

## NCERT Solutions for Class 11 Chemistry : Chapter 4 (Chemical Bonding and Molecular Structure)

Q-1) Explain the formation of a chemical bond

Ans.)

"Chemical bond is an attractive force that bounds the constituents of a chemical species together."

So many theories are suggested for chemical bond formation such as valence shell electron pair repulsion theory, electronic theory, molecular orbital theory, and valence bond theory.

Formation of chemical bond is credited to the tendency of system to achieve stability. It was noticed that inertness of noble gases is direct result of their completely filled outermost orbitals. Consequently, it was proposed that the elements having deficiency of electrons in outermost shells are unstable. Thus, atoms combine with one another and finish their separate octets or duplets to achieve the stable configuration of the closest inert gases. So, this combination may occur either by sharing of electrons. The formed chemical bond is a result of sharing of electrons among atoms is known as a covalent bond. Also, a formed ionic bond is a result of sharing of electrons among atoms.

Q-2) Write Lewis dot symbols for atoms of the following elements :

- Na
- Mg
- O
- N
- B
- Br

Ans.)

a) Na

Sodium atom contains only 1 valence electron. Thus, the Lewis dot symbols for Na is  $\text{Na}\cdot$ .

b) Mg

Magnesium atom contains only 2 valence electrons. Thus, the Lewis dot symbols for Mg is



c) O

Oxygen atom contains only 6 valence electrons. Thus, the Lewis dot symbols for O is



d) N

Nitrogen atom contains only 5 valence electrons. Thus, the Lewis dot symbols for N is



e) B

Boron atom contains only 3 valence electrons. Thus, the Lewis dot symbols for B is



f) Br

Bromine atom contains only 7 valence electrons. Thus, the Lewis dot symbols for Br is



Q-3) Give the Lewis dot symbols for the elements given below:

(a) Al and  $\text{Al}^{3+}$

(b) H and  $\text{H}^-$

(c) S and  $S^{2-}$

Ans.)

(a) Al and  $Al^{3+}$

Aluminium atom contains only 3 valence electrons. Thus, the lewis dot symbols for Al is



The tri-positive charge on aluminium indicates that it has donated 3 electrons.

Thus lewis dot symbol is



(b) H and  $H^{-}$

Hydrogen atom contains only 1 valence electrons. Thus, the lewis dot symbols for H is



The single negative charge on hydrogen indicates that it has gained 1 electron. So, one valence electron plus one gained electron.

Thus lewis dot symbol is



(c) S and  $S^{2-}$

Sulphur atom contains only 6 valence electrons. Thus, the lewis dot symbols for S is



The bi-negative charge on sulphur indicates that it has gained 2 electron. So, six valence electron plus two gained electron.

Thus lewis dot symbol is



Q-4) Draw the Lewis structures for the following molecules and ions :

(a)  $\text{SiCl}_4$

(b)  $\text{H}_2\text{S}$

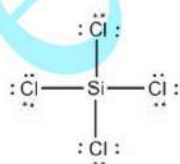
(c)  $\text{HCOOH}$

(d)  $\text{BeF}_2$

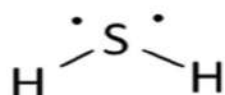
(e)  $\text{CO}_3^{2-}$

Ans.)

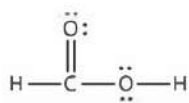
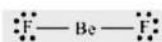
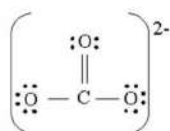
(a)  $\text{SiCl}_4$



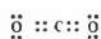
(b)  $\text{H}_2\text{S}$



(c) HCOOH

(d) BeF<sub>2</sub>(e) CO<sub>3</sub><sup>2-</sup>**Q-5) Define octet rule. Write its significance and limitations****Ans.)**

Octet rule says, "atoms can combine either by transfer of valence electrons from one atom to another or by sharing their valence electrons in order to achieve the nearest inert gas configuration by having an octet in their valence shell."



Octet rule explains chemical bond formation depending upon nature of element.

Limitations:

(a) Octet rule fails to predict the relative stability and shape of the molecules.

(b) It is based on inert nature of noble gases. But, some inert gases say, krypton(Kr) and xenon(Xe) form compounds like KrF<sub>2</sub>, XeF<sub>2</sub> etc.

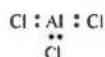
(c) For elements beyond 3<sup>rd</sup> period the octet rule cannot be applied. Elements present beyond 3<sup>rd</sup> period have more than 8 valence electrons surrounding central atom. E.g. SF<sub>6</sub>, PF<sub>6</sub> etc.



(d) For atoms in a molecule having odd number of electrons octet rule is not applied. E.g. For NO<sub>2</sub> and NO octet rule is not applicable.



(e) If a compound is having less than 8 electrons surrounding the central atom then octet rule cannot be applied to that compound. E.g. BeH<sub>2</sub>, AlCl<sub>3</sub>, LiCl etc. is not obeying octet rule.



Q-6) Write the favourable factors for the formation of ionic bond.

Ans.)

Formation of ionic bond takes place by transfer of 1 or more electrons from one atom to another. Thus, ionic bond formation depends on the flexibility of neutral atoms to lose or gain electrons. Formation of ionic bond also depends on the lattice energy of compound which is formed.

The factors that are favourable for ionic bond formation:

- High electron affinity of atoms of non-metal.
- High lattice energy of compound which is formed.
- Low ionization enthalpy of atom of metal.

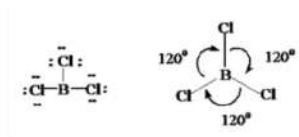
Q-7) Discuss the shape of the following molecules using the VSEPR model:

- $\text{BCl}_3$
- $\text{SiCl}_4$
- $\text{AsF}_5$
- $\text{PH}_3$
- $\text{BeCl}_2$
- $\text{H}_2\text{S}$

Ans.)

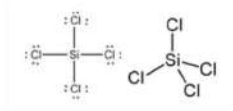
- $\text{BCl}_3$

Central atom is not having any lone pair but have 3 bond pairs. Thus, its shape is  $\text{AB}_3$ . i.e.. Trigonal planar.



- $\text{SiCl}_4$

Central atom is not having any lone pair but have 4 bond pairs. Thus, its shape is  $\text{AB}_4$ . i.e. Tetrahedral.



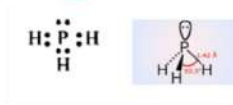
- $\text{AsF}_5$

Central atom is not having any lone pair but have 5 bond pairs. Thus, its shape is  $\text{AB}_5$ . i.e. Trigonalbipyramidal.



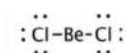
- $\text{PH}_3$

Central atom is having 1 lone pair and is having 3 bond pairs. Thus, its shape is  $\text{AB}_3\text{E}$ . i.e. Trigonalbipyramidal.



- $\text{BeCl}_2$

Central atom is not having any lone pair but have 2 bond pairs. Thus, its shape is  $\text{AB}_2$ . i.e. Linear shape.



(f)  $\text{H}_2\text{S}$

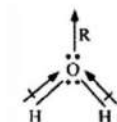
Central atom is having 1 lone pair and is having 2 bond pairs. Thus, its shape is  $\text{AB}_2\text{E}$ . i.e. Bent shape.



**Q-8) Although geometries of  $\text{NH}_3$  and  $\text{H}_2\text{O}$  molecules are distorted tetrahedral, bond angle in water is less than that of Ammonia. Discuss.**

**Ans.)**

The geometry of  $\text{H}_2\text{O}$  and  $\text{NH}_3$ :



Central atom(N) in ammonia is having 1 lone pair and is having 3 bond pairs.

Central atom(O) in water is having 2 lone pair and is having 2 bond pairs.

Thus, these 2 lone pairs on O-atom in water molecule repels the 2 bond pairs. And this repulsion is between lone pair and bond pair on O-atom of  $\text{H}_2\text{O}$  is stronger than the repulsion is between lone pair and bond pair on N-atom of  $\text{NH}_3$ .

Thus, the bond angle in  $\text{H}_2\text{O}$  is less than  $\text{NH}_3$ , even though they are having distorted tetrahedral structure.

**Q-9) How do you express the bond strength in terms of bond order ?**

**Ans.)**

The extent of bonding which occurs between two atoms while forming a molecule is represented by bond strength. As the bond strength increases the bond becomes stronger and the bond order increases.

**Q-10) Define Bond length.**

**Ans.)**

"Bond length is defined as the equilibrium distance between the nuclei of 2 bonded atoms in a molecule."

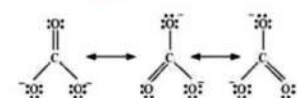
**Q-11) Explain the important aspects of resonance with reference to the  $\text{CO}_3^{2-}$  ion.**

**Ans.)**

Experimental results shows that, all the C-O bond in  $\text{CO}_3^{2-}$  are equivalent.

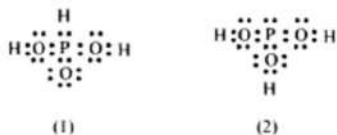
Thus, it is inefficient to represent  $\text{CO}_3^{2-}$  ion by single lewis structure which is having 1 double bond and 2 single bond.

Thus, the resonance structures of  $\text{CO}_3^{2-}$  is :



**Q-12)  $\text{H}_3\text{PO}_3$  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  $\text{H}_3\text{PO}_3$ ? If not, give reasons for the same.**

**Ans.)**

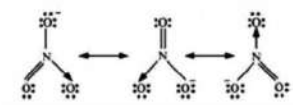


In the given structures the position of atoms are changed, so we cannot take the 2 given structures as canonical form of resonance hybrid which is representing  $H_3PO_3$ .

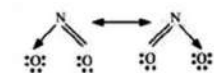
Q-13) Write the resonance structures for  $SO_3$ ,  $NO_2$ , and  $NO_3^-$ .

Ans.)

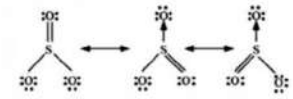
(a)  $NO_3^-$



(b)  $NO_2$



(c)  $SO_3$



Q-14) Use Lewis symbols to show electron transfer between the following atoms to form cations and anions :

(i) O and Ca

(ii) N and Al

(iii) S and K.

Ans.)

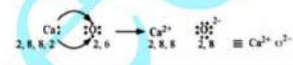
(i) O and Ca

Electronic configurations of O and Ca are:

O: 2, 6

Ca: 2, 8, 8, 2

Here, it clear that Ca has 2 more electrons than nearest inert gas. i.e. Ar, whereas O needs 2 electrons to complete its octet. Thus, the transfer of electrons takes place in following way,



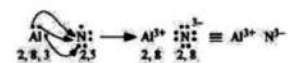
(ii) N and Al

Electronic configurations of N and Al are:

N: 2, 5

Al: 2, 8, 3

Here, it clear that Al has 3 more electrons than nearest inert gas. i.e. Ne, whereas N needs 3 electrons to complete its octet. Thus, the transfer of electrons takes place in following way,



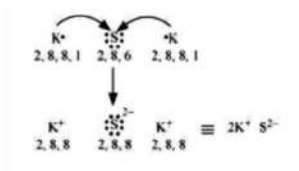
(iii) S and K

Electronic configurations of S and K are:

S: 2, 8, 6

K: 2, 8, 8, 1

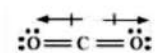
Here, it clear that K has 1 more electrons then nearest inert gas. i.e. Ne, whereas S needs 2 electrons to complete its octet. Thus, the transfer of electrons takes place in following way,



Q-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.

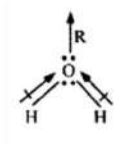
Ans.)

Experimental results show that the dipole moment of  $\text{CO}_2$  is 0. And it is possible only if the shape of molecule is linear as dipole moments of bond between C-O is equal and opposite so, it nullifies each other.



$\therefore$  Resultant,  $\mu = 0$

$\text{H}_2\text{O}$  has 1.84D dipole moment. The value of dipole moments indicates that structure of water molecule is bent as dipole moments of bond between O-H is unequal.



Q-16) Write the significance/applications of dipole moment.

Ans.)

In heteronuclear molecule, there is difference in electro-negativities of constituents of atom, due to which polarisation arises in it. Thus, one end of acquires positive charge and other end acquires negative charge. So, molecules are said to have a dipole moment.

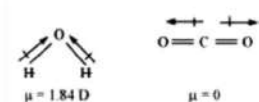
The product of distance between centres of negative and positive charges and magnitude of charge is known as dipole moment ( $\mu$ ). Dipole moment is a vector quantity and it is represented by an arrow in such a way that head of a tail is pointing towards negative centre and tail of arrow is pointing towards positive centre.

$\mu = \text{distance between the centres (r)} \times \text{charge (Q)}$

SI unit of dipole moment is 'esu'.

$$1 \text{ esu} = 3.335 \times 10^{-30} \text{ cm}$$

It is the measure of polarity of a bond. It differentiates the non-polar and polar bonds as all the non-polar molecules have dipole moment equal to 0. It also calculates the ionic character of a molecule.



Q-17) Define electronegativity. How does it differ from electron gain enthalpy ?

Ans.)

"Electronegativity is the ability of an atom in a chemical compound to attract a bond pair of electrons towards itself".

Sr. No	Electronegativity	Electron affinity
1	Tendency to attract the shared pairs of electrons for an atom which is in chemical compound is its electronegativity.	Tendency to gain electrons for an isolated gaseous atom is its electron gain enthalpy.
2	It varies according to element to which it is bounded with.	It does not vary according to element to which it is bounded with.
3	It is not constant for any element.	It is constant for an element.
4	It is not a measurable quantity.	It is a measurable quantity.

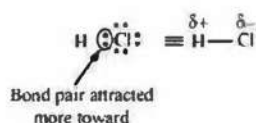
Q-18) Explain with the help of suitable example polar covalent bond.

Ans.)

When two unique atoms having distinct electronegativities join to form a covalent bond, the bond pair of electrons is not shared equally. The nucleus of an atom having greater electronegativity attracts the bond pair. So, the electron distribution gets distorted and an electronegative atom attracts the electron cloud.

Thus, the electronegative element gets slightly negatively charged and on the other hand, the other atom gets slightly positively charged. As a result of this, two opposite poles are developed in a molecule and this type of bond formed is termed as 'polar covalent bond'.

E.g. HCl is having a polar covalent bond. In HCl, Cl- atom is having more electronegativity than H- atom. Thus, bond pair shifts towards Cl- atom and because of that it acquires positive charge.



Q-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>.

Ans.)

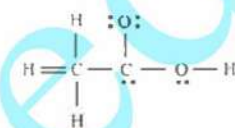
Ionic characteristic of a molecule depends on the difference in electronegativity between constituents atoms. So, higher the difference, the ionic characteristic of a molecule will be higher.

So, the required order of ionic characteristic of the given molecules is

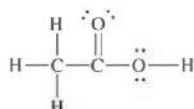


Q-20) The skeletal structure of CH<sub>3</sub>COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.

Ans.)



Correct Lewis structure of CH<sub>3</sub>COOH is given below:

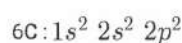


Q-21) Apart from tetrahedral geometry, another possible geometry for CH<sub>4</sub> is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH<sub>4</sub> is not square planar.



Ans.)

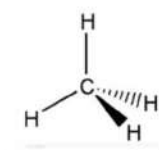
Electronic configuration of C- atom:



Orbital picture of C- atom in excited state is:



Thus, C- atom undergoes  $sp^3$  hybridization in methane molecule and forms tetrahedral structure.



For square planer geometry, C-atom should have  $dsp^2$  hybridization. But as C- atom is not having d- orbital so it cannot undergo  $dsp^2$  hybridization. Thus, Methane's geometry cannot be square planer.

Also in square planar geometry the bond angle is  $90^\circ$  so the stability is not there because of repulsion between

bond pairs. So as per VSEPR theory methane's tetrahedral structure is perfect.

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

Ans.)

Lewis structure of  $\text{BeH}_2$  is:



Central atom is not having any lone pair but have 2 bond pairs. Thus, its shape is  $\text{AB}_2$ . i.e. Linear shape.



Thus, the dipole moment of  $\text{Be-H}$  bond is equal and opposite in direction so it nullifies one another. Thus, the dipole moment of  $\text{BeH}_2$  is 0.

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

Ans.)

N- atom is the central atom of  $\text{NF}_3$  and  $\text{NH}_3$ .

Central atom is having 1 lone pair and is having 3 bond pairs. Thus, for both the shape is  $\text{AB}_3\text{E}$ . i.e. Pyramidal.

As, F-atom is more electronegativity than H- atom,  $\text{NF}_3$  should have higher dipole moment than  $\text{NH}_3$ . But the dipole moment of  $\text{NH}_3$  is 1.46D which is higher than dipole moment of  $\text{NF}_3$  which is 0.24D.

It gets clear from the directions of dipole moments of individual bond in  $\text{NF}_3$  and  $\text{NH}_3$ .



As, both the N-H bond are in same direction it adds to the bond moment of the lone pair, while N-F bond are in opposite direction so they partly cut the bond moment of lone pair.

Thus, dipole moment of  $\text{NH}_3$  is higher than that of  $\text{NF}_3$ .

**Q-24) What is meant by hybridisation of atomic orbitals? Describe the shapes of  $sp$ ,  $sp^2$ ,  $sp^3$  hybrid orbitals.**

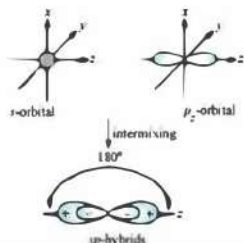
Ans.)

"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".

E.g. 1 s-orbital hybridises with 3 p-orbitals to form 4  $sp^3$  hybrid orbitals.

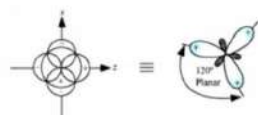
(a) sp hybrid orbital

1 s-orbital hybridises with 1 p-orbital to form 2 sp hybrid orbitals. sp hybrid orbital is having linear shape. The formation of sp orbital is:



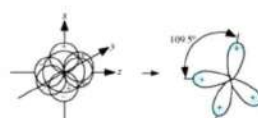
(b)  $sp^2$  hybrid orbital

1 s-orbital hybridises with 2 p-orbitals to form 3  $sp^2$  hybrid orbitals. The shape of  $sp^2$  orbital is trigonal planar.



(c)  $sp^3$  hybrid orbital

1 s-orbital hybridises with 3 p-orbitals to form 4  $sp^3$  hybrid orbitals. The shape of  $sp^3$  orbital is tetrahedron.



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

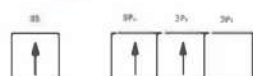


Ans.)

The ground state of valence orbital of Al-atom is:



In excited state the orbital picture of Al-atom is:



Thus, Al-atom in  $AlCl_3$  undergoes  $sp^2$  hybridisation and forms trigonal planar geometry. For the formation of

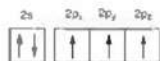
$AlCl_4^-$  the vacant  $3p_z$  orbital will also get involved. Thus,  $sp^2$  hybridisation is converted into  $sp^3$  hybridisation and forms a tetrahedral structure.

Q-26) Is there any change in the hybridisation of B and N atoms as a result of the following reaction?

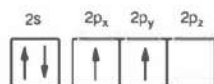


Ans.)

N- atom in  $\text{NH}_3$  is having  $sp^3$  hybridization. Orbital picture of N- atom is shown below:



B- atom in  $\text{BF}_3$  is having  $sp^2$  hybridisation . Orbital picture of B- atom is shown below:



On the reaction of  $\text{NH}_3$  and  $\text{BF}_3$ ,  $\text{F}_3\text{B} \cdot \text{NH}_3$  is obtained as product, as hybridization of B-atom is changed to  $sp^3$ . Although, hybridization of N- atoms remains unchanged.

Q-27) Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in  $\text{C}_2\text{H}_4$  and  $\text{C}_2\text{H}_2$  molecules.

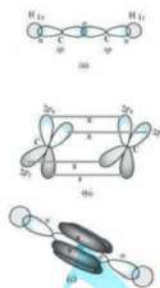
Ans.)

- $\text{C}_2\text{H}_2$

In formation of ethyne ( $\text{C}_2\text{H}_2$ ) molecule, C- atom is having  $sp$  hybridization with 2  $2p$ - orbitals in unhybridized state.

1  $sp$  orbital of each C- atom overlaps the inter-nuclear axis and forms C-C sigma bond. The  $2^{nd}$   $sp$  orbital of each C- atom overlaps half-filled  $1s$  orbital so as to form a sigma bond.

The triple bond between the 2 C- atoms has 1 sigma bond and 2  $\pi$  bonds. This is because 2 unhybridized  $2p$ - orbitals overlaps with the  $2p$ - orbital of other C- atom, thus forming 2  $\pi$  bonds.



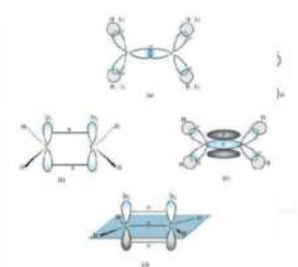
- $\text{C}_2\text{H}_4$

Electronic configuration of carbon atom in excited state is given below:



In the formation of  $\text{C}_2\text{H}_4$  (ethane) molecule 1  $sp^2$  orbital of C- atom overlaps  $sp^2$  orbital of other C- atom. Thus, forming a C-C sigma bond.

The 2 remaining  $sp^2$  orbital of every C- atom forms  $sp^2$ -s  $\sigma$  bond with 2 H- atoms. One c- atom having unhybridized orbital overlaps with the unhybridized orbital of other C- atom and forms a  $\pi$  bond.



Q-28) What is the total number of sigma and pi bonds in the following molecules?

(a)  $C_2H_2$

(b)  $C_2H_4$

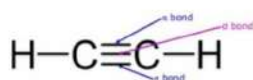
Ans.)

Single bond is formed as the axis of bonding orbital overlaps. Thus, it forms a  $\sigma$  bond. By sidewise overlapping of

orbital double and triple bonds i.e. multiple bonds are formed.  $\pi$  bond is always present in the multiple bonds.

Triple bond consist of 2  $\pi$  and 1  $\sigma$  bond.

(i)  $C_2H_2$

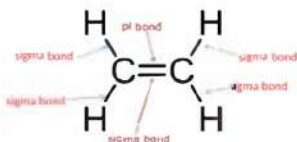


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Thus, there are 2  $\pi$  bonds and 3  $\sigma$  in  $C_2H_2$ .

(ii)  $C_2H_4$

Thus, there are 1  $\pi$  bonds and 5  $\sigma$  in  $C_2H_4$ .



Q-29) Considering x-axis as the internuclear axis which out of the following will not form a sigma bond and why?

(i) 1s and  $2p_x$

(ii) 1s and 1s

(iii) 1s and 2s

(iv)  $2p_y$  and  $2p_y$

Ans.)

(iv)  $2p_y$  and  $2p_y$

$2p_y$  and  $2p_y$  orbitals won't form a  $\sigma$  as it will undergo lateral overlapping and will form a  $\pi$  bond.

Q-30) Which hybrid orbitals are used by carbon atoms in the following molecules?

(i)  $CH_3-CH=CH_2$

(ii)  $CH_3-CH_3$

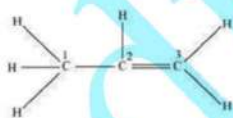
(iii)  $CH_3-CHO$

(iv)  $CH_3COOH$

(v)  $CH_3CH_2-OH$

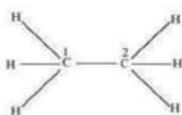
Ans.)

(i)  $CH_3-CH=CH_2$



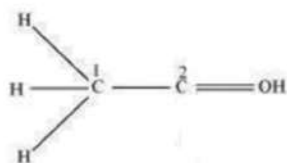
Here,  $C_3$  and  $C_2$  are having  $sp^2$  hybridization and  $C_1$  is having  $sp^3$  hybridization.

(ii)  $CH_3-CH_3$



Here,  $C_1$  and  $C_2$  are having  $sp^3$  hybridization.

(iii)  $CH_3-CHO$



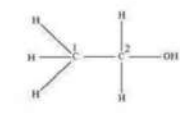
Here,  $C_1$  is having  $sp^3$  hybridization and  $C_2$  is having  $sp^2$  hybridization.

(iv)  $CH_3COOH$



Here,  $C_1$  is having  $sp^3$  hybridization and  $C_2$  is having  $sp^2$  hybridization.

(v)  $CH_3CH_2-OH$



Here,  $C_1$  and  $C_2$  are having  $sp^3$  hybridization.

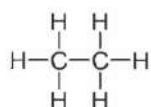
**Q-31) What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving one example of each type.**

**Ans.)**

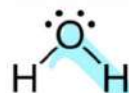
A covalent bond is formed when 2 atoms combine with each other by sharing their valence electrons.

"The shared pairs of electrons present between the bonded atoms are called bond pairs". Each and every electron cannot participate in bonding. "The pairs of electrons which do not participate in bonding are called lone pairs".

E.g. a) Ethane is having 7 bond pairs but zero lone pair.



b) Water is having 2 bond pairs and 2 lone pairs on O-atom.



**Q-32) Distinguish between a sigma and a pi bond.**

**Ans.)**

Sr. No.	Pi bond	Sigma bond
1	Pi bond is formed by lateral overlapping of orbitals.	Sigma bond is formed by end to end overlapping of orbitals.
2	It is comparatively weak bond.	It is comparatively strong bond.
3	There is only one overlapping orbital is p-p.	The overlapping orbitals are s-s, s-p, p-p.
4	Rotation around pi-bond is restricted.	Rotation is possible around sigma bond.
5	Electron cloud is not symmetrical about the line joining 2 nuclei.	Electron cloud is symmetrical about the line joining 2 nuclei.
6	It is having 2 electron clouds one above the plane of atomic nuclei and one below the plane of atomic nuclei.	It is having 1 electron cloud and that is symmetrical about the inter-nuclear axis.

**Q-33) Explain the formation of  $H_2$  molecule on the basis of valence bond theory.**

Ans.)

Assuming 2 H- atoms X and Y with nuclei  $N_X$  and  $N_Y$  and electrons  $e_X$  and  $e_Y$  respectively.

When X and Y are far for each other then there is no interaction between them. As soon as they come closer, the attractive force and repulsive force becomes active.

The repulsive forces are:

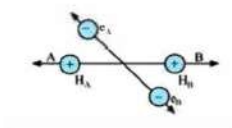
- (i) Between electrons of both the atoms i.e.  $e_X$  and  $e_Y$ .
- (ii) Between nuclei of both the atoms i.e.  $N_X$  and  $N_Y$ .

The attractive forces are:

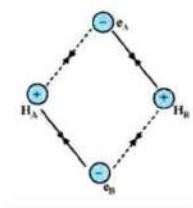
- (i) Between the electron and nucleus of the same atom i.e.  $N_X - e_X$  and  $N_Y - e_Y$ .
- (ii) Between the electron of one atom and nucleus of other atom i.e.  $N_X - e_Y$  and  $N_Y - e_X$ .

The repulsive force push the 2 atoms apart whereas the attractive force tend to bring them together.

Repulsive forces:



Attractive forces:



The values of repulsive forces are less than that of attractive forces. Thus, 2 atoms approach each other. Thus, there is decrease in potential energy. At the end a stage is reached when the repulsive forces balance the attractive forces and the system achieves the minimum energy, which leads to formation of H<sub>2</sub> molecule.

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

Ans.)

The condition that is required for linear combination of an atomic orbitals to form a molecular orbitals are as follows:

- (i) The joining of atomic orbitals must have approximately same energy. This implies in a homo-nuclear molecule, the 1s-orbital of one atom can join with the 1s-orbital of another atom, but cannot join with the 2s-orbital.
- (ii) The joining atomic orbitals must have legitimate orientations to ensure the maximum overlap.
- (iii) The overlapping must be in a large extent.

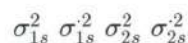
**Q-35) Use molecular orbital theory to explain why the Be<sub>2</sub> molecule does not exist.**

Ans.)

Electronic configuration of Be:



Molecular orbital electronic configuration of Be<sub>2</sub> is:



Thus, bond order of Be<sub>2</sub>:  $0.5(N_b - N_a)$ .

$N_b$ : No. of electrons in the bonding orbitals

$N_b$ : No. of electrons in the anti-bonding orbitals

Therefore, bond order of  $Be_2 = 0.5(4 - 4) = 0$

Zero value of bond order indicates that given molecule is unstable. Thus,  $Be_2$  doesn't exist.

Q-36) Compare the relative stability of the following species and indicate their magnetic properties;

(a)  $O_2$

(b)  $O_2^{2-}$  (peroxide)

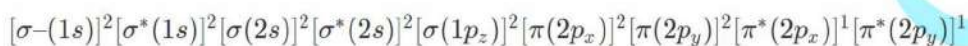
(c)  $O_2^-$  (superoxide)

(d)  $O_2^+$

Ans.)

$O_2$  contain 16 electrons i.e. 8 electrons from each O-atom.

- Electronic configuration of  $O_2$  is:



As 1s-orbital of each O-atom does not involve in the bonding,

No. of bonding electrons =  $N_b = 8$

No. of anti-bonding electrons =  $N_a = 4$

Now,

Bond order =  $0.5(8 - 4) = 2$

- Electronic configuration of  $O_2^{2-}$  (peroxide) is:



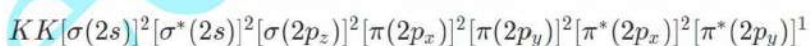
No. of bonding electrons =  $N_b = 8$

No. of anti-bonding electrons =  $N_a = 6$

Now,

Bond order =  $0.5(8 - 6) = 1$

- Electronic configuration of  $O_2^-$  (superoxide) is:



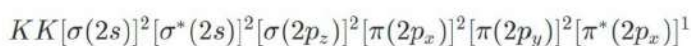
No. of bonding electrons =  $N_b = 8$

No. of anti-bonding electrons =  $N_a = 5$

Now,

Bond order =  $0.5(8 - 5) = 1.5$

- Electronic configuration of  $O_2^+$  is:





No. of bonding electrons =  $N_b = 8$

No. of anti-bonding electrons =  $N_a = 3$

Now,

Bond order =  $0.5(8 - 3) = 2.5$

As, bond dissociation energy  $\propto$  bond order

Hence, higher the bond order, higher stability will be there.

Arrangement according to decreasing order of stability is given as:



**Q-37) Write the significance of a plus and a minus sign shown in representing the orbitals**

**Ans.)**

Generally molecular orbital are represented by the 'wave function'.

Positive (+) sign in representing a molecular orbital indicates positive wave function.

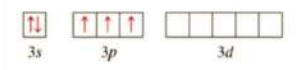
Negative (-) sign in representing a molecular orbital indicates negative wave function.

**Q-38) Describe the hybridisation in case of  $PCl_5$ . Why are the axial bonds longer as compared to equatorial bonds?**

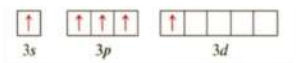
**Ans.)**

The electronic configuration of outer orbital of phosphorus in excited state and in ground state is given below:

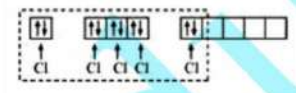
Ground State:



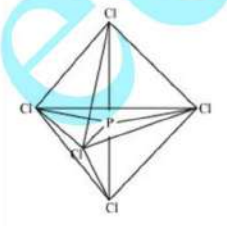
Excited State:



Phosphorus atom is having  $sp^3d$  hybridization. These orbitals are filled due to donation of electron pairs by 5 Cl- atoms:  $PCl_5$



The 5  $sp^3d$  hybrid orbitals present here are directed towards 5 corners of trigonalbipyramidals. Thus, the geometry of  $PCl_5$  is given below:



$PCl_5$  contains 5 P-Cl sigma bonds. Out of which 3 P-Cl bond lie in only 1 plane and they are making  $120^\circ$  with

each other. And as these bonds lie in 1 plane they are known equatorial bonds.

Out of 2 remaining P-Cl bonds one bond lie above the equatorial plane and one bond lie below the equatorial bond.

And they are making  $90^\circ$  with each other. These bonds are called axial bond.

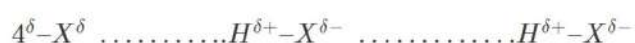
Equatorial bond pairs repel axial bond pairs to a large extent so, equatorial bonds are slightly shorter than axial bonds.

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

**Ans.)**

"H-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".

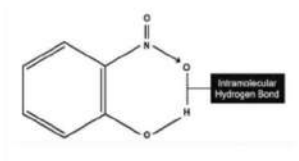
As there is difference in the electronegativities between the atoms, thus, the bond pair electronegative atom and hydrogen atom is drifted away from H-atom. Therefore, hydrogen atom gets electropositive w.r.t. the other atom and procures a positive charge.



The value of H-bond is minimum in gaseous state and maximum in the solid state.

Two types of hydrogen bond are there:

- Intramolecular hydrogen bond e.g., o-nitrophenol
- Inter molecular hydrogen bond e.g., HF, H<sub>2</sub>O etc.



H-bonds are stronger than Van der Waals forces as H-bond are regarded as extreme form of the dipole-dipole interaction.

**Q-40) What is meant by the term bond order? Calculate the bond order of: N<sub>2</sub>, O<sub>2</sub>, O<sub>2</sub><sup>+</sup> and O<sub>2</sub><sup>-</sup>.**

**Ans.)**

Bond Order: It is defined as 0.5 times the difference between the "No. of electrons present in bonding orbitals and No. of electrons present in anti-bonding orbitals" of a molecule.

$$\text{Bond Order} = 0.5(N_b - N_a);$$

N<sub>a</sub>: No. of anti-bonding electrons

N<sub>b</sub>: No. of bonding electrons

O<sub>2</sub> contain 16 electrons i.e. 8 electrons from each O-atom.

- Electronic configuration of O<sub>2</sub> is:

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

As 1s-orbital of each O-atom does not involve in the bonding,

No. of bonding electrons = N<sub>b</sub> = 8

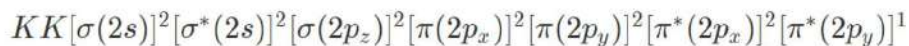
No. of anti-bonding electrons = N<sub>a</sub> = 4

Now,

Now,

$$\text{Bond order} = 0.5(8 - 4) = 2$$

- Electronic configuration of  $O_2^-$  (superoxide) is:



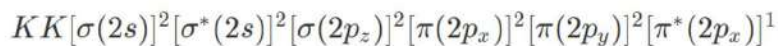
No. of bonding electrons =  $N_b = 8$

No. of anti-bonding electrons =  $N_a = 5$

Now,

$$\text{Bond order} = 0.5(8 - 5) = 1.5$$

- Electronic configuration of  $O_2^+$  is:



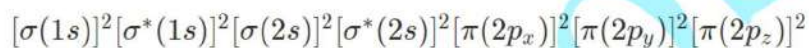
No. of bonding electrons =  $N_b = 8$

No. of anti-bonding electrons =  $N_a = 3$

Now,

$$\text{Bond order} = 0.5(8 - 3) = 2.5$$

- Electronic configuration of  $N_2$  is:



No. of bonding electrons =  $N_b = 10$

No. of anti-bonding electrons =  $N_a = 4$

Now,

$$\text{Bond order} = 0.5(10 - 4) = 3$$