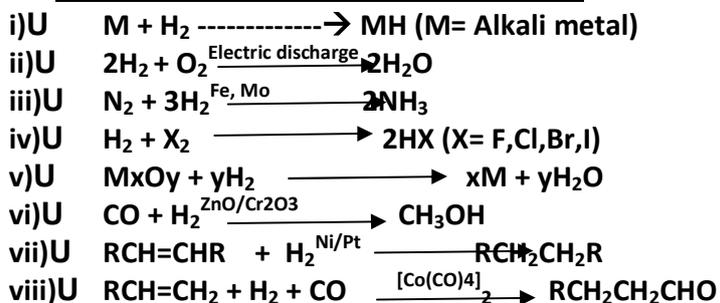
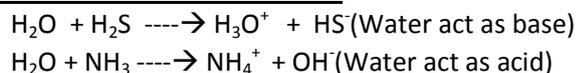


(Hydrogen)

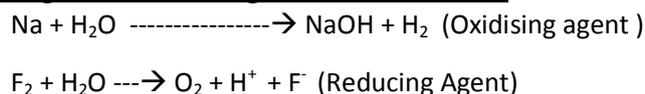
• Chemical Properties of Dihydrogen :



• Amphoteric nature of water

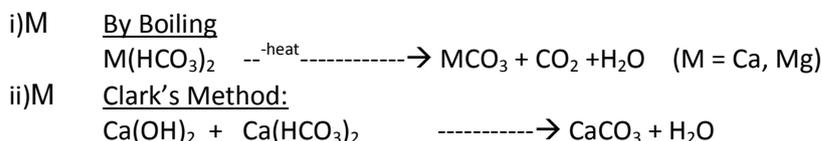


• Oxidising and reducing nature of water

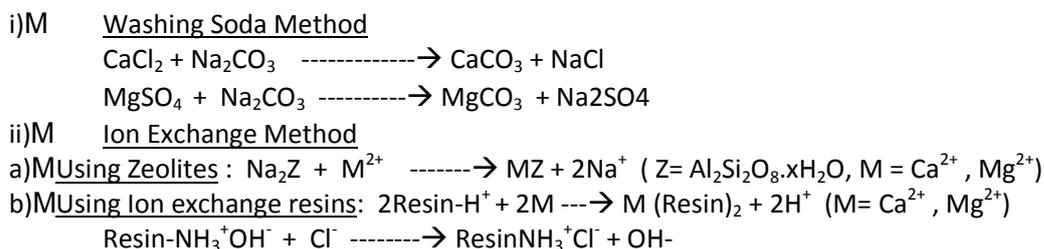


• Softening Of Water:

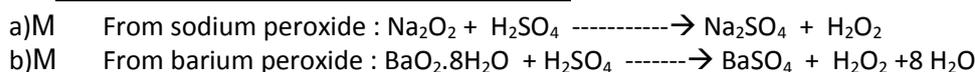
Temporary Hardness:



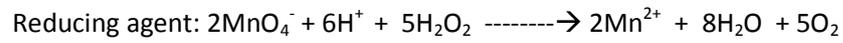
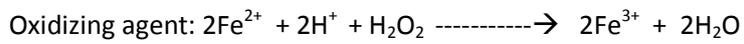
Permanent Hardness:



• Preparation of hydrogen peroxide



• **MOxidising and reducing nature of water**



- **M** Strength of hydrogen peroxide is commonly expressed as volume. This refers to the volume of oxygen which a solution of H₂O₂ will give. Eg. A 20 volume of hydrogen peroxide means that 1 litre of this solution will give 20 litres of oxygen at N.T.P.

Chapter: Environmental Chemistry

1. Name the gas that caused the Bhopal Gas Tragedy.

Ans: Methyl isocyanide

2. List the gases which are responsible for green house effect.

Ans: Carbon dioxide, methane, nitrous oxide, ozone, water vapours and chlorofluorocarbons.

3. What is the role of CO₂ in the 'Green house effect'?

Ans: Heat from the sun after being absorbed by the earth is re-emitted by the earth and absorbed by CO₂ which then radiate it back to the earth, thus, maintaining the constant temperature of the earth.

4. Carbon monoxide gas is more dangerous than carbon dioxide gas. Why?

Ans: Carbon monoxide gas is a poisonous gas as it has more tendency to bind to haemoglobin to form carboxylhaemoglobin than oxygen. Thus, reduces the oxygen carrying capacity of the blood.

CO₂ gas does not combine with blood. Thus, CO₂ is less dangerous.

5. Acid rain is known to contain some acids. Name these acids and from where they come in the rain?

Ans: Sulphuric acid, Nitric acid and carbonic acid.

These acids are formed from the gases like oxides of nitrogen and sulphur which are released in the atmosphere by industries and automobiles by the combustion of coal, petrol, diesel etc.

6. Ozone layer is necessary for life. Why?

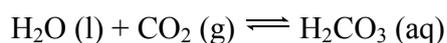
Ans: Ozone absorbs the harmful UV radiations coming from the sun and thus, protects human beings from adverse effect of UV radiations.

7. CO₂ keeps the earth warm. Give reason.

Ans: CO₂ have tendency to absorb most of the heat radiation that are emitted by objects of the earth. Thus, it keeps the earth warm.

8. The pH of normal rain water is 5.6.

Ans: This is because of the dissolution of CO₂ from atmosphere which furnish H⁺ ions to the rain water.



9. Differentiate between Classical smog and Photochemical Smog.

Ans:

CLASSICAL SMOG (London smog)	PHOTOCHEMICAL SMOG (Los Angeles smog)
-------------------------------------	--

1. It is due to the presence of SO ₂ and humidity in the air.	1. It is because of the photochemical reaction of air containing NO ₂ and hydrocarbons.
2. It is reducing in nature.	2. It is oxidizing in nature.
3. It involves smoke and fog.	3. It does not contain smoke and fog.
4. It causes bronchitis irritation	4. It causes irritation in the eyes.

10. Why is photochemical smog so called?

Ans: It is because it is formed as a result of photochemical reaction between oxides of nitrogen and hydrocarbons.

11. What are the secondary pollutants of photochemical smog?

Ans: Nitric oxide (NO), acrolein, formaldehyde and PAN.

12. Expand PAN.

Ans: Polyacrylonitrile

13. Which main compounds are causing damage to ozone layer?

Ans: NO and Chlorofluorocarbons (Freons)

14. What is the nature of London Smog and why?

Ans: Reducing because of the presence of SO₂ and carbon soot which are reducing g in nature.

15. In which part of the atmosphere ozone layer is present?

Ans: Stratosphere

16. What is Biochemical Oxygen Demand?

Ans: It is defined as the total amount of oxygen (in milligram) required by microbes to decompose the organic matter present in certain volume of water sample.

$$\text{BOD} = \frac{\text{Amount of oxygen required (in mg)}}{\text{Volume of water sample (in L)}}$$

17. What are pesticides and herbicides? Give an example of each.

Ans: Pesticides: These substances are used to kill or block the reproductive processes of the pests. Eg: DDT, dieldrin etc.

Herbicides: These are weed killers i.e. they kill unwanted herbs growing along crops. Eg. Sodium Chlorates, Triazines etc.

18. What is green chemistry?

Ans: Green Chemistry is a branch of chemistry in which chemicals of daily needs are produced by using such processes that neither use toxic chemicals nor emit such chemicals into the atmosphere.

19. What is biomagnification?

Ans: Biomagnification is increase in the amount of pesticides per unit weight of organism through food chain from one tropic level to other.

20. What is Eutrophication?

Ans: Eutrophication is the over population of plant in water bodies due to the presence of large amount of nutrients like NO_3^- , PO_4^{3-} which results in killing of animal life.

21. How does H_2O_2 find application in green chemistry and why?

Ans: H_2O_2 is used for bleaching clothes as results obtained by it are much better and less amount of water is utilized in this method.

22. Why is tetrachloroethane replaced by liquid CO_2 in dry-cleaning?

Ans: Tetrachloroethane contaminates the ground water and is also suspected carcinogen.

CHAPTER: STRUCTURE OF AN ATOM

1. Write the spectrum of electromagnetic radiations in the increasing order of their wavelengths or (decreasing frequencies).

Ans: Cosmic rays < γ - rays < X- Rays < Ultraviolet rays < Visible < Infrared < Microwaves < Radiowaves

2. What is the difference between a quantum and a photon?

Ans: The smallest packet of energy of any radiation is called a quantum whereas that of light is called a photon.

3. What is photoelectric effect?

Ans: The phenomenon of ejection of electron from the surface of the metal when a light of suitable frequency strikes on it is called photoelectric effect.

4. Which series of lines of the hydrogen spectrum lie in the visible region?

Ans: Balmer Series

5. Differentiate between orbit and orbital.

Ans:

ORBIT	ORBITAL
1. It is a well defined circular path around the nucleus in which the electron revolves.	1. It is the space around the nucleus within which the probability of finding the electron is maximum.
2. They do not have any directional characteristics.	2. Except s- orbitals, all orbitals have directional characteristics.
3. It can accommodate maximum of $2n^2$ electrons.	3. An orbital can accommodate maximum of 2 electrons.
4. It is not in accordance with Heisenberg's uncertainty principle.	4. It is in accordance with the wave character of electrons and Heisenberg's Uncertainty principle.

6. What is the significance of Ψ and Ψ^2 ?

Ans: Ψ , the wave function of an atom gives the amplitude of the wave while Ψ^2 at any point gives the intensity of the electron i.e. probability of finding the electron at that point.

7. Define Heisenberg's Uncertainty principle.

Ans: It is not possible to determine simultaneously the exact position and exact momentum (or velocity) of an electron.

$$\Delta x \times \Delta p_x \geq \frac{h}{4\pi}$$

8. What are quantum numbers?

Ans: A set of four numbers which give us complete information i.e. shape, energy, orientation of orbital etc. about the electrons in an atom is called quantum numbers.

9. Give the significance of various quantum numbers.

Ans: 1. Principal Quantum Number (n) : It gives the average distance of the electron from the nucleus and the energy associated with it.

2. Azimuthal Quantum Number (l) : It shows the shape of the orbital.

3. Magnetic Quantum Number (m_l) : It informs about the orientation of the orbitals.

4. Spin Quantum Number (m_s) : It tells us about the spin of the electron about its own axis i.e. clockwise or anticlockwise motion.

10. What is the total number of orbitals associated with the principal quantum number $n = 3$?

Ans: For $n = 3$, possible values of l are 0, 1 and 2.

3s ($n = 3, l = 0$) = 1 orbital

3p ($n = 3, l = 1$) = 3 orbitals

3d ($n = 3, l = 2$) = 5 orbitals

Total no. of orbitals: $1 + 3 + 5 = 9$

11. Give the statement of

(i) Aufbau Principle

(ii) Pauli's Exclusion Principle

(iii) Hund's rule of maximum multiplicity

Ans: (i) Aufbau Principle: According to this principle, the ground state of the atoms are filled in the increasing order of their energies.

(ii) Pauli's Exclusion Principle: No two electrons in an orbital can have all the quantum numbers same.

OR

An orbital can accommodate maximum of two electrons but with opposite spin.

(iii) Hund's rule of maximum multiplicity: In degenerate orbitals, electron pairing will not take place unless all the degenerate orbitals are singly filled.

12. Write the configuration of Chromium (Cr = 24) and Copper (Cu = 29).

Ans: Cr = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$ OR [Ar] $3d^5 4s^1$

Cu = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ OR [Ar] $3d^{10} 4s^1$

13. What is the lowest value of n that allows 'g' orbitals to exist?

Ans: For g subshell, $l = 4$. Therefore $n = 5$

14. Nickel atom can lose two electrons and form Ni^{2+} ion. The atomic number is 28. From which orbital will nickel lose two electrons?

Ans: Ni : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

Nickel will lose its two electrons from its outermost 4s orbital.

15. How many subshells are associated with $n = 4$?

Ans: If $n = 4$, $l = 0, 1, 2$ and 3 .

Therefore, subshells will be 4s, 4p, 4d and 4f

16. How many electrons will be present in the subshell having m_s value of $\frac{-1}{2}$ for $n = 4$?

Ans: Subshells present 4s, 4p, 4d and 4f

Subshell	No. of electrons
4s	2
4p	6
4d	10
4f	14
TOTAL	32

No. of electrons with m_s value of $\frac{-1}{2} = 16$

17. The electronic configuration of valence shell of Cu is $3d^{10} 4s^1$ not $3d^9 4s^2$. Why?

Ans: Configurations with half-filled and completely filled orbitals have extra stability due to :

(i) symmetry

ii) Greater ways of exchanging the positions and releasing the energy termed as exchange energy.

18. Which orbital does not have directional characteristic?

Ans: s-orbital

19. What is $(n+l)$ rule or Bohr-Bury rule?

Ans: According to this rule, the lower the value of $(n+l)$ for an orbital, the lower is its energy.

If for two orbitals the value of $(n+l)$ is same, the one with lower value of n will have the lower energy.

20. What is the number of nodes in an orbital?

Ans: Total no. of nodes: $n - 1$

Number of angular nodes: l

Number of radial nodes: $n - l - 1$

21. Write the postulates of Planck's Quantum theory.

Ans: (i) The radiant energy is emitted or absorbed not continuously but discontinuously in the form of small discrete packets of energy, called quantum.

(ii) The energy of each quantum is directly proportional to the frequency of the radiation.

$$E \propto \nu \quad \text{i.e. } E = h\nu \text{ where } h \text{ is proportionality constant called Planck's constant.}$$

(iii) The total amount of energy emitted or absorbed by the body will be some whole number quanta.

$$E = nh\nu, \text{ where } n \text{ is any integer.}$$

22. Give the significance of Heisenberg's uncertainty principle.

Ans: Heisenberg's uncertainty principle is significant only in case of microscopic objects as the energy of photon is sufficient to disturb the position and momentum of the sub-atomic particles and is negligible in case of macroscopic object.

23. Give the series of the line spectra of hydrogen atom.

Ans:

Series	n_1	n_2
Lyman	1	2, 3,
Balmer	2	3, 4,
Paschen	3	4, 5,
Brackett	4	5, 6,
Pfund	5	6, 7,

24. What is Zeeman Effect?

Ans: The splitting of spectral lines in the magnetic field is called Zeeman effect.

25. What is Stark effect?

Ans: The splitting of the spectral lines in the electric field is called Stark effect.

Q1 What do you mean by significant figures?

Ans 1 Significant figures are meaningful digits which are known with certainty plus one which is estimated or uncertain.

Q2 How are 0.50 mol Na₂CO₃ and 0.50 M Na₂CO₃ different?

A 2 0.50 mol Na₂CO₃ means $0.50 \times \text{molecular mass (106)} = 53\text{g}$

0.50 M Na₂CO₃ means 53g Na₂CO₃ are present in 1 ltr of solution

Q3 What is the SI unit of mass? How is it defined?

A3 The SI unit of mass is kilogram (kg). 1 Kilogram is defined as the mass equal to the mass of the international prototype of kilogram

Q4 State Law of Law of Multiple Proportions

A 4 This law was proposed by Dalton in 1803. According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.

Q5 State Avogadro Law

A5 In 1811, Avogadro proposed that equal volumes of all gases at the same temperature and pressure should contain equal number of molecules

Q6 State Dalton Atomic Theory

A6 1. Matter consists of indivisible atoms.

2. All atoms of a given element have identical properties, including identical mass. Atoms of different elements differ in mass.

3. Compounds are formed when atoms of different elements combine in a fixed ratio.

4. Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

Q7 State difference between empirical and molecular formula

A7 An empirical formula represents the simplest whole number ratio of various atoms present in a compound, whereas, the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

Q8 What is Limiting Reagent

A 8 The reactant, which gets consumed first, limits the amount of product formed and is, therefore, called the limiting reagent.

Q9 Define Molarity of a solution

A9 It is the most widely used unit and is denoted by M. It is defined as the number of moles of the solute in 1 litre of the solution. Thus,

Molarity (M) =

No. of moles of solute of solution in litres / Volume of solution in litres

Q10 Name two drugs used in cancer treatment

A 10 cisplatin and taxol

MLL for Class XI

(Chemical Bonding & Molecular Structure)

1.U How do we express Bond Strength in terms of Bond Order and Bond Length?

Ans. Bond strength is directly proportional to Bond Order and inversely proportional to Bond Length i.e. More the bond order, more the bond strength and less the bond length, more is the bond strength.

2.U Although both CO₂ and H₂O are triatomic molecules, the shape of H₂O is bent while CO₂ is linear. Explain on the basis of Dipole Moment.

Ans. The bond moments of two C=O bonds in CO₂ cancel each other indicating the linear structure for CO₂ with dipole moment = 0. The resultant μ for H₂O molecule is $\neq 0$. This indicates a bent structure for H₂O molecule because bond moments of two O-H bonds don't cancel each other with resultant dipole moment = 1.84 D.

3.U Why CH₄ is not square planar but tetrahedral in geometry?

Ans. According to VSEPR theory, the electron pairs around the central atom lie as far apart as possible so as to minimise the repulsive interactions between them. For sq planar arrangement for CH₄, the four C-H bond pairs have to lie at an angle of 90°.

But in tetrahedral arrangement, The C-H bond pairs lie at an angle of 109°28' and therefore have smaller repulsive interactions than those in square planar arrangement.

4.U Which out of NH₃ and NF₃ has higher dipole moment and why ?

Ans. μ (NH₃) = 1.46 D which is greater than μ (NF₃) = 0.25 D, although N-F bonds are more polar, As, in NH₃, the bond moments of three N-H bonds reinforce the lone pair moment while in NF₃ the three N-F bonds oppose the lone pair moment.

5.U What is meant by Hybridisation of Atomic Orbitals? Describe the shapes of sp, sp² and sp³ hybrid orbitals?

Ans. It is the intermixing of different orbitals having slightly different energies so as to redistribute their energies leading to identical shape and equal energies are called Hybridisation.

Shapes:

sp – Linear

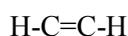
sp² – Trigonal Planar

sp³ – Tetrahedral

6.U What is the total number of sigma and pi bonds in the following molecules ?

(a) C₂H₂ (b) C₂H₄

Ans. H-C≡C-H Sigma Bonds = 3, Pi Bonds = 2



H H

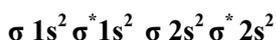
Sigma Bonds = 5, Pi Bonds = 1

7.U Why are axial bonds longer as compared to equatorial bonds in PCl_5 ?

Ans. In the formation of PCl_5 , Phosphorous atom assumes sp^3d hybrid state. The longer nature of axial bonds is due to relative stronger repulsive interactions experienced by axial bond pairs as compared to equatorial bond pairs. As out of 5 HOs, 3 HOs are oriented towards the 3 corners of an equilateral triangle making an angle of 120° between them. Such bonds are called equatorial bonds. The remaining 2 P-Cl bonds are perpendicular to the plane of equatorial bonds, one above and below the equatorial plane called as Axial Bonds.

8.U Use MOT to explain why Be_2 molecule does not exist.

Ans. The atomic number of Be is 4. This means that 8 electrons are to be filled in the molecular orbitals of Be_2 . The MO configuration is :



$$\begin{aligned} \text{Bond Order} &= \frac{1}{2}(N_b - N_a) \\ &= \frac{1}{2}(4 - 4) = 0 \end{aligned}$$

As, the B.O. is zero, the molecule of Be_2 does not exist.

9.U Considering X-Axis as the Internuclear Axis, which out of the following will form a σ bond?

(a) Ms and 1s (b) 1s and $2p_x$ (c) $2p_y$ and $2p_y$ (d) 1s and 2s

Ans. Cases (a), (b) and (c), as it is formed by axial overlap.

10.U Although geometries of NH_3 and H_2O are distorted tetrahedral, bond angle in water is less than in ammonia. Discuss.

Ans. The difference in bond angles is due to different number of lone pairs and bond pairs in two species. In NH_3 , the N atom has one lone pair and 3 bond pairs while in H_2O , the O atom has two lp and 2bp. The repulsive interactions of lone pairs and bond pairs in water are relatively more than those of NH_3 . Hence, Bond angle is smaller.

UNIT - 3 CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

LEVEL 1

1. What would be the IUPAC name and symbol for the element of atomic no.120?

Ans: Unbinilium (Ubn)

2. State Mendeleev's periodic law.

Ans: Physical and chemical properties of elements are the periodic function of their atomic masses.

3. How many groups are there in the p-block of periodic table?

Ans: 6

4. Explain the term electron gain enthalpy.

Ans: The enthalpy change taking place by attracting an electron towards its outermost shell of an isolated gaseous atom.

5. What are lanthanides and actinides?

Ans:

(a) Lanthanides: Elements in which last electron enter in 4f sub-shell of antepenultimate shell.

(b) Actinides: Elements in which last electron enter in 5f sub-shell of antepenultimate shell.

6. What is the general electronic configuration of d-block elements?

Ans: $n - 1 d_{1-10}ns_{0-2}$

7. Arrange the following in increasing order of ionization enthalpy:-

B, C, N, O.

Ans: $B < C < O < N$

8. Arrange the elements in increasing order of their size:- F^- , Li^+ , Na^+ , Cl^- .

Ans: $Li^+ < Na^+ < F^- < Cl^-$

9. Which has the highest value of electron gain enthalpy fluorine or chlorine?

Ans: Chlorine

10. Name the first element having an incomplete 3d-level.

Ans: Na

LEVEL-2

1. What is diagonal relationship?

Ans: Due to similar charge/size ratio element of second period are diagonally related with the third period element like Li is diagonally related to Mg.

2. How are Li and Mg related to each other in the periodic table? Write the name of another pair having such a relationship?

Ans: Diagonal relationship. Be and Al

3. Write down the main characteristic of p block elements?

Ans: In these elements the last electron entering p sub-shell of last shell.

4. Noble gases have zero electron affinity. Why?

Ans: due to completely filled duplet or Octet.

5. Account for the low values of electron affinities of nitrogen and phosphorus?

Ans: Half filled p subshell in N and P.

6. What are f block elements? By which alternative name are they known and why?

Ans: Elements in which last electron enter in f sub-shell of antepenultimate shell. They are called inner transition elements.

7. What are isoelectronic ions? Account for the decrease in size of following species oxide ion > Fluoride ion > sodium ion > magnesium ion

Ans: The ions having same number of electrons are called isoelectronic ions. The above order is due to increasing order of effective nuclear charge.

8. Why ionization of d block elements does remain constant?

Ans: Due to Screening/shielding effect.

LEVEL -3

1. What is modern periodic law? How does it in the removal of some of the anomalies of the Mendeleev's periodic table?

Ans: Modern Periodic law states that "the physical and chemical properties of the elements are periodic functions of their atomic numbers".

When atomic number is taken as fundamental property then abnormalities like position of isotope, anomalous pair itself removed.

2. The ionic radii of alkaline earth metals are smaller than those of alkali metals. Why?

Ans: Effective nuclear charge increases so the radius of alkaline earth metal decreases.

3. Classify the elements into four blocks on the basis of electronic configuration?

Ans: On the basis of electronic configuration

- a) s block elements
- b) p block elements
- c) d block elements
- d) f block elements

4. Describe the characteristics of s-block elements?

Ans: Characteristics of s-block elements:

- i) All the elements are soft metals.
- ii) They have low melting and boiling points.
- iii) They are highly reactive in nature.
- iv) Most of them impart characteristic colours to the flame.
- v) They form ionic compounds by losing their valence electrons and, thus, show oxidation states of +1 and +2 in their compounds.
- vi) The elements are good reducing agents.
- vii) All of them are good conductors of heat and electricity.

5. With reference to periodic table indicate,

- a. Element that is in group 13 and third period both.
- b. First transition element of the fourth period.

Ans: (a) Al (b) Sc

6. What do you mean by atomic radii? Discuss its variation in periodic table?

Ans: Atomic radius may be defined as the distance from the centre of the nucleus to the outermost shell containing the electrons.

The exact determination of the atomic radius of an element is not as simple as expected because:

- Along a period the atomic radii of the elements generally decreases from left to the right. It is due to increased nuclear charge from left to the right.
- The atomic radii of the elements in every group of the p[eriodic table increases as we move downwards.

7. What do you mean by electro negativity? Discuss its variation in periodic table?

Ans: The electronegativity of an element is defined as *the relative electron attractive tendency of its atom for a shared pair of electrons in a chemical bond.*

- It depends upon the nature of the atom with which a particular atom is linked. Different scales of comparison have been given for this purpose.
- It increases along the period and decreases down the group.

8. The electronic configuration of A is $1s^2 2s^2 2p^6 3s^2$

a. To which period does it belong?

b. Write down the formula of the nitrate of A.

c. state whether a solution of nitrate is conductor of electricity or not

Ans: (a) Third period

(b) $A NO_3$

(c) Conductor of electricity

STRUCTURE OF ATOM (NUMERICALS)

- Wavelength, frequency and wave velocity are related to each other by

$$c = \nu\lambda \quad \text{where } c = \text{velocity in light} = 3.0 \times 10^8 \text{ m/s}$$

$$\nu = \text{frequency of } s^{-1} \text{ or Hz}$$

$$\lambda = \text{wavelength in metres}$$

- Wave number ($\bar{\nu}$) is the reciprocal of wavelength $\left(\bar{\nu} = \frac{1}{\lambda}\right)$.

- Energy of a photon: $E = h\nu$ or hc/λ where, h = Planck's constant
 ν = frequency of radiation
 λ = wavelength

- Line Spectrum of Hydrogen atom:

The wave number of lines can be calculated by the following relation :

$$\bar{\nu} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

R = Rydberg's constant = 109677 cm^{-1}

For, Lyman series : $n_1 = 1, n_2 = 2, 3, 4, \dots$

Balmer series : $n_1 = 2, n_2 = 3, 4, 5, \dots$

Paschen series : $n_1 = 3$ and $n_2 = 4, 5, 6, \dots$

Brackett series : $n_1 = 4$ and $n_2 = 5, 6, 7, \dots$

Pfund series : $n_1 = 5$ and $n_2 = 6, 7, 8, \dots$

- The energy of electron in hydrogen atom is given by :

$$E_n = -\frac{2\pi^2 m Z^2 e^4}{n^2 h^2}$$

M = mass of electron, e = charge on electron, Z = atomic number of element

- For hydrogen atom, energy of electron in n^{th} orbit is :

$$E_n = \frac{-1.312 \times 10^6 Z^2}{n^2} \text{ Jmol}^{-1} = \frac{-2.178 \times 10^{-18} Z^2}{n^2} \text{ Jatom}^{-1} = \frac{-13.60 Z^2 eV}{n^2} \text{ atom}^{-1}$$

where Z = atomic number of H or hydrogen like ions.

- The energy absorbed or emitted during electronic transition is given by the difference of the energies of two levels, *i.e.*,

$$E_2 - E_1 = -2.18 \times 10^{-18} \left(\frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \text{ J/atom such that } n_2 > n_1.$$

- The radius of the n^{th} orbit is given by $r_n = \frac{(0.529 \text{ \AA}) n^2}{z}$.

- **Photo-electric Effect:** Kinetic energy of photoelectrons,
 $\frac{1}{2} mv^2 = h(\nu - \nu_0)$ where, ν_0 is threshold frequency
 ν is frequency of incident radiation
 h is Planck's constant
 v is velocity of electron
- **de-Broglie equation:** $\lambda = h/mv = h/p$ where, h is Planck's constant
 p is momentum
- **Heisenberg uncertainty principle:**
 $\Delta x \cdot \Delta p \geq h/4\pi$ where, Δx is uncertainty in position
 Δp is uncertainty in momentum
 h is Planck's constant

NUMERICALS:

1. Calculate the wave number and frequency of radiation having wavelength 5800 Å.

ANS: Formula: Wavenumber ($\tilde{\nu}$) = $1/\lambda = 1/5800 \times 10^{-10} = 1.724 \times 10^4 \text{ cm}^{-1}$

$$\begin{aligned} \text{Frequency } (\nu) &= c/\lambda = 3 \times 10^8 / 5800 \times 10^{-10} \\ &= 5.172 \times 10^{14} \text{ s}^{-1} \end{aligned}$$

2. What is the maximum number of emission lines when the excited electron of a H-atom in $n=6$ drops to the ground state?

ANS: Number of lines produced when electron from n^{th} subshell drops to ground state

$$= n(n-1)/2 = 6(6-1)/2 = 15$$

3. What is the number of photons of light with a wavelength of 4000 pm that provide 1 J of energy?

ANS: If no. of photon is N

$$\text{Energy of } N \text{ photons} = N h \nu \text{ or } N hc/\lambda$$

$$\text{So, in 1 J energy, } N = 1 \times \lambda / hc$$

$$\begin{aligned} &= 1 \times 4000 \times 10^{-12} / 6.626 \times 10^{-34} \times 3 \times 10^8 \\ &= 2.01 \times 10^{16} \end{aligned}$$

4. Calculate the wavelength of an electron moving with a velocity of $2.05 \times 10^7 \text{ ms}^{-1}$.

ANS: Formula: $\lambda = h/mv$

$$\begin{aligned} &= 6.626 \times 10^{-34} / 9.1 \times 10^{-31} \times 2.05 \times 10^7 \\ &= 3.55 \times 10^{-11} \text{ m} \end{aligned}$$

5. What are the frequency and wavelength of a photon emitted during a transition from $n = 5$ state to the $n = 2$ state in the hydrogen atom?

ANS: $n_i = 5$ and $n_f = 2$, so the spectral line is in the visible region of the Balmer series.

$$\begin{aligned}\Delta E &= 2.18 \times 10^{-18} (1/n_i^2 - 1/n_f^2) \\ &= -4.58 \times 10^{-19} \text{ J (emission energy)}\end{aligned}$$

$$\Delta E = h\nu$$

$$\text{So, } \nu = 4.58 \times 10^{-19} \text{ J} / 6.626 \times 10^{-34} = 6.91 \times 10^{14} \text{ Hz}$$

$$\lambda = c/\nu = 434 \text{ nm}$$

6. MA photon of wavelength $4 \times 10^{-7} \text{ m}$ strikes on metal surface, the work function of the metal being 2.13 eV. Calculate (i) the energy of the photon (eV), (ii) the kinetic energy of the emission, and (iii) the velocity of the photoelectron (1 eV = $1.6020 \times 10^{-19} \text{ J}$).

ANS: Energy of photon: $E = h\nu = hc/\lambda$

$$\begin{aligned}\text{Substituting the values in the given expression: } E &= 4.97 \times 10^{-19} \text{ J} \\ &= 4.97 \times 10^{-19} \text{ J} / 1.6020 \times 10^{-19} \\ &= 3.1020 \text{ eV}\end{aligned}$$

K.E.(of emission) = $h\nu - h\nu_0$ or (Energy of photon – Work function)

$$= 3.102 - 2.13 \text{ e} = 0.97 \text{ eV}$$

The velocity of photoelectron: $\frac{1}{2} m v^2 = h\nu - h\nu_0$

Substituting the values in the given expression

$$\begin{aligned}\nu &= \sqrt{0.3418 \times 10^{12} \text{ m}^2 \text{ s}^{-2}} \\ &= 5.84 \times 10^5 \text{ ms}^{-1}\end{aligned}$$

7. MWhat transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum?

ANS: For He^+ , $1/\lambda = R_H \times 2^2 (1/2^2 - 1/4^2)$

Or $1/\lambda = R_H (1/1^2 - 1/2^2)$ (taking $1/2^2$ common from the bracket)

Thus, for hydrogen atom $n = 2$ to $n = 1$ transition would give radiation of the same wavelength as $n = 4$ transition for He^+ .

8. MThe threshold frequency for a metal is s^{-1} . Calculate the kinetic energy of an electron emitted when radiation of frequency $\nu = 1.0 \times 10^{15} \text{ s}^{-1}$ hits the metal.

$$\begin{aligned}\text{ANS: K.E.(of emission)} &= \frac{1}{2} m v^2 = h\nu - h\nu_0 \\ &= 6.63 \times 10^{-34} (1.0 \times 10^{15} - 7.0 \times 10^{14}) \\ &= 1.988 \times 10^{-19} \text{ J}\end{aligned}$$

9. MIf the velocity of the electron in Bohr's first orbit is $2.19 \times 10^6 \text{ ms}^{-1}$. Calculate the de-Broglie wavelength associated with it.

ANS: Formula; $\lambda = h / mv$

$$= 6.626 \times 10^{-34} / 9.11 \times 10^{-31} \times 2.19 \times 10^6$$

$$= 3.32 \times 10^{-10} \text{ m or } 332 \text{ pm}$$

10. The velocity associated with a proton moving in a potential difference of 1000 V is $4.37 \times 10^5 \text{ ms}^{-1}$. If a hockey ball of mass 0.1 kg is moving with this speed, calculate the wavelength associated with this velocity.

ANS: Formula: $\lambda = h / mv$

$$\lambda = 6.63 \times 10^{-34} / 0.1 \times 4.37 \times 10^5$$

$$\lambda = 1.51 \times 10^{-38} \text{ m}$$

11. A microscope using suitable photons is employed to locate an electron in an atom within a distance of 0.1 Å. What is the uncertainty involved in the measurement of its velocity?

ANS: Uncertainty in position (Δx) = 0.1 Å

$$\Delta x \cdot \Delta p = h / 4\pi \text{ or } \Delta x \cdot m\Delta v = h / 4\pi$$

$$\Delta v = h / 4\pi m \cdot \Delta x$$

$$\Delta v = 6.626 \times 10^{-34} / 4 \times 3.14 \times 0.1 \times 10^{-10} \times 9.11 \times 10^{-31}$$

$$= 0.579 \times 10^7 \text{ ms}^{-1}$$

12. An electron has a speed of 40 ms^{-1} (accurate up to 99.99%). What is the uncertainty in locating its position? (Given $m_e = 9.1 \times 10^{-31} \text{ kg}$)

ANS: According to Uncertainty principle: $\Delta x = h / 4\pi m\Delta v$

$$= 6.63 \times 10^{-34} / 4 \times 3.14 \times 9.11 \times 10^{-31} \times 0.04$$

$$= 1.447 \times 10^{-2} \text{ m}$$

S**CHEMICAL BONDING****ONE MARK QUESTIONS**

Q.1. What is the total number of sigma and pi bonds in the following molecules?

(a) C_2H_2 (b) C_2H_4

Ans- there are three sigma and two pi-bonds in C_2H_2 .

there are five sigma bonds and one pi-bond in C_2H_4 .

Q.2. Write the significance of a plus and a minus sign shown in representing the orbitals.

Ans- Molecular orbitals are represented by wave functions. A plus sign in an orbital indicates a positive wave function while a minus sign in an orbital represents a negative wave function.

Q.3. How do you express the bond strength in terms of bond order?

Ans- Bond strength represents the extent of bonding between two atoms forming a molecule. The larger the bond energy, the stronger is the bond and the greater is the bond order.

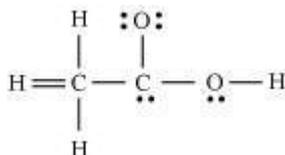
Q.5. Define the bond length.

Ans- Bond length is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule

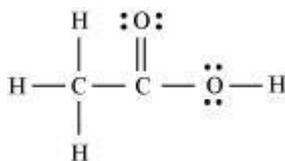
Q.6. Arrange the bonds in order of increasing ionic character in the molecules: LiF, K₂O, N₂, SO₂ and ClF₃.

Ans- N₂ < SO₂ < ClF₃ < K₂O < LiF.

Q.7. The skeletal structure of CH₃COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.



Ans- The correct Lewis structure for acetic acid is as follows:



Q.8. Define octet rule.

Ans- The elements tend to adjust the arrangement of their electrons in such a way that they (except H and He) achieve eight electrons in their outermost shell. This is called octet rule.

Q.9. Define lattice enthalpy.

Ans- The energy required when one mole of an ionic compound in crystalline form is split into the constituent ions is called lattice enthalpy.

Q.10. Which type of bond is formed when the atoms have zero difference in electronegativity?

Ans- Covalent bond.

TWO MARKS QUESTIONS

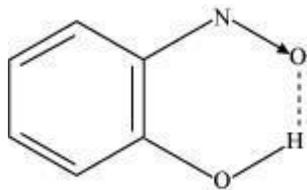
Q.1. Define hydrogen bond. Is it weaker or stronger than the van der Waals forces?

Ans- A hydrogen bond is defined as an attractive force acting between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of a different molecule (may be of the same kind).

There are two types of H-bonds:

(i) **I**ntermolecular H-bond e.g., HF, H₂O etc.

(ii) **I**ntramolecular H-bond e.g., o-nitrophenol



Hydrogen bonds are stronger than Van der Waals forces since hydrogen bonds are regarded as an extreme form of dipole-dipole interaction.

Q.2. Write the favourable factors for the formation of ionic bond.

Ans-(i) Low ionization enthalpy of metal atom.

(ii) **H**igh electron gain enthalpy ($\Delta_{eg} H$) of a non-metal atom.

(iii) **H**igh lattice energy of the compound formed.

Q.3. Although geometries of NH₃ and H₂O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.

Ans- The molecular geometry of NH₃ and H₂O can be shown as:



The central atom (N) in NH₃ has one lone pair and there are three bond pairs. In H₂O, there are two lone pairs and two bond pairs.

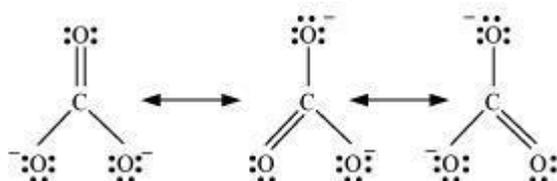
The two lone pairs present in the oxygen atom of H₂O molecule repels the two bond pairs. This repulsion is stronger than the repulsion between the lone pair and the three bond pairs on the nitrogen atom.

Since the repulsions on the bond pairs in H₂O molecule are greater than that in NH₃, the bond angle in water is less than that of ammonia.

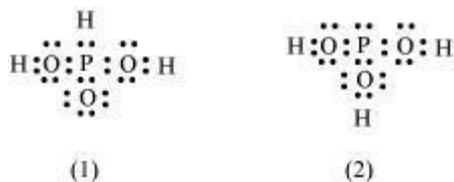
Q.4. Explain the important aspects of resonance with reference to the CO_3^{2-} ion.

Ans- According to experimental findings, all carbon to oxygen bonds in CO_3^{2-} are equivalent. Hence, it is inadequate to represent CO_3^{2-} ion by a single Lewis structure having two single bonds and one double bond.

Therefore, carbonate ion is described as a resonance hybrid of the following structures:



Q.5. H_3PO_3 can be represented by structures 1 and 2 shown below. Can these two structures be taken as the canonical forms of the resonance hybrid representing H_3PO_3 ? If not, give reasons for the same.



Ans- The given structures cannot be taken as the canonical forms of the resonance hybrid of H_3PO_3 because the positions of the atoms have changed.

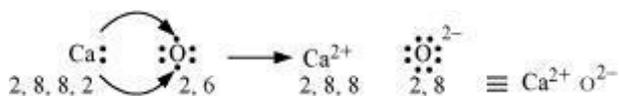
Q.6. Use Lewis symbols to show electron transfer between the following atoms to form cations and anions: (a) Ca and O (c) Al and N.

Ans(a) Ca and O:

The electronic configurations of Ca and O are as follows:

Ca: 2, 8, 8, 2 O: 2, 6

Oxygen requires two electrons more to complete its octet, whereas calcium has two electrons more than the nearest noble gas i.e., Argon. Hence, the electron transfer takes place as:

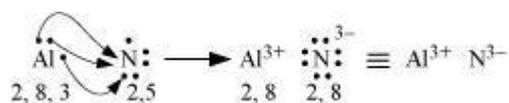


(b) Al and N:

The electronic configurations of Al and N are as follows:

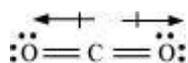
Al: 2, 8, 3 N: 2, 5

Nitrogen is three electrons short of the nearest noble gas (Neon), whereas aluminium has three electrons more than Neon. Hence, the electron transference can be shown as:



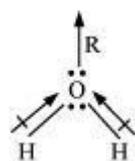
Q.7. Although both CO₂ and H₂O are triatomic molecules, the shape of H₂O molecule is bent while that of CO₂ is linear. Explain this on the basis of dipole moment.

Ans- According to experimental results, the dipole moment of carbon dioxide is zero. This is possible only if the molecule is linear so that the dipole moments of C–O bonds are equal and opposite to nullify each other.



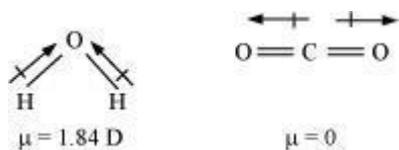
Resultant $\mu = 0$ D

H₂O, on the other hand, has a dipole moment value of 1.84 D (though it is a triatomic molecule as CO₂). The value of the dipole moment suggests that the structure of H₂O molecule is bent where the dipole moment of O–H bonds are unequal.



Q.8. Write the significance/applications of dipole moment.

Ans- Dipole moment is the measure of the polarity of a bond. It is used to differentiate between polar and non-polar bonds since all non-polar molecules (e.g. H₂, O₂) have zero dipole moments. It is also helpful in calculating the percentage ionic character of a molecule.



Q.9. Use molecular orbital theory to explain why the Be₂ molecule does not exist.

Ans- The electronic configuration of Beryllium is $1s^2 2s^2$.

The molecular orbital electronic configuration for Be₂ molecule can be written as:

$$\sigma_{1s}^2 \sigma_{1s}^{*2} \sigma_{2s}^2 \sigma_{2s}^{*2} \quad \text{Hence, the bond order for Be}_2 \text{ is } \frac{1}{2}(N_b - N_a).$$

Where,

N_b = Number of electrons in bonding orbitals

N_a = Number of electrons in anti-bonding orbitals

$$\therefore \text{Bond order of Be}_2 = \frac{1}{2}(4-4) = 0$$

A negative or zero bond order means that the molecule is unstable.

Hence, Be₂ molecule does not exist.

Q.10. Distinguish between a sigma and a pi bond.

Ans- The following are the differences between sigma and pi-bonds:

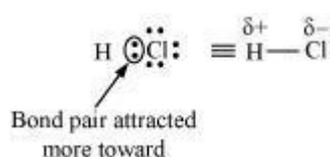
Sigma (σ) Bond	Pi (π) Bond
(a) It is formed by the end to end overlap of orbitals.	It is formed by the lateral overlap of orbitals.
(b) The orbitals involved in the overlapping are $s-s$, $s-p$, or $p-p$.	These bonds are formed by the overlap of $p-p$ orbitals only.
(c) It is a strong bond.	It is weak bond.
(d) The electron cloud is symmetrical about the line joining the two nuclei.	The electron cloud is not symmetrical.
(e) It consists of one electron cloud, which is symmetrical about the internuclear axis.	There are two electron clouds lying above and below the plane of the atomic nuclei.
(f) Free rotation about σ bonds is possible.	Rotation is restricted in case of pi-bonds.

Q.11. Explain with the help of suitable example polar covalent bond.

Ans- When two dissimilar atoms having different electronegativities combine to form a covalent bond, the bond pair of electrons is not shared equally. The bond pair shifts towards the nucleus of the atom having greater electronegativity. As a result, electron distribution gets distorted and the electron cloud is displaced towards the electronegative atom.

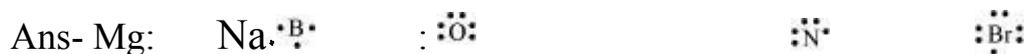
As a result, the electronegative atom becomes slightly negatively charged while the other atom becomes slightly positively charged. Thus, opposite poles are developed in the molecule and this type of a bond is called a polar covalent bond.

HCl, for example, contains a polar covalent bond. Chlorine atom is more electronegative than hydrogen atom. Hence, the bond pair lies towards chlorine and therefore, it acquires a partial negative charge.



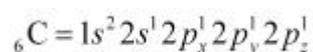
THREE MARKS QUESTIONS

Q.1. Write Lewis dot symbols for atoms of the following elements: Mg, Na, B, O, N, Br.

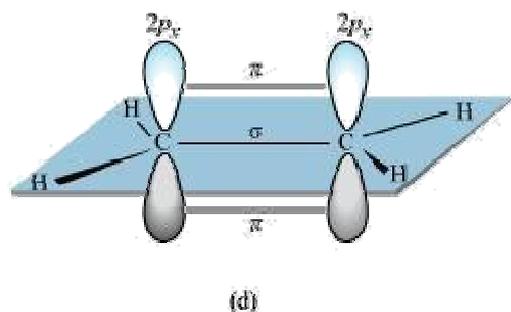
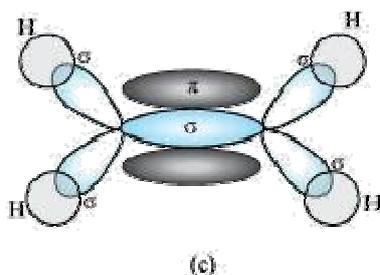
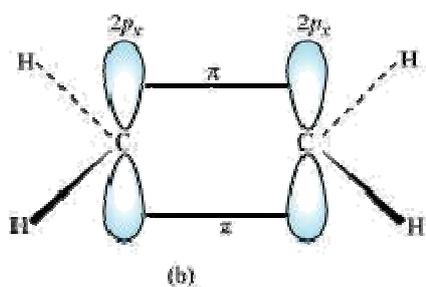
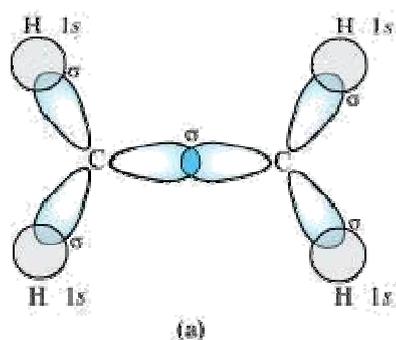


Q.3. Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in C_2H_4 and C_2H_2 molecules.

Ans- C_2H_4 :The electronic configuration of C-atom in the excited state is:



In the formation of an ethane molecule (C_2H_4), one sp^2 hybrid orbital of carbon overlaps a sp^2 hybridized orbital of another carbon atom, thereby forming a C-C sigma bond. The remaining two sp^2 orbitals of each carbon atom form a sp^2-s sigma bond with two hydrogen atoms. The unhybridized orbital of one carbon atom undergoes sidewise overlap with the orbital of a similar kind present on another carbon atom to form a weak π -bond.

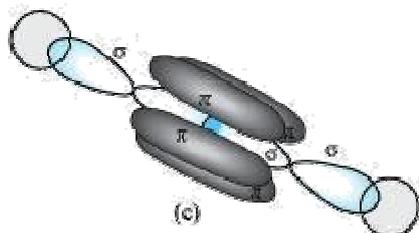
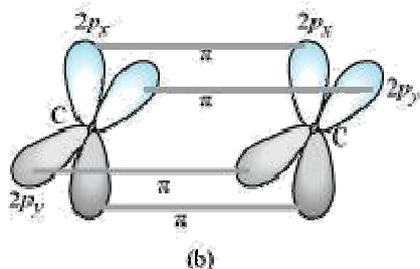
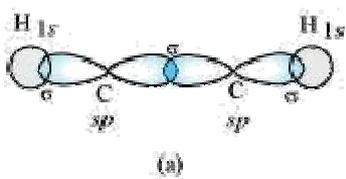


C_2H_2 :

In the formation of C_2H_2 molecule, each C-atom is sp hybridized with two $2p$ -orbitals in an unhybridized state.

One sp orbital of each carbon atom overlaps with the other along the internuclear axis forming a C-C sigma bond. The second sp orbital of each C-atom overlaps a half-filled $1s$ -orbital to form a σ bond.

The two unhybridized $2p$ -orbitals of the first carbon undergo sidewise overlap with the $2p$ orbital of another carbon atom, thereby forming two pi (π) bonds between carbon atoms. Hence, the triple bond between two carbon atoms is made up of one sigma and two π -bonds.



Q.4. Explain the formation of H_2 molecule on the basis of valence bond theory.

Ans- Let us assume that two hydrogen atoms (A and B) with nuclei (N_A and N_B) and electrons (e_A and e_B) are taken to undergo a reaction to form a hydrogen molecule. When A and B are at a large distance, there is no interaction between them. As they begin to approach each other, the attractive and repulsive forces start operating.

Attractive force arises between:

(a) Nucleus of one atom and its own electron i.e., $N_A - e_A$ and $N_B - e_B$.

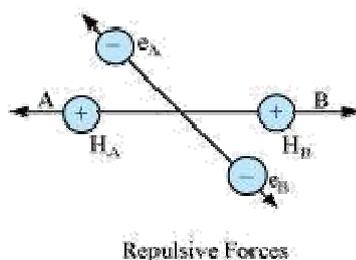
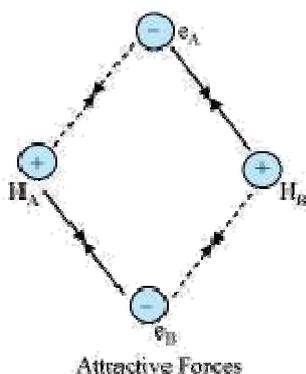
(b) Nucleus of one atom and electron of another atom i.e., $N_A - e_B$ and $N_B - e_A$.

Repulsive force arises between:

(a) Electrons of two atoms i.e., $e_A - e_B$.

(b) Nuclei of two atoms i.e., $N_A - N_B$.

The force of attraction brings the two atoms together, whereas the force of repulsion tends to push them apart.



The magnitude of the attractive forces is more than that of the repulsive forces. Hence, the two atoms approach each other. As a result, the potential energy decreases. Finally, a state is reached when the attractive forces balance the repulsive forces and the system acquires minimum energy. This leads to the formation of a dihydrogen molecule.

Q.5. Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.

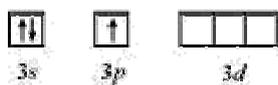
Ans- The given conditions should be satisfied by atomic orbitals to form molecular orbitals:

- (a) The combining atomic orbitals must have the same or nearly the same energy. This means that in a homonuclear molecule, the $1s$ -atomic orbital of an atom can combine with the $1s$ -atomic orbital of another atom, and not with the $2s$ -orbital.
- (b) The combining atomic orbitals must have proper orientations to ensure that the overlap is maximum.
- (c) The extent of overlapping should be large.

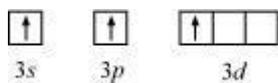
Q.6. Describe the hybridisation in case of PCl_5 . Why are the axial bonds longer as compared to equatorial bonds?

Ans- The ground state and excited state outer electronic configurations of phosphorus ($Z = 15$) are:

Ground state:

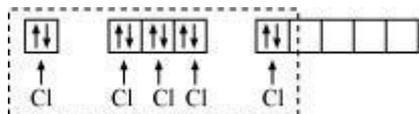


Excited state:

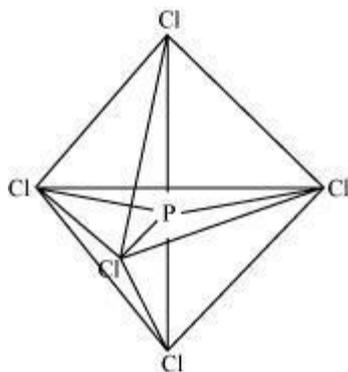


Phosphorus atom is sp^3d hybridized in the excited state. These orbitals are filled by the electron pairs donated by five Cl atoms as:

PCl_5



The five sp^3d hybrid orbitals are directed towards the five corners of the trigonal bipyramidals. Hence, the geometry of PCl_5 can be represented as:



There are five P–Cl sigma bonds in PCl_5 . Three P–Cl bonds lie in one plane and make an angle of 120° with each other. These bonds are called equatorial bonds. The remaining two P–Cl bonds lie above and below the equatorial plane and make an angle of 90° with the plane. These bonds are called axial bonds.

As the axial bond pairs suffer more repulsion from the equatorial bond pairs, axial bonds are slightly longer than equatorial bonds.

Q.7. What is meant by the term bond order? Calculate the bond order of:

N_2 , O_2 , O_2^+ and O_2^- .

Ans- Bond order is defined as one half of the difference between the number of electrons present in the bonding and anti-bonding orbitals of a molecule.

$$\text{Bond order} = \frac{1}{2}(N_b - N_a)$$

Bond order of N_2

$$[\sigma(1s)]^2[\sigma^*(1s)]^2[\sigma(2s)]^2[\sigma^*(2s)]^2[\pi(2p_x)]^2[\pi(2p_y)]^2[\sigma(2p_z)]^2$$

Number of bonding electrons, $N_b = 10$

Number of anti-bonding electrons, $N_a = 4$

$$\text{Bond order of nitrogen molecule} = \frac{1}{2}(10 - 4) = 3$$

Bond order of O_2

$$[\sigma(1s)]^2[\sigma^*(1s)]^2[\sigma(2s)]^2[\sigma^*(2s)]^2[\sigma(2p_z)]^2[\pi(2p_x)]^2[\pi(2p_y)]^2[\pi^*(2p_x)]^1[\pi^*(2p_y)]^1$$

$$\text{Bond order} = \frac{1}{2}(N_b - N_a) = \frac{1}{2}(8 - 4) = 2$$

Hence, the bond order of oxygen molecule is 2.

Similarly, the electronic configuration of O_2^+ can be written as:

$$KK[\sigma(2s)]^2[\sigma^*(2s)]^2[\sigma(2p_z)]^2[\pi(2p_x)]^2[\pi(2p_y)]^2[\pi^*(2p_x)]^1$$

$$\text{Bond order of } O_2^+ = \frac{1}{2}(8 - 3) = 2.5$$

The electronic configuration of O_2^- ion will be:

$$KK[\sigma(2s)]^2[\sigma^*(2s)]^2[\sigma(2p_z)]^2[\pi(2p_x)]^2[\pi(2p_y)]^2[\pi^*(2p_x)]^2[\pi^*(2p_y)]^1$$

$$\text{Bond order of } O_2^- = \frac{1}{2}(8 - 5) = 1.5$$

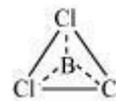
Q.8. Discuss the shape of the following molecules using the VSEPR model:

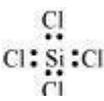
$BeCl_2$, BCl_3 , $SiCl_4$, AsF_5 , H_2S , PH_3

Ans- $BeCl_2$: $Cl:Be:Cl$ The central atom has no lone pair and there are two bond pairs. i.e., $BeCl_2$ is of the type AB_2 . Hence, it has a linear shape.

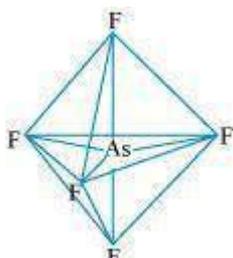
BCl₃:  The central atom has no lone pair and there are three bond pairs.

Hence, it is of the type AB₃. Hence, it is trigonal planar.



SiCl₄:  The central atom has no lone pair and there are four bond pairs.

Hence, the shape of SiCl₄ is tetrahedral being the AB₄ type molecule.



AsF₅: The central atom has no lone pair and there are five bond pairs. Hence, AsF₅ is of the type AB₅. Therefore, the shape is trigonal bipyramidal.

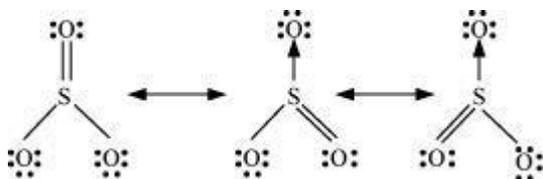
H₂S:  The central atom has one lone pair and there are two bond pairs. Hence, H₂S is of the type AB₂E. The shape is Bent.

PH₃:  The central atom has one lone pair and there are three bond pairs. Hence, PH₃ is of the AB₃E type. Therefore, the shape is trigonal bipyramidal.

Q.9. Write the resonance structures for SO₃, NO₂ and NO₃⁻.

Ans- The resonance structures are:

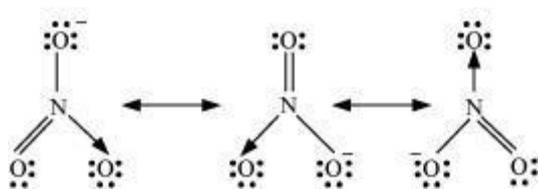
(a) SO₃:



(b) NO₂:



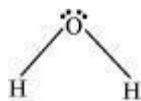
(c) NO_3^- :



Q.10. What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving example.

Ans- The shared pairs of electrons present between the bonded atoms are called **bond pairs**. All valence electrons may not participate in bonding. The electron pairs that do not participate in bonding are called **lone pairs** of electrons.

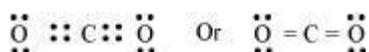
In H_2O , there are two bond pairs and two lone pairs on the central atom (oxygen).



FIVE MARKS QUESTIONS

Q.1. Define octet rule. Write its significance and limitations.

Ans-- The octet rule or the electronic theory of chemical bonding was developed by Kossel and Lewis. According to this rule, atoms can combine either by transfer of valence electrons from one atom to another or by sharing their valence electrons in order to attain the nearest noble gas configuration by having an octet in their valence shell.



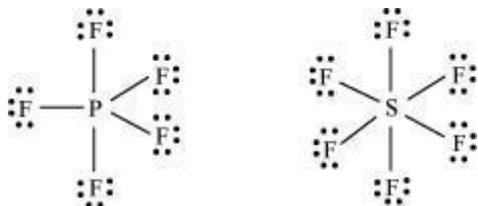
The octet rule successfully explained the formation of chemical bonds depending upon the nature of the element.

Limitations of the octet theory:

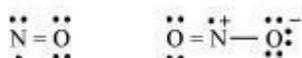
The following are the limitations of the octet rule:

- The rule failed to predict the shape and relative stability of molecules.
- It is based upon the inert nature of noble gases. However, some noble gases like xenon and krypton form compounds such as XeF_2 , KrF_2 etc.

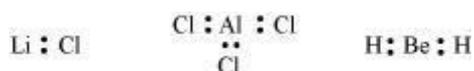
(c) The octet rule cannot be applied to the elements in and beyond the third period of the periodic table. The elements present in these periods have more than eight valence electrons around the central atom. For example: PF_5 , SF_6 , etc.



(d) The octet rule is not satisfied for all atoms in a molecule having an odd number of electrons. For example, NO and NO_2 do not satisfy the octet rule.

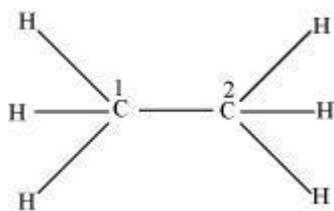


(e) This rule cannot be applied to those compounds in which the number of electrons surrounding the central atom is less than eight. For example, LiCl , BeH_2 , AlCl_3 etc. do not obey the octet rule.



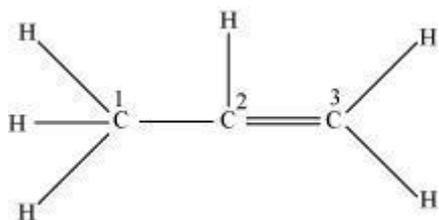
Q.2. Which hybrid orbitals are used by carbon atoms in the following molecules?
 $\text{CH}_3\text{-CH}_3$; (b) $\text{CH}_3\text{-CH=CH}_2$; (c) $\text{CH}_3\text{-CH}_2\text{-OH}$; (d) $\text{CH}_3\text{-CHO}$ (e) CH_3COOH

Ans- (a)



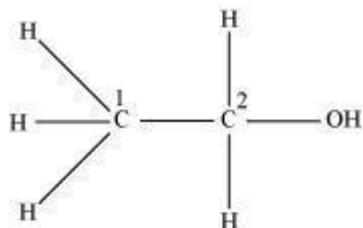
Both C_1 and C_2 are sp^3 hybridized.

(b)



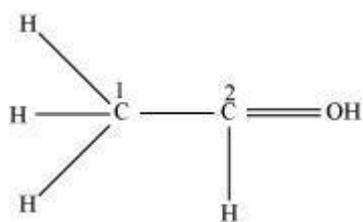
C_1 is sp^3 hybridized, while C_2 and C_3 are sp^2 hybridized.

(c)



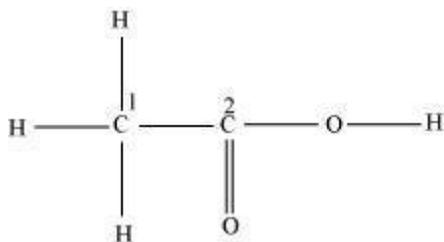
Both C_1 and C_2 are sp^3 hybridized.

(d)



C_1 is sp^3 hybridized and C_2 is sp^2 hybridized.

(e)

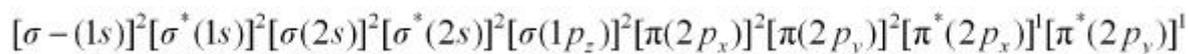


C_1 is sp^3 hybridized and C_2 is sp^2 hybridized.

Q.3. Compare the relative stability of the following species and indicate their magnetic properties;

O_2 , O_2^+ , O_2^- (superoxide), O_2^{2-} (peroxide)

Ans- There are 16 electrons in a molecule of dioxygen, 8 from each oxygen atom. The electronic configuration of oxygen molecule can be written as:



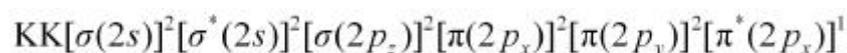
Since the 1s orbital of each oxygen atom is not involved in bonding, the number of bonding electrons = 8 = N_b and the number of anti-bonding orbitals = 4 = N_a .

$$\text{Bond order} = \frac{1}{2}(N_b - N_a)$$

$$= \frac{1}{2}(8 - 4)$$

$$= 2$$

Similarly, the electronic configuration of O_2^+ can be written as:



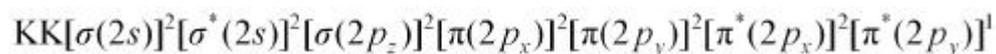
$$N_b = 8$$

$$N_a = 3$$

$$\text{Bond order of } O_2^+ = \frac{1}{2}(8 - 3)$$

$$= 2.5$$

Electronic configuration of O_2^- ion will be:



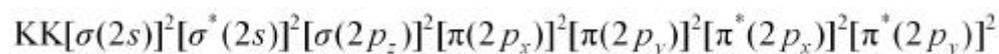
$$N_b = 8$$

$$N_a = 5$$

$$\text{Bond order of } O_2^- = \frac{1}{2}(8 - 5)$$

$$= 1.5$$

Electronic configuration of O_2^{2-} ion will be:



$$N_b = 8$$

$$N_a = 6$$

$$\text{Bond order of } O_2^{2-} = \frac{1}{2}(8-6)$$

$$= 1$$

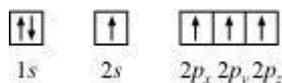
Bond dissociation energy is directly proportional to bond order. Thus, the higher the bond order, the greater will be the stability. On this basis, the order of stability is $O_2^+ > O_2 > O_2^- > O_2^{2-}$.

HOTS

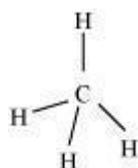
Q.1. Apart from tetrahedral geometry, another possible geometry for CH_4 is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH_4 is not square planar?

Ans- Electronic configuration of carbon atom: ${}_6C: 1s^2 2s^2 2p^2$

In the excited state, the orbital picture of carbon can be represented as:



Hence, carbon atom undergoes sp^3 hybridization in CH_4 molecule and takes a tetrahedral shape.



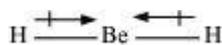
For a square planar shape, the hybridization of the central atom has to be dsp^2 . However, an atom of carbon does not have d -orbitals to undergo dsp^2 hybridization. Hence, the structure of CH_4 cannot be square planar.

Moreover, with a bond angle of 90° in square planar, the stability of CH_4 will be very less because of the repulsion existing between the bond pairs. Hence, VSEPR theory also supports a tetrahedral structure for CH_4 .

Q.2. Explain why BeH_2 molecule has a zero dipole moment although the Be-H bonds are polar.

Ans- The Lewis structure for BeH_2 is as follows: $\text{H}:\text{Be}:\text{H}$

There is no lone pair at the central atom (Be) and there are two bond pairs. Hence, BeH_2 is of the type AB_2 . It has a linear structure.

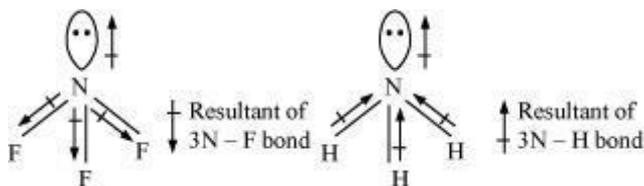


Dipole moments of each H–Be bond are equal and are in opposite directions. Therefore, they nullify each other. Hence, BeH_2 molecule has zero dipole moment.

Q.3. Which out of NH_3 and NF_3 has higher dipole moment and why?

Ans- In both molecules i.e., NH_3 and NF_3 , the central atom (N) has a lone pair electron and there are three bond pairs. Hence, both molecules have a pyramidal shape. Since fluorine is more electronegative than hydrogen, it is expected that the net dipole moment of NF_3 is greater than NH_3 . However, the net dipole moment of NH_3 (1.46 D) is greater than that of NF_3 (0.24 D).

This can be explained on the basis of the directions of the dipole moments of each individual bond in NF_3 and NH_3 . These directions can be shown as:



Thus, the resultant moment of the N–H bonds add up to the bond moment of the lone pair (the two being in the same direction), whereas that of the three N – F bonds partly cancels the moment of the lone pair.

Hence, the net dipole moment of NF_3 is less than that of NH_3 .

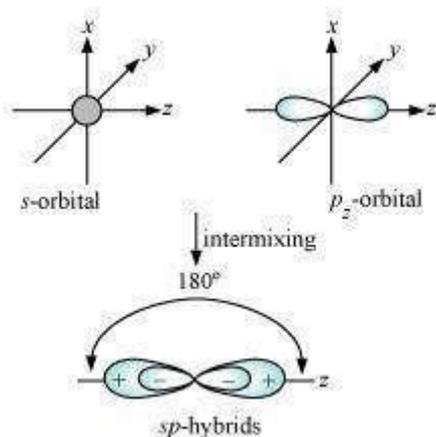
Q.4. What is meant by hybridisation of atomic orbitals? Describe the shapes of sp , sp^2 , sp^3 hybrid orbitals.

Ans- Hybridization is defined as an intermixing of a set of atomic orbitals of slightly different energies, thereby forming a new set of orbitals having equivalent energies and shapes.

For example, one $2s$ -orbital hybridizes with two $2p$ -orbitals of carbon to form three new sp^2 hybrid orbitals.

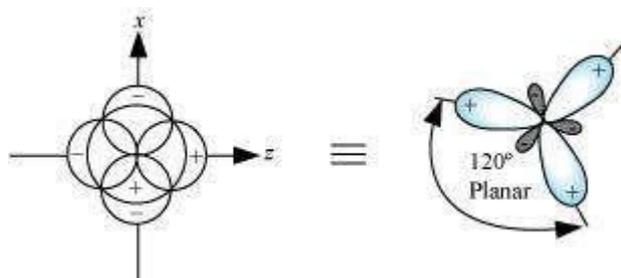
These hybrid orbitals have minimum repulsion between their electron pairs and thus, are more stable. Hybridization helps indicate the geometry of the molecule.

Shape of sp hybrid orbitals: sp hybrid orbitals have a linear shape. They are formed by the intermixing of s and p orbitals as:



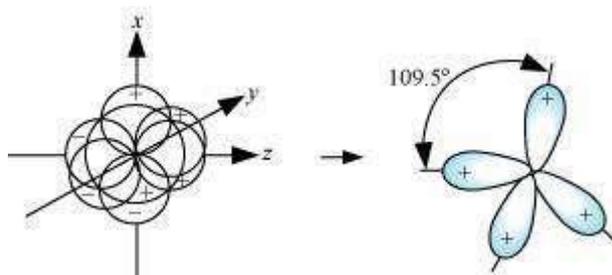
Shape of sp^2 hybrid orbitals:

sp^2 hybrid orbitals are formed as a result of the intermixing of one s -orbital and two $2p$ -orbitals. The hybrid orbitals are oriented in a trigonal planar arrangement as:



Shape of sp^3 hybrid orbitals:

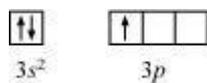
Four sp^3 hybrid orbitals are formed by intermixing one s -orbital with three p -orbitals. The four sp^3 hybrid orbitals are arranged in the form of a tetrahedron as:



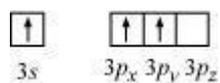
Q.5. Describe the change in hybridisation (if any) of the Al atom in the following reaction.



Ans- The valence orbital picture of aluminium in the ground state can be represented as:



The orbital picture of aluminium in the excited state can be represented as:



Hence, it undergoes sp^2 hybridization to give a trigonal planar arrangement (in AlCl_3).

To form AlCl_4^- , the empty $3p_z$ orbital also gets involved and the hybridization changes from sp^2 to sp^3 . As a result, the shape gets changed to tetrahedral.

THERMODYNAMICS:

EXPECTED QUESTIONS WITH ANSWERS(NUMERICALS)

Q1. Define molar heat capacity. Calculate number of KJ required to raise the temperature of 60g Al from 35° C to 55° Cg

Ans- quantity of heat required to raise the temperature of 1 mol of substance by 1°C or 1 K (C = 24 J/mol K) $n = 60 / 27 = 2.22 \text{ mol}$

$$q = Cn\Delta T$$

$$q = 24 \times 2.22 \times 20 = 1065.6 \text{ J}$$

$$= 1.0656 \text{ KJg}$$

Q2. Enthalpy of combustion of carbon is -393.5 KJ/mol. Find the heat released in the formation of 35.2g of CO₂g

Ans- $C + O_2 \rightarrow CO_2$ (44gm)

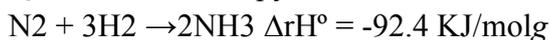
$$\Delta H = -393.5 \text{ KJ/mol}$$

heat released in the formation of 44g CO₂ = -393.5

$$\therefore \text{heat released in the formation of 35.2 g CO}_2 = 393.5 \times 35.2 / 44$$

$$= 314.8 \text{ KJg}$$

Q3. Find the enthalpy of formation of NH₃ (g)



Ans- $\Delta_f H^\circ$ of NH₃ = $\Delta_r H^\circ / 2 = -92.4 / 2 = -46.2 \text{ KJ/molg}$

Q4. The equilibrium constant for the reaction is 10. Calculate the value of ΔG° at 300 K.g

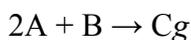
$$\text{Ans- } \Delta G^\circ = -2.303 RT \log K_g$$

$$= -2.303 \times 8.31 \times 300 \times \log 10_g$$

$$= - 55.27 \text{ Jg}$$

$$= - 5.527 \text{ KJ/molg}$$

Q5. At what temperature the reaction will be spontaneous.g



$$\Delta H = 400 \text{ KJ/molg}$$

$$\Delta S = 2.0 \text{ KJ/mol Kg}$$

Ans- $\Delta G = \Delta H - T\Delta S$, $\Delta G = 0$ at equilibriumg

$$T = \Delta H / \Delta S = 400 / 200 = 200 \text{ Kg}$$

HOTS QUESTIONSg

Q1. Specific heat of gas is 0.313 J at constant volume, if molar mass is 40g/mol find atomicity.g

Q2. $A+B \rightarrow C+D$, $\Delta H = -10,000 \text{ J/mol}$, $\Delta S = -33.3 \text{ J/mol K}$

(i) At what temperature reaction is spontaneous?

(ii) At what temperature reaction will reverse?

Q3. In what way internal energy is different from enthalpy?

Q4. $C_2H_4 + C_{12} \rightarrow C_2H_4C_{12}$

$\Delta H = -270.6 \text{ KJ/mol}$, $\Delta S = -139 \text{ J/K/mol}$ is the reaction favored by entropy, enthalpy, both or none?

Find ΔG , if $T = 300 \text{ K}$

Q5. For an isolated system $\Delta U = 0$ what will be ΔS ?

g

SOLUTION TO HOTS QUESTIONS

Answer 1. Molar heat capacity (C_p) = $0.313 \times 40 = 12.52 \text{ J/molg}$

$C_p = R + C_v = 8.31 + 12.52 = 20.83 \text{ J/molg}$

$C_p/C_v = 20.83/12.52 = 1.66$

Hence gas is monoatomic.

Answer 2. $\Delta G = \Delta H - T\Delta S$, $\Delta G = 0$, at equilibrium

$T = \Delta H / \Delta S = -10000 / -33.3 = 300.3 \text{ K}$

Answer 3. Internal energy is sum of all forms of energy stored in a substance. The ΔU refers to heat change in a process if it does not change Δn , ΔT , ΔV . Enthalpy is sum of internal energy and PV energy. Enthalpy refers to heat change in a process if it is carried out at constant temperature and pressure.

Answer 4. As $\Delta H = -ve$ reaction is favored by enthalpy, as ΔS is $-ve$ it is not favored by entropy.

$\Delta H = \Delta H - T\Delta S$

$= -270.6 - 300 \times 139 \times 10^{-3}$

$= -228.9 \text{ KJ/molg}$

Answer 5. For an isolated system with $\Delta U = 0$, the spontaneous change will occur if $\Delta S > 0$.

For eg, in isolated system involving intermixing of gases $\Delta U = 0$ but $\Delta S > 0$ because of increase of disorder.

MLL THERMODYNAMICS

1. Separate out the following into extensive and intensive:

Volume, Temperature, Pressure, Boiling point, Free energy

Ans: Volume and free energy are extensive, other are intensive.

2. Under what condition ΔH becomes equal to ΔU ?

Ans: $\Delta H = \Delta U$ during a process which is carried out in a closed vessel ($\Delta V = 0$) or number of moles of gaseous products = number of moles of gaseous reactants or the reaction does not involve any gaseous reactant or products.

3. For the process to occur under adiabatic conditions, the correct conditions is:

(i) $\Delta T = 0$ (ii) $\Delta p = 0$ (iii) $q = 0$ (iv) $w = 0$

Ans: (iii)

4. A reaction $A + B \rightarrow C + D + q$ is found to have a positive entropy change. The reaction will be

(i) possible at high temperature (ii) possible only at low temperature

(iii) not possible at any temperature (iv) possible at any temperature

Ans: Here $\Delta H = -ve$ and $\Delta S = +ve$. $\Delta G = \Delta H - T\Delta S$. For the reaction to be spontaneous, ΔG should be $-ve$ which will be so at any temperature i.e., option is (iv)

5. In a process, 701 J of heat is absorbed by a system and 394 J of work is done by the system.

What is the change in the internal energy of the process?

Ans: Internal energy of the system increases by 307 J

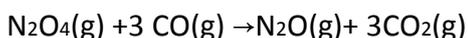
6. Why endothermic reactions are favoured at high temperature?

Ans: If temperature is high, then $T\Delta S$ will be much greater than ΔH in magnitude so that ΔG is highly negative.

7. State the first law thermodynamics and derive a mathematical expression for it.

Ans: Definition and expression for the same i.e., $\Delta U = q + w$ where ΔU is change in internal energy, q is the heat given to the system and w is the work done on the system.

8. Enthalpies of formation of $CO(g)$, $CO_2(g)$, $N_2O(g)$ and $N_2O_4(g)$ are -110, -393, -81 and 9.7 kJ/mol respectively. Find the value of $\Delta_r H$ for the reaction:



Ans: - 777.7 kJ

9. The equilibrium constant for a reaction is 10. What will be the value of ΔG_0 ? $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$, $T = 300 \text{ K}$

Ans: $\Delta G_0 = -2.303RT \log K = -5744.1 \text{ J}$

10. Enthalpy of combustion of carbon to CO_2 is $-393.5 \text{ kJ mol}^{-1}$. Calculate the heat released upon the formation of 35.2 g of CO_2 from carbon and dioxygen gas.

Ans: Heat released when 44g CO_2 is formed = 393.5 kJ

Hence heat released when 35.2 g CO_2 is formed = $393.5 \times 35.2 / 44 = 314.8 \text{ kJ}$

11. Give reasons:

(a) Neither q nor w is a state function but $q+w$ is state function.

(b) The dissolution of ammonium chloride in water is endothermic still it dissolves in water.

Ans:

(a) $q+w = \Delta U$. As ΔU is a state function, hence $q+w$ is a state function.

(b) On dissolution, entropy increases i.e., ΔS is +ve. If $T\Delta S > \Delta H$, then ΔG will be -ve. Hence, the process is spontaneous.

12. Calculate the Standard enthalpy of formation of $\text{CH}_3\text{OH}(\text{l})$ from the following data:

(i) $\text{CH}_3\text{OH}(\text{l}) + 3/2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$; $\Delta_r H_0 = -726 \text{ KJmol}^{-1}$

(ii) $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ $\Delta_c H_0 = -393 \text{ KJmol}^{-1}$

(iii) $\text{H}_2(\text{g}) + 1/2 \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$; $\Delta_f H_0 = -286 \text{ KJmol}^{-1}$

Ans: Aim: $\text{C}(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$ $\Delta_f H_0 = ?$

Eqn. (ii) + $\times 2$ eqn. (iii) - eqn. (i) gives the required Eqn. with $\Delta H = -393 + 2(-286) - (-726) = -239 \text{ KJmol}^{-1}$

MLL FOR CHAPTER 7 EQUILIBRIUM (NUMERICALS)

CLASS XI

1. The following concentrations were obtained for the formation of NH_3 from N_2 and H_2 at equilibrium at 500K. $[\text{N}_2] = 1.5 \times 10^{-2}\text{M}$. $[\text{H}_2] = 3.0 \times 10^{-2}\text{M}$ and $[\text{NH}_3] = 1.2 \times 10^{-2}\text{M}$. Calculate equilibrium constant.

Sol. The equilibrium constant for the reaction

$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ can be written as

$$\begin{aligned}K_c &= \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} \\&= \frac{(1.2 \times 10^{-2})^2}{(1.5 \times 10^{-2})(3.0 \times 10^{-2})^3} \\&= 0.106 \times 10^4 = 1.06 \times 10^3\end{aligned}$$

2. For the equilibrium, $2\text{NOCl}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g})$ the value of the equilibrium constant, K_c is 3.75×10^{-6} at 1069 K. Calculate the K_p for the reaction at this temperature?

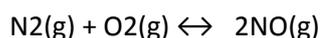
Sol- We know that, $K_p = K_c(RT)^{\Delta n}$ For the above reaction,

$$\Delta n = (2+1) - 2 = 1$$

$$K_p = 3.75 \times 10^{-6} (0.0831 \times 1069)$$

$$K_p = 0.033$$

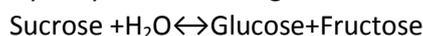
3. At equilibrium, the concentrations of $\text{N}_2 = 3.0 \times 10^{-3}\text{M}$, $\text{O}_2 = 4.2 \times 10^{-3}\text{M}$ and $\text{NO} = 2.8 \times 10^{-3}\text{M}$ in a sealed vessel at 800K. What will be K_c for the reaction



Sol.

$$\begin{aligned}K_c &= \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} \\&= \frac{(2.8 \times 10^{-3})^2}{(3 \times 10^{-3})(4.2 \times 10^{-3})} \\&= 0.622\end{aligned}$$

4. Hydrolysis of sucrose gives



Equilibrium constant K_c for the reaction is 2×10^{13} at 300K. Calculate ΔG^0 at 300K.

$$\begin{aligned}\text{Sol. } \Delta G^0 &= -2.303RT \log K_c \\&= -2.303 \times 8.314 \times 300 \times \log 2 \times 10^{13} \\&= -7.64 \times 10^4 \text{ J/mol}\end{aligned}$$

5. The concentration of hydrogen ion in a sample of soft drink is $3.8 \times 10^{-3}\text{M}$. what is its pH ?

Sol. is its pH ?

$$\text{Solution pH} = -\log[3.8 \times 10^{-3}]$$

$$= -\{\log[3.8] + \log[10^{-3}]\}$$

$= -\{(0.58) + (-3.0)\} = -\{-2.42\} = 2.42$ Therefore, the pH of the soft drink is 2.42 and it can be inferred that it is acidic.

6. M The pKa of acetic acid and pKb of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution.

$$\begin{aligned} \text{Solution pH} &= 7 + \frac{1}{2} [\text{pKa} - \text{pKb}] \\ &= 7 + \frac{1}{2} [4.76 - 4.75] \\ &= 7 + \frac{1}{2} [0.01] = 7 + 0.005 \\ &= 7.005 \end{aligned}$$

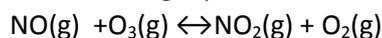
7. M The pKa of acetic acid and pKb of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution.

$$\begin{aligned} \text{Solution} \\ \text{pH} &= 7 + \frac{1}{2} [\text{pKa} - \text{pKb}] \\ &= 7 + \frac{1}{2} [4.76 - 4.75] \\ &= 7 + \frac{1}{2} [0.01] \\ &= 7 + 0.005 = 7.005 \end{aligned}$$

8. M Calculate the solubility of A_2X_3 in pure water, assuming that neither kind of ion reacts with water. The solubility product of A_2X_3 , $K_{sp} = 1.1 \times 10^{-23}$.

$$\begin{aligned} \text{Solution } A_2X_3 &\rightarrow 2A^{3+} + 3X^{2-} = 1.1 \times 10^{-23} \\ S &= \text{solubility of } A_2X_3, \text{ then } [A^{3+}] = 2S; [X^{2-}] = 3S \\ K_{sp} &= (2S)^2(3S)^3 = 108S^5 = 1.1 \times 10^{-23} \text{ thus,} \\ S^5 &= 1 \times 10^{-25} \quad S = 1.0 \times 10^{-5} \text{ mol/L.} \end{aligned}$$

9. M For the following equilibrium, $K_c = 6.3 \times 10^{14}$ at 1000K



Both the forward and reverse reaction in the equilibrium are elementary bimolecular reactions. What is K_c for the reverse reaction?

$$\begin{aligned} \text{Sol } K'_c &= 1/K_c \\ &= 1/6.3 \times 10^{14} = 1.59 \times 10^{-15} \end{aligned}$$

10. Assuming complete dissociation calculate the pH of 0.003M HCl



$$[H^+] = [HCl] = 3 \times 10^{-3} M$$

$$pH = -\log(3 \times 10^{-3}) = 2.52$$

STATES OF MATTER(NUMERICALS)

1. A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 2.27 L volume, up to what volume can the balloon be expanded?

Ans: $P_1V_1=P_2V_2$

$$1 \times 2.27 = 0.2 \times V_2$$

$$V_2 = 2.27 / 0.2$$

$$= 11.35 \text{ L}$$

2. What will be the minimum pressure required to compress 500 dm³ of air at 1 bar to 200 dm³ at 30°C?

Ans: $P_1V_1=P_2V_2$

$$1 \times 500 = P_2 \times 200$$

$$P_2 = 500 / 200$$

$$= 2.5 \text{ bar}$$

3. A vessel of 120 mL capacity contains a certain amount of gas at 35 °C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35 °C. What would be its pressure?

Ans: $P_1V_1=P_2V_2$

$$1.2 \times 120 = P_2 \times 180$$

$$P_2 = 0.8 \text{ bar}$$

4. On a ship sailing in Pacific Ocean where temperature is 23.4 °C, a balloon is filled with 2 L air. What will be the volume of the balloon when the ship reaches Indian Ocean, where temperature is 26.1°C?

Ans: $V_1 / T_1 = V_2 / T_2$

$$2 / 296.4 = V_2 / 299.1$$

$$V_2 = 2.018 \text{ L}$$

5. A student forgot to add the reaction mixture to the round bottomed flask at 27 °C but instead he/she placed the flask on the flame. After a lapse of time he realized his mistake, and using a pyrometer he found the temperature of the flask was 477 °C. What fraction of air would have been expelled out?

Ans: $V_1/T_1 = V_2/T_2$

$$V_1/300 = (V_1+x)/750$$

$$750/300 = (V_1+x)/x$$

$$2.5 = (V_1/x) + 1$$

$$1.5 = V_1/x$$

$$x/V_1 = 0.67$$

$$x = 0.67V_1$$

The amount of air expelled out is 0.67 times the volume of the flask $0.67/1.67 = 0.4011$ fraction of air is expelled.

6. At 25°C and 760 mm of Hg pressure a gas occupies 600 mL volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640 mL.

Ans $P_1V_1/T_1 = P_2V_2/T_2$

$$760 \times 600 / 298 = P_2 \times 640 / 283$$

$$P_2 = 760 \times 600 \times 283 / 298 \times 640$$

$$= 676.6 \text{ mm of Hg}$$

7. At 0°C, the density of a certain oxide of a gas at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?

Ans: $PV = nRT$

$$PV = (m/M) RT$$

$$d = M_1P_1/RT = M_2P_2/RT$$

$$M_1P_1 = M_2P_2$$

$$M_1 = 28 \times 5 / 2$$

$$= 70$$

8. What will be the pressure exerted by a mixture of 3.2 g of methane and 4.4 g of carbon dioxide contained in a 9 dm³ flask at 27 °C ?

Ans: $PV = nRT$

$$n = (3.2/16 + 4.4/44) = 0.3$$

$$P = 0.3 \text{ mole} \times 0.083 \text{ bar dm}^3 / \text{K/mole} \times 300 \text{ K} / 9 \text{ dm}^3$$

$$P = 0.83 \text{ bar.}$$

9. What will be the pressure of the gaseous mixture when 0.5 L of H₂ at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at 27°C?

Ans: $PV=nRT$

$$P_1V_1=P_2V_2 \quad 0.5 \times 0.8 = P_x \cdot 1 \quad \text{pressure of H}_2 \text{ in the flask} = P_x = 0.40 \text{ bar}$$

$$P_1V_1=P_2V_2 \quad 0.7 \times 2 = P_x \cdot 1 \quad \text{pressure of oxygen in the flask} = 1.4 \text{ bar}$$

$$\text{Pressure of the gaseous mixture} = 1.4 + 0.4 = 1.8 \text{ bar.}$$

10. Density of a gas is found to be 5.46 g/dm³ at 27 °C at 2 bar pressure. What will be its density at STP?

Ans: $d = PM/RT$

$$M = dRT/P$$

$$= 5.46 \times 0.083 \times 300 / 2$$

$$= 70$$

$$d = (1.013 \times 70) / (0.083 \times 273)$$

$$= 3.13 \text{ g/dm}^3$$

11. Calculate the temperature of 4.0 mol of a gas occupying 5 dm³ at 3.32 bar. (R = 0.083 bar dm³ K⁻¹ mol⁻¹).

Ans: $PV=nRT$

$$T = PV/nR$$

$$= (3.32 \times 5) / (4 \times 0.083)$$

$$= 50 \text{ K}$$

12. Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm³ at 27°C. R = 0.083 bar dm³ K⁻¹ mol⁻¹.

Ans: $n = n_{O_2} + n_{H_2}$

$$= 8/32 + 4/2$$

$$= 0.25 + 2$$

$$= 2.25 \text{ moles}$$

$$PV = nRT;$$

$$P = nRT/V$$

$$= 2.25 \times 0.083 \times 300 / 1$$

$$= 56.025 \text{ bar}$$

13. Calculate the volume occupied by 8.8 g of CO₂ at 31.1°C and 1 bar pressure. R = 0.083 bar L K⁻¹ mol⁻¹.

Ans: $PV=nRT;$

$$n=8.8/44= 0.2$$

$$V= nRT/P$$

$$= 0.2 \times 0.083 \times 304.1/1$$

$$= 5.05 \text{ L}$$

14. 2.9 g of a gas at 95 °C occupied the same volume as 0.184 g of dihydrogen at 17 °C, at the same pressure. What is the molar mass of the gas?

Ans: $PV= nRT$

$$V=nRT/P= (0.184/2) \times R \times 290/P=(2.9/M) \times R \times 368/P$$

$$0.092 \times 290=368 \times 2.9/M$$

$$M= 368 \times 2.9/0.092 \times 290$$

$$= 40$$

15. A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

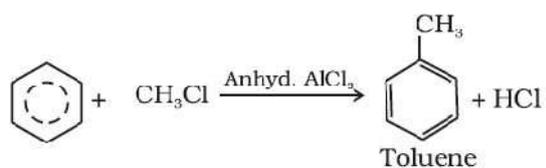
Ans: let 20g H₂ + 80g O₂

$$n= 20/2 + 80/32= 10 + 2.5= 12.5$$

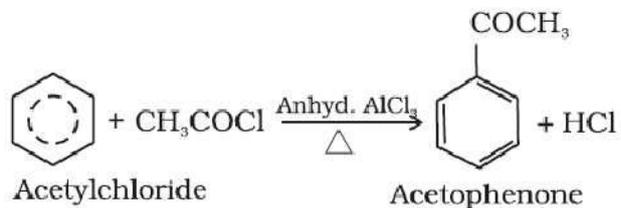
$$\text{Mole fraction of H}_2= 10/12.5= 0.8$$

$$p_{\text{H}_2}= 0.8 \times P$$

(i)

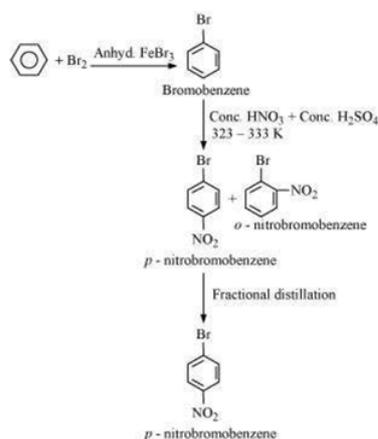


(j)

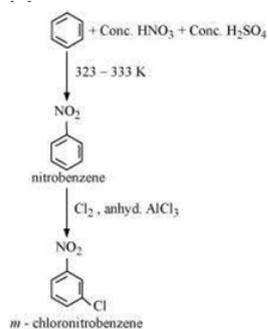


2. How will you convert benzene into (i) p-nitrobromobenzene (ii) m-nitrochlorobenzene (iii) p-nitrotoluene (iv) acetophenone?

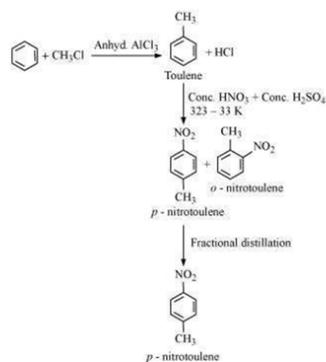
Ans.(i)



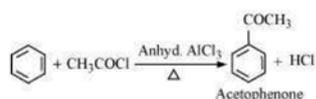
(ii)



(iii)



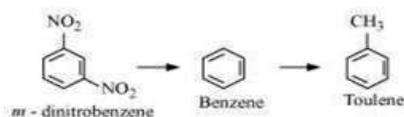
(iv)



3. Out of benzene, m-dinitrobenzene and toluene which will undergo nitration most easily and why? Ans.

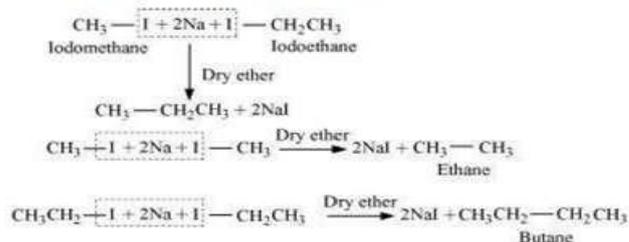
The ease of nitration depends on the presence of electron density on the compound to form nitrates. Nitration reactions are examples of electrophilic substitution reactions where an electron-rich species is attacked by a nitronium ion (NO_2^+).

Now, CH_3 - group is electron donating and NO_2 - is electron withdrawing. Therefore, toluene will have the maximum electron density among the three compounds followed by benzene. On the other hand, m-Dinitrobenzene will have the least electron density. Hence, it will undergo nitration with difficulty. Hence, the increasing order of nitration is as follows:



4. Why is Wurtz reaction not preferred for the preparation of alkanes containing odd number of carbon atoms? Illustrate your answer by taking one example. Ans.

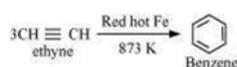
Wurtz reaction cannot be used for the preparation of unsymmetrical alkanes because if two dissimilar alkyl halides are taken as the reactants, then a mixture of alkanes is obtained as the products. Since the reaction involves free radical species, a side reaction also occurs to produce an alkene. For example, the reaction of bromomethane and iodoethane gives a mixture of alkanes.



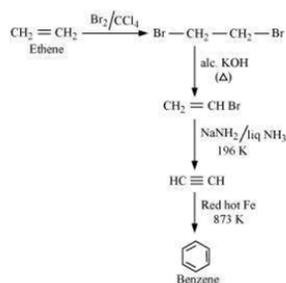
The boiling points of alkanes (obtained in the mixture) are very close. Hence, it becomes difficult to separate them.

5. How would you convert the following compounds into benzene? (i) Ethyne (ii) Ethene (iii) Hexane Ans.

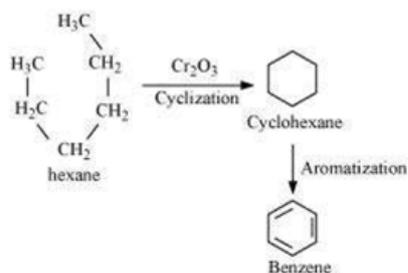
(i)



(ii)



(iii)



6. Why does benzene undergo electrophilic substitution reactions easily and nucleophilic substitutions with difficulty?

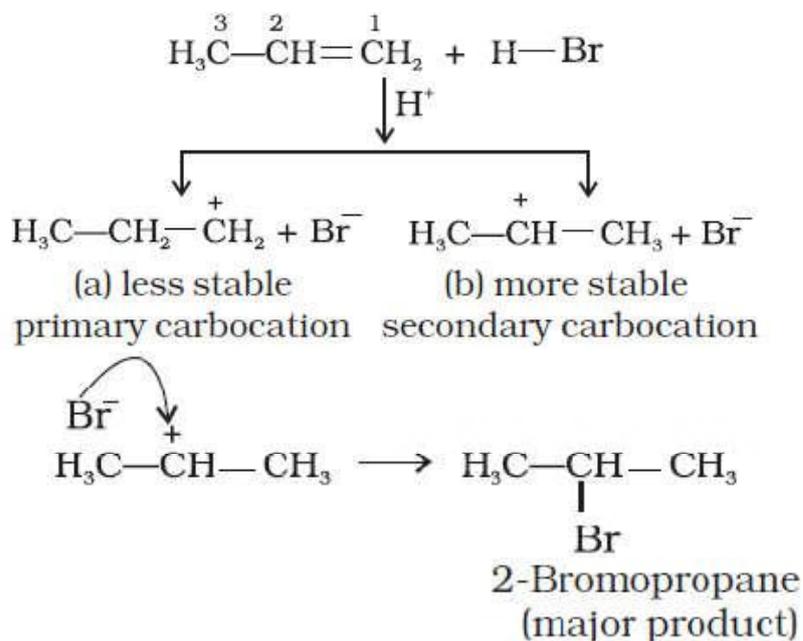
Ans.

Benzene is a planar molecule having delocalized electrons above and below the plane of ring. Hence, it is electron-rich. As a result, it is highly attractive to electron deficient species i.e., electrophiles.

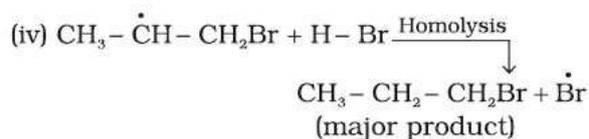
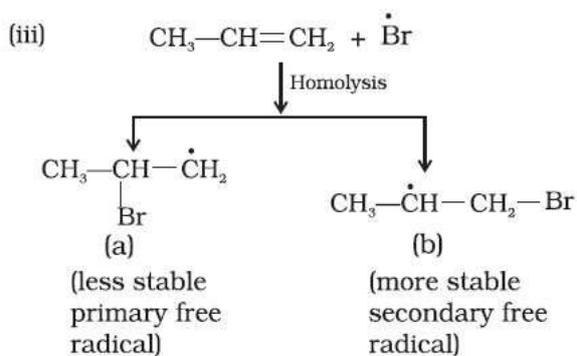
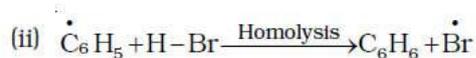
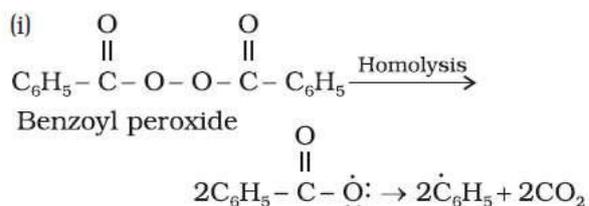
Therefore, it undergoes electrophilic substitution reactions very easily. Nucleophiles are electron-rich. Hence, they are repelled by benzene. Hence, benzene undergoes nucleophilic substitutions with difficulty.

7. Addition of HBr to propene yields 2-bromopropane, while in the presence of benzoyl peroxide, the same reaction yields 1-bromopropane. Explain and give mechanism.

Ans. Addition reaction of HBr to unsymmetrical alkenes follows Markovnikov's rule. Mechanism: Hydrogen bromide provides an electrophile, H^+ , which attacks the double bond to form a carbocation as shown below:



In the presence of peroxide, addition of HBr to unsymmetrical alkenes like propene takes place contrary to the Markovnikov rule. This happens only with HBr but not with HCl and HI. This reaction is known as peroxide or Kharash effector addition reaction anti to Markovnikov rule. Peroxide effect proceeds via free radical chain mechanism as given below:



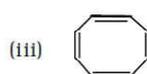
8. What are the necessary conditions for any system to be aromatic?

Ans. (i) Planarity

(ii) Complete delocalisation of the π electrons in the ring

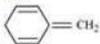
(iii) Presence of $(4n + 2)$ π electrons in the ring where n is an integer ($n = 0, 1, 2, \dots$). This is often referred to as Hückel Rule.

9. Explain why the following systems are not aromatic?



Ans.

(i)



For the given compound, the number of π -electrons is 6.

By Huckel's rule,

$$4n + 2 = 6$$

$$4n = 4$$

$$n = 1$$

For a compound to be aromatic, the value of n must be an integer ($n = 0, 1, 2, \dots$). Since the value of n is an integer, the given compound is aromatic in nature.

(ii)



For the given compound, the number of π -electrons is 4.

By Huckel's rule,

$$4n + 2 = 4$$

$$4n = 2$$

$$n = \frac{1}{2}$$

For a compound to be aromatic, the value of n must be an integer ($n = 0, 1, 2, \dots$), which is not true for the given compound. Hence, it is not aromatic in nature.

(iii)



For the given compound, the number of π -electrons is 8.

By Huckel's rule,

$$4n + 2 = 8$$

$$4n = 6$$

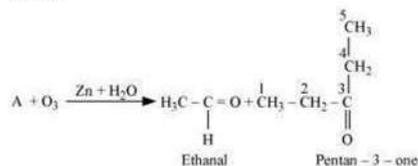
$$n = \frac{3}{2}$$

For a compound to be aromatic, the value of n must be an integer ($n = 0, 1, 2, \dots$). Since the value of n is not an integer, the given compound is not aromatic in nature.

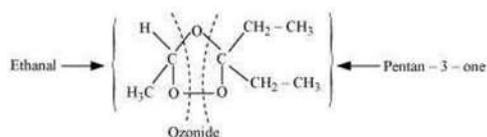
10. An alkene 'A' on ozonolysis gives a mixture of ethanal and pentan-3-one. Write structure and IUPAC name of 'A'.

Ans.

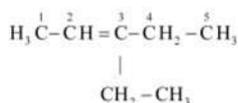
Answer



During ozonolysis, an ozonide having a cyclic structure is formed as an intermediate which undergoes cleavage to give the final products. Ethanal and pentan-3-one are obtained from the intermediate ozonide. Hence, the expected structure of the ozonide is:



This ozonide is formed as an addition of ozone to 'A'. The desired structure of 'A' can be obtained by the removal of ozone from the ozonide. Hence, the structural formula of 'A' is:



The IUPAC name of 'A' is 3-Ethylpent-2-ene.

Unit-10

S-Block Elements

Concept:ALKALI METALS

KNOWLEDGE

1 Mark Questions

Q.1 What happens when KO_2 reacts with water ?

Ans. $2 \text{KO}_2 + 2\text{H}_2\text{O} \rightarrow 2 \text{KOH} + \text{H}_2\text{O}_2 + \text{O}_2$

Q2 Name the chief factor responsible for the anomalous behavior of lithium.

Ans. High polarizing power (charge/size ratio).

Q3 What are the raw material used for the manufacture of washing soda by solvay process ?

Ans. NaCl , CaCO_3 and NH_3

Q.4 Why is sodium metal kept under kerosene oil ?

Ans. Sodium metal is highly reactive and combines with air forming sodium oxide.

Q.5 Name the alkali metal which shows diagonal relationship with Mg?

Ans. Li

Q.6. Name the alkali metal which forms only the normal oxide on heating with air ?

Ans. Li.

2 Mark Questions

Q.1 When a cation is highly polarizing ? Which alkali metal cation has the highest polarizing power ?

Ans. A cation is highly polarizing if its charge/size ratio is very high. Li^+ has the highest polarizing power .

Q.2 why sodium is less reactive than potassium?

Ans. The ionization enthalpy of potassium is less than that of sodium and the electrode potential of

Potassium (-2.95 V) is more negative than that of sodium (-2.71 V) and hence sodium is less reactive than potassium.

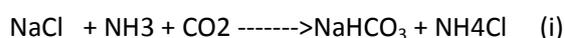
Q.3 Why are potassium and caesium, rather than lithium used in photoelectric cells ?

Ans. Potassium and caesium have much lower ionization enthalpy than that of lithium. As a result, these metals on exposure to light, easily emit electrons but lithium does not. Therefore, K and Cs rather than Li are used in photoelectric cells

3 Mark Questions

Q.1 Discuss the various reactions that occur in the Solvay process.

Ans. In Solvay process, CO_2 is passed through brine saturated with ammonia. When NaHCO_3 being sparingly soluble gets precipitated.



Sodium bicarbonate thus formed is filtered, dried and then heated when sodium carbonate is obtained.



CO_2 needed for the reaction shown in eq. (i) is prepared by heating calcium carbonate and the quick lime, CaO thus formed is dissolved in water to form slaked lime, $\text{Ca}(\text{OH})_2$.

NH_3 needed for the purpose is prepared by heating NH_4Cl

Q.2. In what ways lithium shows similarities to magnesium in its chemical behavior?

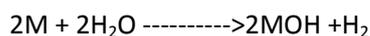
Ans. (a) the hydroxides of both lithium and Mg decompose on heating

(b) Both form ionic nitrides when heated in an atmosphere of nitrogen.

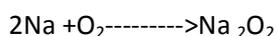
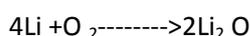
(c) The carbonates of both Li and Mg decompose on heating

Q3. Write the chemical features of alkali metals.

Ans. (I) All alkali metals react with water, liberating H_2 and forming their hydroxides.

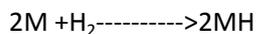


(II) Alkali metals tarnish in air due to formation of an oxide or hydroxide on their surface.



(iii) All alkali metals react with hydrogen to form

ionic hydrides.



Q5. Write two important uses of each of the following: (i) Sodium carbonate (ii) Sodium Hydroxide (iii) Sodium hydrogen carbonate

Ans. (i) Sodium carbonate: for softening of hard water, laboratory reagent.

(ii) Sodium hydroxide: in the manufacture of soap, in the purification of bauxite.

(iii) Sodium hydrogen carbonate: as a mild antiseptic, as an antacid.

UNDERSTANDING

1 Mark Questions

Q1. Which alkali metal will have highest hydration enthalpy?

Ans. Li

Q2. K_2CO_3 cannot be prepared by Solvay process?

Ans. Because $KHCO_3$ is highly soluble in water

2 Mark Questions

Q1 (i) Which alkali metal has the highest ionic radii?

Ans. Cs.

(ii) A solution of sodium in liquid ammonia is strongly reducing in nature. Why?

Ans. Due to presence of ammoniated electrons.

Q2. Why does table salt get wet in rainy season?

Ans. NaCl is not hygroscopic but impurities like $MgSO_4$, $CaSO_4$, $MgCl_2$ and $CaCl_2$, present in it are hygroscopic and absorb moisture from air.

Q3. Why do alkali metals impart color to the flame?

Ans. Alkali metals have low ionization enthalpies. Their valence electrons get excited by absorbing energy from the flame. When these electrons return to the ground state, the energy is emitted

In the form of light.

Q.5 Sodium fire in laboratory should not be extinguished by pouring water. Why?

Ans. Sodium reacts violently with water producing hydrogen gas which also catches fire. Instead of water pyrene should be used.

Q.6. Potassium carbonate cannot be prepared by Solvay process. Why?

Ans. Potassium carbonate cannot be prepared by Solvay process because potassium bicarbonate being more soluble than sodium bicarbonate does not get precipitated when CO_2 is passed through a Conc. Solution of KCl saturated with ammonium.

3 Mark Questions

Q1. Give reason for the following:

- (a) LiCl is more covalent than KCl.
- (b) LiF has lower melting point than NaCl.
- (c) CuCl is more covalent than NaCl.

Ans(a). Due to smaller size and high polarising power of Li^+ ion, LiCl is more covalent than KCl.

(b) Small cation Li^+ ion highly polarizes large anion F^- , LiF is more covalent and hence has lower melting point.

© Due to pseudo inert gas configuration, Cu^+ is more polarizing than Na^+ and hence CuCl is more Covalent NaCl.

Q.2 when an alkali metal dissolves in liquid ammonia, the solution can acquire different colors. Explain.

Ans. The dilute solutions of alkali metals in liquid ammonia exhibit dark blue color because ammoniated electrons absorb energy corresponding to the red region of the visible light. However, if the concentration increases above 3M, the color changes to copper-bronze and the solution acquires metallic luster due to the formation of metal ion clusters.

APPLICATION

2 Mark Questions

Q.1. What is the oxidation state of K in KO_2^- ? Why is KO_2 paramagnetic?

Ans. The oxidation state of K in KO_2^- is +1. It has 1 unpaired electron in π^*2p MO. Therefore it is

Paramagnetic.

Q.2. Why Li_2CO_3 decomposes at a lower temperature whereas Na_2CO_3 at higher temperature ?

Ans. Li_2CO_3 is a salt of a weak acid (CO_2) and a weak base (LiOH). Since the weak base cannot attract CO_2 strongly, therefore, Li_2CO_3 decomposes at a lower temperature. On the other hand, NaOH is a much stronger base than LiOH and hence can attract CO_2 more strongly.

3 Mark Questions

Q1. What happens when

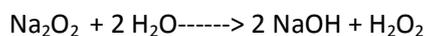
- (i) Sodium metal is dropped in water ?
- (ii) Sodium metal is heated in free supply of air ?
- (iii) Sodium peroxide dissolves in water ?

Ans. (I) $2 \text{Na}(s) + 2 \text{H}_2\text{O} \rightarrow 2 \text{NaOH} + \text{H}_2$

(II) Na_2O_2 is formed



(iii) H_2O_2 is formed



Q.2 Write balanced equation for the reaction between

(a) Na_2O_2 and water

(b) KO_2 and water

(c) Na_2O and CO_2

Ans(a). $\text{Na}_2\text{O} + 2 \text{H}_2\text{O} \rightarrow 2 \text{NaOH} + \text{H}_2\text{O}_2$

(b) $2 \text{KO}_2 + 2 \text{H}_2\text{O} \rightarrow 2 \text{KOH} + \text{H}_2\text{O}_2 + \text{O}_2$

(c) $\text{Na}_2\text{O} + \text{CO}_2 \rightarrow \text{Na}_2\text{CO}_3$

Q.3 How do you account for the strong reducing power of lithium in aqueous solution?

Ans. Electrode potential is a measure of the tendency of an element to lose electrons in the aq. solution. It depends upon following factors :

(i) Sublimation enthalpy.

(II) Ionisation enthalpy

(iii) Hydration enthalpy

Since Li has the smallest size, its enthalpy of hydration is the highest among alkali metals. Although ionization enthalpy of Li is the highest among alkali metals, it is more than compensated by the high enthalpy of hydration. Thus Li has the most negative E° value and hence lithium is the strongest reducing agent.

Q.4. Mention the oxides formed by Li, Na and K.

Ans. $4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O}$

$2\text{Na} + \text{O}_2 \rightarrow \text{Na}_2\text{O}_2$

$\text{K} + \text{O}_2 \rightarrow \text{KO}_2$

Concept: Alkaline Earth Metals ; Anomalous Behaviour & Diagonal Relationship

Knowledge

one mark Questions

Q 1 In second gp which element shows Anomalous Behaviour ?

Ans Be

Q 2 with which element Beryllium shows diagonal relation ship?

Ans : Al

Q3 Name the elements present in the 2nd Group of the Periodic Table:

Ans beryllium, magnesium, calcium, strontium, barium and radium

Two mark Questions

Q 1 Elements present in the 2nd Group of the Periodic Table are also called ?

Ans alkaline earth metals

Q 2. Write electronic Configuration of Be & Mg

Ans Be $1s^2 2s^2 2p^2$

Mg $1s^2 2s^2 2p^6 3s^2 3p^2$

Q 3 What is the diagonal relationship ?

Ans : A close similarity is observed in certain cases between the first element of a group with the second element of the following group. This is referred to as the diagonal relationship and is observed in the following pairs..Li & Mg

Be & Al

Three mark questions

Q 1 Write any three uses of Calcium Hydroxide (Slaked lime)

Ans i) It is used in the preparation of mortar, a building material.

(ii) It is used in white wash due to its disinfectant nature.

(iii) It is used in glass making, in tanning industry, for the preparation of bleaching powder and for purification of sugar.

Q 2 How does the setting of cement takes place & what is the role of gypsum in setting of cement ?

Ans When mixed with water, the setting of cement takes place to give a hard mass. This is due to the hydration of the molecules of the constituents and their Rearrangement. The purpose of adding gypsum is only to slow down the process of setting of the cement so that it gets sufficiently hardened.

Understanding

1 mark question

Q 1 What is the reason for the diagonal relationship ?

Ans Diagonal relationship is due to the similarity in ionic sizes and /or charge/radius ratio of the elements.

Q 2 Beryllium forms covalent compound Why?

Ans Due to high Ionisation Enthalpy

Q 3 Beryllium does not show coordination more than four, Why ?

Ans Due absence of d Orbitals

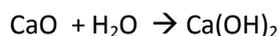
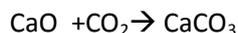
Q 4 Beryllium, the first member of the Group 2 metals, shows different behaviour as compared to magnesium and rest of the members what is this properties

Ans Anomalous Behaviour

2 Mark Question

Q1. What happens when CaO is exposed to atmospheric air? (L2)

Ans On exposure to atmosphere, it absorbs moisture and carbon dioxide to form Calcium hydroxide and Calcium carbonate.



Note; The addition of limited amount of water breaks the lump of lime. This process is called **slaking of lime**.

Q 2. Give reason . the sulphate of Be & Mg are soluble in water .

Ans The sulphates of the alkaline earth metals are all white solids and **stable to heat**. BeSO₄, and MgSO₄ are readily soluble in water; the solubility decreases from CaSO₄ to BaSO₄. The greater hydration enthalpies of Be²⁺ and Mg²⁺ ions overcome the lattice enthalpy factor and therefore their sulphates are soluble in water.

3 Mark Questions

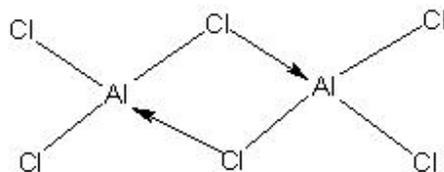
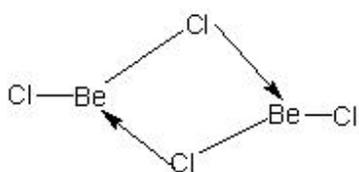
Q 1 What is the Reason for the anomalous behaviour of beryllium ?

Ans The properties of beryllium the first member of the alkaline earth metal, differ from the rest of the member. Its is mainly because of

- M Its small size and high polarizing power.
- M Relatively high electro negativity and ionization energy as compared to other members.
- M Absence of vacant d – orbitals in its valence shell.

Q 2 Give resemblance in structure of chloride of Be & Al .

Ans BeCl_2 and AlCl_3 have bridged chloride polymeric structure.



Q 3. Give reaction for the amphoteric nature of Be & Al.

Ans the oxides and hydroxides of both Be and Al are amphoteric and dissolve in sodium hydroxide as well as in hydrochloric acid.



Application

1 mark question

Q 1 Aluminium, Beryllium is not readily attacked by acids, Why ?(L3)

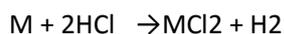
Ans :Because of the presence of an oxide film.

Q 2 The chlorides of both beryllium and aluminium are Lewis acid Why ?

Ans Incomplete octate.

Q 3 Name the gas liberated when alkaline metals react with dil acid?

Ans ; The alkaline earth metals readily react with acids liberating dihydrogen gas .



Q 4 Beryllium chloride, BeCl_2 , melts at 405°C and boils at 520°C . That compares with 714°C and 1412°C for magnesium chloride. Give reason ?

Ans Be forms covalent compounds

Q 5 Give reason .the compounds of alkaline earth metals are less ionic than alkali metals

Ans This is due to increased nuclear charge and smaller size.

2 Mark Question

Q 1 Beryllium shows covalent Character in its compounds ,Why?

Ans Due to the small size, high ionization **energy**, and high sublimation energy of beryllium, its lattice and hydration energies are insufficient to provide complete charge separation and the formation of

simple Be^{2+} ions. In all compounds therefore, even BeO and BeF_2 , there is substantial covalent character in the bonding.

Q 2. Account for the reducing nature of Be

Ans Beryllium has less negative value compared to other alkaline earth metals. However, its reducing nature is due to large hydration energy associated with the small size of Be^{2+} ion and relatively large value of the atomization enthalpy of the metal.

Q 3 How does the atomic and Ionic Radii of alkaline earth metals vary in comparison to alkali metals

Ans The atomic and ionic radii of the alkaline earth metals are smaller than those of the corresponding alkali metals in the same periods. This is due to the increased nuclear charge in these elements. Within the group, the atomic and ionic radii increase with increase in atomic number.

Q 4 How does the of Ionization Enthalpy of alkaline earth metals vary in comparison to alkali metals.

Ans The alkaline earth metals have low ionization enthalpies due to fairly large size of the atoms. Since the atomic size increases down the group, their ionization enthalpy decreases. The first ionisation enthalpies of the alkaline earth metals are higher than those of the corresponding Group 1 metals. This is due to their small size as compared to the corresponding alkali metals. It is interesting to note that the second ionisation enthalpies of the alkaline earth metals are smaller than those of the corresponding alkali metals.

Q 5 . How does the of Hydration Enthalpy of alkaline earth metals vary & compare it with alkali metals.

Ans The hydration enthalpies of alkaline earth metal ions decrease with increase in ionic size down the group. $\text{Be}^{2+} > \text{Mg}^{2+} > \text{Ca}^{2+} > \text{Sr}^{2+} > \text{Ba}^{2+}$ The hydration enthalpies of alkaline earth metal ions are larger than those of alkali metal ions. Thus, compounds of alkaline earth metals are more extensively hydrated than those of alkali metals, e.g., MgCl_2 and CaCl_2 exist as $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ and $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$ while NaCl and KCl do not form such hydrates.

3 Mark Questions

Q 1 Mention the anomalous behaviour of beryllium

Ans Beryllium, the first member of the Group 2 metals, shows anomalous behaviour as compared to magnesium and rest of the members

- i)M Beryllium has exceptionally small atomic and ionic sizes and thus does not compare well with other members of the group. Because of high ionisation enthalpy and small size it forms compounds which are largely covalent and get easily hydrolysed.

- ii)M Beryllium does not exhibit coordination number more than four as in its valence shell there are only four orbitals. The remaining members of the group can have a coordination number of six by making use of d-orbitals.
- iii)M The oxide and hydroxide of beryllium, unlike the hydroxides of other elements in the group, are amphoteric in nature.

Q 2 Mention the Diagonal Relationship between Beryllium and Aluminium

The ionic radius of Be^{2+} is estimated to be 31 pm; the charge/radius ratio is nearly the same as that of the Al^{3+} ion. Hence beryllium resembles aluminium in some ways. Some of the similarities are:

- (i)M Like aluminium, beryllium is not readily attacked by acids because of the presence of an oxide film on the surface of the metal.
- (ii)M Beryllium hydroxide dissolves in excess of alkali to give a beryllate ion, $[\text{Be}(\text{OH})_4]^{2-}$ just as aluminium hydroxide gives aluminate ion, $[\text{Al}(\text{OH})_4]^-$.
- (iii)M The chlorides of both beryllium and aluminium have Cl^- bridged chloride structure in vapour phase. Both the chlorides are soluble in organic solvents and are strong Lewis acids. They are used as Friedel Craft catalysts.
- (iv)M Beryllium and aluminium ions have strong tendency to form complexes, BeF_4^{2-} , AlF_6^{3-} .

Q 3 What is the colour imparted to the flame by Calcium, Strontium and Barium?

Ans Calcium, strontium and barium impart characteristic brick red, crimson and apple green colours respectively to the flame. In flame the electrons are excited to higher energy levels and when they drop back to the ground state, energy is emitted in the form of visible light.

Q 4 Why are Be & Mg inert to O_2 & H_2O ? (L3)

Ans Beryllium and magnesium are chemically inert to oxygen and water because of the formation of an oxide film on their surface. However, powdered beryllium burns brilliantly on ignition in air to give BeO and Be_3N_2 .

Magnesium is more electropositive and burns with dazzling brilliance in air to give MgO and Mg_3N_2 .

Calcium, strontium and barium are readily attacked by air to form the oxide and nitride. They also react with water with increasing vigour even in cold to form hydroxides.

HYDROCARBONS

1. Why carbon does have a larger tendency of catenation than silicon although they have same number of electrons?

Ans .It is due to the smaller size C-C bond which is stronger (335 KJ mol^{-1}) than in Si-Sibond ($225.7 \text{ KJ mol}^{-1}$).

2.Methane does not react with chlorine in dark. Why?

Ans .Chlorination of methane is a free radical substitution reaction. In dark, chlorine is unable to be converted into free radicals, hence the reaction does not occur.

3.Why do alkynes not show geometrical isomerism?

Ans.Alkynes have linear structure. So they cannot show geometrical isomerism.

4.Although benzene is highly unsaturated; it does not undergo addition reactions. Give reason.

Ans .Unlike olefins, π -electrons of benzene are delocalized (resonance) and hence these are uncreative towards addition reactions.

5.The boiling point of hydrocarbons decreases with increase in branching. Give reason.

Ans. Branching result into a more compact (nearly spherical) structure. This reduces the effective surface area and hence the strength of the Vander wall's forces, thereby leading to a decrease in the boiling point.

6. Unsaturated compounds undergo addition reactions. Why?

Ans. Unsaturated hydrocarbon compounds contain carbon – carbon double or triple bonds. The π -bond is multiple bond is unstable and therefore addition takes place across the multiple bonds.

7. Cyclobutane is less reactive than cyclopropane. Justify.

Ans . In cyclobutane molecule, the C-C-C bond angle is 90° while it is 60° in cycloprpane. shows that the deviation from the tetrahedral bond angle (109°) in cyclobutane is less than in cyclopropane. In other words, cyclopropane is under great strain compared with cyclobutane and is therefore more reactive.

8. The boiling point of alkanes shows a steady increase with increase in molecular mass. Why?

Ans .This is due to the fact that the intermolecular van der walls forces increase with increase of the molecular size or the surface area of the molecule.

9. Pentane has three isomers i.e; pentane, 2-methyl butane and 2,2-dimethyl propane .

The b.p of pentane is 309.1K whereas 2,2-dimethyl propane shows a b.p of 282.5k . Why?

Ans .With the increase in number of branched chains, the molecule attains the shape of a sphere. This results in smaller area of contact and therefore weak inter molecular forest between spherical molecules, which are overcome a relatively lower temperatures.

10. All the four C-H bonds in methane are identical. Give reasons.

Ans .The four C-H bonds of methane are identical because all of these are formed by the overlapping of the same type of orbital's i.e; hybrid orbital's of carbon and s-orbital of hydrogen.

11. When alkanes are heated, the C-C bonds rather than the C-H bonds break. Give

reason.

Ans. When alkanes are heated, the C-C bonds rather than the C-H bonds break because the C-C bond has a lower bond energy ($\Delta H=83\text{ K Cal/mol}$) than the C-H bond ($\Delta H=99\text{ K Cal/mol}$).

12. Although acetylene is acidic in nature, it does not react with NaOH or KOH. Give reason?

Ans. Acetylene is a very weak acid ($pK_a=25$) and hence only an extremely strong base like amide ion (NH_2^-) can successfully remove a proton.

13. n-pentane has higher boiling point than neopentane but the melting point of neopentane is higher than that of n-pentane.

Ans. Because of the presence of branches in neo-pentane the surface area and van der Waals forces of attraction are very weak in neopentane than in n-pentane. Therefore the b.p of neopentane is lower than that of n-pentane.

M.P depends upon the packing of the molecules in the crystal lattice. Since neopentane are more symmetrical than n-pentane therefore, it packs much more closely in the crystal lattice than n-pentane and hence neopentane has much higher m.p than n-pentane.

14. The dipole moment of trans 1,2-dichloroethane is less than the cis-isomer. Explain.

Ans. The structure of trans isomer is more symmetrical as compared to the cis-isomer. In the trans-isomer, the dipole moments of the polar C-Cl bonds are likely to cancel effect of each other and the resultant dipole moment of the molecule is nearly zero. But in the cis-isomer, these do not cancel. Therefore, the cis isomer has a specific moment but is zero in case of trans isomer.

15. Ethyne is acidic in nature in comparison to ethene and ethane. Why is it so?

Ans. Hydrogen atoms in ethyne are attached to the sp hybridised carbon atoms whereas they are attached to sp² hybridized carbon atoms in ethene and sp³ hybridised carbons in ethane. Due to the maximum percentage of s-character (50%), the sp hybridized orbital's of carbon atoms in ethyne molecules have highest electronegativity: Which attracts the shared pair of the C-H bond of ethyne to a greater extent than that of the sp² hybridized orbital's of carbon in ethene and the sp³ hybridized orbital of carbon in ethane. Thus in ethyne molecule, hydrogen atoms can be liberated as protons more easily as compared to ethene and ethane.

16. p-chloro nitro benzene has less dipole moment (2.4 D) than p-nitro toluene (4.4 D).

Why?

Ans. In p-chloro nitro benzene the individual moments are in opposite directions and hence partially cancel. When in p-nitro toluene, both moments are in the same direction and hence add each

17. Lassaigne's test is not shown by diazonium salts. Why?

Ans. Diazonium salts usually leave N₂ on heating much before they have a chance to react with the fused sodium metal. Therefore, diazonium salts do not show positive Lassaigne's test for nitrogen.

18. Why is sp hybrid orbital more electronegative than sp² or sp³ hybridized orbitals?

Ans. The greater the s-character of the hybrid orbital's, the greater is the electronegativity. Thus, a carbon atom having an sp hybrid orbital with 50% s-character is more electronegative than that possessing sp² or sp³ hybridized orbital's

19. Explain why is $(\text{CH}_3)_3\text{C}^+$ more stable than CH_3CH_2^+ and CH_3^+ is the least stable cation.

Ans. Hyper conjugation interaction in $(\text{CH}_3)_3\text{C}^+$ is greater than in CH_3CH_2^+ as $(\text{CH}_3)_3\text{C}^+$ has nine C-H bonds. In CH_3^+ , the C-H bond is in the nodal plane of the vacant 2p orbital and hence cannot overlap with it. Thus, CH_3^+ lacks hyper conjugate stability.

20. Will CCl_4 give white precipitate of AgCl on heating it with AgNO_3 ?

Ans. CCl_4 does not give white precipitate with silver nitrate solution.

$\text{CCl}_4 + \text{AgNO}_3 \rightarrow$ No reaction.

Carbon tetrachloride contains chlorine but it is bonded to carbon by a covalent bond.

Therefore it is not in ionic form. Hence, it does not combine with AgNO_3 solution.

21. Lassaigne's test is not shown by diazonium salt. Why?

Ans: On heating diazonium salt, loses Nitrogen and could not fuse with the Sodium metal therefore diazonium salt does not show Positive Lassaigne's test for Nitrogen.

22. Why are alcohols weaker acids than water?

Ans: The alkyl group in alcohol has a +I effect due to which electron density increases on the oxygen atom which makes the release of H^+ ion more difficult from alcohol $\text{R}-\text{O}-\text{H}$

23. Why is nitric acid added to Sodium Extract before adding Silver nitrate for testing halogens?

Ans: Nitric acid is added to decompose NaCN & Na_2S .

$\text{NaCN} + \text{HNO}_3 \rightarrow \text{NaNO}_3 + \text{HCN}$

$\text{Na}_2\text{S} + \text{HNO}_3 \rightarrow 2\text{NaNO}_3 + \text{H}_2\text{S}$

24. Why is benzene extraordinarily stable though it contains three double bonds?

Ans: Due to resonance in benzene the π -electron cloud gets delocalized resulting in the stability of the molecule.

25. Why do alkanes not dissolve in water but dissolve in benzene?

Ans: Because alkanes are non-polar and water is a polar solvent.

MLL ----- P BLOCK (13,14,15) CALSS –XI

1. Name the groups which called p block element.
2. What is the general electronic configuration of p block elements?
3. Boiling point of NH_3 is more than PH_3 .
4. White phosphorus is more reactive than red phosphorus.
5. Write the reaction between phosphorus and sodium hydroxide.
6. Boron is not able to form BF_6^{3-} .
7. Graphite can be used as dry lubricant . how ?
8. How water gas can be produced ?
9. How CO is dangerous for human ?
10. What are electron deficient compounds ? Are BCl_3 and SiCl_4 electron deficient species ? Explain.

Answers

1. Group no 13-18
2. ns^2np^{1-6}
3. Due to hydrogen bonding .
4. The bond angle in p-p bond is 60° in white phosphorus while 100° in red phosphorus (bond strain)
5. $\text{P}_4 + 3\text{NaOH} \rightarrow \text{PH}_3 + \text{NaH}_2\text{PO}_2$
6. Unavailability of d – orbitals.
7. Graphite is made of layers that can be slip on each others.
8. $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{H}_2(\text{g})$ (473 1273K)
9. CO get react with haemoglobin and form complex compound .it can be fatal for human.
10. Electron deficient compound are having incomplete octet.
 BCl_3 is electron deficient compound while SiCl_4 is not.

1. What is inert pair effect .

Ans. The inability of valence shell s-electrons to participate in the bond formation is known as inert pair effect.

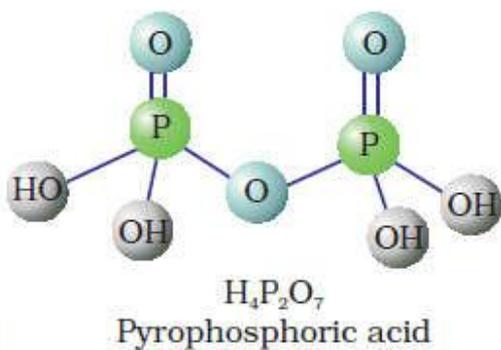
2. Name the ore of Tin.

Ans. Cassiterite, SnO₂.

3. How many oxygen atoms are shared per tetrahedron are shared.

Ans. Two oxygen atoms per tetrahedron are shared .

4. Give the structure of pyrophosphoric acid ,H₄P₂O₇.



Ans. acid

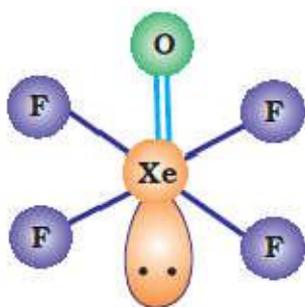
5. Why does sulphur in vapour state exhibit paramagnetic behaviour ?

Ans. Sulphur in gaseous state exist as S₂ molecule and it has got Unpaired electrons .

6. Arrange in increasing order of Bond energy: F₂ , Cl₂ , Br₂ , I₂,

Ans. I₂ < F₂ < Br₂ < Cl₂

7. Write the structure of XeOF₄



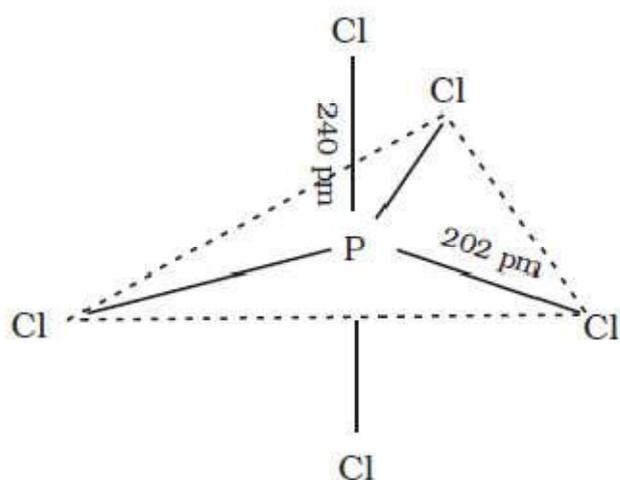
(d) Square pyramidal

8. Why are interhalogen compounds more reactive than halogens ?

Ans. It is because interhalogen compounds are polar in nature.

9. Are all five bonds in PCl_5 molecule equivalent ? Justify your answer.

Ans



(No Axial bonds are larger than equatorial bonds)

Reasoning (P BLOCK 13,14,15 GROUP) -----

1. PCl_5 can be formed but NCl_5 cannot be formed. why?
2. How you can say this is a disproportionation reaction.
 $3\text{HNO}_2 \rightarrow \text{HNO}_3 + \text{H}_2\text{O} + 2\text{NO}$.
3. PH_3 has lower boiling point than NH_3 . Why?
4. Why is BiH_3 the strongest reducing agent amongst all the hydrides of Group 15 elements ?
5. Why is N_2 less reactive at room temperature?
6. Why does NH_3 act as a Lewis base ?
7. Why does NO_2 dimerise ?
8. Bond angle in PH_4^+ is higher than that in PH_3 . Why?
9. Why does PCl_3 fume in moisture ?
10. How do you account for the reducing behavior of H_3PO_2 on the basis of its structure ?
11. White fumes appear around the bottle of anhydrous aluminium chloride. Give reason.
12. Why boric acid is considered as weak acid ?
13. $[\text{SiF}_6]^{2-}$ is known whereas $[\text{SiCl}_6]^{2-}$ not. why.
14. Diamond is covalent, yet it has high melting point. Why?
15. Conc. HNO_3 can be transported in aluminium container. why?

Answer

1. Due to absence of d – orbitals .
2. The oxidation no of N is not equal in all molecules.
3. Unlike NH₃, PH₃ molecules are not associated through hydrogen bonding in liquid state. That is why the boiling point of PH₃ is lower than NH₃.
4. due to less bond dissociation enthalpy.
5. due to high bond dissociation enthalpy.
6. Nitrogen atom in NH₃ has one lone pair of electrons which is available for donation. Therefore, it acts as a Lewis base.
7. Due to presence of an odd electron on the molecules.
8. Less repulsion in bond pair (absence of lone pair of electron) in PH₄⁺
9. PCl₃ hydrolyses in the presence of moisture giving fumes of HCl.
$$\text{PCl}_3 + 3\text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_3 + 3\text{HCl}$$
10. In H₃PO₂, two H atoms are bonded directly to P atom which imparts reducing character to the acid.
11. Anhydrous aluminium chloride is partially hydrolysed with atmospheric moisture to liberate HCl gas. Moist HCl appears white in colour.
12. It cannot release H⁺ ions its own. It can receive OH⁻ ions from water to complete its octet and in turn it release H⁺ ion.
13. The main reasons are :
 - (i) Six large chloride ions cannot be accommodated around Si⁴⁺ due to limitation of its size.
 - (ii) Interaction between lone pair of chloride ion and Si⁴⁺ is not very strong.
- 14 . diamond have three dimensional strong covalent bonding .
- 15 . concentrated nitric acid renders aluminium passive by forming a protective oxide layer on the surface.

S Block Elements

1. Which alkali metals form superoxides? Give reason.

Solution: Potassium, Rubidium and caesium form superoxides. Li and Na do not form superoxide. In KO (potassium superoxide) larger anions are stabilized by bigger cation through lattice energy.

2. Why LiOH is a weak base while CsOH is a strong base?

Solution: Li ion due to its small size, is able to attract valence electrons from OH ion to form an ionic bond containing covalency (Fajan's rule) Being covalent, it is sparingly soluble and less basic. For CsOH, size of Cs ion is bigger. Hence, ionic bond strength is maximum and more soluble in water. Hence is a strong base.

3. Why is LiF sparingly soluble in water while other halides of Li are soluble in organic solvents?

Solution: This can be explained through covalency in ionic compounds. As per Fajan's rule (i) small size of cation (2) larger charge of cation (3) large size of anion causes more polarization in ionic bond of ionic crystals. For LiF, size of Li is small. Hence covalency in the ionic bond is possible.

4. Be and Mg salts do not give color to the flame. Give reasons.

Solution: The electrons of nS orbital in Ca, Sr and Ba are excited to higher energy electrons and returns to their original energy level (s orbital). Be and Mg are made of atoms of small size. They have high ionisation energy. Hence, two paired electrons cannot be excited to higher energy level by the energy of the flame.

5. Stability of carbonates of alkaline earth metal increases on moving down the group. Give reason.

Solution: Be CO₃ on heating gives BeO and CO₂. BeO is more stable as the size of Be and O are nearly same. As BeO is more stable than BeCO₃.

6. Although Li ion is the smallest among the alkali metal ions, its conductivity in aqueous solution is less than that of Cs ion. Give reasons.

Solution: Smallest Li ion forms biggest hydrated Li ion. Cs ion forms only smaller hydrated ions than that of Li ion.

7. why Li salts are hydrated?

Li⁺ has high degree of hydration and for this reason lithium salts are mostly hydrated, e.g., LiCl·2H₂O

8. Why alkali metals are highly electro positive & they are not found in free state ?

The loosely held s-electron in the outermost valence shell of these elements makes them the most electropositive metals. They have low ionization enthalpy.

9. Give reason .the melting point and boiling point of alkali metals are low

The melting and boiling points of the alkali metals are low indicating weak metallic bonding due to the presence of only a single valence electron in them.

10. Give reason for the colour imparted to the flame by alkali metals

The alkali metals and their salts impart characteristic colour to an oxidizing flame. This is because the heat from the flame excites the outermost orbital electron to a higher energy level. When the excited electron comes back to the ground state, there is emission of radiation in the visible region.

11. Why Solvay process cannot be extended for the manufacture of potassium carbonate?

Solvay process cannot be extended to the manufacture of potassium carbonate because potassium hydrogencarbonate is too soluble to be precipitated by the addition of ammonium hydrogencarbonate to a saturated solution of potassium chloride.

12. Give reason . the sulphate of Be & Mg are soluble in water .

The greater hydration enthalpies of Be^{2+} and Mg^{2+} ions overcome the lattice enthalpy factor and therefore their sulphates are soluble in water.

13. Why do the the alkali metals give blue solution ,when treated with liq NH_3 ?

The alkali metals dissolve in liquid ammonia giving deep blue solutions which are conducting in nature The blue colour of the solution is due to the ammoniated electron which absorbs energy in the visible region of light and thus imparts blue colour to the solution.

REASONING QUESTIONS OF CLASS XI

CHAPTER: 3 CLASSIFICATION OF ELEMENTS AND PERIODICITY IN PROPERTIES

Q.1. Why do elements in the same group have similar physical and chemical properties?

Ans: Similar valence shell electronic configuration of the elements of the same group.

Q.2. Why are there 32 elements in the sixth period of the periodic table?

Ans: Sixth period begins with filling of the principal quantum number, $n = 6$. According to the Aufbau principle, in the ground state of the atoms, the orbitals are filled in order of their increasing energies. Therefore, in sixth period, electrons enter in 6s, 4f, 5d and 6p subshells.

Q.3. Why are cations smaller and anions are larger in radii than the parent atoms?

Ans: Cations are smaller in radii than their parent atoms because by the loss of one or more electrons, effective nuclear charge increases. Due to this, forces of attraction of the nucleus for electrons increases and hence, the size decreases.

Anions are larger in radii than their parent atoms because by the addition of one or two electrons, nuclear charge experienced per electron decreases. Due to this, forces of attractions between the nucleus and the valence shell electrons decreases and hence, the ionic radius increases.

Q.4. The first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium?

Ans: Na: $1s^2 2s^2 2p^6 3s^1$

Mg: $1s^2 2s^2 2p^6 3s^2$

The first ionization enthalpy of sodium is lower than that of magnesium because the electron to be removed in both cases is from 3s – orbital but the nuclear charge is lower in Na than Mg.

After the removal of first electron, Na^+ acquires inert gas configuration ($\text{Na}^+ : 1s^2 2s^2 2p^6$) and hence, the removal of second electron becomes difficult. While in case of Mg, after the removal of first electron, the electronic configuration of Mg^+ is $1s^2 2s^2 2p^6 3s^1$. In this case, $3s^1$ electron is easy to remove in comparison to the removal of electron from inert gas configuration. Therefore, I.E.₂ of Na is higher than Mg.

Q.5. The first ionization enthalpy of oxygen is lower than nitrogen. Give reason.

Ans: Oxygen has lower first ionization enthalpy than nitrogen because in case of oxygen ($1s^2 2s^2 2p^4$), the loss of electron is to take place from 2p orbital which is not symmetrical.

While in case of nitrogen ($1s^2 2s^2 2p^3$), the removal of electron is to be carried from the stable half filled configuration.

Q.6. Out of Be and B, which has higher first ionization enthalpy and why?

Ans: Be has higher $\Delta_i H_1$ than boron.

Be: $1s^2 2s^2$

B: $1s^2 2s^2 2p^1$

In Be the electron to be removed during the ionization is an s-electron while the electron to be removed during the ionization of B is a p-electron. The penetration power of 2s electrons to the nucleus is more than that of the 2p electron. As a result, 2s electrons are more attracted by the nucleus than 2p electron. Therefore, Be has higher first ionization enthalpy than B.

Q.7. Chlorine has higher electron gain enthalpy than fluorine. Justify.

Ans: The less negative value of electron gain enthalpy of fluorine is due to its small size which has strong inter-electronic repulsions. Thus, fluorine has less tendency to accept an electron and its electron gain enthalpy is less than chlorine.

Q.8. Out of O and F, which has higher electron gain enthalpy and why?

Ans: F has more negative electron gain enthalpy than O because it has smaller size and greater urge to have a noble gas configuration.

Q.9. Why is atomic number a better basis for the classification of the elements than atomic mass?

Ans: The atomic mass of the element relates to the nucleus which is present in the centre of the atom. But the properties of the elements depend upon the electronic configuration which is linked with atomic number. Therefore, atomic number is a better basis for the classification of the elements than the atomic mass.

Q.10. Why is $\Delta_i H_1$ of Mg more than that of Al?

Ans: The $\Delta_i H_1$ of Al is expected to be more than that of Mg because of smaller size and greater nuclear charge. But actually it is less since the electronic configuration of Mg is more symmetrical than that of Al.

Q.11. Why is $\Delta_i H_2$ value of an element more than its $\Delta_i H_1$ value?

Ans: With the loss of an electron, the effective nuclear charge experienced per unit electron increases because of which electrons are held with greater nuclear force. Therefore, greater energy is required to remove the second electron from the gaseous atom. The $\Delta_i H_2$ value of an element is always more than its $\Delta_i H_1$ value.

Q.12. Why does the ionization enthalpy increase as we move from left to right across the given period in a periodic table?

Ans: On moving across a period from left to the right, the effective nuclear charge increases because of decrease in atomic size. As a result, electrons experience more and more attraction towards the nucleus. Therefore, the first ionization enthalpy increases from left to the right in a period.

Q.13. Explain why do ionization enthalpies decrease down a group of the periodic table.

Ans: The ionization enthalpies of the elements decrease down the group because:

- (i) there is increase in the number of the main energy shells (n) in moving from one element to other.
- (ii) There is an increase in the magnitude of the screening effect because of increase in the number of inner electrons.

As a result, electrons experience lesser attraction towards the nucleus down the group and ionization enthalpy decreases.

Q.14. Why do halogens have the maximum negative electron gain enthalpies in their respective periods?

Ans: The atoms of the group 17 have $ns^2 np^5$ electronic configuration of the valence shells and they need only one electron to achieve the configuration of the nearest noble gas atom ($ns^2 np^6$). Thus, the members of the halogens have the maximum negative electron gain enthalpies.

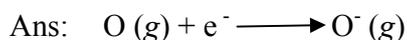
Q.15. Noble gas elements have positive electron gain enthalpies. Why?

Ans: The members of the noble gas family have highly symmetrical electronic configuration and have completely filled orbitals. Therefore, their atoms have hardly any urge to take up extra electrons. In case, these are made to accept electrons under special conditions, enthalpy will be required and their electron gain enthalpies will be positive.

Q.16. Would you expect the first ionization enthalpies for two isotopes of the same element to be same or different? Justify your answer.

Ans: Isotopes are atoms of the same element which have same atomic number but different mass number. Therefore, they have same number of electrons and nuclear charge (protons). Thus, they will have almost same first ionization enthalpies.

Q.17. Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer.



When an electron is added to oxygen atom to form ion, energy is released. Hence, first electron gain enthalpy of oxygen is negative.

But when another electron is added to O^- ion to form O^{2-} ion, it feels stronger force of repulsion. Hence, addition of second electron takes place with the absorption of energy, Thus, the second ionization enthalpy of oxygen is positive.

Q.18. First member of each group of representative elements shows anomalous behaviour. Why?

Ans: First member of each group of representative elements shows anomalous behaviour due to the following reasons:

(i) small size (ii) high ionization enthalpy (iii) high electronegativity

(iv) absence of d- orbitals

Q.19. As we move along the period, the metallic character decreases while the non- metallic character increases. Why?

Ans: As we move along the period, the number of valence electrons increases by one at each succeeding element but the number of shells remain the same. Due to this, effective nuclear charge increases.

More the effective nuclear charge more is the attraction between nuclei and electron. Hence, tendency of the element to lose electrons decreases. Thus, metallic character decreases.

Tendency of an element to gain electrons increases with increase in effective nuclear charge, so non-metallic character increases along the period.

Q.20. Why does N have very low electron gain enthalpy?

Ans: Nitrogen has a symmetrical stable half filled configuration. Therefore, it has less urge to accept an electron. Hence, N has very low electron gain enthalpy.

Q.21. Along the period, the atomic size generally decreases with increase in atomic number but at the end of each period, it increases abruptly. Give reason.

Ans: Along the period, generally the atomic size decreases because the electrons are added to the same shell which do not screen each other from the increased nuclear charge. Thus, the effective nuclear charge increases and the outer electrons are attracted by the nucleus more strongly which results in decreased radii.

There is abrupt increase in the atomic size at the end of each period i.e. in case of noble gases because they involve vander Waal's radius while rest of the elements have covalent radius.

Q.22. Why is Vander Waal's radius of an element always larger than the covalent radius?

Ans: The Vander Waal's forces existing between atoms in the solid state are weak and the atoms are held at larger distances. Therefore, the internuclear distances in case of atoms held by vander Waals forces are larger than that between covalently bonded atoms. Consequently, vander Waals radius is always larger than the covalent radius.

Q.23. Li and Mg resembles in their properties although they do not belong to the same group. Why?

Ans: Li and Mg resembles in their properties due to diagonal relationship. The cause of diagonal relationship is due to opposing trends in periodic properties along a period and down the group. On moving diagonally, the two different trends mutually cancel and therefore, the elements are expected to have similar characteristics.

Q.24. Why is it not possible to determine the atomic radius of an element experimentally?

Ans: (i) It is not possible to measure the exact distance since an atom does not have a well defined boundary and the probability of an electron is never zero even at larger distance from the nucleus.

(ii) It is not possible to isolate a single atom.

(iii) The atomic radius also changes from one bonded state to the other.

Q.25. In isoelectronic ions, the ionic radii decrease with increase in atomic number. Give reason.

Ans: Isoelectronic species have same no. of electrons but different atomic number. The species with increased atomic number will have greater strength of nuclear charge for the same no. of electrons. Thus, the strength of attractive forces increases and the ionic radii decreases.

Q.26. Ionisation enthalpy decreases in a group down the group. Explain.

Ans: The nuclear charge increases due to addition of electrons but since the electrons are added to the new energy shell, the distance between the valence shell electrons and the nucleus increases. At the same time, the shielding effect on the outermost electrons increases due to increase in the number of inner electrons.

The increased shielding effect overweighs the increasing nuclear charge and thus, a less amount of energy is required to remove the outermost electrons. Hence, ionization enthalpy decreases down the group.

Q.27. Electronegativity of elements increases on moving from left to right in a periodic table. Assign reason.

Ans: Electronegativity of elements increases on moving along the period from left to right in a periodic table. This is due to the decrease in atomic size and increase in effective nuclear charge. As a result, the attraction between the outer electrons and the nucleus increases. Ultimately, electronegativity increases.

CHAPTER – 7 EQUILIBRIUM

Q.1. Why pure liquids and solids can be ignored while writing the equilibrium constant expression?

Ans: Molar concentration of pure solid and liquid (if in excess) is constant. That's why pure liquids and solids can be ignored while writing the equilibrium constant expression.

Q.2. Ice melts slowly at higher altitudes. Why?

Ans: The melting of ice is favoured at high pressure because there is decrease in volume in the forward reaction. Since, at high altitudes, atmospheric pressure is low and therefore, melting of ice slows down.

Q.3. Why solution of sugar in water does not conduct electricity whereas that of common salt in water does?

Ans: Common salt is an electrolyte which in aqueous solution gives Na^+ and Cl^- ions. Hence, it conducts electricity. Sugar is a non-electrolyte which does not dissociate to give ions in solution. Hence, it does not conduct electricity.

Q.4. Why is ammonia termed as a base though it does not contain OH^- ions?

Ans: The basic nature of ammonia is due to its tendency to donate the electron pair. Therefore, it behaves as Lewis base.

Q.5. BF_3 does not have proton but still acts as an acid. Why?

Ans: BF_3 is an electron deficient compound and hence, acts as Lewis acid by accepting a pair of electrons from electron rich species.

Q.6. How does common ion affect the solubility of electrolyte?

Ans: Solubility of an electrolyte decreases due to common ion effect.

Q.7. Will AgCl be more soluble in aqueous solution or in NaCl solution and why?

Ans: In NaCl solution, the Cl^- ions will increase. Since, the solubility product, $K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-]$ remains constant, $[\text{Ag}^+]$ will decrease. Therefore, the solubility of AgCl will be less in NaCl solution than in water.

Q.8. Why the pH of blood remains almost constant though we quite often eat spicy food?

Ans: Blood is a buffer containing H_2CO_3 and HCO_3^- . Small amounts of an acid or base do not disturb the pH.

Q.9. Would you expect the equilibrium constant for the reaction $\text{I}_2 (\text{g}) \rightleftharpoons 2\text{I} (\text{g})$ to increase or decrease as temperature increases? Assign reason.

Ans: In the forward reaction, energy is needed to bring about the dissociation of I_2 molecules which is endothermic reaction. Increase in temperature favours endothermic reaction i.e. the forward reaction. Therefore, the equilibrium constant will increase with increase in temperature.

Q.10. Why is there a fizz when a soda water bottle is opened?

Ans: Soda water bottle contains $\text{CO}_2 (\text{g})$ dissolved in it under pressure. When the bottle is opened, the pressure of the gas inside the bottle decreases. As a result, the solubility of the gas decreases and escapes from the bottle producing fizz.

Q.11. Acetic acid is highly soluble in water but still a weak electrolyte. Why?

Ans: The strength of the electrolyte does not depend upon its amount present in the solution but on its ionization in solution. Since, acetic acid is a weak acid, it is ionized only to a small extent. Therefore, it is a weak electrolyte.

Q.12. The solubility of CO_2 in water decreases with increase in temperature. Why?

Ans: The dissolution of CO_2 gas in water is exothermic because the gas contracts in volume. According to Le-Chatelier's principle, increase in temperature favours endothermic reaction which is the reverse reaction. Therefore, the solubility of CO_2 in water will decrease with increase in temperature.

Q.13. Will the ionic product of water increase or decrease if temperature is increased?

Ans: The ionic product of water will increase upon increasing the temperature because more of the ions will be formed.

Q.14. An aqueous solution of CuSO_4 is acidic while that of Na_2SO_4 is neutral. Why?

Ans: CuSO_4 is a salt of weak base and strong acid. Upon hydrolysis, it will form a weak base and a strong acid. Thus, the solution contains free H^+ ions exhibiting acidic character.

Na_2SO_4 is a salt of strong acid and strong base. The solution will have same no. of H^+ and OH^- ions exhibiting neutral behaviour.

Q.15. Water is rather neutral but becomes a strong base when HCl is dissolved in water. Explain.

Ans: The basic character of water is explained on the basis of Bronsted – Lowry concept. HCl is a strong acid which readily loses H^+ ions. All the ions are accepted by H_2O molecules to form H_3O^+ ions. Thus, H_2O is a strong base in contact with strong acid.

Q.16. Why do we add dilute HCl in group II of mixture analysis?

Ans: The cations of group II are precipitated as sulphides by passing H_2S gas through the aqueous mixture solution. But before passing the gas, the solution must be acidified with dilute HCl so as to suppress the concentration of S^{2-} ions in solution due to common ion effect.



(weak electrolyte)



This is done in order to avoid the precipitation of any cation of group IV if present in the mixture as its sulphide.

Q.17. The salt of strong acid and strong base does not undergo hydrolysis. Comment.

Ans: In this case, neither the cation nor the anion undergoes hydrolysis. Therefore, the solution remains neutral. For eg. In the aqueous solution of NaCl, its ions Na^+ and Cl^- have no tendency to react with the H^+ or OH^- ions of water. This is because the possible products of such reaction are NaOH and HCl, which are themselves completely dissociated. As a result, there is no change in the concentration of H^+ and OH^- ions.

CHAPTER- HYDROGEN

Q.1. Why does hydrogen occur in diatomic form rather than in monoatomic form under normal conditions?

Ans: In monoatomic form, hydrogen atom has only one electron in K- shell ($1s^1$) while in diatomic form, K- shell is complete ($1s^2$) acquiring the stable noble gas configuration of Helium.

Q.2. Among NH_3 , H_2O and HF which would you expect to have highest magnitude of hydrogen bonding and why?

Ans: The strength of Hydrogen bonding depend on the magnitude of polarity of the bond. Since, the electronegativity of F is the highest, H – F would be maximum polar showing higher magnitude of hydrogen bonding.

Q.3. Hydrogen peroxide can act both as an oxidizing as well as reducing agent. Explain.



Since, oxygen atom in H_2O_2 can undergo an increase as well as decrease in oxidation number. Therefore, it can act both as reducing as well as oxidising agent.

Q.4. How does H_2O_2 behave as a bleaching agent?

Ans:

EASY AND SCORING AREAS

OXIDATION NUMBER, REACTIVITY

1. Define oxidation and reduction in terms of oxidation number.
Ans Increase in Oxidation Number is Oxidation and decrease in Oxidation Number is called reduction.
2. What is meant by disproportionation? Give one example.
Ans : In a disproportionation reaction an element simultaneously oxidized and reduced.
$$P_4 + 3OH^- + 3H_2O \rightarrow PH_3 + 3H_2PO_2^-$$
3. What is O.N. of sulphur in H_2SO_4 ? Ans: +6
4. Identify the central atom in the following and predict their O.S.
 HNO_3
Ans: central atom:- N; O.S. +5
5. Out of Zn and Cu which is more reactive?
Ans: Zn.
6. What is galvanization?
Ans: Coating of a less reactive metal with a more reactive metal e.g.- coating of iron surface with Zn to prevent rusting of iron.
7. How is standard cell potential calculated using standard electrode potential?
Ans: $E^0_{cell} = E^0_{cathode} - E^0_{anode}$
8. What is O.S. of oxygen in H_2O_2 ?
Ans: - -1.
9. The formation of sodium chloride from gaseous sodium and gaseous chloride is a redox process justify.
Ans: Na atom get oxidize and Cl is reduced.
10. What is SHE? What is its use?

Ans :Standard Hydrogen Electrode (SHE) has been selected to have zero standard potential at all temperatures. It consists of a platinum foil coated with platinum black (finely divided platinum) dipping partially into an aqueous solution in which the activity (approximate concentration 1M) of hydrogen ion is unity and hydrogen gas is bubbled through the solution at 1 bar pressure. The potential of the other half cell is measured by constructing a cell in which reference electrode is standard hydrogen electrode. The potential of the other half cell is equal to the potential of the cell.