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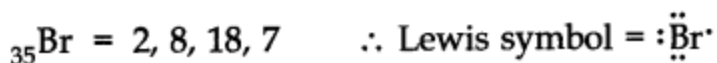
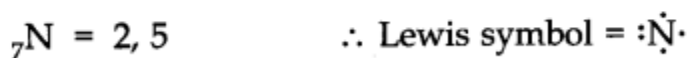
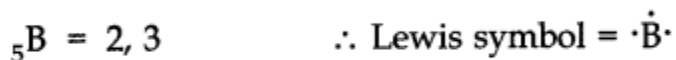
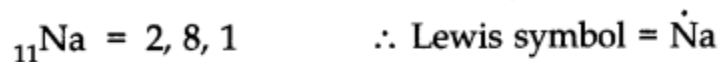
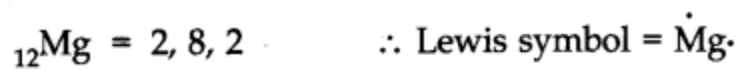
## Chapter 4 Chemical Bonding and Molecular Structure

**Question 1. Explain the formation of a chemical bond.**

**Answer:** 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.

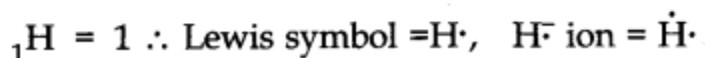
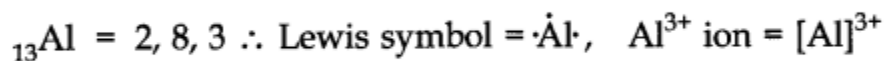
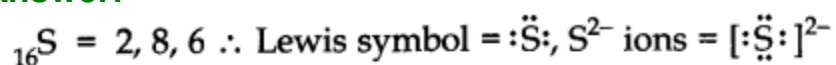
**Question 2. Write Lewis dot symbols for atoms of the following elements: Mg, Na, B, O, N, Br.**

**Answer:**



wis symbols for the following atoms and ions: S and S<sup>2-</sup> ; Al and Al<sup>3+</sup>; H and H<sup>-</sup>

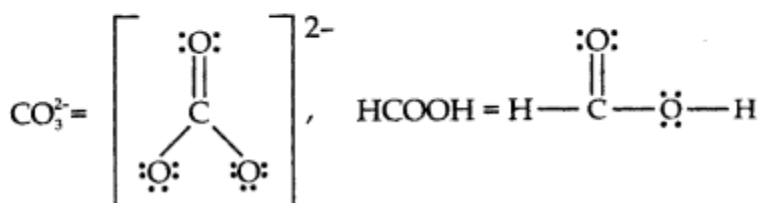
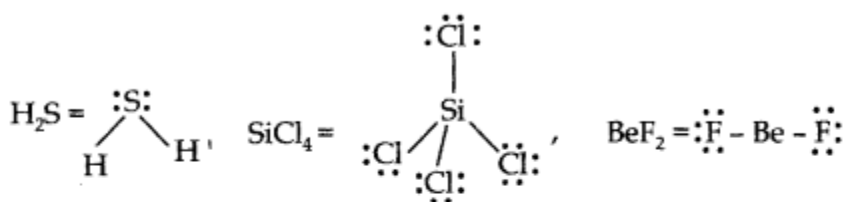
**Answer:**



**Question 4. Draw the Lewis structures for the following molecules and ions:**

**H<sub>2</sub>S, SiCl<sub>4</sub>, BeF<sub>2</sub>, CO<sub>3</sub><sup>2-</sup>, HCOOH**

**Answer:**



**Question 5. Define Octet rule. Write its significance and limitations.**

**Answer: 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:**

- 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: XeF<sub>2</sub>, XeF<sub>4</sub> etc. Therefore, validity of the octet rule has been challenged.
- This theory does not account for shape of molecules.

**Question 6. Write the favourable factors for the formation of ionic bond.**

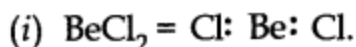
**Answer:**

1. Low ionization enthalpy of metal atoms
2. High electron gain enthalpy of non-metal atoms
3. High lattice enthalpy of compound formed.

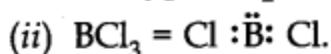
**Question 7. Discuss the shape of the following molecules using the VSEPR model:**

**BeCl<sub>2</sub>, BCl<sub>3</sub>, SiCl<sub>4</sub>, AsF<sub>5</sub>, H<sub>2</sub>S, PH<sub>3</sub>**

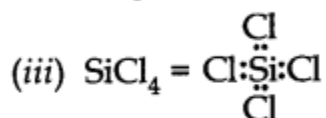
**Answer:**



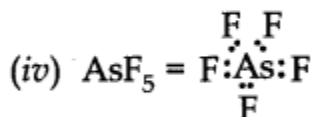
The central atom has only two bond pairs and there is no lone pair, *i.e.*, it is of the type AB<sub>2</sub>. Hence, shape is **linear**.



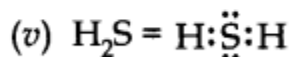
The central atom has only 3 bond pairs and no lone pair, *i.e.*, it is of the type AB<sub>3</sub>. Hence, shape is **triangular planar**.



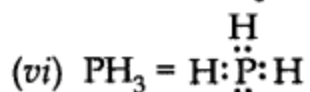
Bond pairs = 4, lone pairs = 0, *i.e.*, it is of the type AB<sub>4</sub>. Hence, shape is **Tetrahedral**.



Bond pairs = 5, lone pairs = 0, *i.e.*, it is of the type AB<sub>5</sub>. Hence, shape is **Trigonal bipyramidal**.



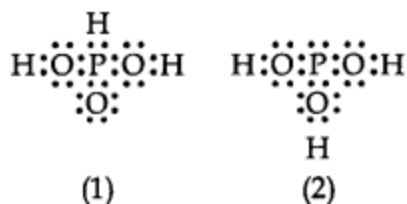
Bond pairs = 2, lone pairs = 2, *i.e.*, it is of the type AB<sub>2</sub>L<sub>2</sub>. Hence, shape is **Bent/V-shaped**.



Bond pairs = 3, lone pair = 1, *i.e.*, it is of the type AB<sub>3</sub>L. Hence, shape is **Trigonal**.

**Question 8. Although geometries of NH<sub>3</sub> and H<sub>2</sub>O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.**

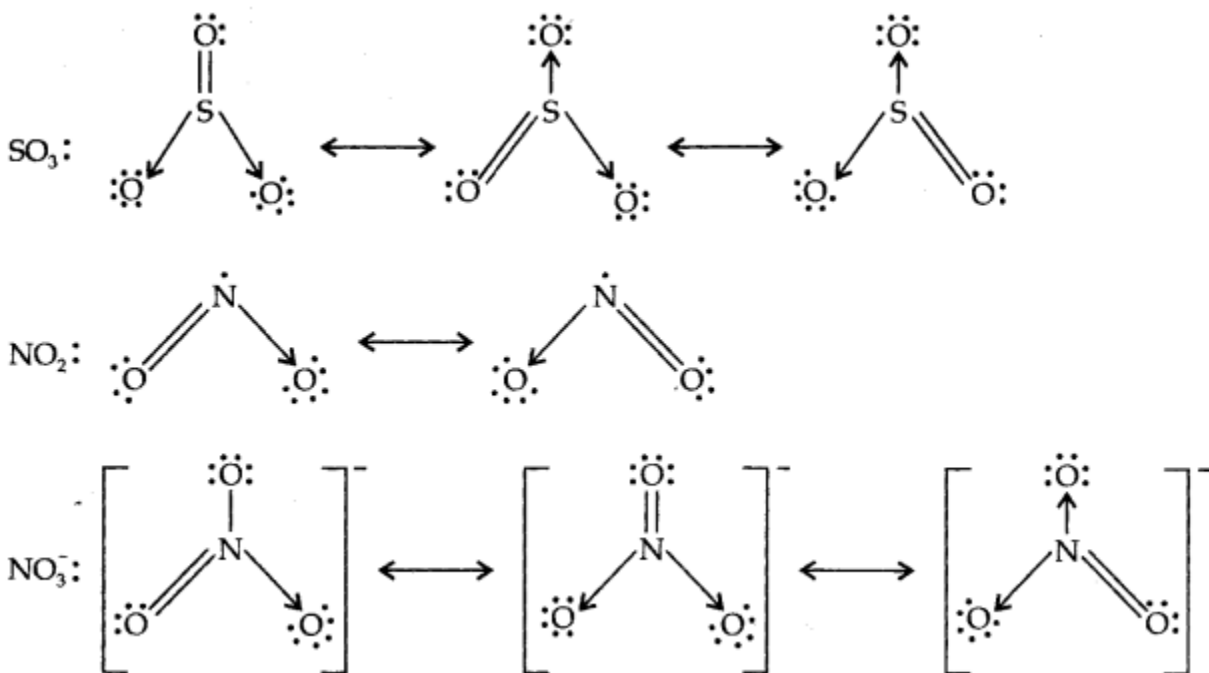




**Answer:** No, these cannot be taken as canonical forms because the positions of atoms have been changed.

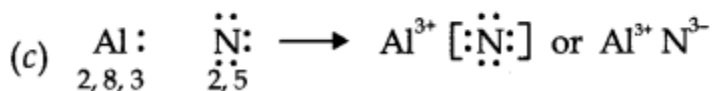
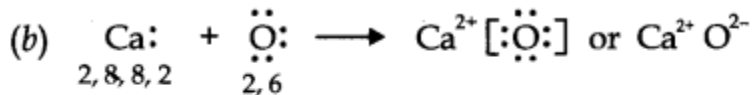
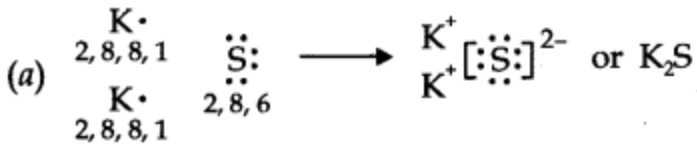
**Question 13. Write the resonance structures for  $\text{SO}_3$ ,  $\text{NO}_2$  and  $\text{NO}_3^-$**

**Answer:**



**Question 14. 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.**

**Answer:**



**Question 15.** Although both  $\text{CO}_2$  and  $\text{H}_2\text{O}$  are triatomic molecules, the shape of  $\text{H}_2\text{O}$  molecule is bent while that of  $\text{CO}_2$  is linear. Explain this on the basis of dipole moment.

**Answer:** In  $\text{CO}_2$ , there are two  $\text{C}=\text{O}$  bonds. Each  $\text{C}=\text{O}$  bond is a polar bond. The net dipole moment of  $\text{CO}_2$  molecule is zero. This is possible only if  $\text{CO}_2$  is a linear molecule. ( $\text{O}=\text{C}=\text{O}$ ). The bond dipoles of two  $\text{C}=\text{O}$  bonds cancel the moment of each other.

Whereas,  $\text{H}_2\text{O}$  molecule has a net dipole moment (1.84 D).  $\text{H}_2\text{O}$  molecule has a bent structure because here the  $\text{O}-\text{H}$  bonds are oriented at an angle of  $104.5^\circ$  and do not cancel the bond moments of each other.

**Question 16.** Write the significance/applications of dipole moment.

**Answer:**

- 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.
- In the determination of shapes of molecules.
- In calculating the percentage ionic character.

**Question 17.** Define electronegativity. How does it differ from electron gain enthalpy?

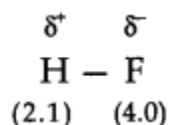
**Answer: Electronegativity:** Electronegativity is the tendency of an atom to attract shared pair of electrons. It is the property of bonded atom.

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

**Question 18.** Explain with the help of suitable example polar covalent bond.

**Answer:** 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 fluoride molecule, fluoride has greater electronegativity than hydrogen. Thus, the shared electron pair is displaced more towards the fluorine atom, the latter will acquire a partial negative charge ( $\delta^-$ ). At the same time hydrogen atom will have a partial positive charge ( $\delta^+$ ). Such a covalent bond is known as polar covalent bond or simply polar bond.

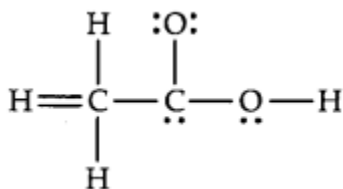
It is represented as



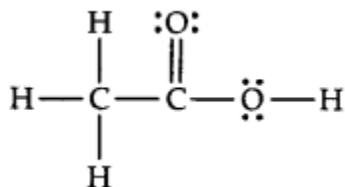
**Question 19.** Arrange the bonds in order of increasing ionic character in the molecules: LiF, K<sub>2</sub>O, N<sub>2</sub>, SO<sub>2</sub> and ClF<sub>3</sub>.

**Answer:** N<sub>2</sub> < SO<sub>2</sub> < ClF<sub>3</sub> < K<sub>2</sub>O < LiF

**Question 20.** The skeletal structure of CH<sub>3</sub>COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.



**Answer:**



**Question 21.** Apart from tetrahedral geometry, another possible geometry for CH<sub>4</sub> is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH<sub>4</sub> is not square planar?

**Answer:** According to VSEPR theory, if CH<sub>4</sub> were square planar, the bond angle would be 90°. For tetrahedral structure, the bond angle is 109°28'.

Therefore, in square planar structure, repulsion between bond pairs would be more and thus the stability will be less.

**Question 22. Explain why  $\text{BeH}_2$  molecule has a zero dipole moment although the  $\text{Be—H}$  bonds are polar.**

**Answer:**  $\text{BeH}_2$  is a linear molecular ( $\text{H—Be—H}$ ), the bond angle =  $180^\circ$ .  $\text{Be—H}$  bonds are polar due to difference in their electronegativity but the bond polarities cancel each other. Thus, molecule has resultant dipole moment of zero.

**Question 23. Which out of  $\text{NH}_3$  and  $\text{NF}_3$  has higher dipole moment and why?**

**Answer:** In  $\text{NH}_3$  and  $\text{NF}_3$ , the difference in electronegativity is nearly same but the dipole moment of  $\text{NH}_3 = (1.46\text{D})$  For Example,  $\text{NH}_3 = (0.24\text{D})$   
In  $\text{NH}_3$ , the dipole moments of the three  $\text{N—H}$  bonds are in the same direction as the lone pair of electron. But in  $\text{NF}_3$ , the dipole moments of the three  $\text{N—F}$  bonds are in the direction opposite to that of the lone pair. Therefore, the resultant dipole moment in  $\text{NH}_3$  is more than in  $\text{NF}_3$ .

**Question 24. What is meant by hybridisation of atomic orbitals? Describe the shapes of  $\text{sp}$ ,  $\text{sp}^2$ ,  $\text{sp}^3$  hybrid orbitals.**

**Answer: 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:**

**$\text{sp}$  hybridisation:** When one s-and one p-orbital, intermix then it is called  $\text{sp}$ -hybridisation. For example, in  $\text{BeF}_2$ , Be atom undergoes  $\text{sp}$ -hybridisation. It has linear shape. Bond angle is  $180^\circ$ .

**$\text{sp}^2$  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.

**$\text{sp}^3$  hybridisation:** One s-and three p-orbitals get hybridised to form four equivalent hybrid orbitals. These orbitals are directed towards the four corners of a regular tetrahedron.

**Question 25. Describe the change in hybridisation (if any) of the Al atom in the following reaction.  $\text{AlCl}_3 + \text{Cl}^- \longrightarrow \text{AlCl}_4^-$ .**

**Answer:** Electronic configuration of  $_{13}\text{Al} = 1s^2 2s^2 2p^6 3s^1 3p_x^1 3p_y^1$   
(excited state)



Hence, hybridisation will be  $sp^2$

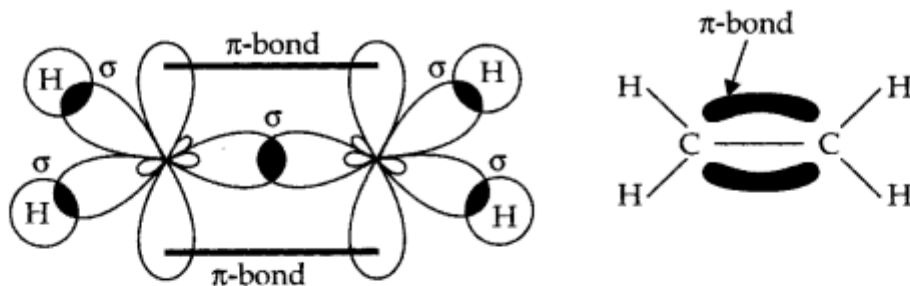
In  $AlCl_4^-$ , the empty  $3p_z$  orbital is also involved. So, the hybridisation is  $sp^3$  and the shape is tetrahedral.

**Question 26.** Is there any change in the hybridisation of B and N atoms as a result of the following reaction ?  $BF_3 + NH_3 \longrightarrow F_3B.NH_3$

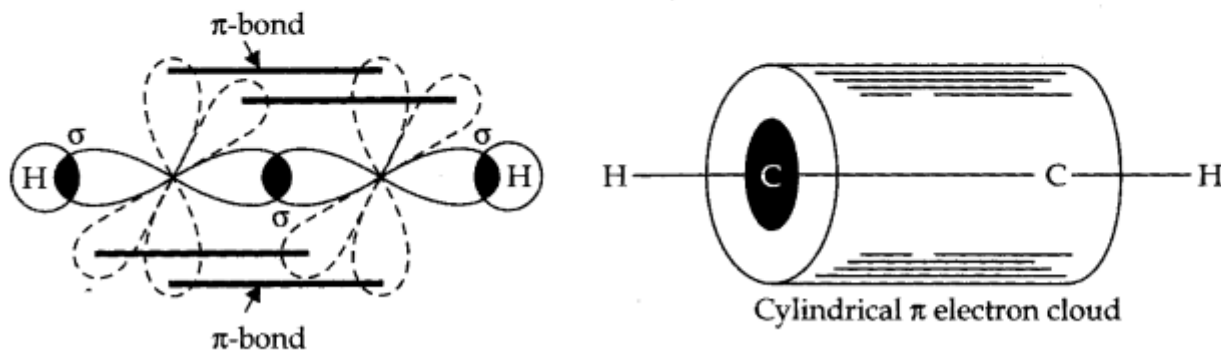
**Answer:** In  $BF_3$ , B atom is  $sp^2$  hybridised. In  $NH_3$ , N is  $sp^3$  hybridised. After the reaction, hybridisation of B changes from  $sp^2$  to  $sp^3$ .

**Question 27.** 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.

**Answer:**



*Orbital picture of ethene molecule*



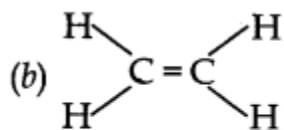
*Orbital picture of ethyne molecule*

**Question 28.** What is the total number of sigma and pi bonds in the following molecules?

(a)  $C_2H_2$  (b)  $C_2H_4$

**Answer:** (a)  $H-C \equiv C-H$

Sigma bond = 3  $\Pi$  bonds = 2



Sigma bond = 5

$\pi$  bonds = 1

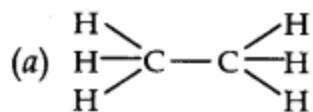
**Question 29.** Considering X-axis as the internuclear axis which out of the following will not form a sigma bond and why? (a)  $1s$  and  $1s$  (b)  $1s$  and  $2p_x$  (c)  $2p_y$  and  $2p_y$  (d)  $1s$  and  $2s$

**Answer:** (c) It will not form a  $\sigma$ -bond because taking x-axis as the internuclear axis, there will be lateral overlap between the two  $2p_y$  orbitals forming a  $\Pi$ -bond.

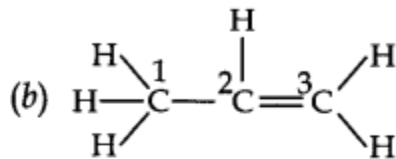
**Question 30.** Which hybrid orbitals are used by carbon atoms in the following molecules?

(a)  $\text{CH}_3\text{-CH}_3$  (b)  $\text{CH}_3\text{-CH}=\text{CH}_2$  (c)  $\text{CH}_3\text{-CH}_2\text{-OH}$  (d)  $\text{CH}_3\text{-CHO}$  (e)  $\text{CH}_3\text{COOH}$ .

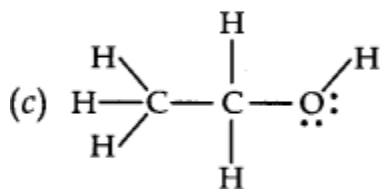
**Answer:**



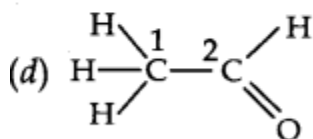
Both C-atoms use  $sp^3$  hybrid orbitals.



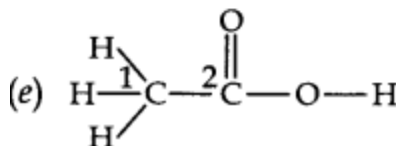
$C_1 = sp^3$ ,  $C_2 = sp^2$ ,  $C_3 = sp^2$



Both C-atoms use  $sp^3$  hybrid orbitals.



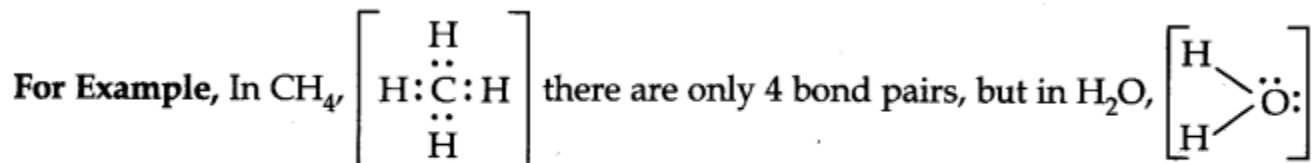
$C_1 = sp^3$ ,  $C_2 = sp^2$



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

**Question 31. What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving one example of each type.**

**Answer:** The electron pair involved in sharing between two atoms during covalent bonding is called shared pair or bond pair. At the same time, the electron pair which is not involved in sharing is called lone pair of electrons.



there are two bond pairs and two lone pairs.

**Question 32. Distinguish between a sigma bond and a pi bond.**

**Answer:**

<i>Sigma (<math>\sigma</math>) Bond</i>	<i>pi (<math>\pi</math>) Bond</i>
(1) $\sigma$ -bond is formed by the axial overlap of the atomic orbitals.	(1) $\pi$ -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.

**Question 33. Explain the formation of  $H_2$  molecule on the basis of valence bond theory.**

**Answer:** Let us consider the combination between atoms of hydrogen  $H_A$  and  $H_B$  and  $e_A$  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

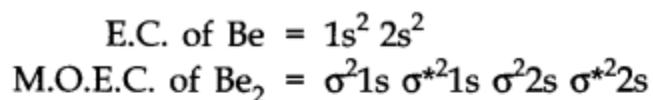
and vice versa. The nuclei of the atoms as well as their electrons repel each other. Energy is needed to overcome the force of repulsion. Although the number of new attractive and repulsive forces is the same, but the magnitude of the attractive forces is more. Thus, when two hydrogen atoms approach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

**Question 34. Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.**

**Answer:**

- The combining atomic orbitals should have comparable energies. For example, 1s orbital of one atom can combine with 1s atomic orbital of another atom, 2s can combine with 2s.
- The combining atomic orbitals must have proper orientations. So that they are able to overlap to a considerable extent.
- The extent of overlapping should be large.
- **Question 35. Use molecular orbital theory to explain why the  $Be_2$  molecule does not exist.**

**Answer:**



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

Hence,  $\text{Be}_2$  does not exist.

- **Question 36. Compare the relative stability of the following species and indicate their magnetic properties:  $\text{O}_2$ ,  $\text{O}_2^-$ ,  $\text{O}_2^-$  (Superoxide),  $\text{O}_2^{2-}$  (peroxide)**

**Answer:**  $\text{O}_2$ — Bond order = 2, paramagnetic

- $\text{O}_2^+$ — Bond order = 2.5, paramagnetic  
 $\text{O}_2^-$ — Bond order = 1.5, paramagnetic  
 $\text{O}_2^{2-}$ — Bond order = 1, diamagnetic  
Order of relative stability is  
 $\text{O}_2^+ > \text{O}_2 > \text{O}_2^- > \text{O}_2^{2-}$   
(2.5) (2.0) (1.5) (1.0)

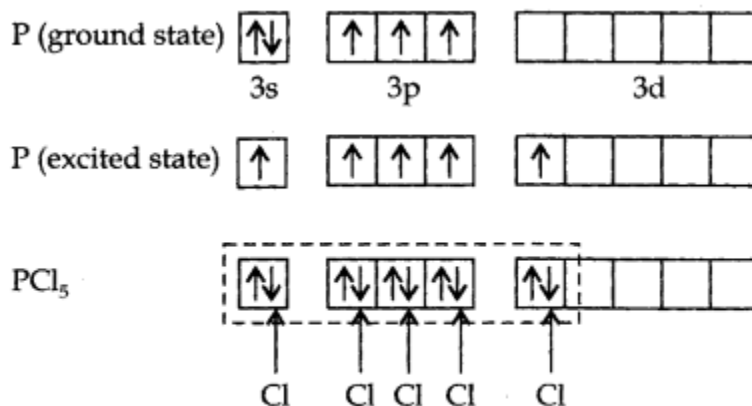
- **Question 37. Write the significance of plus and minus sign in representing the orbitals,**

**Answer:** Plus and minus sign is used to identify the nature of electrons wave. Plus (+ve) sign denotes crest, while (-ve) sign denotes trough.

- **Question 38. Describe the hybridisation in case of  $\text{PCl}_5$ . Why are the axial bonds longer as compared to equatorial bonds?**

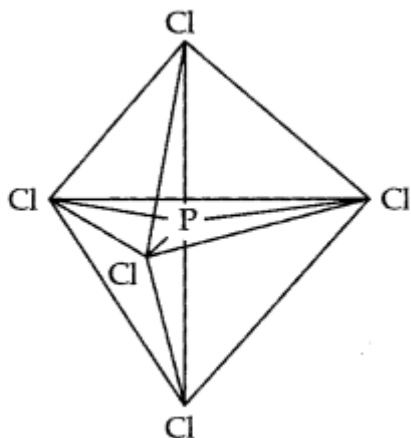
**Answer:** The ground state E.C. and the excited state E.C. of

phosphorus are represented as:



*sp<sup>3</sup> hybrid orbitals filled by electron pairs donated by five Cl atoms*

The one s, three-p and one d-orbitals hybridise to yield five sets of SP<sup>3</sup>d hybrid orbitals which are directed towards the five corners of a trigonal bipyramidal as in Fig.



*Trigonal bipyramidal geometry of PCl<sub>5</sub> molecule*

Because axial bond pairs suffer more repulsive interaction from the equatorial bond pairs, therefore axial bonds have been found to be slightly longer and hence slightly weaker than equatorial bonds.

**Question 39. Define hydrogen bonds. Is it weaker or stronger than the van der Waals forces?**

**Answer:** 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, a partially positively charged hydrogen atom forms a bond with the other more electronegative atom. This bond is known as a hydrogen bond. Hydrogen bond is stronger than the van der Waals forces.

**Question 40. What is meant by the term bond order? Calculate the bond order of  $N_2$ ,  $O_2$ ,  $O_2^+$ ,  $O_2^-$**

**Answer:** Bond order is defined as the half of the difference between the number of electrons present in bonding and antibonding molecular orbitals.

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

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

$$(i) \text{ M.O. configuration of } N_2 = [\sigma 1s]^2 [\sigma^* 1s]^2 [\sigma 2s]^2 [\sigma^* 2s]^2 [\pi 2p_x]^2 [\pi 2p_y]^2 [\sigma 2p_z]^2$$

$$\begin{aligned} \text{Bond order (B.O.)} &= \frac{1}{2}(N_b - N_a) \\ &= \frac{1}{2}[8 - 2] = 3 \end{aligned}$$

$$(ii) \text{ M.O. configuration of } O_2 = [\sigma 1s]^2 [\sigma^* 1s]^2 [\sigma 2s]^2 [\sigma^* 2s]^2 [\sigma 2p_z]^2$$

$$\begin{aligned} \text{B.O} &= \frac{1}{2}[N_b - N_a] \\ &= \frac{1}{2}[8 - 4] = 2 \end{aligned}$$

$$(iii) \text{ M.O. configuration of } O_2^+ = KK[\sigma 2s]^2 [\sigma^* 2s]^2 [\sigma 2p_z]^2 [\pi 2p_x]^2 [\pi 2p_y]^2 [\pi^* 2p_x]^1$$

$$\text{B.O.} = \frac{1}{2}[8 - 3] = 2.5$$

$$(iv) \text{ M.O. configuration of } O_2^-$$

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

$$\text{B.O.} = \frac{1}{2}[8 - 5] = 1.5$$

**Question 39. Define hydrogen bonds. Is it weaker or stronger than the van der Waals forces?**

**Answer:** 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, a partially positively charged hydrogen

atom forms a bond with the other more electronegative atom. This bond is known as a hydrogen bond. Hydrogen bond is stronger than the van der Waals forces.

**Question 40. What is meant by the term bond order? Calculate the bond order of  $N_2$ ,  $O_2$ ,  $O_2^+$ ,  $O_2^-$**

**Answer:** Bond order is defined as the half of the difference between the number of electrons present in bonding and antibonding molecular orbitals.

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

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

$$(i) \text{ M.O. configuration of } N_2 = [\sigma 1s]^2 [\sigma^* 1s]^2 [\sigma 2s]^2 [\sigma^* 2s]^2 [\pi 2p_x]^2 [\pi 2p_y]^2 [\sigma 2p_z]^2$$

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

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

$$(ii) \text{ M.O. configuration of } O_2 = [\sigma 1s]^2 [\sigma^* 1s]^2 [\sigma 2s]^2 [\sigma^* 2s]^2 [\sigma 2p_z]^2$$

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

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

$$(iii) \text{ M.O. configuration of } O_2^+ = KK[\sigma 2s]^2 [\sigma^* 2s]^2 [\sigma 2p_z]^2 [\pi 2p_x]^2 [\pi 2p_y]^2 [\pi^* 2p_x]^1$$

$$\text{B.O.} = \frac{1}{2}[8 - 3] = 2.5$$

$$(iv) \text{ M.O. configuration of } O_2^-$$

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

$$\text{B.O.} = \frac{1}{2}[8 - 5] = 1.5$$