NCERT Solutions for Class-XI Chemistry

Chapter-4
NCERT Chemistry Class 11

1. Explain the formation of a chemical bond.

1. A chemical bond is defined as an attractive force that holds the constituents (atoms, ions etc.) together in a chemical species.

Various theories have been suggested for the formation of chemical bonds such as the electronic theory, valence shell electron pair repulsion theory, valence bond theory, and molecular orbital theory.

A chemical bond formation is attributed to the tendency of a system to attain stability. It was observe that the inertness of noble gases was because of their fully filled outermost orbitals. Hence, it was postulated that the elements having incomplete outermost shells are unstable (reactive). Atoms, therefore, combine with each other and complete their respective octets or duplets to attain the stable configuration of the nearest noble gases.

This combination can occur either by sharing of electrons or by transferring one or more electrons from one atom to another. The chemical bond formed as a result of sharing of electrons between atoms is called a covalent bond. An ionic bond is formed as a result of the transference of electrons from one atom to another.

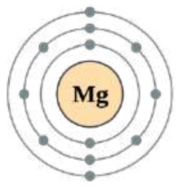
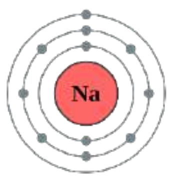
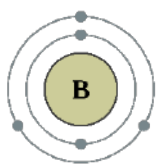
2. Write Lewis dot symbols for atoms of the following elements: Mg, Na, B, O, N, Br.

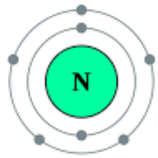
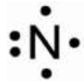
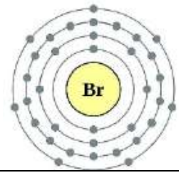
2. Steps to draw Lewis symbols:

⇒ Step 1: Count the number of valence electrons (electrons in the outermost shell)

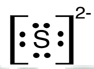
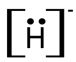
⇒ Step 2: Write the symbol of element in the centre.

⇒ Step 3: Draw dots around the symbol according to the number of valence electrons.

Element	Atomic number	Electronic configuration	No. of valence electrons	Representation	Lewis dot symbol
Mg	12	2,8,2	2		
Na	11	2,8,1	1		
B	5	2,3	3		

O	8	2,6	6		
N	7	2,5	5		
Br	35	2,8,18,7	7		

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

3. (i) S and S^{2-}
The number of valence electrons in sulphur is 6.
The Lewis dot symbol of sulphur (S) is .
The dinegative charge infers that there will be two electrons more in addition to the six valence electrons. Hence, the Lewis dot symbol of S^{2-} is . (ii) Al and Al^{3+}
The number of valence electrons in aluminium is 3.
The Lewis dot symbol of aluminium (Al) is .
The tripositive charge on a species infers that it has donated its three electrons. Hence, the Lewis dot symbol is $[Al]^{3+}$.
(iii) H and H^-
The number of valence electrons in hydrogen is 1.
The Lewis dot symbol of hydrogen (H) is .
The uninegative charge infers that there will be one electron more in addition to the one valence electron.
Hence, the Lewis dot symbol is .

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

4. (a) H_2S:
The atomic number of S is 16 and the no. of valence electrons is 6. The atomic number of H is 1 and no. of valence electron is 1. Hence, the total number of valence electrons is:
Two hydrogen (2 × 1 valence electron) + sulphur (6 valence electrons) = Total 8 valence electrons

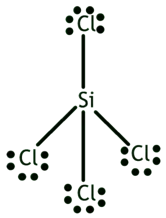
[Each line represents a pair of electrons]

The central atom S shares its two valence electrons out of six valence electrons to complete its octet and each hydrogen atom shares its one valence electron to complete the octet. The remaining four electrons act as a lone pair (electrons which do not participate in bonding) which is placed on the sulphur atom.

(b) $SiCl_4$

The atomic number of Si is 14 and the no. of valence electrons is 4. The atomic number of Cl is 17 and no. of valence electron is 7. Hence, the total number of valence electrons is:

Four chlorine (4 × 7 valence electrons) + silicon (4 valence electrons) = Total 32 valence electrons

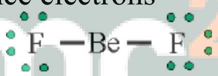


The central atom Si shares its all valence electrons and each chlorine atom shares its one valence electron out of seven valence electrons to complete the octet.

(c) BeF_2:

The atomic number of Be is 4 and the no. of valence electrons is 2. The atomic number of F is 9 and no. of valence electron is 7. Hence, the total number of valence electrons is:

Four fluorine (4 × 7 valence electrons) + Beryllium (2 valence electrons) = Total 30 valence electrons

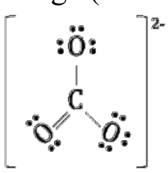


The central atom Be shares its all valence electrons and each fluorine atom shares its one valence electron out of seven valence electrons to complete the octet.

(d) CO_3^{2-}

The atomic number of C is 6 and the no. of valence electrons is 4. The atomic number of O is 8 and no. of valence electron is 6. Hence, the total number of valence electrons is:

Three oxygen (3 × 6 valence electrons) + Carbon (4 valence electrons) + negative charge (2 extra electrons) = Total 24 valence electrons



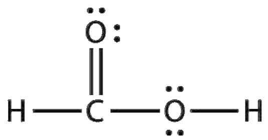
The central atom C shares its all valence electrons with each oxygen atom. Carbon requires only two electrons to complete its octet hence, only one of the oxygen atom shares its two valence electrons out of six valence electrons which are represented by a double bond.

(e) HCOOH:

The atomic number of C is 6 and the no. of valence electrons is 4. The atomic number of O is 8 and no. of valence electron is 6.

The atomic number of hydrogen is 1 and no. of valence electron is 1. Hence, the total number of valence electrons is:

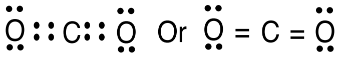
two oxygen (2 × 6 valence electrons) + Carbon (4 valence electrons) + two hydrogen (2 × 1 valence electron) = Total 18 valence electrons



The central carbon atom shares its all 6 valence electrons with one hydrogen and two oxygen atoms to complete its octet. One of the oxygen atoms shares its two valence electrons with carbon which is represented by a double bond.

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

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



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

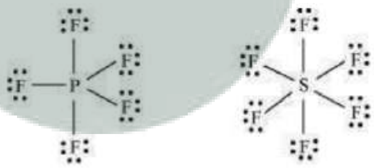
Limitations of the octet theory:

The following are the limitations of the octet rule:

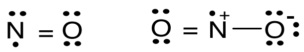
(a) The rule failed to predict the shape and relative stability of molecules.

(b) It is based upon the inert nature of noble gases. However, some noble gases like xenon and krypton form compounds such as XeF_2, KrF_2 etc.

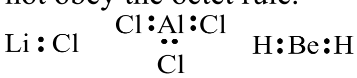
(c) The octet rule cannot be applied to the elements in and beyond the third period of the periodic table. The elements present in these periods have more than eight valence electrons around the central atom. For example: PF_5, SF_6, etc.



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



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



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

6. The favourable factors for the formation of the ionic bond are given below:

⇒ Ionization Enthalpy: It is defined as the amount of energy required to remove the most loosely bound electron from an isolated gaseous atom of an element. The lesser the ionization energy, the greater is the ease of the formation of a cation. Alkali metals and alkaline earth metals have low ionization energies and, therefore, they form metals cations very easily.

⇒ Electron Affinity: It is defined as the amount of energy released when an electron is added to an isolated gaseous atom of an element. The formation of an anion occurs with the addition of one or more electrons to an atom. The greater the energy released during this process, the easier will be the formation of an anion. Thus, high electron affinity of a non-metals favours the formation of an anion.

⇒ Lattice Energy: It is defined as the amount of energy released when cations and anions are brought from infinity to their respective equilibrium sites in the crystal lattice to form one mole of the ionic compound. The higher the magnitude of the lattice energy, the greater is the tendency of the formation of an ionic bond.

Thus, the lower ionization energy of an atom (generally a metal atom), the high electron affinity of another atom (generally a non-metal atom) and a high lattice energy of the compound formed to facilitate the formation of an ionic bond.

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

7. Cl : Be : Cl

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

BCl_3:

```
 Cl       Cl
 ..       ..
Cl : B : ClCl : B : Cl
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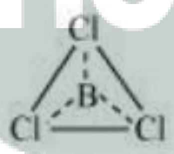
The central atom has no lone pair and there are three bond pairs. Hence, it is of the type AB_3. Hence, it is trigonal planar.

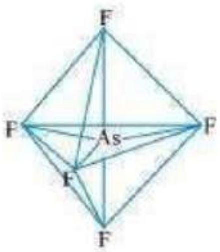


$SiCl_4$:

```
    Cl            Cl
    ..            ..
Cl : Si : ClCl : Si : Cl
    ..            ..
    Cl            Cl
```

The central atom has no lone pair and there are four bond pairs. Hence, the shape of $SiCl_4$ is tetrahedral being the AB_4 type molecule.

AsF_5:



The central atom has no lone pair and there are five bond pairs. Hence, AsF_5 is of the type AB_5. Therefore, the shape is trigonal bipyramidal.

H_2S:

H:S:H

The central atom has one lone pair and there are two bond pairs. Hence, H_2S is of the type AB_2E. The shape is Bent.

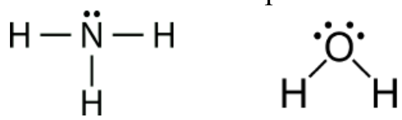
PH_3:

H:P:H
H

The central atom has one lone pair and there are three bond pairs. Hence, PH_3 is of the AB_3E type. Therefore, the shape is trigonal bipyramidal.

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

8. In NH_3, there is only one lone pair on nitrogen atom to repel the bond pairs whereas in H_2O there are two lone pairs on O-atom to repel the bond pairs as shown below:



Hence, the repulsion on bond pairs are greater in H_2O than in NH_3 and hence the bond angle is less.

Note More the no. of lone pairs on an atom, greater will be the repulsion and as a result, less will be the bond angle.

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

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

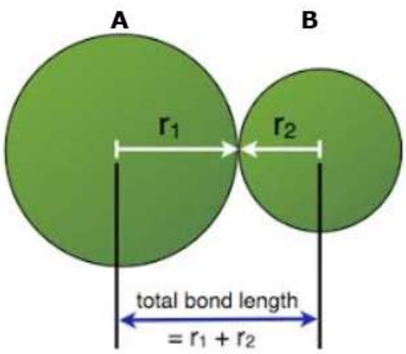
10. Define the bond length.

10. The bond length is defined as the average distance between the nuclei of two bonded atoms in a molecule.

Example:

Let us consider a diatomic molecule. The atoms in this molecule are always vibrating with respect to each other. The question of any fixed distance between them, therefore, does not arise. We can, however, think of an average distance between the nuclei of the two atoms bonded to each other. This is called bond length or bond distance.

Thus,



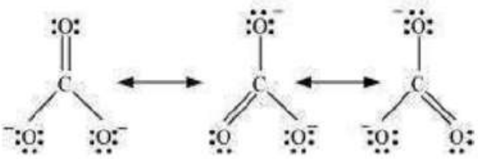
The bond length in a covalent molecule AB.

$R = r_1 + r_2$

Where R is the bond length
R_1 and r_2 are the covalent radii of atoms A and B respectively
Note: The bond length decreases with a multiplicity of the bond formed between the two atoms. Thus, the C Ξ C bond is shorter than C=C bond which in turn, is shorter than C—C bond.

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

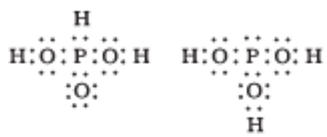
11. According to experimental findings, all carbon to oxygen bonds in CO_3^{2-} are equivalent.
Hence, it is inadequate to represent CO_3^{2-} ion by a single Lewis structure having two single bonds and one double bond.
Therefore, carbonate ion is described as a resonance hybrid of the following structures:



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

(1) (2)

12.

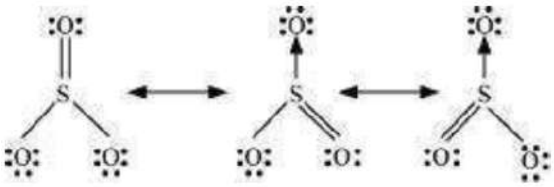


(1) (2)

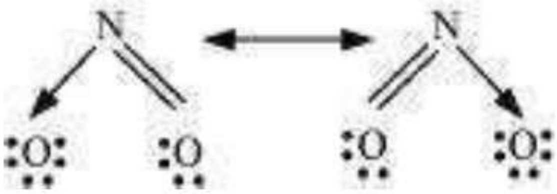
No, the structures 1 and 2 cannot be taken as canonical forms because the position of atoms has been changed.
While making resonating structures, following conditions must be followed:
⇒ We cannot change the position of any atom. Atoms must be in their own position.
⇒ We can only shift the bonds from one atom to another.
We can only change the position of electrons from one atom to another.

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

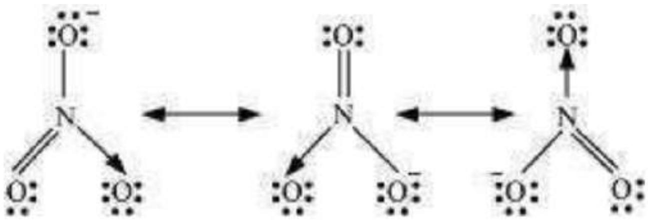
13. The resonance structures are:
(a) SO_3:



(b) NO_2

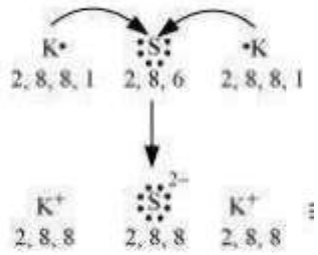
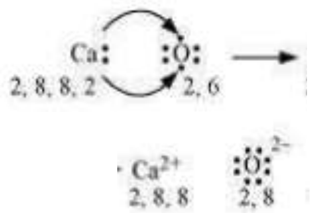
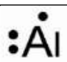


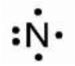
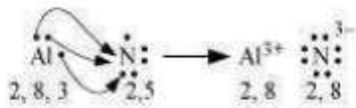
(c) NO_3^-



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.

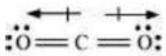
14.

Atoms	Atomic number	Electronic Configuration	No. of valence e-	Transfer of electrons	Cations and anions formed
K S	19 16	2,8,8,1 2,8,6	1⟶  6⟶ 	 (K requires one electron and S need 2 electrons to complete its octet)	$2K^+ S^{2-}$
Ca O	20 8	2,8,8,2 2,6	2⟶  6⟶ 	 (Oxygen requires two more elections to complete its octet and Ca has two electrons more than noble gas)	$Ca^{2+}O^{2-}$
Al	13	2,8,3	3⟶ 		

N	7	2,5	5⟶ 	 (Nitrogen requires three electrons to complete its octet and Al has 3 more electrons than the noble gas)	$Al^{3+}N^{3-}$

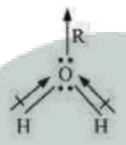
15. Although both CO_2 and H_2O are triatomic molecules, the shape of H_2O molecule is bent while that of CO_2 is linear. Explain this on the basis of dipole moment.

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



Resultant $\mu = 0$ D

H_2O, on the other hand, has a dipole moment value of 1.84 D (though it is a triatomic molecule as CO_2). The value of the dipole moment suggests that the structure of H_2O molecule is bent where the dipole moment of O–H bonds are unequal.



16. Write the significance/applications of dipole moment.

16. When a covalent bond is formed between two dissimilar atoms or we can say, heteronuclear molecules like HF, the electron pair shared between two atoms displaced towards the more electronegative atom(fluorine) since electronegativity of fluorine is far greater than that of hydrogen. The resultant covalent bond is a polar covalent bond. This is called polarisation.

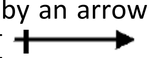
As a result of polarization, the molecule acquires the dipole moment which can be defined as the product of the magnitude of the charge and the distance between the centers of positive and negative charge.

Dipole moment (u) = Charge (Q) × distance of separation (r)

Significances/Applications:

⇒ The SI unit of Dipole moment is Debye. [1Debye = 3.35×10^{-30}m]

⇒ Dipole moment is a vector quantity. It has a magnitude and a direction.

⇒ It is represented by an arrow with tail on a negative charge and head pointing towards the positive charge. []

⇒ It is measure of the polarity of a bond. For example, in HF molecule, the dipole moment for the H^+F^- when an electron is completely transferred from hydrogen atom to fluorine atom is given by 4.48D.

⇒ The dipole moment also used to calculate the per cent ionic character.

$$\text{Per cent ionic character} = \frac{\text{Actual dipole moment of the bond}}{\text{Dipole momnet of the pure ionic bond}} \times 100$$

⇒ On the basis of dipole moment, we can also find out the structure of molecule

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

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

Electronegativity of any given element is not constant. It varies according to the element to which it is bound. It is not a measurable quantity. It is only a relative number. On the other hand, electron gain enthalpy is the enthalpy change that takes place when an electron is added to a neutral gaseous atom to form an anion. It can be negative or positive depending upon whether the electron is added or removed. An element has a constant value of the electron gain enthalpy that can be measured experimentally.

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

18. As we know that, if the bond is formed by the mutual sharing of electrons between the atoms, then it is called the covalent bond.

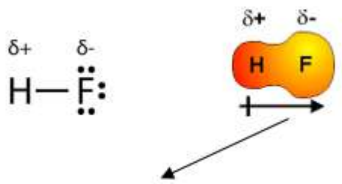
A covalent bond formed between two dissimilar atoms, the bond formed is said to be a polar covalent bond. A polar covalent bond is more stable than a pure covalent bond

The bond is formed between two dissimilar atoms means both having different electronegativities and which means the bond pair of electrons is not shared equally. The atom which is having high electronegativity, its bond pair shift towards the nucleus. As a result, the atom with higher electronegativity generally has a greater tendency to attract the electron towards itself.

This unsymmetric distribution of electrons leads to partial charge separation. The electronegative atom acquires partial negative charged and the other atom acquires a slightly positive charge.

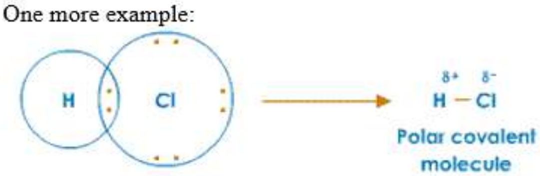
For example, HF, which contains a polar covalent bond (having different electronegativities). As we know that fluorine is more electronegative atom than the hydrogen. The electron pair shared between hydrogen and fluorine atoms, is, therefore, remains closer to the fluorine atom. As a result, fluorine acquires a partial negative charge and hydrogen acquires a partial positive charge.

To denote the partial charge on an atom, we use the symbol "δ".



Electron pair is shifting towards more electronegative atom, i.e., F

One more example:

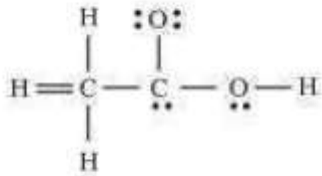


Transfer of electron pair towards more electronegative chlorine.

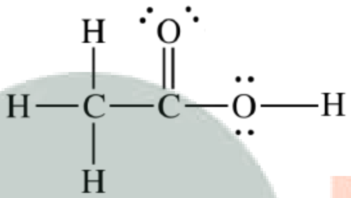
19. Arrange the bonds in order of increasing ionic character in the molecules: LiF, K_2O, N_2, SO_2 and ClF_3.

19. The ionic character in a molecule is dependent upon the electronegativity difference between the constituting atoms. The greater the difference, the greater will be the ionic character of the molecule.
On this basis, the order of increasing ionic character in the given molecules is
$N_2 < SO_2 < ClF_3 < K_2O < LiF$.

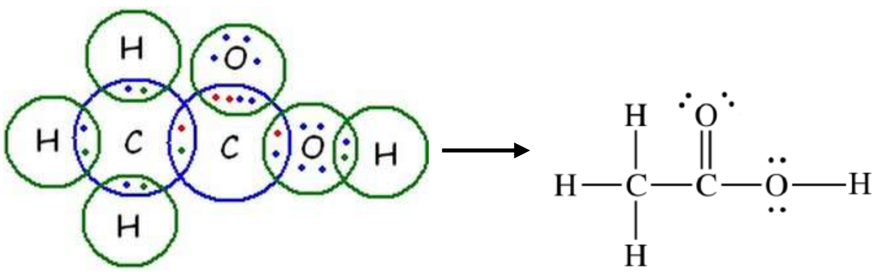
20. The skeletal structure of CH3COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.



20. The correct structure of CH_3COOH is:

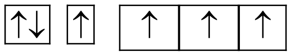


Explanation:
The atomic number of C is 6 and no. of valence electrons is 4.
The atomic number of O is 8 and no. of valence electrons is 6.
The atomic number of H is 1 and no. of valence electron is 1.
Hence, the total number of valence electrons in the element is:
Two carbon (2 × 4 valence electrons) + two oxygen (2 × 6 valence electrons) + 4 hydrogen (4 × 1 valence electron) = 24 total valence electrons. Hence, the skeletal structure of CH_3COOH is:



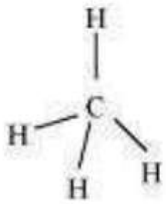
21. Apart from tetrahedral geometry, another possible geometry for CH_4 is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH_4 is not square planar?

21. Electronic configuration of carbon atom:
6C: $1s^2\ 2s^2\ 2p^2$
In the excited state, the orbital picture of carbon can be represented as:



1s 2s $2p_x$ $2p_y$ $2p_z$

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



For a square planar shape, the hybridization of the central atom has to be dsp^2.
However, an atom of carbon does not have d-orbitalsto undergo dsp^2 hybridization. Hence, the structure of CH_4 cannot be square planar.
Moreover, with a bond angle of 90° in square planar, the stability of CH_4 will be very less because of the repulsion existing between the bond pairs. Hence, VSEPR theory also supports a tetrahedral structure for CH_4.

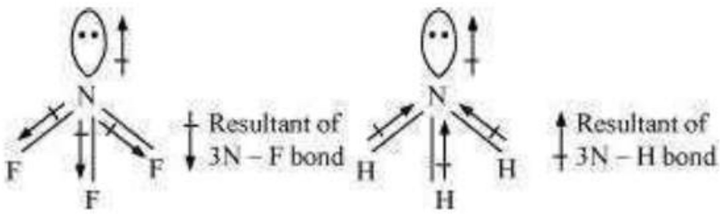
22. Explain why BeH_2 molecule has a zero dipole moment although the Be–H bonds are polar.

22. The Lewis structure for BeH_2 is as follows:
H:Be:H
Here there is no lone pair at the central atom (Be) and there are two bond pairs on either side. Hence, BeH_2 is of the type AB_2 having linear structure.

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

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

23. In both molecules i.e., NH_3 and NF_3, the central atom (N) has a lone pair electron and there are three bond pairs. Hence, both molecules have a pyramidal shape. Since fluorine is more electronegative than hydrogen, it is expected that the net dipole moment of NF_3 is greater than NH_3. However, the net dipole moment of NH_3 (1.46 D) is greater than that of NF_3 (0.24 D).
This can be explained on the basis of the directions of the dipole moments of each individual bond in NF_3 and NH_3. These directions can be shown as:



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

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

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

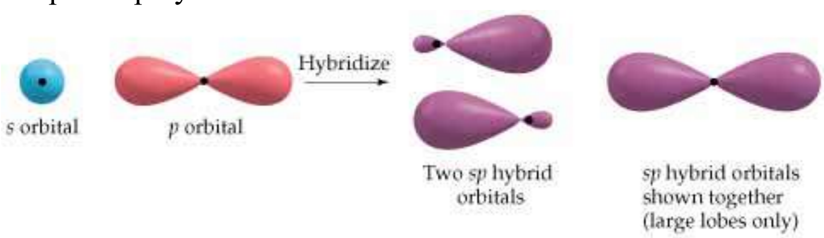
24. Hybridization can be defined as an intermixing of a set of atomic orbitals of slightly different energies, thereby forming a new set of orbitals having equivalent energies and shapes.

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

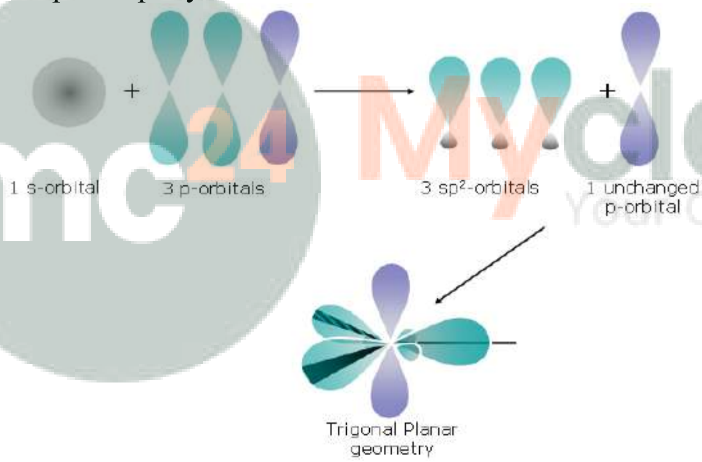
These hybrid orbitals have minimum repulsion between their electron pairs and thus, are more stable. Hybridization helps indicate the geometry of the molecule.

Shape of sp hybrid orbitals:

Shape of sp hybrid orbitals:



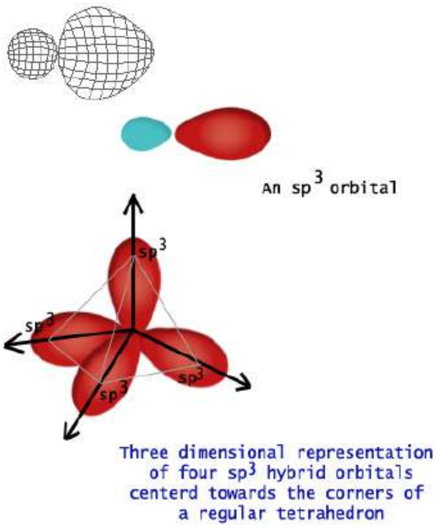
Shape of sp^2 hybrid orbitals:



sp^2 hybrid orbitals are formed because of the intermixing of one s-orbital and two p orbitals.

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

The shape of sp^3 hybrid orbitals:

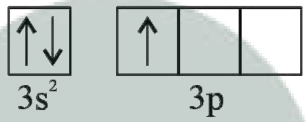


Four sp^3 hybrid orbitals are formed by intermixing one s-orbital with three p-orbitals.

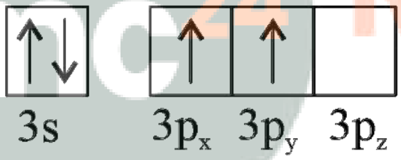
25. Describe the change in hybridisation (if any) of the Al atom in the following reaction.

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

25. The valence orbital picture of aluminium in the ground state can be represented as:



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

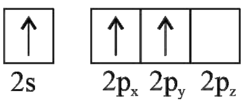


Hence, it undergoes sp^2 hybridization to give a trigonal planar arrangement (in $AlCl_3$). To form $AlCl_4^-$, the empty 3pz orbital also gets involved and the hybridization changes from sp^2 to sp^3. As a result, the shape gets changed to tetrahedral.

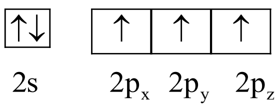
26. Is there any change in the hybridisation of B and N atoms as a result of the following reaction?

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

26. Boron atom in BF3 is sp^2 hybridized. The excited state of boron can be represented by:



While the nitrogen atom in NH3 is sp^3 hybridized. The excited state can be shown as:



On completion of this reaction, an adduct $F_3B.NH_3$ is formed as the hybridization of Boron changes to sp^3, while the hybridization of nitrogen remains intact.

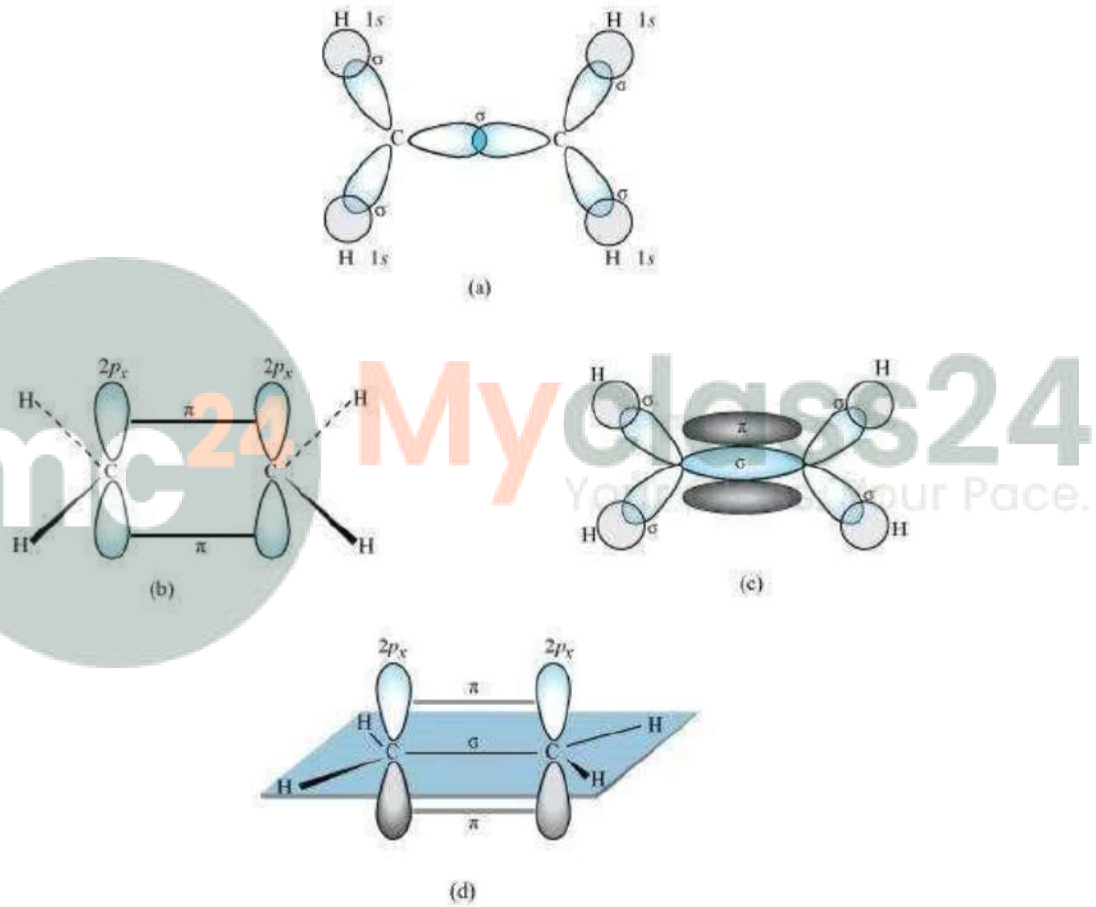
27. Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in C_2H_4 and C_2H_2 molecules.

27. C_2H_4 :

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

$_6C = 1s^2 2s^1 2p_x^1 2p_y^1 2p_z^1$

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



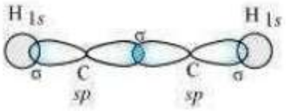
C_2H_2 :

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

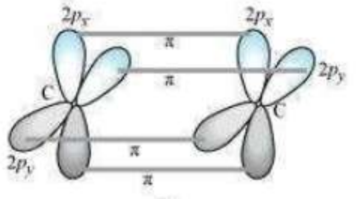
One sp orbital of each carbon atom overlaps with the other along the internuclear axis forming a C–C sigma bond. The second sp orbital of each C–atom overlaps a half-filled 1sorbital to form a σ bond.

The two unhybridized 2p-orbitals of the first carbon undergo sidewise overlap with the 2p orbital of another carbon atom, thereby forming two pi (π) bonds between carbon atoms.

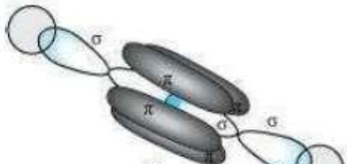
Hence, the triple bond between two carbon atoms is made up of one sigma and two π-bonds.



(a)



(b)



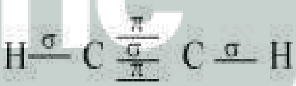
(c)

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

(a) C_2H_2

(b) C_2H_4

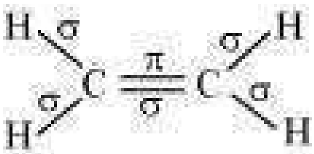
28.



There are three sigma and two pi-bonds in C_2H_2.

Explanation: A single bond is a result of the axial overlap of bonding orbitals. Hence, it contributes a sigma bond. A multiple bond (double or triple bond) is always formed as a result of the sidewise overlap of orbitals. A pi-bond is always present in it. A triple bond is a combination of two pi-bonds and one sigma bond.

B. C_2H_4



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

Explanation: A single bond is a result of the axial overlap of bonding orbitals. Hence, it contributes a sigma bond. A multiple bond (double or triple bond) is always formed as a result of the sidewise overlap of orbitals. A pi-bond is always present in it. A triple bond is a combination of two pi-bonds and one sigma bond.

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 2px (c) 2py and 2py (d) 1s and 2s.

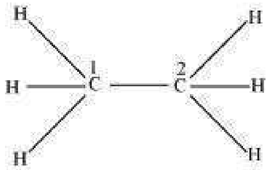
29. 2py and 2py orbitals will not a form a sigma bond. Taking x-axis as the internuclear axis, 2py and 2py orbitals will undergo lateral overlapping, thereby forming a pi(π) bond.

30. Which hybrid orbitals are used by carbon atoms in the following molecules?
$CH_3–CH_3$; (b) $CH_3–CH=CH_2$; (c) CH_3-CH_2-OH; (d) CH_3-CHO (e) CH_3COOH

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

A. $CH_3–CH_3$;

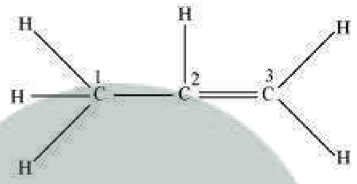
Ans:



Both C_1 and C_2 are sp_3 hybridized.

B. $CH_3–CH=CH_2$;

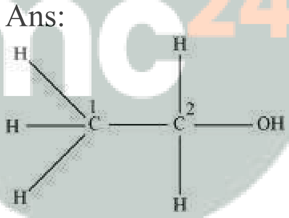
Ans:



C_1 is sp3 hybridized, while C_2 and C_3 are sp_2 hybridized.

C. CH_3-CH_2-OH;

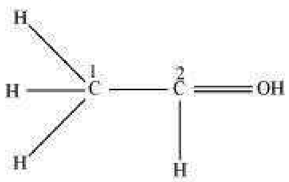
Ans:



Both C_1 and C_2 are sp^3 hybridized

D. CH_3-CHO

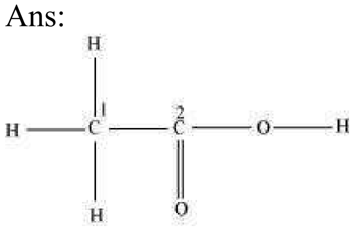
Ans:



C_1 is sp3 hybridized and C_2 is sp^2 hybridized.

E. CH_3COOH

Ans:



C_1 is sp3 hybridized and C_2 is sp^2 hybridized.

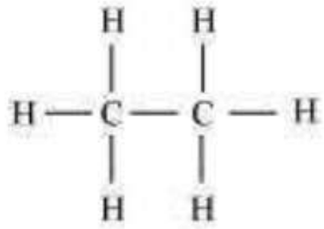
31. What do you understand by bond pairs and lone pairs of electrons? Illustrate by giving one example of each type.

31. When two atoms combine by sharing their one or more valence electrons, a covalent bond is formed between them.

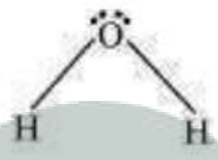
The shared pairs of electrons present between the bonded atoms are called bond pairs.

All valence electrons may not participate in bonding. The electron pairs that do not participate in bonding are called lone pairs of electrons.

For example, in C_2H_6 (ethane), there are seven bond pairs but no lone pair present.



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



32. Distinguish between a sigma and a pi bond.

32. The differences of each type are listed below:

Sigma Bond:

1. The covalent bond formed by the overlap of atomic orbitals along the internuclear axis is called sigma bond

2. The overlapping orbitals are oriented along the internuclear axis.

3. s as well as p orbitals can form this type of bonds.

4. It is stronger than pi bond

Pi Bond:

1. The covalent bond formed by the lateral overlap of two p orbitals which are mutually parallel but oriented perpendicular to the internuclear axis is called the pi bond.

2. The overlapping orbitals are oriented perpendicular to the inter nuclear axis.

3. The bond is not rotationally symmetrical around the internuclear axis.

Only p orbitals can form this type of bond.

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

33. Let us assume that two hydrogen atoms (A and B) with nuclei (N_A and N_B) and electrons (e_A and e_B) are taken to undergo a reaction to form a hydrogen molecule.

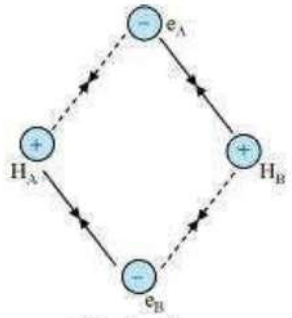
When A and B are at a large distance, there is no interaction between them. As they begin to approach each other, the attractive and repulsive forces start operating.

Attractive force arises between:

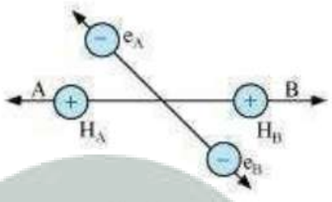
(a) Nucleus of one atom and its own electron i.e., $N_A - e_A$ and $N_B - e_B$.

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

Repulsive force arises between:
(a) Electrons of two atoms i.e., $e_A - e_B$.
(b) Nuclei of two atoms i.e., $N_A - N_B$.
The force of attraction brings the two atoms together, whereas the force of repulsion tends to push them apart.



Attractive Forces



Repulsive Forces

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

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

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

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

35. The electronic configuration of Beryllium is $1s^2 2s^2$.
The molecular orbital electronic configuration for Be_2 molecule can be written as:
$\sigma_{1s}^2\ \sigma_{1s}^{*2} \sigma_{2s}^2\ \sigma_{2s}^{*2}$
Hence, the bond order for Be_2 is $\frac{1}{2}(N_b - N_a)$
Where,
N_b = Number of electrons in bonding orbitals
N_a = Number of electrons in anti-bonding orbitals

$\therefore$ Bond order of Be_2 $\frac{1}{2}(4-4) = 0$

A negative or zero bond order means that the molecule is unstable. Hence, Be_2 molecule 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)

36. The electronic configuration of oxygen molecule can be written as:

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

Here number of bonding electrons is 8 (N_a) and number of antibonding electrons is 4 (N_b)

Therefore, bond order can be found out by,

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

= 1/2(8-4)

= 2

The electronic configuration of O_2^+ can be written as:

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

$N_b = 8$

$N_a = 3$

Bond order is ½(8-3)

= 2.5

The electronic configuration of O_2^- will be:

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

$N_b = 8$

$N_a = 5$

Bond order is ½ (8-5)

=1.5

The electronic configuration of O_2^{2-} will be:

$N_b = 8$

$N_a = 6$

Bond order is ½ (8-6)

=1

Bond dissociation energy is directly proportional to bond order. Thus, the higher the bond order, the greater will be the stability. On this basis, the order of stability is:

$O_2^+ > O_2 > O_2^- > O_2^{2-}$

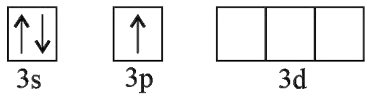
37. Write the significance of a plus and a minus sign shown in representing the orbitals.

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

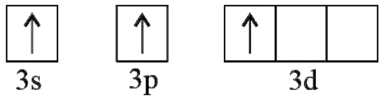
38. Describe the hybridisation in case of PCl_5. Why are the axial bonds longer as compared to equatorial bonds?

38. The ground state and excited state outer electronic configurations of phosphorus (Z = 15) are:

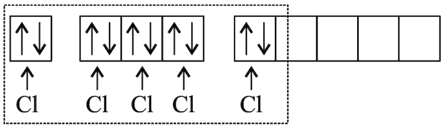
Ground state:



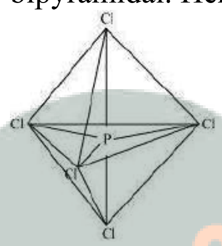
Excited state:



Phosphorus atom is sp^3d hybridized in the excited state. These orbitals are filled by the electron pairs donated by five Cl atoms as: PCl_5.



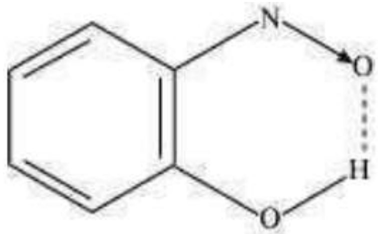
The five sp^3d hybrid orbitals are directed towards the five corners of the trigonal bipyramidal. Hence, the geometry of PCl_5 can be represented as:



There are five P–Cl sigma bonds in PCl_5. Three P–Cl bonds lie in one plane and make an angle of 120° with each other. These bonds are called equatorial bonds.
The remaining two P–Cl bonds lie above and below the equatorial plane and make an angle of 90° with the plane. These bonds are called axial bonds.
As the axial bond pairs suffer more repulsion from the equatorial bond pairs, axial bonds are slightly longer than equatorial bonds.

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

39. A hydrogen bond is defined as an attractisve force acting between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of a different molecule (may be of the same kind).
Due to a difference between electronegativities, the bond pair between hydrogen and the electronegative atom gets drifted far away from the hydrogen atom. As a result, a hydrogen atom becomes electropositive with respect to the other atom and acquires a positive charge.
$4\delta - X^{\delta-}$ …….$H^{\delta+}$ $-X^{\delta-}$ …….$H^{\delta+}$ $-X^{\delta-}$
The magnitude of H-bonding is maximum in the solid state and minimum in the gaseous state.
There are two types of H-bonds:
(i) Intermolecular H-bond e.g., HF, H_2O etc.
(ii) Intramolecular H-bond e.g., o-nitrophenol



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

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

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

If N_a is equal to the number of electrons in an anti-bonding orbital, then N_b is equal to the number of electrons in a bonding orbital.

Bond Order:

$\frac{1}{2}(N_b - N_a)$

If $N_b > N_a$, then the molecule is said be stable. However, if $N_b \le N_a$, then the molecule is considered to be unstable.

Bond order of N_2 :

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

Number of bonding electrons, N_b = 10

Number of anti-bonding electrons, N_a = 4

Bond order of nitrogen molecule= ½ (10-4)=3

Bond order of O_2:

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

Number of bonding electrons, N_b = 8

Number of anti-bonding electrons, N_a = 4.

Bond order of Oxygen molecule = ½ (8-4) = 2

Bond order of O_2^+ :

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

Number of bonding electrons, N_b = 8

Number of anti-bonding electrons, N_a = 3

Bond order of O_2^- = ½ (8-3) = 2.5

Bond order of O_2^- :

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

Number of bonding electrons, N_b = 8

Number of anti-bonding electrons, N_a = 5

Bond order of O_2^- = ½ (8-5) = 1.5