Unit 1 Basic Concepts of Chemistry and Chemical Calculations

Matter is defined as anything that has mass and occupies space. All matter is composed of atoms.

Physical Classification of Matter:

Matter can be classified as solids, liquids and gases based on their physical state.

Chemical Classification:

Matter can be classified into mixtures and pure substances based on chemical compositions.

Mixtures classified as homogeneous or heterogeneous mixtures based on their physical appearance.

An **element** consists of only one type of atom.

Monatomic unit – Gold (Au), Copper (Cu) & Inert gases

Polyatomic unit - Hydrogen (H_2), Phosphorous (P_4) and Sulphur (S_8)

Compounds are made up of molecules which contain two or more atoms of different elements.

Example : Carbon dioxide (CO_2), Glucose ($C_6H_{12}O_6$), Hydrogen Sulphide (H_2S), Sodium Chloride (NaCl)

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

Example:

${}_{17}Cl^{35}$ and ${}_{17}Cl^{37}$ in the ratio 77 : 23, the average relative atomic mass of chlorine is

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

= 35.46 u

Empirical formula

Empirical formula of a compound is the formula written with the **simplest ratio** of the number of different atoms present in one molecule of the compound as subscript to the atomic symbol.

Molecular formula

Molecular formula of a compound is the formula written with the **actual number** of different atoms present in one molecule as a subscript to the atomic symbol.

Example:

The molecular formula of acetic acid (CH_3COOH) is $C_2H_4O_2$

The ratio of C : H : O is 1 : 2 : 1 and hence the empirical formula is CH_2O.

Oxidation Number: (oxidation state)

It is defined as the imaginary charge left on the atom when all other atoms of the compound have been removed in their usual oxidation states that are assigned according to set of rules.

General rules of oxidation Number:

1) The oxidation state of a free element (i.e. in its uncombined state) is zero.

Example : each atom in H_2, Cl_2, Na, S_8 have the oxidation number of zero.

2) For a monatomic ion, the oxidation state is equal to the net charge on the ion.

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Example : The oxidation number of sodium in Na^+ is +1.
The oxidation number of chlorine in Cl^- is –1.
3) The algebraic sum of oxidation states of all atoms in a molecule is equal to zero, while in ions, it is equal to the net charge on the ion.
Example:
In H_2SO_4, = 0 . In SO_4^{2-} =+ 4
4) Hydrogen has an oxidation number of +1 in all its compounds except in metal hydrides where it has – 1 value.
Example:
Oxidation number of hydrogen in hydrogen chloride (**H**Cl) is + 1.
Oxidation number of hydrogen in sodium hydride (Na**H**) is –1.
5) **Fluorine** has an oxidation state of – 1 in all its compounds.
6) The oxidation state of **oxygen** in most compounds is **–2**. Exceptions are peroxides, super oxides and compounds with fluorine.
Example : Oxidation number of oxygen, i) in **water** (H_2O) is **–2**.
ii) in **hydrogen peroxide** (H_2O_2) is **–1**.
iii) **in super oxides** such as KO_2 is **–1/2**
iv) in **oxygen difluoride** (OF_2) is **+ 2**.
7) Alkali metals have an oxidation state of + 1 and alkaline earth metals have an oxidation state of + 2 in all their compounds.
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.
(**Limiting Reagent-** The reactant which gets consumed first or limits the amount of product formed is known as **limiting reagent**)
The other reagents which are in excess are called the **excess reagents**.
Example: $\mathbf{S + 3F_2 \rightarrow SF_6}$
Sulphur is the limiting reagent and **fluorine** is the excess reagent.
A reaction in which oxidation number of the element **increases** is called oxidation.
A reaction in which oxidation number of the element **decreases** is called reduction.

Redox reactions in which

- two substances combine to form compound(s) are called combination reaction.
- a compound breaks down into two (or) more components is called decomposition reaction
- a compound is replaced by an ion (or atom) of another element are called displacement reactions
- the same compound can undergo both oxidation and reduction and the oxidation state of one and the same element is both increased and decreased called disproportionate reactions.
- competition for electrons occurs between various metals is called competitive electron transfer reaction.

The equation of redox reaction is balanced either by oxidation number method or by ion-electron method.

I. Choose the best answer.

1. 40 ml of methane is completely burnt using 80 ml of oxygen at room temperature. The volume of gas left after cooling to room temperature is

(a) 40 ml CO_2 gas (b) 40 ml CO_2 gas and 80 ml H_2O gas
(c) 60 ml CO_2 gas and 60 ml H_2O gas (d) 120 ml CO_2 gas

2. 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

(a) 201 u (b) 202 u (c) 199 u **(d) 200 u**

3. Assertion : Two mole of glucose contains 12.044×10^{23} molecules of glucose
Reason : Total number of entities present in one mole of any substance is equal to 6.02×10^{22}

(a) both assertion and reason are true and the reason is the correct explanation of assertion
(b) both assertion and reason are true but reason is not the correct explanation of assertion
(c) Assertion is true but reason is false (d) both assertion and reason are false

4. Carbon forms two oxides, namely carbon monoxide and carbon dioxide. The equivalent mass of which element remains constant?

(a) Carbon **(b) oxygen** (c) both carbon and oxygen (d) neither carbon nor oxygen

5. The equivalent mass of a trivalent metal element is 9 g eq^{-1} the molar mass of its anhydrous oxide is **(a) 102 g** (b) 27 g (c) 270 g (d) 78 g

6. The number of water molecules in a drop of water weighing 0.018 g is

(a) 6.022×10^{26} (b) 6.022×10^{23} **(c) 6.022×10^{20}** (d) 9.9×10^{22}

7. 1 g of an impure sample of magnesium carbonate (containing no thermally decomposable impurities) on complete thermal decomposition gave 0.44 g of carbon dioxide gas. The percentage of impurity in the sample is

(a) 0 % (b) 4.4 % **(c) 16 %** (d) 8.4 %

8. When 6.3 g of sodium bicarbonate is added to 30 g of acetic acid solution, the residual solution is found to weigh 33 g.The number of moles of carbon dioxide released in the reaction is (a) 3 (b) 0.75 **(c) 0.075** (d) 0.3

9. When 22.4 litres of H_2 (g) is mixed with 11.2 litres of Cl_2 (g), each at 273 K at 1 atm the moles of HCl (g), formed is equal to

(a) 2 moles of HCl(g) (b) 0.5 moles of HCl(g) (c) 1.5 moles of HCl(g) **(d) 1 moles of HCl(g)**

10. How concentrated sulphuric acid is a moderately strong oxidising agent. Which of the following reactions does not show oxidising behaviour?

(a) $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$ (b) $C + 2H_2SO_4 \rightarrow CO_2 + 2SO_2 + 2H_2O$
(c) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$ (d) none of the above

11. Choose the disproportionation reaction among the following redox reactions.

(a) $3Mg_{(s)} + N_{2(g)} \rightarrow Mg_3N_{2(s)}$ **(b) $P_4(s) + 3NaOH + 3H_2O \rightarrow PH_{3(g)} + 3NaH_2PO_{2(aq)}$**
(c) $Cl_{2(g)} + 2KI_{(aq)} \rightarrow 2KCl_{(aq)} + I_2$ (d) $Cr_2O_{3(s)} + 2Al_{(s)} \rightarrow Al_2O_{3(s)} + 2Cr_{(s)}$

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12.The equivalent mass of potassium permanganate in alkaline medium is

$MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$

(a) 31.6 **(b) 52.7** (c) 79 (d) None of these

13.Which one of the following represents 180g of water?

(a) 5 Moles of water (b) 90 moles of water

(c) $\frac{6.022 \times 10^{23}}{180}$ moleculesof water **(d) $6.022 x 10^{24}$ molecules of water**

14.7.5 g of a gas occupies a volume of 5.6 litres at 0^o C and 1 atm pressure. The gas is

(a) NO (b) N_2O (c) CO (d) CO_2

15.Total number of electrons present in 1.7 g of ammonia is

(a) 6.022×10^{23} (b) $\frac{6.022 \times 10^{22}}{1.7}$ (c) $\frac{6.022 \times 10^{24}}{1.7}$ (d) $\frac{6.022 \times 10^{23}}{1.7}$

16.The correct increasing order of the oxidation state of sulphur in the anions SO_4^{2-}, SO_3^{2-} , $S_2O_4^{2-}$, $S_2O_6^{2-}$ is

(a) $SO_3^{2-} < SO_4^{2-} < S_2O_4^{2-} < S_2O_6^{2-}$ (b) $SO_4^{2-} < S_2O_4^{2-} < S_2O_6^{2-} < SO_3^{2-}$

(c) $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-} < SO_4^{2-}$ (d) $S_2O_6^{2-} < S_2O_4^{2-} < SO_4^{2} SO_3^{2-}$

17.The equivalent mass of ferrous oxalate is

(a) $\frac{Molar\ mass\ of\ ferrous\ oxalate}{1}$ b) $\frac{Molar\ mass\ of\ ferrous\ oxalate}{2}$ **c) $\frac{Molar\ mass\ of\ ferrous\ oxalate}{3}$**

(d) none of these

18.If Avogadro number were changed from 6.022×10^{23} to $6.022 x 10^{20}$, this would change

(a) the ratio of chemical species to each other in a balanced equation

(b) the ratio of elements to each other in a compound

(c) the definition of mass in units of grams **(d) the mass of one mole of carbon**

19.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

(a) Number of molecules in A and B are same

(b) Number of molecules in B is more than that in A.

(c) The ratio between the number of molecules in A to number of molecules in B is 2:1

(d) Number of molecules in B is three times greater than the number of molecules in A.

20.What is the mass of precipitate formed when 50 ml of 8.5 % solution of $AgNO_3$ is mixed with 100 ml of 1.865 % potassium chloride solution?

(a) 3.59 g (b) 7 g (c) 14 g (d) 28 g

21.The mass of a gas that occupies a volume of 612.5 ml at room temperature and pressure (250 °C and 1 atm pressure) is 1.1g. The molar mass of the gas is

(a) 66.25 g mol^{-1} **(b) 44 g mol^{-1}** (c) 24.5 g mol^{-1} d) 662.5 g mol^{-1}

22.Which of the following contain same number of carbon atoms as in 6 g of carbon-12.

(a) 7.5 g ethane (b)8 g methane (c) both (a) and (b) (d) none of these

23. Which of the following compound(s) has /have percentage of carbon same as that in ethylene (C_2H_4) **(a) propene** (b) ethyne (c) benzene (d) ethane

24.Which of the following is/are true with respect to carbon -12.

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(a) relative atomic mass is 12 u (b) oxidation number of carbon is +4 in all its compounds.
(c) 1 mole of carbon-12 contain 6.022×10^{22} carbon atoms. (d) all of these

25. Which one of the following is used as a standard for atomic mass.

(a) $_6C^{12}$ (b) $_7C^{12}$ (c) $_6C^{13}$ (d) $_6C^{14}$

Answer the following

26. Define relative atomic mass

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

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

27. What do you understand by the term mole.

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.

28. Define equivalent Mass

Equivalent mass of an element, compound or ion is the number of parts of mass of an element combines or displaces 1.008 g hydrogen or 8 g oxygen or 35.5g chlorine.

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

When all other atoms of the compound have been removed in their usual oxidation states that are assigned according to set of rules.

30. Distinguish between oxidation and reduction

	Oxidation	**Reduction**
(i)	Addition of oxygen and removal of hydrogen	Additional of hydrogen and removal of oxygen
(ii)	This process involves loss of electrons $Fe^{2+} \rightarrow Fe^{3+} + e^-$	This process involves gain electrons. $Cu^{2+} + 2e- \rightarrow Cu$
(iii)	Oxidation number increases	Oxidation number decreases
(iv)	Addition of non-metal	Removal of non-metal $2HgCl_2 + SnCl_2 \rightarrow Hg_2Cl_2 + SnCl_4$
(v)	Removal of metal $2KI + H_2O_2 \rightarrow 2KOH + I_2$	Addition of metal $HgCl_2 + Hg \rightarrow Hg_2Cl_2$

31. Calculate the molar mass of the following compounds (i) urea [$CO(NH_2)_2$] (ii) acetone [CH_3COCH_3] (iii) boric acid [H_3BO_3] (iv) sulphuric acid [H_2SO_4]

(i) urea [$CO(NH_2)_2$]

C : 1 × 12.01 = 12.01
O : 1 × 16 = 16.00
N : 2 × 14.01 = 28.02
H : 4 × 1.01 = 4.04
60.07 g mol^{-1}

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(ii) acetane [CH_3COCH_3]

C : 3 × 12.01 = 36.03
H : 6 × 1.01 = 6.06
O : 1 × 16 = 16.00
58.09 g mol^{-1}

(iii) boric acid [H_3BO_3]

H : 3 × 1.01 = 3.03
B : 1 × 11 = 11.00
O : 3 × 16 = 48.00
62.03 g mol^{-1}

(iv) sulphuric acid [H_2SO_4]

H : 2 × 1.01 = 2.02
S : 1 × 32.06 = 32.06
O : 4 × 16 = 64.00
98.08 g mol^{-1}

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

Given : The density of CO_2 at 273K and 1 atm pressure = 1.965 kgm^{-3} At 273 K and 1 atm pressure, 1 mole of CO_2 occupies a volume of 22.4L

Mass of 1 mole of CO_2 $= \frac{1.965Kg}{1m^3} \times 22.4L = \frac{1.965 \times 10^3\ g \times 22.4 \times 10^{-3} m^3}{1m^3}$

Molar mass of CO_2 = **44 gmol^{-1}**

ANOTHER METHOD:

Molecular mass = Density x Molar mass

Molar volume of CO_2 = 2.24 x 10^{-2} m^3

Density of CO_2 = 1.965kgm^{-3}

Molecular mass of CO_2 = 1.965 x 10^3gm^{-3} x 2.24 x 10^{-2}m^3
=1.965 x 10^1 x 2.24
= 44g

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**

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	$1 \times 6.022 \times 10^{23}$

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**

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Mg^{26}	**24.99**	**10.00**
Mg^{25}	**25.98**	**11.01**

Average atomic mass $= \frac{(78.99 \times 23.99)+(10 \times 24.99)+(11.01 \times 25.98)}{100}$

$= \frac{2430.9}{100} = 24.31u.$

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

Q	**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	200atoms	50 molecules	50 atoms	50atoms	50 molecules	**z_2**
b	1 mol	1mol	3 mol	1 mol	1 mol	1mol	**x and y**
(c	50atom	25atom	50 molecules	25atom	25atom	25 molecules	**y**
d	2.5 mol	5 mol	5 mol	2.5mol	2.5 mol	2.5 mol	**x**

(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 + $z_2 \rightarrow xyz_2$

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 $= 6.645 \times 10^{-23}$ g $\times 6.022 \times 10^{23}$

$= 40g$

$\therefore$ number of moles of element in 0.320 kg **(number of moles = mass/atomic mass)**

$= \frac{1mole}{40g} \times 0.320kg$

$= \frac{1\ mol \times 320g}{40g} = 8$ 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**
1	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.
2	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

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3	Its unit is u or amu	Its unit is g mol^{-1}
4	Molecular mass of CO: (1 × at.mass of C) + (1 × at. Mass of O) 1× 12.01 amu + 1 × 16 amu = 28. 01 am	Molar mass of CO: 1×12.01+ 1× 16 = 28.01g mol^{-1}

38. What is the empirical formula of the following ?

(i) Fructose ($C_6H_{12}O_6$) found in honey

(ii) Caffeine ($C_8H_{10}N_4O_2$) a substance found in tea and coffee.

Compound	**Molecular Formula**	**Empirical Formula**
Fructose	$C_6H_{12}O_6$	CH_2O
Caffeine	$C_8H_{10}N_4O_2$	$C_4H_5N_2O$

39. The reaction between aluminium and ferric oxide can generate temperatures up to 3273 K and is used in welding metals. (Atomic mass of Al = 27u Atomic mass of O = 16u) 2Al + Fe_2O_3 →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?

Given: $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

	Reactants		**Products**	
	Al	**Fe_2O_3**	**Al_2O_3**	**Fe**
Amount of reactant allowed to react	324 g	1.12 kg	-	-
Number of moles allowed to react	$\frac{324}{27} = 12\ mol$	$\frac{1.12 \times 10^3}{160}$= 7mol	-	-
Stoichiometric Co-efficient	2	1	1	2
Number of moles consumed during reaction	12 mol	6 mol	-	-
Number of moles of reactant unreacted and number of moles of product formed	-	1mol	6mol	12mol

Molar mass of Al_2O_3 formed

= 6 mol× 102 g mol^{-1} $\left[\begin{matrix} Al_2O_3 \\ (2\times 27)+(3\ \times 16) \\ 54+48=102 \end{matrix}\right]$ = 612g

Excess reagent = Fe_2O_3

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

= 160 g $\left[\begin{matrix} Fe_2O_3 \\ (2\times 56)+(3\ \times 16) \\ 112+48=160 \end{matrix}\right]$ = 160 g

40. **How many moles of ethane is required to produce 44 g of CO_2(g) after combustion**.

Balanced equation for the combustion of ethane

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$$C_2H_6 + 3\frac{1}{2}O_2 \rightarrow 2CO_2 + 3H_2O$$

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

To produce 4 moles of CO_2, 2 moles of ethane is required

To produce 1 mole (44 g) of CO_2 required number of moles of ethane

$$= \frac{2\ mol\ ethane}{4\ mol\ CO_2} \times 1\ mol\ CO_2 = ½ \text{ mole of ethane} = 0.5 \text{ mole of ethane}$$

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

$H_2O_2 \rightarrow$ Oxidising agent

-1 -2

$H_2O_2 + Fe^{2+} \rightarrow Fe^{3+} + H_2O$ (Acidic medium)

($1e^- \times 2$) $1e^-$

Ferrous ion is oxidized by H_2O_2 to Ferric ion.

The balanced equation is $Fe^{2+} \rightarrow Fe^{3+} + e^- \times 2$

$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$

$\underline{2Fe^{2+} \rightarrow 2Fe^{3+} + 2e^-}$

$H_2O_2 + 2Fe^{2+} + 2H^+ \rightarrow Fe^{3+} + 2H_2O$

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.

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
O	17.02	16	$\frac{17.02}{16} = 1.06$	$\frac{1.06}{1.06} = 1$	1

Empirical Formula = C_6H_6O

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

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

since Molar mass = 2 × Vapour density

Molecular formula n × n empirical formula

Molecular formula $(C_6H_6O) \times 1 = C_6H_6O$

43.A Compound o 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 oxygen as water 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	$\frac{69.5}{16} = 4.34$	$\frac{4.34}{0.31} = 14$

Empirical formula = $Na_2SH_{20}O_{14}$

$$n = \frac{molar\ mass}{Calculated\ empirical\ formula\ mass}$$

$$= \frac{322}{322} = 1$$

$$\left[\begin{array}{c} Na_2SH_{20}O_{14} \\ = (2 \times 23) + (1 \times 32) + (20 \times 1) + 14(16) \\ = 46 + 32 + 20 + 224 \\ = 322 \end{array}\right]$$

Molecular formula = $Na_2\,S\,H_{20}\,O_{14}$

Since all the hydrogen in the compound present as water

Molecular formula is $Na_2\,SO_4.10H_2O$.

44. Balance the following equations by oxidation number method

i) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

ii) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

iii) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

iv) $KMnO_4 + H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

+6 -1 +3 0

i) $K_2Cr_2O_7 + KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

$2 \times 3e^-$ $1e^-$

$K_2Cr_2O_7 + 6KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + I_2 + H_2O$

$K_2Cr_2O_7 + 6KI + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + 3I_2 + H_2O$

$K_2Cr_2O_7 + 6KI + 7H_2SO_4 \rightarrow 4K_2SO_4 + Cr_2(SO_4)_3 + 3I_2 + 7H_2O$

+7 +4 +4 +6

ii) $KMnO_4 + Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

$3e^-$ $2e^-$

⇨ $2KMnO_4 + 3Na_2SO_3 \rightarrow MnO_2 + Na_2SO_4 + KOH$

⇨ $2KMnO_4 + 3Na_2SO_3 \rightarrow 2MnO_2 + 3Na_2SO_4 + KOH$

⇨ $2KMnO_4 + 3Na_2SO_3 \rightarrow 2MnO_2 + 3Na_2SO_4 + 2KOH$

0 +5 +2 +4

iii) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

$2e^-$ $1e^-$

$Cu + 2HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

$Cu + 2HNO_3 + 2HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$

iv) $\overset{+7}{KMnO_4} + \overset{+3}{H_2C_2O_4} + H_2SO_4 \rightarrow K_2SO_4 + \overset{+2}{MnSO_4} + \overset{+4}{CO_2} + H_2O$

$5e^-$ $\quad 1e^- \times 2$

$2KMnO_4 + 5H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + CO_2 + H_2O$

$2KMnO_4 + 5H_2C_2O_4 + H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + H_2O$

$2KMnO_4 + 5H_2C_2O_4 + 3H_2SO_4 \rightarrow K_2SO_4 + 2MnSO_4 + 10CO_2 + 8H_2O$

45. Balance the following equations by ion electron method.

i) $KMnO_4 + SnCl_2 + HCl \rightarrow MnCl_2 + SnCl_4 + H_2O + KCl$

ii) $C_2O_4^{2-} + Cr_2O_7^{2-} \rightarrow Cr^{3+} + CO_2$ (in acid medium)

iii) $Na_2S_2O_3 + I_2 \rightarrow Na_2S_4O_6 + NaI$ (in acid medium)

iv) $Zn + NO_3^- \rightarrow Zn^{2+} + NO$

(i) Half reaction are:

$\overset{+7}{MnO_4^-} \rightarrow Mn^{2+}$ ---- (1)

And $Sn^{2+} \rightarrow Sn^{4+}$ ---- (2)

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

(2) ➔ $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

(1)× 2 ➔ $2MnO_4^- + 16H^- + 10e^- \rightarrow 2Mn^{2+} + 8H_2O$

(2) × 5➔ $5Sn^{2+} \rightarrow 5Sn^{4+} + 10e^-$

➔ $2MnO_4^- + 5Sn^{2+} + 16H^+ \rightarrow 2Mn^{2+} + 5Sn^{4+} + 8H_2O$

(ii) $\overset{+3}{C_2O_4^{2-}} \rightarrow \overset{+4}{CO_2}$ ---- (1)

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ ---- (2)

(1) ➔ $C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$ ---- (3)

$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$ ---- (4)

(2) × 3 ➔ $3C_2O_4^{2-} \rightarrow 6CO_2 + 6e-$ ---- (5)

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ ---- (4)

➔ $Cr_2O_7^{2-} + 3C_2O_4^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 6CO_2 + 7H_2O$

(iii) $S_2O_3^{2-} \rightarrow S_4O_6^{2-}$ ---- (1)

Half reaction ➔ $I_2 \rightarrow I$ ---- (2)

(1) ➔ $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$ ---- (3)

(2) ➔ $I_2 + 2e^- \rightarrow 2I^-$ ---- (4)

(3) + (4) ➔ $2S_2O_3^{2-} + I_2 \rightarrow S_4O_6^{2-} + 2I^-$

(iv) $Zn \rightarrow Zn^{2+}$ ---- (1)

$NO_3^- \rightarrow NO$ ---- (2)

(1) ➔ $Zn \rightarrow Zn^{2+} + 2e^-$ ---- (3)

(2) ➔ $NO_3^- + 3e^- + 4H^+ \rightarrow NO + 2H_2O$ ---- (4)

(3) × 3 ➔ $3Zn \rightarrow Zn^{2+} + 6e^-$ ---- (5)

(4) × 2 ➔ $2NO_3^- + 6e^- + 8H^+ \rightarrow 2NO + 4H_2O$ ---- (6)

$3Zn + 2NO_3^- + 8H^+ \rightarrow 3Zn^{2+} + 2NO + 4H_2O$

2. QUANTUM MECHANICAL MODEL OF ATOM

Important Formulae

1. $\Delta E = h\upsilon$ or $\upsilon = \dfrac{E_2 - E_1}{h}$
2. $r_n = \dfrac{(0.529)n^2}{Z}$ Å
3. $E_n = \dfrac{(-13.6)Z^2}{n^2}$ eV atom^{-1} (or) $\dfrac{(-1312.8)Z^2}{n^2}$ KJ mol^{-1}
4. $E = h\upsilon$, $E = mc^2$
5. $\lambda = \dfrac{h}{mv}$ or $\dfrac{h}{p}$
6. $2\pi = nh/mv$ (or) $mvr = nh/2\pi$
7. $\Delta x.\Delta p \geq \dfrac{h}{4\pi}$, (or) $\Delta x.\ m.\Delta v \geq \dfrac{h}{4\pi}$,
8. $\hat{H}\Psi = E\Psi$
9. $\dfrac{\partial^2\Psi}{\partial x^2} + \dfrac{\partial^2\Psi}{\partial y^2} + \dfrac{\partial^2\Psi}{\partial z^2} + \dfrac{8\pi^2 m}{h^2}(E - V)\Psi = 0$
10. Angular momentum = $\sqrt{l(1+1)}\dfrac{h}{2\pi}$

Important Points to Remember

- **Atom** – The basic unit that makes up all matter
- **Planck's constant (h)** – 6.626 x 10^{-34} Js.
- **Angular momentum** – mvr ; mvr = nh/2π
- **Bohr's atomic model** – The electron is revolving around the nucleus in a certain fixed circular path called stationary orbit. The energies of electrons are quantized. Electron can revolve only in orbits in which the angular momentum (mvr) of the electron must be equal to an integral multiple of h/2π
- **Radius of the nth orbit** – $r_n = \dfrac{(0.529)n^2}{Z}$ A
- **Energy of the nth orbit** - $E_n = \dfrac{(-13.6)Z^2}{n^2}$ eV atom^{-1} (or) $\dfrac{(-1312.8)Z^2}{n^2}$ KJ mol^{-1}
- **Louis de Broglie** – All forms of matter showed dual character such as particle and wave.
- **de Broglie** – All forms of matter showed dual character such as particle and wave.
- **de Broglie equation** - λ = h/mv
- **Mass of an electron** – 9.11 x 10^{-31} kg.
- **λ** - de Broglie wavelength

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Quantization of de Broglie concept - $2\pi r = nh/mv$ (or) $mvr = nh/2\pi$

- **Heisenberg's uncertainty principle -** $\Delta x.\Delta p \geq \frac{h}{4\pi}$, where Δx = uncertainty in position.
  Δp = uncertainty in momentum.
- **Bohr's radius of 1st orbital** – 0.529 Å
- **Schrödinger equation** – $\hat{H}\,\Psi$ - $E\,\Psi$
  $\hat{H}$ = Hamiltonian operator.
  E = wave function
  Ψ = Energy of the system
- **Schrödinger wave equation -** $\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}+\frac{8\pi^2 m}{h^2}(E-V)\Psi=0$
- **Quantum numbers** – The set of four quantum number which can characterize an electron in an atom. They represent the energy level, sub, shell, number of electron in an orbit, orientation of orbitals in space and the spin of the electron.
- **Atomic orbitals** – Permitted energy values called eigen values which are known as atomic orbitals.
- **$[\Psi]^2$ – It** is the probability of finding the electrons.
- **Node** – The region where the probability density function reduces to zero is called nodal surface (or) a radial node.
- **Shapes of orbital -**s-orbital - spherical shape, p-orbital - dumb bell shape
  d-orbital – clover leaf
- **Effective nuclear charge** – The net charge experienced by the electron.
- **Order of E.N.C** – s > p > d > f in an orbital.
- **Energy order in orbital** – s < p < d < f.
- **Aufbau's principle** – "In the ground state of the atoms, the orbitals are filled in the order of their increasing energies".
- **Pauli's exclusion principle** – "No two electrons in an atom can have the same set of values for all four quantum numbers".
- **Types of quantum numbers** – Principal quantum (n), azimuthal quantum number (*l*), magnetic quantum number (m) and spin quantum number (s).

	Principal quantum number (n):	**Azimuthal quantum number (*l*)**	**Magnetic quantum number (m):**	**Spin quantum number (s):**
1	It represents the energy level in which electron is present	It represents the subshell in which electron is present	It represents different orientation of orbitals in space	It represents that spin of the electron
2	Maximum number of electrons that can be accommodated in the shell is $2n^2$.	Maximum number of electrons that can be accommodated in subshell is 2(21+1)	The magnitude of angular momentum is determined.	Electron revolves in clockwise and anti-clockwise directions.

3	$E_n = \frac{-(1312.8)Z^2}{n^2}$ $kJ\ mol^{-1}$.	It is used to calculate the orbital angular momentum by the formula $\sqrt{l(l+1)}\,\frac{h}{2\pi}$	If l = 1, m = -1, 0, + 1	S values are $+^1/_2$ (or) $-^1/_2$.
4	It represents the distance of the electron from the nucleus.	l = 0, 1, 2, 3 Subshell = s, p, d, f No. of electrons = 2, 6, 10, 14		
	n = 1, 2, 3, 4 K L M N			

- **Hund's rule of maximum multiplicity** – It states that electron pairing in the degenerate orbitals does not take place until all the available orbitals contains one electron each.
- **Electronic configuration** – The distribution of electrons into various orbitals of an atom is called its electronic configuration.
- **Exchange energy** – During the exchanging process of two or more electrons with the same spin present in degenerate orbitals, the amount of energy released in called exchange energy.

Unit 2 Quantum Mechanical Model of Atom

Choose the best answer

1. Electronic configuration of species M^{2+} is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$ and its atomic weight is 56. The number of neutrons in the nucleus of species M is

a) 26 b) 22 **c) 30** d) 24

2. The energy of light of wavelength 45 nm is

a) 6.67×10^{15}J b) 6.67×10^{11}J **c) 4.42×10^{-18}J** d) 4.42×10^{-15}J

3. The energies E_1 and E_2 of two radiations are 25 eV and 50 eV respectively. The relation between their wavelengths ie λ_1 and λ_2 will be

a) $\frac{\lambda_1}{\lambda_2} = 1$ **b) $\lambda_1 = 2\lambda_2$** c) $\lambda_1 = \sqrt{25 \times 50}\lambda$ d) $2\lambda_1 = \lambda_2$

4. Splitting of spectral lines in an electric field is called

a) Zeeman effect b) Shielding effect c) Compton effect **d) Stark effect**

5. Based on equation $E = -2.178 \times 10^{-18}\ J\ \frac{z^2}{n^2}$, certain conclusions are written. Which of them is not correct ? (NEET)

a) Equation can be used to calculate the change in energy when the electron changes orbit **b) For n = 1, the electron has a more negative energy than it does for n = 6 which means that the electron is more loosely bound in the smallest allowed orbit**

c) The negative sign in equation simply means that the energy of electron bound to the nucleus is lower than it would be if the electrons were at the infinite distance from the nucleus.

d) Larger the value of n, the larger is the orbit radius.