

1. Explain the formation of a chemical bond?

Ans. The atoms of different elements combine with each other in order to complete their respective octets or duplet to attain nearest noble gas configuration.

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

Ans. $_{12}\text{Mg} = 1s^2 2s^2 2p^6 3s^2 \therefore$ Lewis symbol = Mg :

$_{11}\text{Na} = 1s^2 2s^2 2p^6 3s^1 \therefore$ Lewis symbol = $\dot{\text{N}}\text{a}$

$_2\text{B} = 1s^2 2s^2 2p^1 \therefore$ Lewis symbol = $\cdot\dot{\text{B}}\cdot$

$_8\text{O} = 1s^2 2s^2 2p^4 \therefore$ Lewis symbol = $:\ddot{\text{O}}:$

$_7\text{N} = 1s^2 2s^2 2p^3 \therefore$ Lewis symbol = $:\dot{\text{N}}\cdot$

$_{35}\text{Br} = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5 \therefore$ Lewis symbol = $:\ddot{\text{Br}}\cdot$

3. Write Lewis symbols for the following atoms and ions : S and S²⁻ ; Al and Al³⁺ ; H and H⁻

Ans. $_{16}\text{S} = 1s^2 2s^2 2p^6 3s^2 3p^4 \therefore$ Lewis symbol = $:\ddot{\text{S}}:$, S²⁻ ions = $\left[:\ddot{\text{S}}: \right]^{2-}$

$_{13}\text{Al} = 1s^2 2s^2 2p^6 3s^2 3p^1 \therefore$ Lewis symbol = $\cdot\dot{\text{Al}}\cdot$, Al³⁺ ions = $\left[\text{Al} \right]^{3+}$

$_1\text{H} = 1s^1 \therefore$ Lewis symbol = H, H⁻ ions = H.

4. Draw the Lewis structures for the following molecules and ions: H₂S, SiCl₄, BeF₂, CO₃²⁻, HCOOH

Ans. $\text{H}_2\text{S} = \begin{array}{c} \text{:S:} \\ \diagup \quad \diagdown \\ \text{H} \quad \text{H} \end{array}$ $\text{SiCl}_4 = \begin{array}{c} \text{:Cl:} \\ | \\ \text{:Cl:} \text{---} \text{Si} \text{---} \text{Cl} \\ | \\ \text{:Cl:} \end{array}$ $\text{BeF}_2 = \text{:}\ddot{\text{F}}\text{---Be---}\ddot{\text{F}}\text{:}$

$\text{CO}_3^{2-} = \left[\begin{array}{c} \text{:O:} \\ || \\ \text{C} \\ / \quad \backslash \\ \text{:O:} \quad \text{:O:} \end{array} \right]^{2-}$ $\text{HCOOH} = \text{H} \text{---} \text{C} \begin{array}{c} \text{:O:} \\ || \\ \text{---}\ddot{\text{O}}\text{---H} \end{array}$

5. Define octet rule. Write its significance and limitations.

Ans. Octet Rule: Elements combine with each other in order to complete their respective octets i.e. , 8 electrons in their outer most shell.

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

Limitations of octet rule: The octet rule fails to explain

- (i) the formation of molecules in which the central atom has less than eight electrons in the valence shell such as BeCl_2 , BF_3 etc.
- (ii) the formation of molecules in which the central atom has more than eight electrons in the valence shell such as PF_5 , SF_6 etc.
- (iii) the formation of compounds of noble gases especially xenon and krypton such as XeF_2 , XeF_6 etc.
- (iv) odd electron molecules like NO , NO_2 .

6. Write the favorable factors for the formation of ionic bond.

- Ans.**
- (i) Low ionisation enthalpy of the metal atom.
 - (ii) High electron gain enthalpy of the non-metal atom.
 - (iii) High lattice enthalpy of the compound formed.

7. Discuss the shapes of the following molecules using the VSEPR model: BeCl_2 , BCl_3 , SiCl_4 , AsF_5 , H_2S , PH_3

- Ans.**
- (i) $\text{BeCl}_2 = \text{Cl}:\text{Be}:\text{Cl}$. The central atom has only two bond pairs and no lone pair. Hence linear.
 - (ii) $\text{BCl}_3 = \text{Cl}:\overset{\text{Cl}}{\underset{\cdot\cdot}{\text{B}}}:\text{Cl}$. The central atom has only 3 bond pairs and no lone pair. Hence triangular planar.
 - (iii) $\text{SiCl}_4 = \text{Cl}:\overset{\text{Cl}}{\underset{\cdot\cdot}{\underset{\cdot\cdot}{\text{Si}}}}:\text{Cl}$. Bond pairs = 4, lone pair = 0, shape = Tetrahedral
 - (iv) $\text{AsF}_5 = \text{F}:\overset{\text{F}}{\underset{\cdot\cdot}{\underset{\cdot\cdot}{\underset{\cdot\cdot}{\text{As}}}}}\text{F}$. Bond pairs = 5, lone pair = 0, shape = Trigonal bipyramidal
 - (v) $\text{H}_2\text{S} = \text{H}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{S}}}:\text{H}$. Bond pairs = 2, lone pairs = 2. Shape = Bent/V-shaped
 - (vi) $\text{PH}_3 = \text{H}:\overset{\cdot\cdot}{\underset{\cdot\cdot}{\underset{\cdot\cdot}{\text{P}}}}:\text{H}$. Bond pairs = 3, lone pair = 1. Shape = pyramidal

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

- Ans.** In NH_3 , there is only one lone pair on N-atom to repel the bond pairs whereas in H_2O , there are two lone pairs on O-atom to repel the bond pairs. Hence, the repulsions on bond pairs are greater in H_2O than in NH_3 and hence, the bond angle is less.

9. How do you express the bond length in terms of bond order?

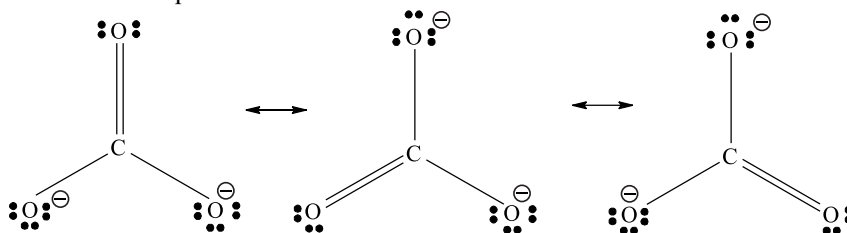
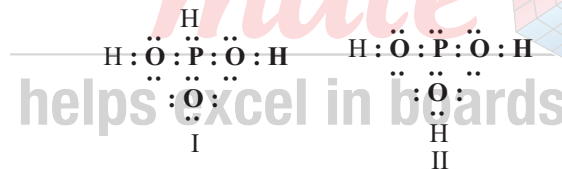
- Ans.** Greater the bond order, shorter is the bond length.

10. Define the bond length.

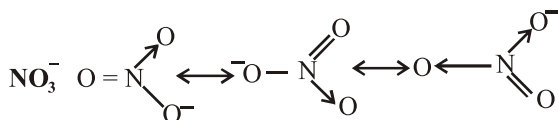
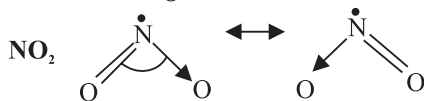
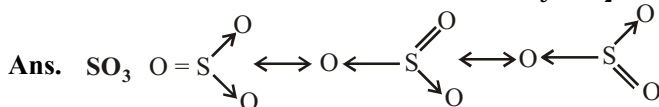
Ans. The equilibrium distance between the centres of nuclei of the two bonded atoms is called its bond length.

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

Ans. CO_3^{2-} has resonance in its structure due to which the bond lengths between C–O are equal.

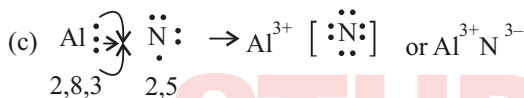
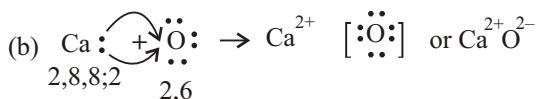
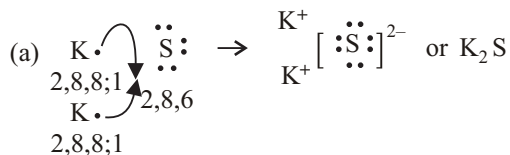
**12. H_3PO_3 can be represented by the structures I and II shown below. Can these two structures be taken as the canonical forms of the resonance hybrid of H_3PO_3 ? If not, give reason for the same**

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

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

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.

Ans.



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

Ans. The dipole moment studies show that net dipole moment of CO₂ molecule is zero. This is possible only if CO₂ is a linear molecule ((O=C=O)) so that dipole moments of C–O are equal and opposite and hence cancel out. On the other hand, H₂O molecule is found to have a net dipole moment (1.84 D) though it contains 2 O–H bonds. This shows that it is a bent molecule.

16. Write the significance/applications of dipole moment.

Ans. Applications of dipole moment are: (a) in determining polarity of bonds, (b) calculation of percentage ionic character, (c) distinguish between cis and trans isomer, (d) distinguish between ortho, meta and para isomers.

17. Define electronegativity. How does it differ from electron affinity?

Ans. Electronegativity of an element is the tendency of its atom to attract the shared pair of electrons towards itself in a covalent bond.

Electronegativity denotes the relative attraction of covalently bonded atoms for the bonding electron pairs while electron affinity is the energy released when an electron is added to a gaseous atom or ion to form a gaseous ion.

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

Ans. When two dissimilar atoms having different electronegativity combine together to form a covalent bond, the shared pair of electrons does not lie at equal distances from the nuclei of both the bonded atoms but shifts towards

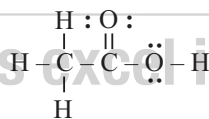
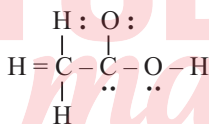
the atom having greater electronegativity. Since the more electronegative atom attracts the electrons more strongly, the distribution of electrons gets distorted. As a result, one end of the molecule, having more electronegative atom becomes slightly negatively charged while the other acquires slightly positive charge. Thus, positive and negative poles are developed and this type of bond is called polar covalent bond.

For example in HCl, chlorine being more electronegative than hydrogen will pull bonded electrons towards itself and develop a negative charge while hydrogen will have a partial positive charge.

19. Arrange the following molecules in order of increasing ionic character of their bonds LiF, K_2O , N_2 , SO_2 , ClF_3

Ans. $N_2 < SO_2 < ClF_3 < K_2O < LiF$

20. 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 an acetic acid.



Ans.

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

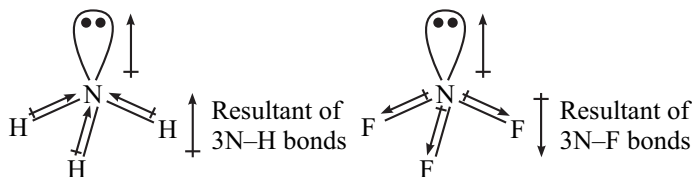
Ans. Electronic configuration of C-atom is $1s^2 2s^2 2p_x^1 2p_y^1$. Hence, it undergoes sp^3 hybridisation which gives a tetrahedral shape. For square planar, the hybridisation required is dsp^2 which is not possible for C-atom in CH_4 .

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

Ans. This is because BeH_2 molecule is linear (H-Be-H) so that the two Be-H dipole moments are equal and opposite and hence cancel out.

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

Ans.



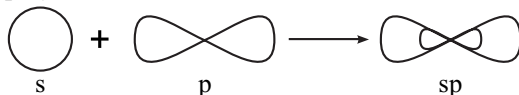
NH_3 has higher dipole moment than NF_3 . This is explained on the basis of the following two reasons.

- (1) The dipole formed between the lone pair and N atom has to be taken into consideration which is in the direction of the lone pair.
- (2) F is more electronegative than N, therefore direction of bond is from N to F whereas N is more electronegative than H, the direction of the bond is from H to N. Thus (see diagram) resultant moment of N–H bonds add up to the bond moment of lone pair, whereas that of 3 N–F bonds partly cancels the resultant moment of lone pair. Hence, the net dipole moment of NF_3 is less than that of NH_3 .

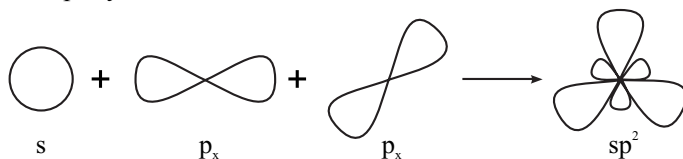
24. What is meant by hybridisation of atomic orbitals? Describe the shapes of sp, sp^2 and sp^3 hybrid orbitals.

Ans. Hybridisation is defined as the mixing of the atomic orbitals belonging to the same atom but having slightly different energies so that a redistribution of energy takes place between them resulting in the formation of new orbitals of equal energies and identical shapes. The new orbitals thus formed are known as hybrid orbitals.

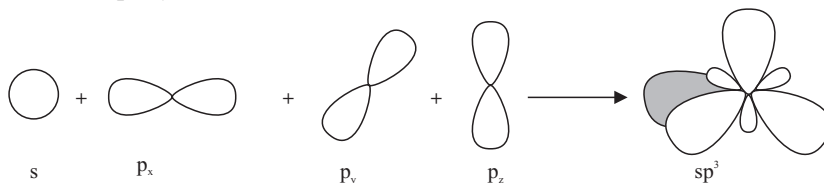
sp-hybridisation:— when one s and one p orbital belonging to the same main shell of an atom mix together to form two new equivalent orbitals type of is called sp hybridisation.



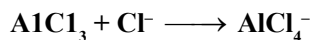
sp^2 -hybridisation:— when one s and two p orbitals of the same shell of an atom mix to form three new equivalent orbitals, the type of hybridisation is called sp^2 hybrid orbitals.



sp³-hybridisation:- when one s and three p orbitals of the same shell of an atom mix to form four new equivalent orbitals, the type of hybridisation is called sp³ hybrid orbitals.



25. Describe the change in hybridisation (if any) of the Al atom in the following reaction:



Ans. E.C. of $_{13}\text{Al} = 1s^2 2s^2 2p^6 3s^2 3p^1$ (Ground state) or $1s^2 2s^2 2p^6 3s^1 3p_x^1 3p_y^1$ (Excited state). Hence, it undergoes, sp² hybridisation to give it planar triangular structure. To form AlCl_4^- , the empty 3p_z orbital is also involved so that hybridisation is sp³ and the shape is tetrahedral.

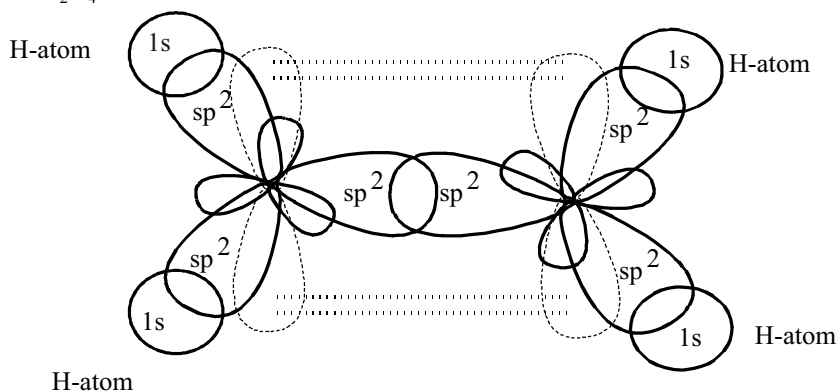
26. Is there any change in the hybridisation of B and N atoms as a result of the reaction,

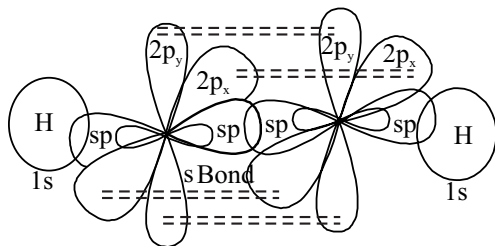


Ans. In BF_3 , B is sp² hybridised and in NH_3 , N is sp³ hybridised. After the reaction, hybridization of B changes to sp³ but that of N remains unchanged.

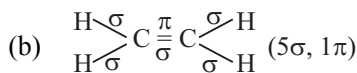
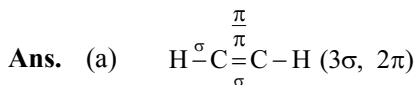
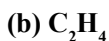
27. Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in C_2H_4 and C_2H_2 molecules.

Ans. C_2H_4 :-





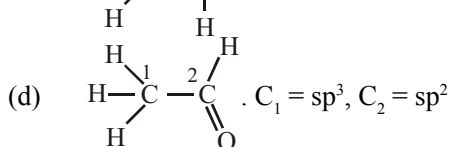
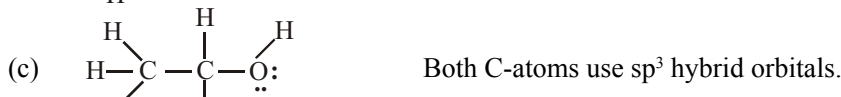
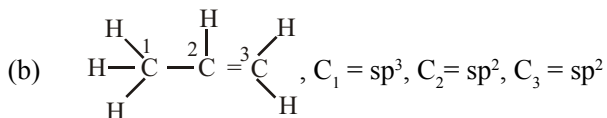
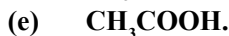
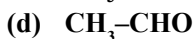
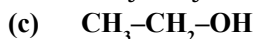
28. What is total number of sigma and pi bonds in the following molecules?

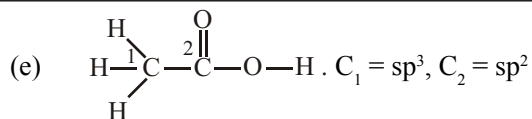


29. Considering x -axis as the internuclear axis which out of the following will not form a sigma bond and why? (a) $1s$ and $1s$ (b) $1s$ and $2p_x$ (c) $2p_y$ and $2p_y$ (d) $1s$ and $2s$.

Ans. Only (c) will not form a σ -bond because taking x -axis as the intermolecular axis, there will be lateral (sideway) overlap between the two $2p_y$ orbitals forming a π -bond.

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





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

Ans. Covalent bonds are formed by mutual sharing of electrons between the two atoms. The shared pairs of electrons thus present between the bonded atoms are called bond pairs.

All the electrons of an atom may not participate in the bonding. The electron pairs present on the atoms which do not take part in bonding are called

lone pairs of electrons. For example, in CH_4 $\left(\begin{array}{c} \text{H} \\ \vdots \\ \text{H} : \text{C} : \text{H} \\ \vdots \\ \text{H} \end{array} \right)$, there are only 4 bond pairs but in H_2O $\left(\begin{array}{c} \cdot\cdot \\ \cdot\cdot \\ \text{O} \\ \cdot\cdot \\ \cdot\cdot \\ \text{H} \quad \text{H} \end{array} \right)$, there are two bond pairs and two lone pairs.

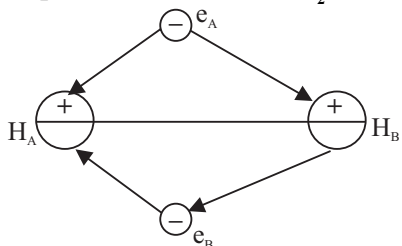
32. Distinguish between a sigma bond and a pi bond.

Ans.

<i>Sigma Bond</i>	<i>Pi Bond</i>
1. It is formed by overlap of orbitals along internuclear axis	1. It is formed by sideways overlapping of orbitals
2. It is formed by overlapping of s-s orbital, s- p_z orbital and p_z - p_z orbital	2. It is formed by overlap of p_x - p_x and p_y - p_y orbitals.
3. The electron cloud of a sigma bond is symmetrical about the internuclear axis and the overlapping is quite large, thus sigma bonds are strong	3. As the electron clouds overlap above and below the plane of internuclear axis, the overlapping is small, thus pi bonds are weak
4. Free rotation about a sigma bond is possible	4. Free rotation about a pi bond is not possible, since on rotating the overlapping vanishes and the bond breaks
5. Bond length is shorter	5. Bond length is longer
6. No preset conditions are required for the formation of bond	6. Pi bonds are always formed after sigma bonds have formed

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

Ans.



Consider two hydrogen atoms A and B with electron e_A and e_B , respectively. H_A represents the nucleus of hydrogen atom 'A' and H_B for hydrogen atom 'B'. When the two hydrogen atoms approach each other, the following two forces come into existence:

- Attractive forces between (i) electron e_A and nucleus H_B , (ii) electron e_B and nucleus H_A .
- Repulsive forces between (i) electron e_A and electron e_B , (ii) nucleus H_A and nucleus H_B .

The attractive forces tend to bring the atom close to each other, whereas repulsive forces tend to push them apart. Since the magnitude of attractive forces is greater than the repulsive forces, as a result, the energy of the system decreases and a molecule of hydrogen is formed.

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

Ans. The main points are

- When two atomic orbitals combine they lose their original identity and form new orbitals which are termed as molecular orbitals.
- Only those atomic orbitals combine to form molecular orbitals which have comparable energies and proper orientation. For example, $1s$ would only combine with $1s$ and not with $2s$. $2p_x$ would only combine with $2p_x$ of other atom and so on.
- The number of molecular orbitals formed is equal to the number of atomic orbitals which combine with one another.
- The two atomic orbitals combine to give bonding molecular orbital and anti-bonding molecular orbital. The bonding molecular orbital has lower energy than anti-bonding molecular orbital and hence, greater stability than other.
- The filling up of molecular orbitals takes place according to the same rules as those of atomic orbitals.

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

Ans. E.C. of ${}_4\text{Be} = 1s^2 2s^2$

M.O. E.C. of $\text{Be}_2 = \sigma_{1s}^2 \sigma_{1s}^* 2 \sigma_{2s}^2 \sigma_{2s}^* 2$

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

Hence, Be_2 does not exist.

36. 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. The electronic configuration of

$\text{O}_2 = \text{KK}[\sigma(2s)]^2 [\sigma^*(2s)]^2 [\sigma(2p_z)]^2 [\pi(2p_x)^2 = \pi(2p_y)^2] [\pi^*(2p_x)^1 = \pi^*(2p_y)^1]$

The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 4

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

It is paramagnetic due to presence of 2 unpaired electrons in the $\pi^*(2p_x)$ and $\pi^*(2p_y)$.

The electronic configuration of

$\text{O}_2^+ = \text{KK}\sigma(2s^2) (\sigma^*2s^2) (\sigma(2p_z)^2) (\pi(2p_x)^2 = \pi(2p_y)^2) (\pi^*2p_x^1 = \pi^*2p_y^0)$

The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 3

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = \frac{1}{2}(8 - 3) = 2\frac{1}{2} = 2.5$$

It is paramagnetic due to presence of 1 unpaired electrons in the $\pi^*(2p_x)$.

The electronic configuration of O_2^{-1} (super oxide) = $\text{KK}[\sigma(2s)]^2 [\sigma^*(2s)]^2 (\sigma(2p_z)^2) (\pi(2p_x)^2 = \pi(2p_y)^2) (\pi^*2p_x^2 = \pi^*2p_y^1)$

The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 5

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = \frac{1}{2}(8 - 5) = 1\frac{1}{2} = 1.5$$

It is paramagnetic due to presence of 1 unpaired electrons in the $\pi^*(2p_y)$.

The electronic configuration of O_2^{-2} (peroxide) = $\text{KK}[\sigma(2s)]^2 [\sigma^*(2s)]^2 (\sigma(2p_z)^2) (\pi(2p_x)^2 = \pi(2p_y)^2) (\pi^*2p_x^2 = \pi^*2p_y^2)$

The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 6

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = 1/2 (8 - 6) = 1$$

It is dia-magnetic due to presence of no unpaired electrons.

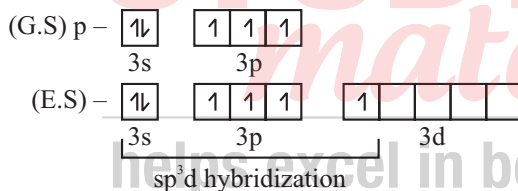
The relative stability would be $O_2^+ > O_2 > O_2^{-1} > O_2^{-2}$ since higher the bond order more is the stability.

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

Ans. As orbitals are represented by wave functions, a plus sign in an orbital represents a +ve wave function and a minus sign represents a -ve wave function.

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

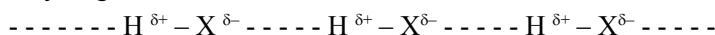
Ans. In PCl_5 , the hybridisation is sp^3d , thus, it has a trigonal bi-pyramidal structure. This is because one of the 3s electron of 'p' gets promoted to 3d orbital giving the electronic configuration $1s^2 2s^2 2p^6 3s^1 3p_x^1 3p_y^1 3p_z^1 3d^1$, thus it involves sp^3d hybridisation.



The equatorial and the axial bond length in PCl_5 are different. The axial bond lengths are longer than the equatorial bond length. This is due to minimise bond-pair repulsion for the large size Chlorine atoms, as axial bonds face more repulsion than equatorial.

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

Ans. Hydrogen bond – Whenever a molecule contains a hydrogen atom linked to a highly electronegative atom (F, O, N), this atom attracts the shared pair of electrons more and so this end of the molecule becomes slightly negative while the other H-end becomes slightly positive. The negative end attracts the positive end as a result a weak bond is formed between them. This bond is called hydrogen bond.

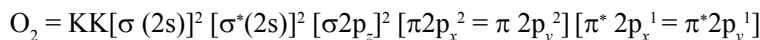


Hydrogen bond is stronger than the van der Waal's forces.

40. What is meant by the term bond order? Calculate the bond order of: O_2 , O_2^+ , O_2^- and N_2 ?

Ans. Bond order is defined as half of the difference between the number of electrons present in the bonding and the anti-bonding orbitals i.e. B.O. = $1/2 (N_b - N_{ab})$.

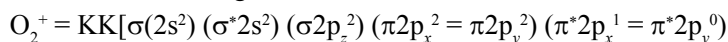
The electronic configuration of



The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 4

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = 1/2 (8 - 4) = 2,$$

The electronic configuration of



The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 3

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = 1/2 (8 - 3) = 2\frac{1}{2},$$

The electronic configuration of O_2^{-1} (super oxide) = $KK[\sigma(2s)]^2 [\sigma^*(2s)]^2 (\sigma 2p_z^2) (\pi 2p_x^2 = \pi 2p_y^2) (\pi^* 2p_x^2 = \pi^* 2p_y^1)$

The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 5

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = 1/2 (8 - 5) = 1\frac{1}{2}.$$

The electronic configuration of $N_2 = KK [\sigma(2s)]^2 [\sigma^*(2s)]^2 [\pi(2p_x)]^2 = [\pi(2p_y)]^2 [\sigma(2p_z)]^2$

The number of bonding electrons (N_b) = 8, number of anti-bonding electrons (N_{ab}) = 2

$$\text{Bond order} = \frac{N_b - N_{ab}}{2} = 1/2 (8 - 2) = 3,$$

The order of stability will be $N_2 > O_2^+ > O_2 > O_2^{-1}$.