1.Basic Concepts of Chemistry and Chemical Calculations

1) 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)= Average mass of an atom / Unified atomic mass

2) Define relative molecular mass

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

Relative molecular mass = Average mass of a molecule / Unified atomic mass

3) What do you understand by the term mole.

The total number of atoms present in 12 g of carbon -12 isotope.

1 Mole = 6.023×10^{23} entities

4) Define Avogadro Number:

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

5) Define Molar Mass:

Molar mass is defined as the mass of one mole of a substance. The molar mass of a compound is equal to sum of the relative atomic masses of its constituents expressed in g mol^{-1}.

6) Define 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.

7) Define equivalent mass.

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

Equivalent mass has no unit but gram equivalent mass has the unit g eq^{-1}

8) Define 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.

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

9) Define 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. The molecular formula of acetic acid is $C_2H_4O_2$

Molecular formula of a compound is a whole number multiple of the empirical formula. The whole number can be calculated from the molar mass of the compound using the following expression,

Whole number (n) = Molar mass of the compound / Calculated empirical formula mass

10) What is stochiometry?

Stoichiometry gives the numerical relationship between chemical quantities in a balanced chemical equation.

11) What is Limiting Reagents , Excess 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. It limits the further reaction from taking place and is called as the limiting reagent. The other reagents which are in excess are called the excess reagents.

12) What do you understand by the term oxidation number.

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.

13) Distinguish between oxidation and reduction.

Oxidation	Reduction
Addition of oxygen is takes place	Addition of Hydrogen is takes place
Removal of Hydrogen is takes place	Removal of Oxygen is takes place
Removal of electron is takes place	Gain of electron is takes place
Increasing of oxidation number.	Decreasing of oxidation number.

14) Calculate the molar mass of the following compounds.

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

C = 12 x 1 = 12
O = 16 x 1 = 16
N = 14 x 2 = 28
H = 1 x 4 = 4
60 g / mol

ii) acetone [CH_3COCH_3]

C = 12x 3 = 36
H = 1 x 6 = 6
O = 16 x 1 = 16
58 g / mol

iii) boric acid [H_3BO_3]

H = 1 x 3 = 3
B = 11 x1 = 11
O = 16 x 3 = 48
62 g / mol

iv) sulphuric acid [H_2SO_4]

H = 1 x 2 = 2
S = 32 x 1 = 32
O = 16 x 4 = 64
98 g / mol

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

Given :

Density = 1.965 kg m^{-3} at 273 K and 1 atm

Molar Mass of CO_2 = ?

Solution :

At 273 K and 1 atm , Molar volume of CO_2 = 22.4 L

Mass of 1 mol of CO_2 = 1.965 kg / m^3 x 22.4 L

= 1.965 x $\mathbf{10^3}$ g / $\mathbf{m^3}$ x 22.4 x $\mathbf{10^{-3}}$ $\mathbf{m^3}$

Mass of 1 mol of CO_2 = 44 g mol^{-1}

16) 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

Ans: i) 1 mol of ethanol = C_2H_5OH = 1 x 6.023 x 10^{23} = 6.023 x 10^{23} oxygen atoms.

ii) 1 mol of formic acid = HCOOH = 2 x 6.023 x 10^{23} = 12.046 x 10^{23} oxygen atoms.

iii) 1 mol of H_2O = H_2O = 1 x 6.023 x 10^{23} = 6.023 x 10^{23} oxygen atoms.

Hence **Formic acid** contains the greatest number of moles of oxygen atoms.

17) Calculate the average atomic mass of naturally occurring magnesium using the following data, Isotope, Isotopic atomic mass, Abundance (%) are Mg^{24} : 23.99, 78.99, Mg^{26} :24.99 ,10.00 Mg^{25} : 25.98 , 11.01.

Solution: Isotope	Isotopic Atomic Mass	Abundance
Mg^{24}	23.99	78.99
Mg^{26}	24.99	10.00
Mg^{25}	25.98	11.01

Average Atomic Mass = (78.99 x 23.99) + (10 x 24.99) + (11.01 x 25.98) / 100

= 2430.9 / 100 = 24.31 u

18) In a reaction x + y + z_2 ⟶ 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) 1mol 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

Solution:	Given			Consumed			Limiting reagent
	x	+ y	+ z_2	x	+ y	+ z_2	
	200	200	50	50	50	50	z_2
	1	1	3	1	1	1	x and y
	50	25	50	25	25	25	y
	2.5	5	5	2.5	2.5	2.5	x

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

Mass of one atom = 6.645×10^{-23} g

Mass of 1 mole of atom = 6.645×10^{-23} g x 6.023×10^{23} = 40 g

Number of moles of element in 0.320 Kg = 1 mole / 40g x 0.320 kg

= 1 mole x 320 g / 40 g = 8 mol.

20) 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. Ratio of mass of molecule to the unified mass	1. Mass of 1 mole of substance
2. Relative Molecular Mass of any compound is Calculated by adding the relative atomic Masses of its constituent atoms.	2. Molar Mass of a compound is equal to the sum of the relative atomic masses of its constituent atoms.
3. Its Unit is u or amu	3. Its Unit is g / mol
4. **Molecular Mass of CO :** (1 x at.mass of C) + (1x at.mass of O) (1 x 12) + (1x 16) = 28 amu	4. **Molar Mass of CO:** (1 x at.mass of C) + (1x at.mass of O) (1 x 12) + (1x 16) = 28 g / mol

21) 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.

Ans: i) Fructose ($C_6H_{12}O_6$) = Empirical formula CH_2O

ii) Caffeine ($C_8H_{10}N_4O_2$) = Empirical formula $C_4H_5N_2O$

22) The reaction between aluminium and ferric oxide can generate temperatures up to 3273 K and is used in welding metals. (Atomic mass of Al = 27 u Atomic mass of 0 = 16 u)
$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$; 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?

Ans:

	Reactants			Products	
	2Al +	**Fe_2O_3**	$\longrightarrow$	**Al_2O_3 +**	**2Fe**
Amount of substance reacted	324 g	1.12 Kg		-	-
Number of moles reacted	324/27= **12**	1.12×10^{-3} / 160=**7**		-	-
Stoichiometric amount	2	1		1	2
Number of moles consumed	12 moles	6 moles		-	-
Number of moles remains	-	1 mol		6 mol	12 mol

* The mass of Al_2O_3 formed = 6 x 102 = 612 g. [Al_2O_3 = { (2 x 27) + (3 x 16)} = 102]
* The excess reagent is left at the end of the reaction = 1 x 160 = 160 g.[Fe_2O_3 ={ (2 x56) +(3 x 16) }= 160]

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

Balanced equation for the combustion of ethane ,

$C_2H_6 + 7/2\ O_2 \longrightarrow 2\ CO_2 + 3H_2O$; $2\ C_2H_6 + 7\ O_2 \longrightarrow 4\ CO_2 + 6\ H_2O$

4 moles of CO_2 produced by = 2 moles of ethane

1 mole of CO_2 (44g) produced by = 2 moles of ethane / 4 moles of CO_2 x 1 mole of CO_2

= 0.5 mol of ethane.

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

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

25) An organic compound present in vinegar has 40 % carbon, 6.6 % hydrogen and 53.4 % oxygen. Find the empirical formula of the compound.

Element	Percentage	Atomic Mass	Relative no. of moles	Simplest ratio	Simplest ratio (in whole no)
C	40	12	$\frac{40}{12} = 3.3$	$\frac{3.3}{3.3} = 1$	1
H	6.6	1	$\frac{6.6}{1} = 6.6$	$\frac{6.6}{3.3} = 2$	2
O	53.4	16	$\frac{53.4}{16} = 3.3$	$\frac{3.3}{3.3} = 1$	1

26. Explain the general rules for assigning oxidation number of an atom.

1. The oxidation number of the element in the free (or) elementary state is always Zero.
2. The oxidation number of the element in monoatomic ion is equal to the net charge on the ion.
3) The algebric 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.
4) Hydrogen has an oxidation number of +1 in all its compounds except in metal hydrides where it has – 1 value.
5. The oxidaton number of fluorine is always - 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.
7. Alkali metals have an oxidation state of + 1 and alkaline earth metals have an oxidation state of + 2 in all their compounds.

27. Explain the types of Redox Reactions.

1. Combination reactions:

Redox reactions in which two substances combine to form a single compound are called combination reaction.

Example:

$$C + O_2 \longrightarrow CO_2$$

Oxidation; Reduction

2. Decomposition reaction:

Redox reactions in which a compound breaks down into two or more components are called decomposition reactions. The oxidation number of the different elements in the same substance is changed.

Example:

$$2K\overset{+5}{Cl}\overset{-2}{O}_3 \xrightarrow{\Delta} 2K\overset{-1}{Cl} + 3\overset{0}{O}_2$$

Oxidation; Reduction

3. Displacement reactions:

Redox reactions in which an ion (or an atom) in a compound is replaced by an ion (or atom) of another element are called displacement reactions.

(i) Metal displacement reactions:

Example:

$$\overset{0}{Zn} + 2\overset{+1}{H}Cl \longrightarrow \overset{+2}{Zn}Cl_2 + \overset{0}{H}_2$$

Oxidation; Reduction

ii) Non-metal displacement

Example:

$$\overset{+2}{Cu}SO_4\,(aq) + \overset{0}{Zn}(s) \longrightarrow \overset{0}{Cu}(s) + \overset{+2}{Zn}SO_4\,(aq)$$

Oxidation; Reduction

4. Disproportionation reaction (Auto redox reactions)

In redox reactions, the same compound can undergo both oxidation and reduction. In such reactions, the oxidation state of one and the same element is both increased and decreased. These reactions are called disproportionation reactions.

Example:

$$2H_2\overset{-1}{O}_2 \longrightarrow 2H_2\overset{-2}{O} + \overset{0}{O}_2$$

Oxidation; Reduction

2. Quantum Mechanical Model of Atom

IMPORTANT NOTES:

1. Atom: Atom is a smallest particle of all matter.

2. Bohr atom model:The energy of electromagnetic radiation is quantised in units of hν (where ν is the frequency of radiation and h is Planck's constant 6.626×10^{-34} Js).

3. De Broglie equation: $\lambda = h / mv$

4. Circumference of the orbit, $2\pi r = n\lambda$

5. Heisenberg's uncertainty principle : $\Delta x \, . \, \Delta p \geq h / 4\pi$

6. The time independent Schrodinger equation :$H\Psi = E\Psi$

7. Orbital: Region of space around the nucleus within which the probability of finding an electron of given energy is maximum.

8. Shape of orbitals: s- orbital : Symmetrically spherical
   P- orbital : Dumb-bell
   d- orbital: clover leaf (dz^2 is dumb-bell shape with a doughnut)

9. The energy of the electron in the n^{th} orbit: $E_n = -(1312.8)\, Z^2 / n^2$ ($KJ\ mol^{-1}$)

10. The energies of various orbitals: 1s < 2s = 2p < 3s = 3p = 3d <4s = 4p = 4d = 4f < 5s = 5p = 5d = 5f < 6s = 6p = 6d = 6f < 7s

11. Order of energies of various orbitals are based on the (n+l) rule.

12. The effective nuclear charge = The net charge experienced by the electron is called.

13. The order of the effective nuclear charge : s > p > d > f.

14. The energy of the orbitals : s < p < d < f.

15. Filling of orbitals based on Aufbau principle, Pauli exclusion principle and Hund's rule.

16. Aufbau principle: The orbitals are filled in the order of their increasing energies.

17. Pauli Exclusion Principle : "No two electrons in an atom can have the same set of values of all four quantum numbers."

18. Hund's rule : It states that electron pairing in the degenerate orbitals does not take place until all the available orbitals contains one electron each.

19. Electronic configuration : The distribution of electrons into various orbitals of an atom.

2o. Stability of orbitals:

(i) Completely filled	:	s^2	p^6	d^{10}	f^{14}	- More stable
(ii) Half filled	:	s^1	p^3	d^5	f^7	- Stable
(iii) Partially filled	:	-	p^1, p^2 p^4, p^5	d^1, d^2, d^3, d^4 d^6, d^7, d^8, d^9	$f^1, f^2, f^3, f^4, f^5, f^6$ $f^8, f^9, f^{10}, f^{11}, f^{12}, f^{13}$	- Unstable

VERY SHORT ANSWER

1. Which quantum number reveal information about the shape, energy, orientation and size of orbitals?

Principal Quantum number - Size and energy of orbitals.

Magnetic Quantum number – Orientation of orbitals.

Angular Quantum number – Shape of orbitals.

2. The stabilisation of a half filled d - orbital is more pronounced than that of the p-orbital why?

Half-filled d-orbital are having maximum possible exchange (energy) than half filled p-orbital. Hence half-filled d- orbitals are more stable.

3. Define orbital? what are the n and l values for $3p_x$ and $4d_{x^2-y^2}$ electron?

Orbital: An orbital is the region of space around the nucleus within which the probability of finding an electron of given energy is maximum.

Orbital	n	l
$3p_x$	3	1
$4d_{x^2-y^2}$	4	2

4. Bohr Quantum condition:

The circumference of the orbit should be an integral multiple of the wavelength of the electron wave. $2\pi r = n\lambda$

5. Aufbau principle:

The orbitals are filled in the order of their increasing energies.

6. Pauli Exclusion Principle :

"No two electrons in an atom can have the same set of values of all four quantum numbers."

7. Hund's rule

Electron pairing in the degenerate orbitals does not take place until all the available orbitals contains one electron each.

8. Exchange energy:

If two or more electrons with the same spin are present in degenerate orbitals, there is a possibility for exchanging their positions. During exchange process the energy is released and the released energy is called exchange energy.

9. What is Zeeman effect and Stark effect?

The splitting of spectral lines in the presence of magnetic field is called Zeeman effect.
The splitting of spectral lines in an electric field is called Stark effect.

SHORT ANSWER

1. Consider the following electronic arrangements for the d5 configuration.

(a)

↑↓	↑↓	↑		

(b)

↑	↑	↑	↑↓	

(c)

↑	↑	↑	↑	↑

(i) which of these represents the ground state

Ans: (c) , represents the ground s

↑	↑	↑	↑	↑

(ii) which configuration has the maximum exchange energy.

Ans: (

↑	↑	↑	↑	↑

(C) has the maximum of 10 Possible exchanges . (Has maximum exchange energy).

2. How many orbitals are possible for n =4?

n = 4 , l = 0, 1, 2, 3

Where l = 0, s – orbital (4s) = 1

l = 1, p- orbital ($4p_x$,$4p_y$,$4p_z$) = 3

l = 2 , d – orbital (4dxy ,4dyz, 4dzx,$4dx^2$-y^2 ,$4dz^2$) = 5

l = 3 , f- orbital [$4fz^3$, $4fxz^2$, $4fyz^2$,4fxyz, $4fz(x^2-y^2)$,$4fx(x^2-3y^2)$,$4fy(3x^2-y^2)$] = 7

16

Totally there are 16 0rbitals are possible .

3. How many radial nodes for 2s, 4p, 5d and 4f orbitals exhibit? How many angular nodes?

Orbitals	n	l	Radial nodes (n-l-1)	Angular nodes (l value)
2s	2	0	1	0
4p	4	1	2	1
5d	5	2	2	2
4f	4	3	0	3

4. State and explain pauli's exclusion principle.

Pauli's exclusion principle: "No two electrons in an atom can have the same set of values of all four quantum numbers."

It means that, each electron must have unique values for the four quantum numbers (n, *l*, m and s).

Consider the $2s^1$, $2s^2$ electrons in Be atom. The 4 quantum numbers are,

Electron	n	l	m_l	m_s
$2s^1$	2	0	0	+ ½
$2s^2$	2	0	0	-½

5. Explain briefly the time independent schrodinger wave equation?

$$H\Psi = E\Psi \longrightarrow 1$$

$$H = \left\{ \frac{-h^2}{8\Pi^2 m} \left[\frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2} \right] + V \right\}$$

$$(1) \Rightarrow \left\{ \frac{-h^2}{8\Pi^2 m} \left[\frac{\partial^2 \Psi}{\partial x^2} + \frac{\partial^2 \Psi}{\partial y^2} + \frac{\partial^2 \Psi}{\partial z^2} \right] + V\Psi \right\} = E\Psi$$

Multiply by - $8\Pi^2 m / h^2$ and rearrange,

$$\frac{\partial^2 \Psi}{\partial x^2} + \frac{\partial^2 \Psi}{\partial y^2} + \frac{\partial^2 \Psi}{\partial z^2} + \frac{8\Pi^2 m}{h^2} (E - V)\Psi = 0$$

This is the time independent Schrodinger wave equation.

6. Determine the values of all the four quantum numbers of the 8th electron in O- atom and 15th electron in Cl atom and the last electron in chromium.

Answer:

$_8$**O =** $1s^2\ 2s^2\ 2p^4$, the last electron is $\mathbf{2p^4}$.

Quantum numbers: $n = 2$, $l = +1$, $m_l = +1$, $m_s = -\frac{1}{2}$

$_{17}$ **Cl =** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^5$, the last electron is $\mathbf{3p^5}$.

Quantum numbers: $n = 3$, $l = +1$, $m_l = 0$, $m_s = +\frac{1}{2}$

$_{24}$Cr = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 4s^1\ 3d^5$, the last electron is $4s^1$.

Quantum numbers: $n = 4$, $l = 0$, $m_l = 0$, $m_s = +\frac{1}{2}$.

7. For each of the following, give the sub level designation, the allowable m values and the number of orbitals

i) n = 4, l =2, ii) n =5, l = 3 iii) n=7, l=0

Answer :

S.NO	n	l	Sub-shell designation	m	Total
I	4	2	4 d	-2,-1,0,1,2	5 orbitals
II	5	3	5f	-3,-2,-1,0,1,2,3	7 orbitals
III	7	0	7s	0	1 orbital

8. Give the electronic configuration of Mn^{2+} and Cr^{3+}

$_{25}$**Mn =** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 4s^2\ 3d^5$, $\mathbf{Mn^{2+}}$= $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 3d^5$ **(loss 2 e-)**

$_{24}$**Cr =** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 4s^1\ 3d^5$, $\mathbf{Cr^{3+}}$= $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 3d^3$ **(loss 3 e-)**

9. An atom of an element contains 35 electrons and 45 neutrons. Deduce

i) the number of protons, ii) the electronic configuration for the element

iii) All the four quantum numbers for the last electron

Answer:

i) Number of protons : 35

ii) Electronic configuration : $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4P^5$.

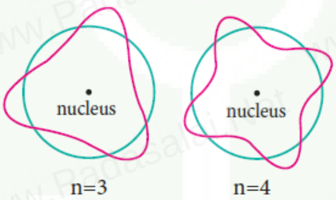
iii) Four quantum numbers ($4P^5$) : $n = 4$, $l = 1$, $m_l = 0$, $m_s = -\frac{1}{2}$

10. Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wave length associated with the electron revolving around the nucleus.

According to the de Broglie concept, the electron that revolves around the nucleus exhibits both particle and wave character.

The circumference of the orbit should be an integral multiple of the wavelength of the electron wave.

$$2\pi r = n\lambda$$



11. Davison and Germer experiment :

The wave nature of electron was experimentally confirmed by Davisson and Germer. They allowed the accelerated beam of electrons to fall on a nickel crystal and recorded the diffraction pattern. The resultant diffraction pattern is similar to the x-ray diffraction pattern.

12. Heisenberg's uncertainty principle

It is impossible to **accurately** determine both the position as well as the momentum of a microscopic particle **simultaneously**'.

$$\Delta x.\Delta p \geq h/4\pi$$

where, Δx uncertainties in the position

Δp uncertainties in the momentum.

13. What are the Limitations of Bohr atom model?

1. It is applicable only to species having one electron such as hydrogen, Li^{2+} etc...
2. It was unable to explain the splitting of spectral lines.
3. It was unable to explain why the electron is revolve around the nucleus in a fixed orbit.

14. Which has the stable electronic configuration? Ni^{2+} or Fe^{3+} .

$Ni^{2+} = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 3d^8$, Partially filled , configuration . Unstable.

$Fe^{3+} = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3s^2\ 3p^6\ 3d^5$, Half filled configuration . More stable.

ANSWER IN DETAIL:

1. Describe the Aufbau principle

Aufbau principle:

In the ground state of the atoms, the orbitals are filled in the order of their increasing energies.
* The electrons first occupy the lowest energy orbital available to them.
* Once the lower energy orbitals are completely filled, then the electrons enter the next higher energy orbitals.
* The order of filling of various orbitals in accordance with (n+/) rule.
* The increasing order of energies of orbitals is as follows:
1s < 2s < 2p < 3s < 3p < 4s <3d <4p <5s<4d < 5p < 6s<4f< 5d < 6p < 7s < 5f < 6d….

2. Postulates of Bohr atom modal.

1. The energies of electrons are quantised
2. The electron is revolving around the nucleus in a circular path called stationary orbit.
3. Electron revolve only in those orbits in which the angular momentum (mvr) of the electron must be equal to an integral multiple of h/2п.

   i.e. mvr = nh/2п --------1 , where n = 1,2,3,...etc.,
4. As long as an electron revolves in the fixed stationary orbit, it doesn't lose its energy. However, when an electron jumps from higher energy state (E_2) to a lower energy state (E1), the excess energy is emitted as radiation.
5. The frequency of the emitted radiation is,

$$E_2 - E_1 = hv$$

and $$v = \frac{(E_2 - E_1)}{h} \text{ -------- } 2$$

Conversely, when suitable energy is supplied to an electron, it will jump from lower energy orbit to a higher energy orbit.

3. Derive De-Broglie's relation.

Planck's quantum hypothesis:

$$E = hv \text{ -------- } 1$$

Einsteins mass-energy relationship

$$E = mc^2 \text{ ---------} 2$$

From 1 and 2

$$h v = mc^2$$

$$h c/\lambda = mc^2$$

$$\lambda = h / mc \text{ -------- } 3$$

c is replaced by v (velocity of particle of matter)

$$\boxed{\lambda = h / mv}$$

4. Explain quantum numbers:

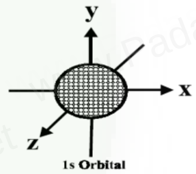
Principal quantum number	Azimuthal quantum number	Magnetic quantum number	Spin quantum number
It explains the electron in different energy levels called orbital or shell.	It explains the electron in different sub-shells or sub-orbital.	It explains the orientation of orbital	It explains the direction of spin of electron.
Denoted by 'n'	Denoted by 'l'	Denoted by 'm'	Denoted by 's'
n= 1,2,3,4,5,………	Value of 'l' depends on the value of 'n' l = 0 to n-1: l=0, s- sub shell, l=1, p- sub shell, l=2, d- sub shell, l=3, f- sub shell,	Value of 'm' depends on the value of 'l'. m= -l to +l	Value of s = + 1/2 or -1/2
Where n=1 , it denotes the first orbit or K- shell; n=2 , it denotes the second orbit or L shell. Etc	n=1 , l = 0; 1s orbital; n= 2, l =0, 1(2s,2p) n = 3, l= 0,1,2(3s,3p, 3d)	l = 0 , m= 0 (one orientation) ; l =1 , m= -1,0 ,+1 (3 orientation) l=2 (5 0rientations)	s = + 1/2 clockwise rotation, s =-1/2 anti - clockwise rotation.

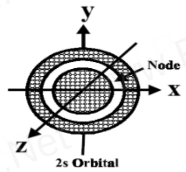
5. Explain the shape of orbital:

s-orbital :

* l=0, m=0, one orientation.
* Symmetrically spherical shape.
* Number of node is found by (n-1).
* The number of radial nodes are equal to (n-l-1).
* Shape of orbital depends on principal quantum number (n).

Diagram:

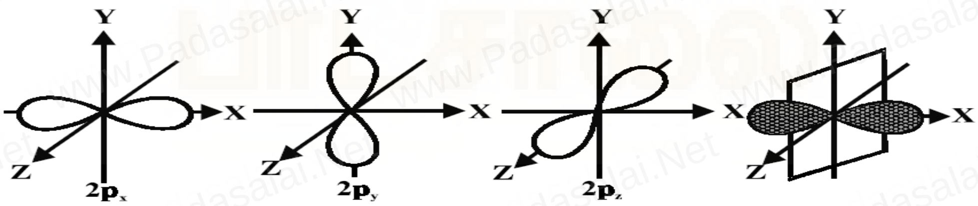




p-orbital:

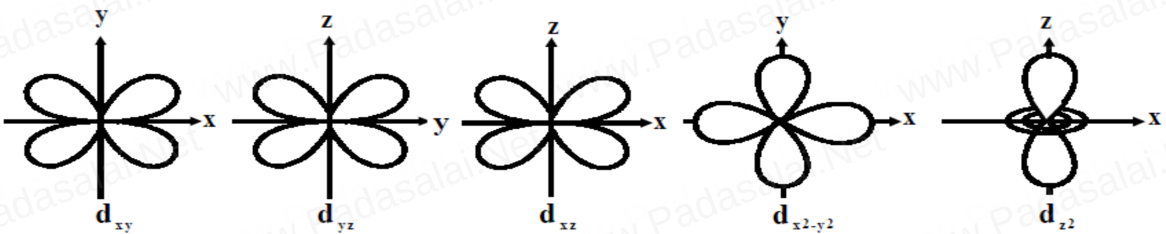
* l=1, m=-1,0,+1, three orientations. p_x,p_y,p_x.
* Dumb-bell spherical shape.
* Each d-orbital contain one node at the centre of the two lobes.

Diagram:



d-orbital:

* l=2, m=-2,-1, 0,+1,+2. 5 orientations. d_{xy}, d_{yz}, d_{xz}, d_{x2-y2}, d_z^2
* clover leaf shape. (dz^2 is dumb-bell shape with a doughnut)



PROBLEMS

1. Calculate the uncertainty in position of an electron, if Δv = 0.1% and u = 2.2 x10^{6}ms^{-1}

Given :

Δv = 0.1%

V = 2.2 x10^{6}ms^{-1}

Δx = ?

Solution:

log 4	0.6021
log 3.14	0.4969
log 9.1	0.959
log 2.2	0.3424
	2.4004
log 6.626	– 0.8213
	– 2.4004
	2.4209
Antilog	2.636 x 10-2

$$\Delta x.\Delta p = h/4\pi$$

$$\Delta x.m\,\Delta v = h/4\pi$$

$$\Delta x = h/4\pi .m.\,\Delta v$$

$$= \frac{6.626 \times 10^{-34}\ Kg.\ m^2.s^{-1}}{4 \times 3.14 \times 9.1 \times 10^{-31}\ Kg \times 2.2 \times 10^{3}\ m\ .\ s^{-1}}$$

$$= \frac{6.626 \times 10^{-6}\ m}{4 \times 3.14 \times 9.1 \times 2.2}$$

$$= 2.636 \times 10^{-8}\ m$$

Answer: $\Delta x = 2.636 \times 10^{-8}$ m

2. The quantum mechanical treatment of the hydrogen atom gives the energy value: E_n = -13.6 / n^2 eV atom^{-1}. i) use this expression to find ΔE between n = 3 and n=4
ii) Calculate the wavelength corresponding to the above transition.

Solution:

(i) when n=3 , $E_3 = -13.6 / 3^2 = -13.6 / 9 = -1.511$ eV atom^{-1} .

when n=4 , $E_4 = -13.6 / 4^2 = -13.6 / 16 = -0.85$ eV atom^{-1} .

$$\Delta E = E_4 - E_3$$

$$= -0.85 - (-1.511) = +0.661 \text{ eV atom}^{-1} .$$

(ii) Wave length , (λ)

$$\Delta E = hC / \lambda$$

$$\lambda = h\,C / \Delta E$$

$$= \frac{6.626 \times 10^{-34}\ Kg.\ m^2.s^{-1} \times 3 \times 10^{8}\ m.\ s^{-1}}{0.661}$$

$$= 10.02 \times 10^{-34} \times 3 \times 10^{8}\ m$$

Answer: $\lambda = 30 \times 10^{-26}$ m

3. How fast must a 54g tennis ball travel in order to have a de Broglie wavelength that is equal to that of a photon of green light 5400A^0?

Given :

m = 54g = 54 x 10 $^{-3}$ Kg

λ = 5400A^0 = 5400 x 10 $^{-10}$ m

V = ?

Solution:

λ = h / mv , v= h / m. λ

$$v = \frac{6.626 \times 10^{-34}\ Kg.\ m^2.s^{-1}}{54 \times 10^{-3}\ Kg \times 5400 \times 10^{-10}\ m}$$

$$= \frac{6.626 \times 10^{-21}\ m\ s^{-1}}{291600}$$

$= 2.273 \times 10^{-26}$ m s $^{-1}$

Answer: $V = 2.273 \times 10^{-26}$ m s $^{-1}$

4. Calculate the energy required for the process.

$He^+_{(g)} \rightarrow He^{2+}_{(g)} + e^-$

The ionisation energy for the H atom in its ground state is - 13.6 eV atom^{-1}.

Solution:

Ionisation energy = -13.6 Z ^{2} / n ^{2} eV atom $^{-1}$

Ionisation energy of hydrogen atom = -13.6 x 1^2 / 1 ^{2} = -13.6 e V atom $^{-1}$

For He$^+$, Z=2 , n=1,

Ionisation energy = -13.6 x 2^2 / 1 ^{2} = - 13.6 x 4 = - 54.4 e V atom $^{-1}$

5. An ion with mass number 37 possesses unit negative charge. If the ion contains 11.1% more neutrons than electrons. Find the symbol of the ion.

Solution:

	Atom	Uni negative ion
Number of electron	x – 1	x
Number of protons	x – 1	x –1
Number of neutrons	y	y

Given that , y = x + 11.1 % of x

$= [x + \frac{11.1 x}{100}] = x + 0.111$

Y = 1.111 x

Mass number = 37

(number of protons + number of neutrons = 37)

(x – 1) + 1.111 x = 37

x + 1.111x = 38

$2.111x = 38$

$x = \frac{38}{2.11} = 18.009$

x = 18 (whole number)

Atomic number = x – 1

= 18 –1 = 17

Mass number = 37

Symbol of the ion $_{17}^{37}Cl^{-}$

6. The Li^{2+} ion is a hydrogen like ion that can be described by the Bohr model. Calculate the Bohr radius of the third orbit and calculate the energy of an electron in 4 th orbit.

Solution :

Bohr radius, r_n = $(0.529)\ n^2 / Z$ A^0 where n= shell , Z = atomic number

Bohr radius of the third orbit r_3 = $(0.529)\ 3^2 / 3$

for Li^{2+} (z = 3) = $(0.529)\ (9) / 3$

$r_3 = 1.587\ A^0$

Energy of an electron E_n = $(-13.6)\ Z^2 / n^2$

Energy of an electron in the fourth orbit (E_4) = $(-13.6)\ 3^2 / 4^2$

(E_4) = – 7.65 eV $atom^{-1}$

7. Protons can be accelerated in particle accelerators. Calculate the wavelength (in A) of such accelerated proton moving at $2.85 \times 10^8\ ms^{-1}$ (the mass of proton is 1.673×10^{-27} Kg).

Given:

$V = 2.85 \times 10^8\ ms^{-1}$

$m = 1.673 \times 10^{-27}$ Kg

$\lambda = ?$

Solution:

$\lambda = h / mv$

$= \frac{6.626 \times 10^{-34}\ Kg.\ m^2.s^{-1}}{1.673 \times 10^{-27}\ Kg \times 2.85 \times 10^8\ ms^{-1}} = 1.390 \times 10^{-15}\ m = 1.390 \times 10^{-5}\ A^o.$

Answer : **$\lambda = 1.390 \times 10^{-5}\ A^o$**

8. What is the de Broglie wavelength (in cm) of a 160g cricket ball travelling at 140 Km hr^{-1}.

Given : $m = 160g = 160 \times 10^{-3}$ Kg

$V = 140\ Km\ hr^{-1} = 140 \times 10^3 / 60 \times 60 = 0.0388 \times 10^3\ m\ s^{-1}$

$\lambda = ?$

Solution :

$\lambda = h / mv = \frac{6.626 \times 10^{-34}\ Kg.\ m^2.s^{-1}}{160 \times 10^{-3}\ Kg \times 0.0388 \times 10^3\ m\ s^{-1}}$

$= 6.626 \times 10^{-34} / 6.2222$

$= 1.0659 \times 10^{-34}$ m $= 1.0659 \times 10^{-36}$ c m

Answer: $\lambda = 1.0659 \times 10^{-36}$ c m

9. Suppose that the uncertainty in determining the position of an electron in an orbit is 0.6 A ^{o}. What is the uncertainty in its momentum?

Given :

$\Delta x = 0.6$ A^{O} $= 0.6 \times 10^{-10}$ m

$\Delta p = ?$

Solution :

$\Delta x.\Delta p = h/4\pi$

$\Delta p = h/4\pi . \Delta x$ $= 6.626 \times 10^{-34}$ Kg. m^2.s^{-1} / $4 \times 3.14 \times 0.6 \times 10^{-10}$ m

$= 6.626 \times 10^{-2}$ Kg. m s^{-1} / 7.536 $= 8.794 \times 10^{-3}$ Kg. m s^{-1}

Answer: $\Delta p = 8.794 \times 10^{-3}$ Kg. m s^{-1}

10. Show that if the measurement of the uncertainty in the location of the particle is equal to its de Broglie wavelength, the minimum uncertainty in its velocity is equal to its velocity / 4п.

Given :

$\Delta x = \lambda$

$\Delta V = ?$

Solution :

$\Delta x.\Delta p = h / 4\pi$

$\Delta x.\, m \Delta V = h / 4\pi$

$\Delta V = h / 4\pi . \Delta x. M$

$= h / 4\pi . \lambda . m$ (since $\Delta x = \lambda$)

$= h / 4\pi . (h/mv) . m$

$= V / 4\pi$

Answer:

$\Delta V = V / 4\pi$ Hence proved.

11. What is the de Broglie wave length of an electron, which is accelerated from the rest, through a potential difference of 100V ?

Given :

Potential difference = 100 V = $100 \times 1.6 \times 10^{-19}$ J

λ = ?

Solution :

$$\lambda = h / \sqrt{2 m e v}$$

$$= \frac{6.626 \times 10^{-34} \text{ Kg. m}^2.\text{s}^{-1}}{\sqrt{2 \times 9.1 \times 10^{-31} \text{ Kg} \times 100 \times 1.6 \times 10^{-19} \text{ J}}}$$

$$= 1.2 \times 10^{-10} \text{ m}$$

Answer: $\lambda = 1.2 \times 10^{-10}$ m

12. Identify the missing quantum numbers and the sub energy level

n	l	m	Sub energy level
?	?	0	4d
3	1	0	?
?	?	?	5p
?	?	-2	3d

Answer:

n	l	m	Sub energy level
4	***2***	0	4d
3	1	0	**3p**
5	1	-1 or 0 or +1	5p
3	2	-2	3d

3. PERIODIC CLASSIFICATION OF ELEMENTS

1. Define modern periodic law.

"The physical and chemical properties of the elements are periodic functions of their atomic numbers."

2. What are isoelectronic ions? Give examples.

An ions with the similar electronic configurations are called as isoelectronic ions.

Ex: Li^{+} : $1s^{2}$ Na + : $1s^{2}\ 2s^{2}\ 2p^{6}$

Be^{2+} : $1s^{2}$ F- : $1s^{2}\ 2s^{2}\ 2p^{6}$

3. What is effective nuclear charge ?

The net nuclear charge experienced by valence electrons in the outermost shell is called the effective nuclear charge.

$Z_{eff} = Z - S$ Z = Atomic number , S = Screening constant

4. Is the definition given below for ionisation enthalpy is correct? "Ionisation enthalpy is defined as the energy required to remove the most loosely bound electron from the valence shell of an atom"

Yes , the definition is correct.

5. Magnesium loses electrons successively to form Mg^{+}, Mg^{2+} and Mg^{3+} ions. Which step will have the highest ionisation energy and why?

Mg + Energy (IE_1) $\rightarrow$ Mg^{+} + e-

($3s^2$) ($3s^1$)

Mg^{+} + Energy (IE_2) $\rightarrow$ Mg^{2+} + e-

($3s^1$) ($2p^6$)

Mg^{2+} Energy (IE_3) $\rightarrow$ Mg^{3+} + e-

($2p^6$) ($2p^5$)

Generally each successive IE is greater than that of the previous one .

$IE_1 < IE_2 < IE_3$. Hence Mg^{2+} to Mg^{3+} requires more IE.

6. Define electronegativity.

It is defined as the relative tendency of an element present in a covalently bonded molecule, to attract the shared pair of electrons towards itself.

7. How would you explain the fact that the second ionisation potential is always higher than first ionisation potential?

The effective nuclear charge of the cation is higher than the corresponding neutral atom. The cation require more energy to release one electron then neutral atom. **$IE_1 < IE_2$.**

8. Energy of an electron in ground state of the hydrogen atom is -2.8×10^{-18} J. Calculate the ionisation enthalpy of atomic hydrogen in terms of kJ mol^{-1}.

Solution :

The ground state of 1 hydrogen atom = -2.8×10^{-18} J

The ground state of 1 mole hydrogen atom = -2.8×10^{-18} J x 6.023×10^{23}

= - 1312 KJ / Mole

Ionisation energy = $E_\infty - E_1$

= 0 - (-1312)

= +1312 KJ / Mole

9. The electronic configuration of atom is one of the important factor which affects the value of ionisation potential and electron gain enthalpy. Explain

The electronic configuration of atoms are 3 types,

(i) Completely filled : s^2 p^6 d^{10} f^{14} - More stable (More IE) , Zero EA

(ii) Half filled : s^1 p^3 d^5 f^7 - Stable (Less IE) , More EA

(iii) Partially filled : - p^1, p^2 d^1, d^2, d^3, d^4 - Unstable (Very low IE) , Less EA
p^4, p^5 d^6, d^7, d^8, d^9

* Ionisation Energy : Ne : $1s^2\ 2s^2\ 2p^6$ [⇅][⇅][⇅] Completely filled - More IE
N : $1s^2\ 2s^2\ 2p^3$ [1][1][1] Half filled - Less IE
O : $1s^2\ 2s^2\ 2p^4$ [⇅][1][1] Partially filled - very low IE

* Electron Affinity : Ne : $1s^2\ 2s^2\ 2p^6$ [⇅][⇅][⇅] Completely filled - Zero EA
N : $1s^2\ 2s^2\ 2p^3$ [1][1][1] Half filled - More EA
O : $1s^2\ 2s^2\ 2p^4$ [⇅][1][1] Partially filled - Less EA

10. In what period and group will an element with Z = 118 will be present?

Element with Z= 118 will occupy in period number 7 and Group number 18 (Noble gas) .
Z= 118 (Og) : [Rn] ^{86} $5f^{14}\ 6d^{10}\ 7s^2\ 7p^6$.

11. Justify that the fifth period of the periodic table should have 18 elements on the basis of quantum numbers.

According to Aufbau principle , 5 th period elements filled in the orbitals like s,d,p ($5s^2$, $4d^{10}$, $6p^6$) , totally 18 electron accommodate in these orbitals . Hence 5 th period contains 18 elements.

12. Elements a, b, c and d have the following electronic configurations:
a) $1s^2, 2s^2, 2p^6$ b) $1s^2, 2s^2, 2p^6, 3s^2, 3p^1$ c) $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ d) $1s^2, 2s^2, 2p^1$
Which elements among these will belong to the same group of periodic table?

Elements with equal electron in outermost shell belongs to same group.
Hence a & c belongs to same group. And b & d belongs to same group.

13. Give the general electronic configuration of lanthanides and actinides?

The lanthanides ($4f^{1-14}$, $5d^{0-1}$, $6s^2$) and the actinides ($5f^{0-14}$, $6d^{0-2}$, $7s^2$)

14. Why halogens act as oxidising agents?

Halogens have high electron negativity and electron affinity values.
Halogens have a unstable np^5 electronic configuration. So they easily gaining one electron it becomes a Stable Fully filled np6 electronic Configuration.
Hence it act as a strong oxidising agents.

15. Mention any two anomalous properties of second period elements.

Lithium and Beryllium form covalent compounds but other forms ionic compounds
The second period has only 2 orbitals (2s and 2p) ,the maximum valency is 4. But others have more orbital in the valence shell and higher valencies.

Example Boron forms BF^{-4} but Aluminium forms AlF_6^{-3}

16. Explain the pauling method for the determination of ionic radius.

Pauling assumed that ions present in a crystal lattice are perfect spheres, and they are in contact with each other therefore,

i) The sum of the radii will be equal to the inter nuclear distance between them.

$$r(C^+) + r(A^-) = d\,(C^+ - A^-) \qquad (1)$$

where $r(C^+)$ = radius of cation, $r(A^-)$: radius of anion,

$d(C^+ - A^-)$ = internuclear distance between C^+ and A^- ions .

ii) Pauling also assumed that the radius of the ion having noble gas electronic configuration (Na^+ and Cl^- having $1s^2\ 2s^2,\ 2p^6$ configuration) is inversely proportional to the effective nuclear charge felt at the periphery of the ion.

$$\text{i.e. } r_{C^+} \alpha \frac{1}{(Z_{eff})_{C+}} \text{ ——— } (1)$$

and

$$r_{A^-} \alpha \frac{1}{(Z_{eff})_{A^-}} \text{ ——— } (3)$$

Where Z_{eff} is the effective nuclear charge and $Z_{eff} = Z - S$

Dividing the equation 1 by 3

$$\frac{r_{C^+}}{r_{A^-}} = \frac{(Z_{eff})_{A^-}}{(Z_{eff})_{C^+}} \text{ -----}(4)$$

On solving equation (1) and (4) the values of r_{C^+} and r_{A^-} can be obtained

Let us explain this method by calculating the ionic radii of Na+ and F- in NaF crystal whose interionic distance is equal to 231 pm .

17. Explain the periodic trend of ionisation potential.

Ionization energy is the amount of energy required to remove an loosely bounded electron from the outermost shell of an atom. Unit is eV

<u>Along the Group</u> : It decreases along the group.

Reason: As we move down the group the valence electrons are added into new shells. As a result the distance between the nucleus and the valence electrons increases. Hence the nuclear charge decreases and the ionization energy also decreases.

<u>Along the period</u> : It increases along the period

Reason: As we move along the period the valence electrons are added to the same shell. So the Nuclear charge increases, And the attraction between the valence electron and the nucleus increases. Hence more energy is required to remove the valence electron, so Ionization energy increases.

18. Explain the diagonal relationship.

On moving diagonally across the periodic table, the second and third period elements show certain similarities.

Even though the similarity is not same as we see in a group, it is quite pronounced in the following pair of elements.

Li ↘	Be ↘	B ↘	C
Na	Mg	Al	Si

The similarity in properties existing between the diagonally placed elements is called 'diagonal relationship'.

19. Why the first ionisation enthalpy of sodium is lower that that of magnesium while its second ionisation enthalpy is higher than that of magnesium?

Why the IE_1 value of sodium is lower than Mg. but the IE_2 value is higher than Mg.

Na11 = $1s^2\ 2s^2\ 2p^6\ 3s^1$

Mg12 = $1s^2\ 2s^2\ 2p^4\ 3s^2$

- First IE value of Mg > Na . Because Mg is smaller in size and has high nuclear charge.
- But in second IE value Mg < Na. because Sodium has fully filled Stable np^6 electronic configuration. And has higher nuclear charge.

20. By using paulings method calculate the ionic radii of K+ and Cl- ions in the potassium chloride crystal. Given that dK+-Cl- = 3.14 A

Solution

$$r(K^+) + r(Cl^-) = d(K^+\text{–}Cl^-) = 3.14\ \text{Å} \qquad (1)$$

K^+ and Cl^- ions have Ar (Z=18) type configuration. The effective nuclear charge for K^+ and Cl^- can be calculated as follows.

K^+	=	$(1s^2)$ innershell	$(2s^2\,2p^6)$ $(n-1)^{th}$ shell	$(3s^2\,3p^6)$ n^{th} shell

$$Z^*(K^+) = Z - S$$
$$= 19 - [(0.35 \times 7) + (0.85 \times 8) + (1 \times 2)]$$
$$= 19 - 11.25 = 7.75$$
$$Z^*(Cl^-) = 17 - [(0.35 \times 7) + (0.85 \times 8) + (1 \times 2)]$$
$$= 17 - 11.25 = 5.75$$

$$\therefore \frac{r(K^+)}{r(Cl^-)} = \frac{Z^*(Cl^-)}{Z^*(K^+)} = \frac{5.75}{7.75} = 0.74$$

$$\therefore r(K^+) = 0.74\, r(Cl^-) \qquad (2)$$

Substitute (2) in (1)

$$0.74\, r(Cl^-) + r(Cl^-) = 3.14\ \text{Å}$$
$$1.74\, r(Cl^-) = 3.14\ \text{Å}$$
$$r(Cl^-) = \frac{3.14\text{Å}}{1.74} = 1.81\text{Å}$$

From (2)

$$r(K^+) = 0.74\, r(Cl^-)$$
$$= 0.74 \times 1.81\ \text{Å}$$
$$= 1.33\ \text{Å}$$
$$r(K^+) = 1.33\ \text{Å}$$
$$r(Cl^-) = 1.81\ \text{Å}$$

21. Explain the following, give appropriate reasons.

(i) Ionisation potential of N is greater than that of O

N_7 = 1s2 2s2 2p3

O_8 = 1s2 2s2 2p4

Nitrogen has high Nuclear Charge

Nitrogen has Stable Half filled np^3 electronic configuration.

Hence Ionisation energy of N > O

(ii) First ionisation potential of C-atom is greater than that of B atom, where as the reverse is true is for second ionisation potential.

$B_5 = 1s^2 2s^2 2p^1$

$C_6 = 1s^2 2s^2 2p^2$

First IE value of C >B . Because Carbon is smaller in size and has high nuclear charge. So Carbon has more Ionisation energy Boron.

But in second IE value C < B .Because Boron has fully filled Stable ns^2 electronic configuration. And has high nuclear charge.

(iii) The electron affinity values of Be, Mg and noble gases are zero and those of N (0.02 eV) and P (0.80 eV) are very low

$Mg_{12} = 1s^2 2s^2 2p^6 3s^2$

$Ne_{10} = 1s^2 2s^2 2p^6$

$Be_4 = 1s^2 2s^2$

- Be, Mg and Noble gases are having fully filled stable electronic configuration.
- These elements will not accept electrons. Hence they have zero EA.

(iv) The formation of $F^-_{(g)}$ from $F_{(g)}$ is exothermic while that of $O^{2-}_{(g)}$ from $O_{(g)}$ is endothermic.

$F_{(g)} \longrightarrow F^-_{(g)}$ (exothermic)

$2s^2 2p^5$ **$2s^2 2p^6$**

$O_{(g)} \longrightarrow O^{2-}_{(g)}$ (endothermic)

$2s^2 2p^4$ **$2s^2 2p^6$**

45. What is screening effect? Briefly give the basis for pauling's scale of electronegativity.

The inner shell electrons act as a shield between the nucleus and the valence electrons. This effect is called shielding effect.

The net nuclear charge experienced by valence electrons in the outermost shell is called the effective nuclear charge.

Zeff = Z – S Z = Atomic number , S = Screening constant.

One such method was developed by Pauling, he assigned arbitrary value of electronegativities for hydrogen and fluorine as 2.2 and 4.0 respectively. Based on this the electronegativity values for other elements can be calculated using the following expression

$X_A - X_B$) = 0.182 $\sqrt{E_{AB} - (E_{AA} \times E_{BB})}$.

Where E_{AB}, E_{AA} and E_{BB} are the bond dissociation energies of AB, A_2 and B_2 molecules respectively.

The electronegativity of any given element is not a constant and its value depends on the element to which it is covalently bound. The electronegativity values play an important role in predicting the nature of the bond.

46. State the trends in the variation of electronegativity in group and periods.

Electro negativity is a tendency of a element present is covalent molecule to attract the shared pair of electrons towards itself.

Along the Group : It decreases along the group.

Reason

- As we move down the group the nuclear charge decreases
- The atomic size increases.

- **Along the period :** It increases along the period

Reason

- As we move along the period the nuclear charge increases
- The atomic size decreases.
- The attraction between the valence electron and the nucleus Increases

47. Define Atomic radius, Electron Affinity.

Atomic radius

Atomic radius of an atom is defined as the distance between the centre of its nucleus and the outermost shell containing the valence electron.

Electron Affinity

It is defined as the amount of energy released (required in the case noble gases) when an electron is added to the valence shell of an isolated neutral gaseous atom in its ground state to form its anion. It is expressed in kJ mol-1

$$A + 1\ e^- \longrightarrow A^- + EA$$

4. Hydrogen

1. What are the similarity of hydrogen with Alkali metals

Hydrogen have $1s^1$ and Alkali metals have n s^1 electronic configuration
Like Alkali metals, Hydrogen forms unipositive ions. Eg H+ and Na+
Like Alkali metals, Hydrogen forms Oxides(H2O) and sulphides(H2S).
It also acts as a reducing agent.

2. Define isotopes and explain the types of isotopes of Hydrogen.

Elements having same atomic number but different mass number is called as isotopes.
Hydrogen has 3 isotopes.
Protium = $1H^1$, Deuterium = $1H^2$,Tritium = $1H^3$

3. Explain Ortho and Para hydrogen

In hydrogen molecule, if the two nuclei rotates in the same direction is called as Ortho hydrogen
In a hydrogen molecule, if the two nuclei rotates in the opposite direction is called as Para hydrogen

4. Give the difference between Ortho and para hydrogen

S.No	ORTHO HYDROGEN	PARA HYDROGEN
1.	Both the nuclei rotates in the same direction	Both the nuclei rotates in the opposite direction
2.	75% at room temperature	25% at room temperature
3.	It is more stable	It is less stable
4.	Its melting point is 13.95K	Its melting point is 13.83K
5.	Its boiling point is 20.39K	Its boiling point is 20.26K
6.	It has a net magnetic moment	It has Zero magnetic moment

5. How will you convert Para hydrogen to Ortho hydrogen

- By using catalyst like Iron
- By passing electric discharge
- By heating at 8000C
- By mixing with paramagnetic molecules like oxygen
- By mixing with atomic hydrogen

6. Predict the gas and solid for HCl and NaH. Give reasons.

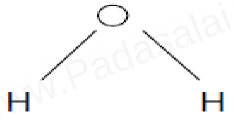
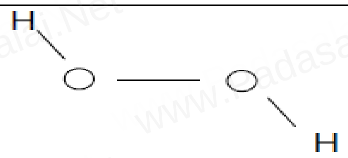
- HCl is a gas because of absence of intermolecular hydrogen bonding
- NaH is a solid because of electron transfer.

7. NH_3, H_2O and HF arrange in the order of hydrogen bonding

- The order of increasing strength of Hydrogen bonding is HF > H2O > NH3
- The strength of hydrogen bonding increases with the increase in Electro negativity.

- The order of electro negativity is F > O > N.
- Hence HF has the strongest Hydrogen bonding.

8. Compare the structure of water and peroxide.

S.Nm	H_2O	H_2O_2
1	It is bent structure	It is Open book structure
2	The bond angle is 104.5^0	The bond angle is 90.2^0
3		

9. Give the preparation of Tritium

Tritium is prepared by bombarding Lithium with slow neutrons.

$$3Li^6 + 0n^1 \longrightarrow 1T^3 + 2He^4$$

10. Show that Tritium is a Beta emitter. (Give the Properties)

$$1T^3 \longrightarrow 2He^4$$

The Half- life period of Tritium is 12.3 years.

11. What is Syngas or Water gas or synthetic gas and give its use

When Carbon reacts with steam at 1000^0 C, it gives a mixture of Carbon monoxide and hydrogen. This mixture is called as Water gas or Syngas.

$$C + H_2O \xrightarrow{1000\,^oC} H_2 + CO$$

12. Explain the water gas shift reaction.

When carbon monoxide is reacts with steam at 4000C in the presence of iron catalyst it gives hydrogen.

$$CO + H_2O \xrightarrow{400\,^o C} H_2 + CO_2$$

13. List the uses of Heavy water

It is used as Moderators in Nuclear reactor.
It is used as tracer element to study the mechanisms of organic reactions.
It is used as coolant in nuclear reactors to absorb the heat.

14. List the uses of Deuterium

It is used as tracer element to study the mechanisms of organic reactions.
High speed deuterium is used in Artificial radio activity
It is used in the preparation of heavy water which used as Moderators in nuclear reactor

15. List the uses of Hydrogen

Liquid hydrogen is used as Rocket Fuel
Atomic hydrogen is used for cutting and welding metals
Hydrogen is used for preparing Fertilizer and explosives
Hydrogen is used in Fuel cells for generating electricity.
It is used as catalyst for the preparation of Vanaspathi.
It is used for the preparation of Methanol and industrial solvent

16. What is Soft water and Hard water

Water free from calcium and magnesium salts is called soft water.

Water containing chlorides and sulphate of magnesium & calcium ions is called as Hard water.

17. What is Temporary hardness and how it is removed

Temporary hardness is due to the presence of Bicarbonates of Magnesium and Calcium.

It can be removed by boiling and filtration.

18. What is permanent Hardness and how it is removed

Temporary hardness is due to the presence of Chlorides and Sulphates of Magnesium and Calcium. It is removed by adding washing soda.

19. What is 30% solution of hydrogen peroxide? Or what is 100 volume of hydro peroxide?

30 % solution of hydrogen peroxide is called as '100-volume' hydrogen peroxide. At STP on heating 1 ml of hydrogen peroxide liberated 100 ml of oxygen.

20. Why hydrogen peroxide is store in Plastic bottles not in Glass bottles

It dissolves the Alkali metals present in glass.

It undergoes a Catalyzed disproportionation reaction. So it is stored in Plastic bottles.

21. List the uses of Hydrogen peroxide

It is used as Antiseptic

It is used to bleach paper and textile

It is used in water treatment to oxidize pollutant in water.

22. How Hydrogen peroxide is used to restore the colour of old paintings

Hydrogen peroxide is used to restore the white colour of the old paintings

Hydrogen sulphide in air reacts with the white pigment to form a black colored lead sulphide.

Hydrogen peroxide reacts lead sulphide to give white coloured lead sulphate.

23. What is hydrogen bonding

When hydrogen is covalently bonded to a highly electronegative atom such as fluorine, oxygen and nitrogen, the bond is polarized.

24. What are Types of hydrogen bonding?

Intra molecular hydrogen bonding :

The hydrogen bonding which occur within the molecule.

Eg. Ortho Nitro Phenol

Inter molecular hydrogen bonding:

The hydrogen bonding which occur between two or more molecules

Eg. Water.

25. What are the importance of Hydrogen bonding

It plays an important role in bio molecules like proteins.

It plays a important role in the structure of DNA,

It holds the two helical Nucleic acid chains of the DNA together.

7. THERMODYNAMICS

1. State the first law of thermodynamics.

The total energy of an isolated system remains constant though it may change from one form to another.

2. Define Hess's law of constant heat summation.

The enthalpy change of a reaction either at constant volume or constant pressure is the same whether it takes place in a single or multiple steps provided the initial and final states are same.

3. Explain intensive properties with two examples

The property that is independent of the mass or the size of the system is called an intensive property.

Examples : Refractive index, density, temperature, Boiling point, Freezing point, molar volume, etc.,

4. Define the following terms:

a. isothermal process: The temperature of the system remains constant, during the change from its initial to final state.

b. adiabatic process: There is no exchange of heat (q) between the system and surrounding during the process.

c. isobaric process: The pressure of the system remains constant during its change from the initial to final state.

d. isochoric process: The volume of system remains constant during its change from initial to final state.

5. What is the usual definition of entropy? What is the unit of entropy?

Entropy is a measure of the molecular disorder (randomness) of a system. It is defined as, $dS = dq_{rev} / T$

6. Predict the feasibility of a reaction when

i) both ΔH and ΔS positive : Non-feasible

ii) both ΔH and ΔS negative :Non- feasible

iii) ΔH decreases but ΔS increases : Feasible

7. Define is Gibb's free energy.

Gibbs free energy is defined as below G = H – TS , H= Enthalpy , T= Temperature, S= Entropy.

8. Define enthalpy of combustion.

The change in the enthalpy when one mole of the substance is completely burnt in excess of air is called as Heat of Combustion.

9. Define molar heat capacity. Give its unit.

The amount of heat absorbed by one mole of a substance to raise its temperature by one Kelvin is called as Molar heat capacity.

Its Unit = J K-1 mol-1

10. Define the calorific value of food. What is the unit of calorific value?

The amount of heat produced when one gram of a substance is completely burnt, is called as Calorific value of food. Its unit = J kg -1

11. Define enthalpy of neutralization.

The change in the enthalpy when one gram equivalent of acid is completely neutralized by one gram equivalent of a base.

12. What is lattice energy?

The amount of energy required to completely remove the ions from its crystal lattice to an infinite distance is called as lattice energy.

13. What are state and path functions? Give two examples.

state function:

A state function is a thermodynamic property of a system whose value does not depends on the path by which the system changes from its initial to final state. Eg. **P, V and T**

Path function:

A Path function is a thermodynamic property of a system whose value depends on the path by which the system changes from its initial to final state. Eg. **Work and Heat.**

14. Give Kelvin statement of second law of thermodynamics.

It is impossible to construct a machine that absorbs heat from a hot source and converts it completely into work without transferring a part of heat to a cold sink.

15. The equilibrium constant of a reaction is 10, what will be the sign of ΔG? Will this reaction be spontaneous?

The relationship between Free energy and **equilibrium constant k is,**

$\Delta G^{O} = - RT \ln K$

K , T are positive , then ΔG^{O} will be negative . When ΔG^{O} negative the process is spontaneous.

16. Enthalpy of neutralization is always a constant when a strong acid is neutralized by a strong base: account for the statement.

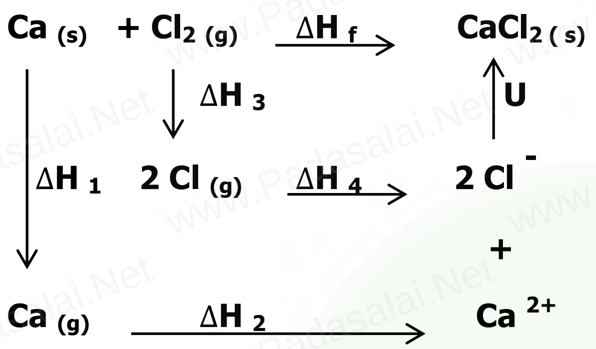
$H^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow H_2O_{(l)} \quad \Delta H = -57.32\ kJ$

The heat of neutralisation of a strong acid and strong base is around – 57.32 kJ.

17. State the third law of thermodynamics.

Thus the third law of thermodynamics states that the entropy of pure crystalline substance at absolute zero is zero.

18. Write down the Born-Haber cycle for the formation of $CaCl_2$.



$\Delta H_f = \Delta H_1 + \Delta H_2 + \Delta H_3 + 2\,\Delta H_4 + U$

19. Identify the state and path functions out of the following:

a)Enthalpy b)Entropy c) Heat d) Temperature e) Work f)Free energy.

Ans: Heat ,Work

20. State the various statements of second law of thermodynamics.

i) Kelvin- Planck statement

"It is impossible to construct an engine which operated in a complete cycle will absorb heat from a single body and convert it completely to work without leaving some changes in the working system".

ii) Clausius statement:

"It is impossible to transfer heat from a cold body to a hot body by a machine without doing some work".

iii) Entropy statement:

'A process accompanied by increase in entropy tends to be spontaneous".

iv) "Efficiency of a machine can never be cent percent".

$$\% \text{ Efficiency} = \frac{(T_1 - T_2)}{T_1} \times 100,$$

By II law, $T_2 < T_1$ % efficiency less than 100.

21.What are spontaneous reactions? What are the conditions for the spontaneity of a process?

A reaction that occurs under the given set of conditions without any external driving force is called a spontaneous reactions.

* All the natural processes are spontaneous.
* It occur its own and does not need to be induced.
* A spontaneous process is accompanied by increase in the Entropy.
* For a spontaneous process, the enthalpy change at constant pressure will be negative.
* Combining negative ΔH and positive ΔS, ΔG get negative sign.

Essential conditions **:** $\Delta G < 0$, $\Delta H < 0$, $\Delta S > 0$

22. List the characteristics of internal energy.

1. It is an extensive property.
2. It is a state function
3. Change in internal energy is . $\Delta U = U_f - U_i$
4. In cyclic process , $\Delta U = 0$.
5. $\Delta U = U_f - U_i$ = -ve ($U_f < U_I$)
6. $\Delta U = U_f - U_i$ = +ve ($U_f > U_I$).

23. List the characteristics of Gibbs free energy.

1) Free energy is defined as G = (H-TS). 'G' is a state function.

2) G-Extensive property. ΔG- become intensive property , when the system is closed.

3) G has a single value for the thermodynamic state of the system.

4) $\Delta G<0$ - spontaneous , $\Delta G = 0$ – equilibrium, $\Delta G > 0$ - non- spontaneous

5) $\Delta G = \Delta H - T\Delta S$. $\Delta H = \Delta E + P\Delta V$ and $\Delta E = q - w$. But $T\Delta S = q$

$\Delta G = q - w + P\Delta V - q$. **$\Delta G = - w + P\Delta V$ = network.**

24. Define Zeroth law of thermodynamics

If two systems are in thermal equilibrium with a third one, then they tend to be thermal equilibrium with themselves.

25. List the application of the bomb calorimeter.

It is used to determine the amount of heat released in a combustion reaction
It is used to determine the calorific value of food
It is used in food processing and explosive testing industries.

HARD WORK NEVER FAILS

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