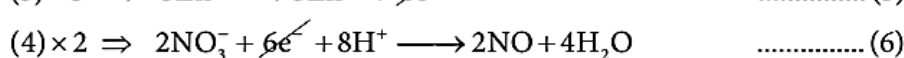
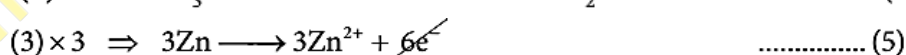
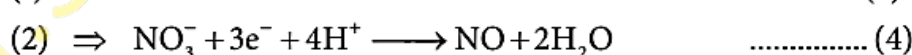
(ii)



(iii)



(iv) Half reactions are



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EVALUATE YOURSELF

1. By applying the knowledge of chemical classification, classify each of the following into elements, compounds or mixtures.

Sugar, Sea water, Distilled water, Carbon dioxide, Copper wire, Table salt, Silver plate, Naphthalene balls

(i) **Element:** Copper wire, Silver plate

(ii) **Compound:** Sugar, distilled water, carbon dioxide, Table salt, Naphthalene balls

(iii) **Mixture:** Sea water.

2. Calculate the relative molecular mass of the following.

(i) Ethanol [C₂H₅OH],

(ii) Potassium permanganate [KMnO₄],

(iii) Potassium dichromate [K₂Cr₂O₇],

(iv) Sucrose [C₁₂H₂₂O₁₁]

(i) C₂H₅OH = (2 x 12) + (5 x 1) + (1 x 16) + (1 x 1) = **46g**

(ii) KMnO₄ = (1 x 39) + (1 x 55) + (4 x 16) = **158g**

(iii) K₂Cr₂O₇ = (2 x 39) + (2 x 52) + (7 x 16) = **294g**

(iv) C₁₂H₂₂O₁₁ = (12 x 12) + (22 x 1) + (11 x 16) = **342g**

3. a) Calculate the number of moles present in 9g of ethane.

b) Calculate the number of molecules of oxygen gas that occupies a volume of 224ml at 272K and 3atm pressure.

a) Molar mass of ethane, C₂H₆ = (2 x 12) + (6 x 1) = **30g mol⁻¹**

No. of moles = $\frac{\text{mass}}{\text{molar mass}} = \frac{9\text{g}}{30\text{g mol}^{-1}} = \mathbf{0.3\text{ mole}}$

b) At 273K and 1 atm pressure 1 mole of a gas occupies a volume of 22.4L.

∴ number of moles of oxygen, that occupies a volume of 224ml at 273K and 3atm pressure

= $\frac{1\text{ mole}}{273\text{K} \times 1\text{ atm} \times 22.4\text{L}} \times 0.224\text{L} \times 273\text{K} \times 3\text{ atm} = \mathbf{0.03\text{ mole}}$

1 mole of oxygen contains 6.022 x 10²³ molecules

∴ 0.03 mole of oxygen contains = 6.022 x 10²³ x 0.03 = **1.807 x 10²²** molecules of oxygen

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

b) Calculate the equivalent mass of potassium dichromate. The reduction half reaction in acid medium is, Cr₂O₇²⁻ + 14H⁺ + 6e⁻ → 2Cr³⁺ + 7H₂O

a) Mass of the metal = W₁ = 0.456g

Mass of the metal chloride = W₂ = 0.606g

∴ Mass of chlorine = W₂ – W₁ = 0.606 – 0.456 = **0.15g**

0.15g of chlorine combine with 0.456g of metal

∴ 35.5g of chlorine will combine with $\frac{0.456}{0.15} \times 35.5 = \mathbf{107.92\text{g eq}^{-1}}$

b) Equivalent mass of a oxidising agent = $\frac{\text{Molar mass}}{\text{No. of moles of electrons gained by one mole of the reducing agent}}$

= $\frac{294.18\text{ mol}^{-1}}{6\text{eq mol}^{-1}} = \mathbf{49.03\text{g eq}^{-1}}$

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5. A Compound on analysis gave the following percentage composition C= 54.55%, H= 9.09%, O= 36.36%. Determine the empirical formula of the compound.

Element	Percentage	Atomic mass	Relative no. of atoms $= \frac{\text{Percentage}}{\text{Atomic mass}}$	Simple ratio	Ratio (whole Nos)
C	54.55%	12	$54.55 / 12 = 4.55$	$4.55 / 2.27 = 2$	2
H	9.09%	1	$9.09 / 1 = 9.09$	$9.09 / 2.27 = 4$	4
O	36.36%	16	$36.36 / 16 = 2.27$	$2.27 / 2.27 = 1$	1

∴ Empirical formula = **C₂H₄O**

6. Experimental analysis of a compound containing the elements x, y, z on analysis gave the following data. x = 32%, y = 24%, z = 44%. The relative number of atoms of x, y and z are 2, 1 and 0.5 respectively. (molecular mass of the compound is 400g) Find out.
i) The atomic masses of the element x, y, z. ii) Empirical formula of the compound and iii) Molecular formula of the compound.

Element	Percentage	Relative no. of atoms $= \frac{\text{Percentage}}{\text{Atomic mass}}$	Atomic mass = $\frac{\text{Percentage}}{\text{Relative no. of mass}}$	Simple ratio	Ratio (whole Nos)
X	32%	2	$32 / 2 = 16$	4	4
Y	24%	1	$24 / 1 = 24$	2	2
Z	44%	0.5	$44 / 0.5 = 88$	1	1

∴ Empirical formula = **X₄Y₂Z**

Calculate empirical formula mass = $(16 \times 4) + (24 \times 2) + (88 \times 1)$
 $= 64 + 48 + 88 = 200$

$$n = \frac{\text{molar mass}}{\text{calculated empirical formula mass}} = \frac{400}{200} = 2$$

∴ Molecular formula $(X_4Y_2Z)_2 = X_8Y_4Z_2$

7. The balanced equation for a reaction is given below $2x + 2y \rightarrow 4l + m$
 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 reagent left at the end of the reaction.

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

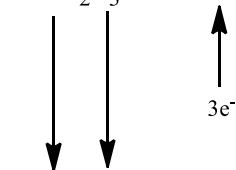
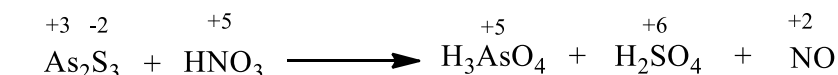
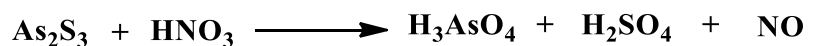
Product formed = 16 moles of l & 4 moles of m

Amount of excess reactant = 3 moles of y

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8. Balance the following equation using oxidation number method



$$\begin{array}{r} 2 \times 2e^- \quad 3 \times 8e^- \\ \hline 4e^- \quad + \quad 24e^- \\ \hline 28e^- \end{array}$$

Equate the total no. of electrons in the reactant side by cross multiplying,



Based on reactant side, balance the products



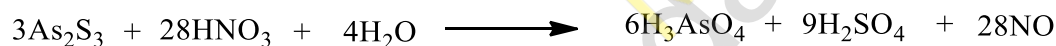
Product side: 36 hydrogen atoms & 88 oxygen atoms

Reactant side: 28 hydrogen atoms & 84 oxygen atoms

Difference is 8 hydrogen atoms & 4 oxygen atoms

∴ Add 4 H₂O molecule on the reactant side.

Balanced equation,



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GOVERNMENT QUESTIONS AND ANSWERS

1. Statement 1 : Two mole of glucose contains 12.044×10^{23} molecules of glucose

Statement 2 : Total number of entities present in one mole of any substance is equal to 6.02×10^{22} . [GMQP-2018]

Whether the above statements are true? Is there any relation between these two statements?

Ans. The statements 1 & 2 are true. But there is no relation between statement 1 and statement 2.

2. How many moles of hydrogen is required to produce 10 moles of ammonia ? [HY-2018]



To produce 2 moles of ammonia, 3 moles of hydrogen are required

To produce 10 moles of ammonia

$$= \frac{3 \text{ moles of } H_2}{2 \text{ moles of } NH_3} \times \frac{10 \text{ moles of } NH_3}{1}$$

= 15 moles of hydrogen are required.

3. Define limiting reagent. [GMQP-2018; QY-2018]

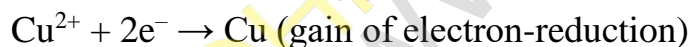
Ans. Limiting reagent : When a reaction is carried out using non-stoichiometric quantities of the reactants, the product yield will be determined by the reactant that is completely consumed. It limits the further reaction from taking place and is called as the limiting reagent.

4. Write the electronic concept of oxidation and reduction reactions. [QY. & HY. 2018]

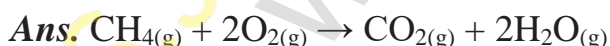
Ans. The process can be explained on the basis of electrons. The reaction involving loss of electron is termed oxidation



The reaction involving gain of electron is termed reduction.



5. Calculate the amount of water produced by the combustion of 32 g of methane. [QY-2018]

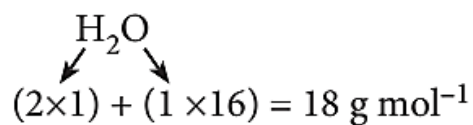
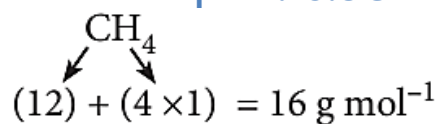


As per the stoichiometric equation,

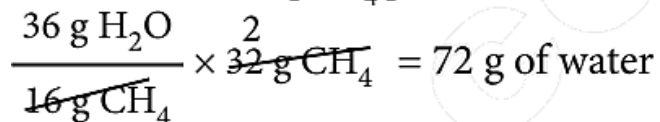
Combustion of 1 mole (16 g) CH_4 produces 2 moles ($2 \times 18 = 36$ g) of water.

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Combustion of 32 g CH_4 produces



1.