4 Chemical Bonding and Molecular Structure

LESSON AT A GLANCE

- **Chemical Bond:** The force that holds different atoms in a molecule is called chemical bond.
- **Octet Rule:** Atoms of different elements take part in chemical combination in order to complete their octet or to attain the noble gas configuration.
- **Valence Electrons:** It is the outermost shell electron which takes part in chemical combination.
- **Ionic or Electrovalent Bond:** Ionic or Electrovalent bond is formed by the complete transfer of electrons from one atom to another atoms. Generally it is formed between metals and non-metals. We can say that it is the electrostatic force of attraction which holds the oppositely charged ions together.
- **Characteristics of Ionic Compound**

  (*i*) **Physical state:** They generally exist as crystalline solids, known as crystal lattice. Ionic compounds do not exist as single molecules like other gaseous molecules like, H_2, N_2, O_2 Cl_2 etc.

  (*ii*) **Melting and boiling points:** Since ionic compounds contains high interionic force between them they generally have high melting and boiling points.

  (*iii*) **Solubility:** They are soluble in polar solvents such as water but do not dissolve in organic solvents like benzene, CCl_4 etc.

  (*iv*) **Electrical conductivity:** In solid state they are poor conductor of electricity but in molten state or when dissolved in water they conduct electricity.

(*v*) **Ionic reactions:** Ionic compounds produce ions in the solution which gives very fast reaction with oppositely charged ions.

For example,

$$Na^+Cl^-(aq) + Ag^+NO_3^-(aq) \longrightarrow AgCl^-(s) + NaNO_3(aq)$$

- **Formal Charge:** In polyatomic ions, the net charge is the charge on the ion as a whole and not by particular atom. However charges can be assigned to individual atoms or ions. These are called formal charges.
- **Bond Length:** It is defined as the equilibrium distance between the centres of the nuclei of the two bonded atoms. It expressed in terms of A°. Experimentally it can be defined by X-ray diffraction or electron diffraction method.
- **Bond Angle:** It is defined as the angle between the lines representing the orbitals containing the bonding electrons.
- **Bond Enthalpy:** It is defined as the amount of energy required to break one mole of bonds of a particular type to separate them into gaseous atoms.

  Bond Enthalpy is also known as bond dissociation enthalpy or simple bond enthalpy unit of bond enthalpy = kJ mol^{-1}.

  Greater the bond enthalpy stronger is the bond. For *e.g.*, the H—H bond enthalpy in hydrogen is 435.8 kJ mol^{-1}.
- **Valence Bond Theory:** Valence bond theory was introduced by Heilter and London (1927) and developed by Pauling and others. It is based on the concept of atomic orbitals and the electronic configuration of the atoms.
- **Hybridisation:** Hybridisation is the process of intermixing of the orbitals of slightly different energies so as to redistribute their energies resulting in the formation of new set of orbitals of equivalent energies and shape.
- **Hydrogen Bonding:** When highly electronegative elements like nitrogen, oxygen, flourine are attached to hydrogen to form covalent bond, the electrorns of the covalent bond are shifted towards the more electronegative atom. Thus partial positive charge develop on hydrogen atom which forms a bond with the other electronegative atom. This bond is known as hydrogen bond and is weaker than the covalent bond. For example, in HF molecule hydrogen bond exists between hydrogen atom of one molecule and fluorine atom of another atom.

TEXTBOOK QUESTIONS SOLVED

Q1. *Explain the formation of a chemical bond.*

Ans. According to Kossel and Lewis, atoms combine together in order to complete their respective octets so as to acquire the stable inert gas configuration. This can occur in two ways by transfer of one or more electrons from one atom to other or by sharing of electrons between two or more atoms.

Q2. *Write Lewis dot symbols for atoms of the following elements: Mg, Na, B, O, N, Br.*

Ans. $_{12}Mg$ = 2, 8, 2 ∴ Lewis symbol = $\dot{Mg}\cdot$

$_{11}Na$ = 2, 8, 1 ∴ Lewis symbol = $\dot{Na}$

$_{5}B$ = 2, 3 ∴ Lewis symbol = $\cdot\dot{B}\cdot$

$_{8}O$ = 2, 6 ∴ Lewis symbol = $:\ddot{O}:$

$_{7}N$ = 2, 5 ∴ Lewis symbol = $:\dot{\underset{\cdot}{N}}\cdot$

$_{35}Br$ = 2, 8, 18, 7 ∴ Lewis symbol = $:\ddot{\underset{\cdot\cdot}{Br}}\cdot$

Q3. *Write Lewis symbols for the following atoms and ions: S and S^{2-}; Al and Al^{3+}; H and H^-*

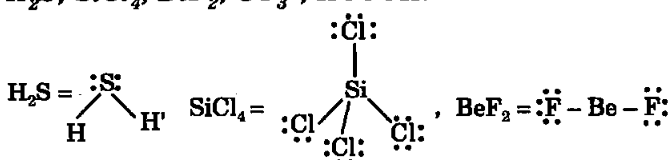
Ans. $_{16}S = 2, 8, 6$ ∴ Lewis symbol = $:\ddot{S}:$, S^{2-} ions = $[:\ddot{\underset{\cdot\cdot}{S}}:]^{2-}$

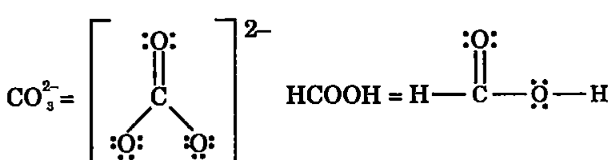
$_{13}Al = 2, 8, 3$ ∴ Lewis symbol = $\cdot\dot{Al}\cdot$, Al^{3+} ion = $[Al]^{3+}$

$_{1}H = 1$ ∴ Lewis symbol = $H\cdot$, H^- ion = $\dot{H}\cdot$

Q4. *Draw the Lewis structures for the following molecules and ions: H_2S, $SiCl_4$, BeF_2, CO_3^{2-}, HCOOH.*

Ans.





Q5. *Define octet rule. Write its significance and limitations.*

Ans. **Octet rule:** Atoms of elements combine with each other in order to complete their respective octets so as to acquire the stable gas configuration.

Significance: It helps to explain why different atoms combine with each other to form ionic compounds or covalent compounds.

Limitations of Octet rule:

(*i*) According to Octet rule, atoms take part in chemical combination to achieve the configuration of nearest noble gas elements. However, some of noble gas elements like Xenon have formed compounds with fluorine and oxygen. For example: $X_e F_2$, $X_e F_4$ etc.
Therefore, validity of the octet rule has been challenged.

(*ii*) This theory does not account for shape of molecules,

Q6. *Write the favourable factors for the formation of ionic bond.*

Ans. (*i*) Low ionisation enthalpy of the element.

(*ii*) High electron gain enthalpy of the element involved in is forming an anion, *i.e.*, a non-metal.

(*iii*) The difference in the electronegativity between the two elements forming cations and anions should be large.

Q7. *Discuss the shape of the following molecules using the VSEPR model:*

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

Ans. (*i*) $BeCl_3$ = Cl:Be:Cl. The central atom has only two bond pairs and no lone pair, *i.e.*, it is of the type AB_2. Hence, shape is **linear.**

(*ii*) BCl_3 = Cl $\ddot{:B:}$ Cl. The central atom has only 3 bond pairs and no lone pair, *i.e.*, it is of the type AB_3. Hence, shape is **triangular planar**

(*iii*) $SiCl_4$ = Cl $\overset{Cl}{\underset{Cl}{:\ddot{\underset{..}{Si}}:}}$ Cl. Bond pairs = 4, lone pairs = 0, *i.e.*, it is of the type AB_4.

Shape = **Tetrahedral.**

(*iv*) AsF_5 = F $\overset{F\;\;F}{:\underset{..}{As}:}$ F, with F below. Bond pairs = 5, lone pairs = 0, *i.e.*, it is of the type AB_5. Shape
= **Trigonal bipyamidal.**

(*v*) H_2S = H : S̤̈ : H. Bond pairs = 2, lone pairs = 2, *i.e.*, it is of the type AB_2L_2.
Shape = **Bent/V-shaped.**

(*vi*) PH_3 = H : P̤̈ : H. Bond pairs = 3, lone pair = 1, *i.e.*, it is of the type AB_3L.
Shape = **Trigonal.**

Q8. *Although geometries of NH_3 and H_2O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.*

Ans. In N̈H_3

H_2Ö:

Because of two low pairs of electrons on O-atom repulsion on bond pairs one greater in H_2O in comparison to NH_3. Thus the bond angle is less is H_2O molecules.

Q9. *How do you express the bond strength in terms of bond order?*

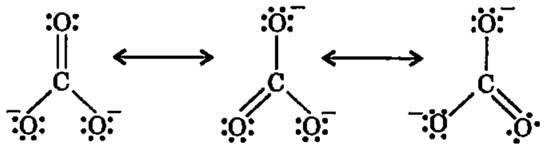
Ans. Bond strength is directly proportional to the bond order. Greater the bond order more in the bond strength.

Q10. *Define the bond length.*

Ans. **Bond Length:** It is the equilibrium distance between the nuclei of two bonded atoms in a molecule. Bond Lengths are measured by spectroscopic methods.

Q11. *Explain the important aspects of resonance with reference to the CO_3^{2-} ion.*

Ans.



I II III

Resonance in CO_3^{2-}, I, II and III represent the three canonical forms

(*i*) In these structures, the position of nuclei are same.

(*ii*) All the three forms have almost equal energy.

(*iii*) Same number of paired and unpaired electrons they differ only in their position.

Q12. *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.*

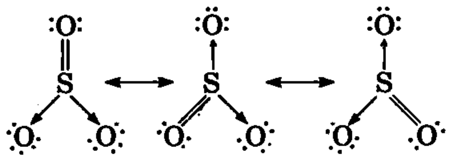
(1) (2)

Ans. These two structures do not represent the resonance structures because the position of atoms is different in these two structures.

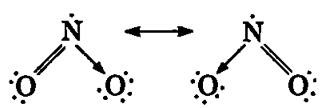
Q13. *Write the resonance structures for SO_3, NO_2 and NO_3^-.*

Ans. Resonance structures are as under:

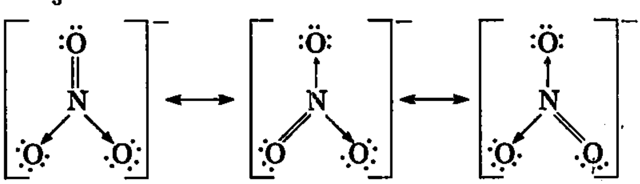
(*i*) SO_3



(*ii*) NO_2



(*iii*) NO_3^-



Q14. *Use Lewis symbols to show electron transfer between the following atoms to form cations and anions: (a) K and S (b) Ca and O (c) Al and N.*

Ans. (*a*) K and S combine as K_2S with K^+ as the cation and S^{2-} as the anion. S atom is short of two electrons for its octet to complete, hence it accepts two electrons from two K atoms. K atom after losing one electron becomes K^+ and attains noble gas electronic configuration. So for K and S atoms, we have

$$K^{\bullet} \xrightarrow{-e^-} K^+$$

$$:\ddot{\underset{\cdot\cdot}{S}} \xrightarrow{+2e^-} \left(:\ddot{\underset{\cdot\cdot}{S}}:\right)^{2-} \text{ and } 2K^+ + S^{2-} \longrightarrow (K^+)_2S^{2-}$$

(*b*) In the similar way for CaO formation

$$Ca^{:} + \ddot{\underset{\cdot\cdot}{O}}: \longrightarrow \overset{2+}{Ca} + :\ddot{\underset{\cdot\cdot}{O}}:^{2-} \longrightarrow \overset{2+\,2-}{CaO}$$

(*c*) $\dot{\underset{\cdot}{Al}}\cdot + \dot{\underset{\cdot\cdot}{N}}: \longrightarrow Al^{3+} + \left(:\ddot{\underset{\cdot\cdot}{N}}:\right)^{3-} \longrightarrow Al^{3+}N^{3-}$

Q15. *Although both CO_2 and H_2O are triatomic molecules, the shape of H_2O molecule is bent while that of CO_2 is linear. Explain this on the basis of dipole moment.*

Ans. In CO_2 there are two C=O bonds. Each C=O bond is a polar bond.

The net dipole moment of CO_2 molecule is zero. This is possible only if CO_2 is a linear molecule. (O=C=O). The bond dipoles of two C=O bond cancel the moment with each other.

Whereas H_2O molecule has a net dipole moment (1.84 D) H_2O molecule has a bent structure because here the O—H bonds are oriented at an angle of 104.5° and do not cancel the bond moments of each other.

Q16. *Write the significance / applications of dipole moment.*

Ans. (*i*) **In predicting the nature of the molecules:** Molecules with specific dipole moments are polar in nature and those of zero dipole moments are non-polar in nature.

(*ii*) In the determination of shapes of molecules.

(*iii*) In calculating the percentage ionic character.

Q17. *Define electronegativity. How does it differ from electron gain enthalpy?*

Ans. **Electronegativity:** Electronegativity is the tendency of an atom to attract shared pair of electrons. It is the property of bonded atom.

Whereas electron gain enthalpy is the tendency of an atom to attract outside electron. It is the property of an isolated atom.

Q18. *Explain with the help of suitable example polar covalent bond.*

Ans. When two atoms with different electronegativity are linked to each other by covalent bond, the shared electron pair will not be in the centre because of the difference in electronegativity. For example, in hydrogen flouride molecule, flouride has greater electronegativity than hydrogen. Thus the shared electron pair is displaced more towards flourine atom, the later will acquire a partial negative charge (d^-). At the same time hydrogen atom will have a partial positive charge (d^+). Such a covalent bond is known as polar covalent bond or simply polar bond.

It is represented as

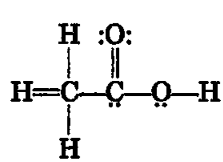
$$\overset{\delta^+}{\underset{(2.1)}{\text{H}}} \; — \; \overset{\delta^-}{\underset{(4.0)}{\text{F}}}$$

Q19. *Arrange the bonds in order of increasing ionic character in the molecules: LiF, K_2O, N_2, SO_2 and ClF_3.*

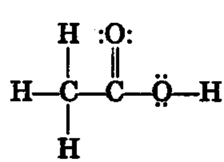
Ans. On the basis of the difference in electronegativity values, the order is

$$LiF > K_2O > ClF_3 > SO_2 > N_2$$

Q20. *The skeletal structure of CH_3COOH 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



Q21. *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. According to VSEPR theory, if CH_4 where square planar, the bond angle would be 90°. For tetrahedral structure, the bond angle is 109°28[1]. Therefore, in square planar structure, repulsion between bond pairs would be more and thus the stability will be less.

Q22. *Explain why BeH_2 molecule has a zero dipole moment although the Be—H bonds are polar.*

Ans. BeH_2 molecule is linear H—Be—H, the bond moments due to Be—H bonds are acting opposite to each other, so cancel their effect. Therefore, the BeH_2 has zero dipole moment.

Q23. *Which out of NH_3 and NF_3 has higher dipole moment and why?*

Ans. In NH_3 and NF_3, the difference in electronegativity is nearly same but the dipole moment of NH_3 = (1.46D)

For example, NF_3 = (0.24D)

In NH_3 the dipole moments of the three N—H bonds are in the same direction lone pair of electron. But in NF_3, the dipole moments of the three N—F bonds are in the direction opposite to that of the lone pair. Therefore, the resultant dipole moment in NH_3 is more than its NF_3.

Q24. *What is meant by hybridisation of atomic orbitals? Describe the shapes of sp, sp^2, sp^3 hybrid orbtials.*

Ans. **Hybridisation:** It is defined as the process of intermixing of atomic orbitals of slightly different energies to give rise to new hybridized orbitals having equivalent energy and identical shapes.

Shapes of Orbitals:

sp hybridisation: Ones and one p orbital. Intermix is called sp-hybridisation. For example, in BeF_2, Be atom undergoes sp-hybridisation. It has linear shape. Bond angle is 180°.

sp^2 hybridisation: One *s* and two *p*—orbitals get hybridised to form three equivalent hybrid orbitals. The three hybrid orbitals directed towards three corners of an equilateral triangle. It is, therefore, known as trigonal hybridisation.

sp^3 hybridisation: One S and three P^- orbitals get hybridised to form four equiva-lent hybrid orbitals. These orbitals are directed towards the four corners of a regular tetrahedron.

Q25. *Describe the change in hybridisation (if any) of the Al atom in the following reaction.*

$$AlCl_3 + Cl^- \longrightarrow AlCl_4^-$$

Ans. On the basis of the electronic configuration of $_{13}$Al atom, its valence shell configuration is $3s^2 3p^1$ in the ground state which changes to $3s^1\ 3p_x^1\ 3p_y^1$ in the excited state. This leads to sp^2 type of hybridisation with one $2p_z$ orbital empty. This $2p_z$ orbital is used to accommodate a lone pair of electrons from the Cl^- ion. This ways the sp^2 hybridisation also changes to sp^3, *i.e.*,

$$\underset{sp^2}{AlCl_3} + :\ddot{\underset{..}{Cl}}:^{(-)} \longrightarrow \underset{sp^3}{[AlCl_4]^-}$$

Q26. *Is there any change in the hybridisation of B and N atoms as a result of the following reaction?*

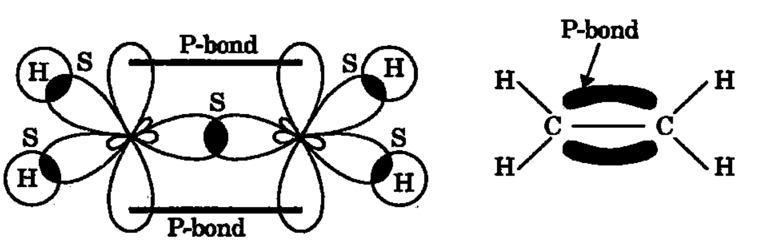
$$BF_3 + NH_3 \longrightarrow F_3B.NH_3$$

Ans. B in BF_3 is sp^2 hybridised with an empty $2p_z$ orbital. $:NH_3$ molecule donates its lone pair of electrons on N atom into $2p_z$ orbital and the hybridisation changes to sp^3. N in $:NH_3$ has unchanged sp^3 hybridisation:

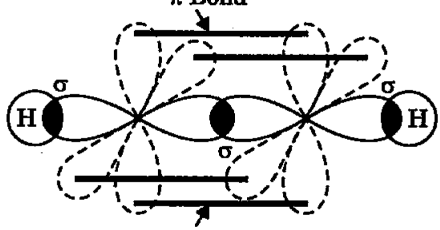
$$\underset{\underset{sp^2}{\uparrow}}{BF_3} + \underset{\underset{sp^3}{\uparrow}}{:NH_3} \longrightarrow \underset{\underset{sp^3}{\uparrow}}{F_3B} \leftarrow \underset{\underset{sp^3}{\uparrow}}{NH_3}$$

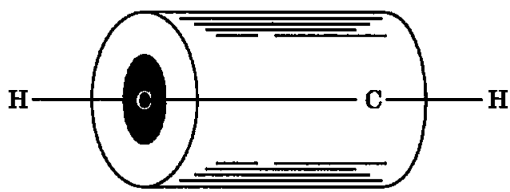
Q27. *Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in C_2H_4 and C_2H_2 molecules.*

Ans.



Orbital picture of ethene molecule





Cylindrical π electron cloud

Q28. *What is the total number of sigma and pi bonds in the following molecules?*

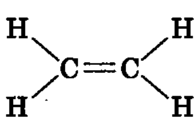
(a) C_2H_2 *(b)* C_2H_4

Ans. (*a*) H—C°C—H

Sigma bond = 3

π bonds = 2

(*b*)



Sigma bond = 5

π bonds = 1

Q29. *Considering x-axis as the internuclear axis, which out of the following will not form a sigma bond and why? (a) 1s and 1s (b) 1s and $2p_x$ (c) $2p_y$ and $2p_y$ (d) 1s and 2s.*

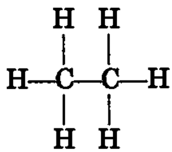
Ans. (*c*). $2p_y$ and $2p_y$ orbitals can only form π bond and not σ when *x*-axis is taken. In this case, only sideways overlapping is possible.

Q30. *Which hybrid orbitals are used by carbon atoms in the following molecules?*

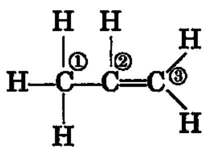
(*a*) $CH_3—CH_3$ (*b*) $CH_3—CH{=}CH_2$

(*c*) $CH_3—CH_2—OH$ (*d*) $CH_3—CHO$

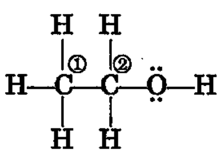
(*e*) CH_3COOH

Ans. (*a*) 

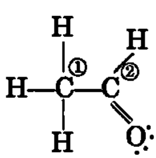
Both C atoms use sp^3 hybrid orbitals.

(*b*) 

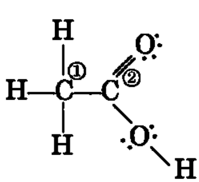
$C_1 = sp^3$; $C_2 = sp^2$; $C_3 = sp^2$

(*c*) 

$C_1 = sp^3$; $C_2 = sp^3$

(*d*) 

$C_1 = sp^3$; $C_2 = sp^2$

(*e*) 

$C_1 = sp^3$; $C_2 = sp^2$

Q31. *What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving one example of each type.*

Ans. When two atoms form a covalent bond by the sharing of valence electrons, the shared pair is called a bond pair or bonding pair. The electron pair or pairs not participating in the bond formation are called lone pairs or non-bonding pairs of electrons. Examples are:

H
(*i*) CH_4; H :C̤̈: H
H

Methane molecule has only four bond pairs of electrons.

(*ii*) NH_3; H :N̤̈:
H

Ammonia has one lone pair and three bond pairs of electrons.

Q32. *Distinguish between a sigma and a pi bond.*

Ans.

Sigma (σ) Bond	*Pi (π) Bond*
(1) σ-bond is formed by the axial overlap of the atomic orbitals.	(1) π-bond is formed by the sidewise overlap of atomic orbitals.
(2) The bond is quite strong.	(2) Comparatively weak bond.
(3) Only one lobe of the p-orbitals is involved in the overlap.	(3) Both lobes of the p-orbitals are involved in the overlap.
(4) Electron cloud of the molecular orbital is symmetrical around the internuclear axis.	(4) The electron cloud is not symmetrical.

Q33. *Explain the formation of H_2 molecule on the basis of valence bond theory.*

Ans. Let us consider the combination between atoms of hydrogen H_A and H_B. If e_A and e_B be their respective electrons.

As they tend to come closer, two different forces operate between the nucleus and the electron of the other and vice versa. The nuclei of the atom as well as their electrons repel each other. Energy is needed to overcome the force of repulsion. Although the number of new attractive and repulsive forces is the same, but the magnitude of the attractive forces is more, Thus when two hydrogen atoms approach each other the overall potential energy of the system decreases. Thus, a stable molecule of hydrogen is formed.

Q34. *Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.*

Ans. Following are the conditions required for the linear combination of atomic orbitals to form molecular orbitals:

(*i*) The combining atomic orbitals should have same symmetry for the effective overlap. For example, $2p_z$

orbital of one atom can combine with $2p_z$ orbital of the other atom when z-axis is taken as the internuclear axis. $2p_y$ or $2p_x$ orbital of one atom cannot combine with $2p_z$ orbital of another.

(*ii*) The combining atomic orbitals should have similar energies. For example, 1*s* orbital of one atom can combine with 1*s* orbital of another but neither 2*s* nor 3*s* for the formation of a homonuclear molecule such as A_2. This combination, however, is possible for the formation of a heteronuclear molecule, AB.

(*iii*) The combining atomic orbitals must overlap to the maximum extent so as to increase the electron density between the nuclei of a molecular orbital.

Q35. *Use molecular orbital theory to explain why the Be_2 molecule does not exist.*

Ans. E.C of Be = $1S^2\ 2S^2$

M.O.E.C of $Be_2 = \sigma^2 1s\ \sigma^{*2} 1s\ \sigma^2 2S\ s^{*2} 2S$

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

Hence Be_2 does not exist.

Q36. *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. O_2— Bond order = 2 paramagnetic

O_2^+— Bond order = 2.5 paramagnetic

O_2— Bond order = 1.5 paramagnetic

O_2^{2-}— Bond order = 1 Diamagnetic

Order of relative Stability is

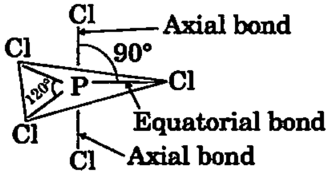
$$\underset{(2.5)}{O_2^+} > \underset{(2.0)}{O_2} > \underset{(1.5)}{O_2^-} > \underset{(1.0)}{O_2^{2-}}$$

Q37. *Write the significance of a plus and a minus sign shown in representing the orbitals.*

Ans. Plus (+) and minus (–) signs represent the signs of wave functions of the orbitals.

Q38. *Describe the hybridisation in case of PCl_5. Why are the axial bonds longer as compared to equatorial bonds?*

Ans. Phosphorus atom (P), in PCl_5 possess sp^3d type of hybridisation and the geometry of the molecule is trigonal bipyramidal. d_{z^2} orbital is used in this hybridisation.



In this molecule, axial bond pairs experience more repulsion from equatorial bond pairs. This way axial bonds become slightly longer and hence slightly weaker than the equatorial bonds, thus, making the molecule more reactive.

Q39. *Define hydrogen bond. Is it weaker or stronger than the van der Waals' forces?*

Ans. When hydrogen is attached with highly electronegative element in a covalent bonding the electrons of the covalent bond are shifted towards the more electronegative atom. Thus partially positively charged hydrogen atom forms a bond with the other more electronegative atom This bond is known as hydrogen bond. Hydrogen bond is stronger than the van der Waals forces

Q40. *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 the half of the difference between the number of electrons present in bonding and antibonding molecular orbitals.

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

$$\text{E.C of } N_2 = 1s^2\, 2s^2\, 2p_x^{\;1} 2p_y^{\;1} 2p_z^{\;1}$$

M.O. Configuration

$$\text{of } N_2 = [\sigma 1s]^2 [\sigma * 1s]^2 [\sigma 2s]^2 [\sigma * 2s]^2 [\pi 2p_x]^2 [\pi 2p_y]^2 [\sigma 2p_z]^2$$

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

$$= \frac{1}{2}[8-2] = 3$$

B.O of O_2

M.O. configuration of O_2

$$= [\sigma 1s]^2 [\sigma * 1s]^2 [\sigma 2s]^2 [\sigma * 2s]^2 [\sigma 2p_z]^2$$

$$B.O. = \frac{1}{2}[N_b - N_a]$$

$$= \frac{1}{2}[8-4] = 2$$

M.O. configuration of O_2^+

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

$$= \frac{1}{2}[8-3] = 2.5$$

M.O. configuration of O_2^-

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

$$= \frac{1}{2}[8-5] = 1.5$$

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