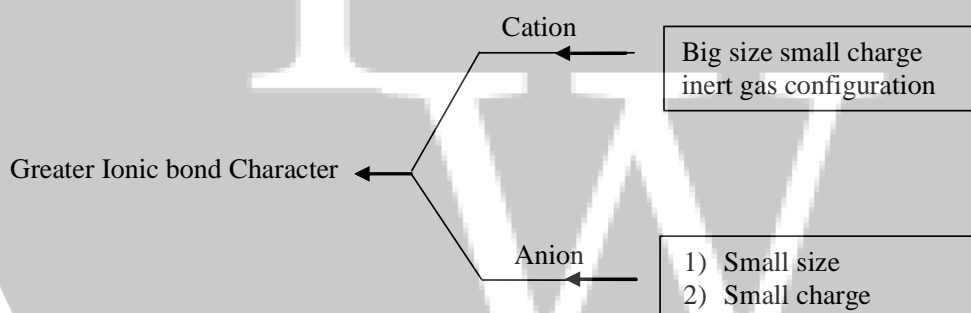# CHAPTER 04 CHEMICAL BONDING & MOLECULAR STRUCTURE

- An atom (or) group of atoms having independent existence in nature is defined as molecule.
- Force of attraction that holds two atoms (or) oppositely charged ions in molecule is chemical bond.
- An atom must have eight electrons is its outermost energy level for its stability is called as octet rule.
- The number of electrons an atom loses (or) gains in an ionic bond formation is called electro valency

- According to Fajan's Rules

- The no. of electrons an atom of the element losses (or) gains is known as electrovalency of that element.
- Favourable conditions for cation formation.
  1) Low I.P
  2) Low charge on ion
  3) Big atomic size
  4) Inert gas configuration
- Favourable conditions for cation formation.
  1) High E.N & E.A
  2) Small atomic size
  3) Low charge on ion

Cation ← Big size small charge inert gas configuration

Greater Ionic bond Character ←

Anion ← 1) Small size 2) Small charge

- LiCl is more covalent in nature and so dissolved in non-polar solvents like alcohol (or) ether.

- Ions which have 18 electrons $\left[ns^2 np^6 nd^{10} \right]$ in their outermost main energy level are stable and it is called as pseudo inert gas configuration (or) nickel group configuration.

- The energy liberated when an ionic crystal is formed from an assembly of isolated gaseous ions is called lattice energy

- Lattice energy

$$\left(\cup \right) = \left[-No \frac{AZ^+ .Z^- e^2}{r} + No \frac{Be^2}{r^n} \right]$$

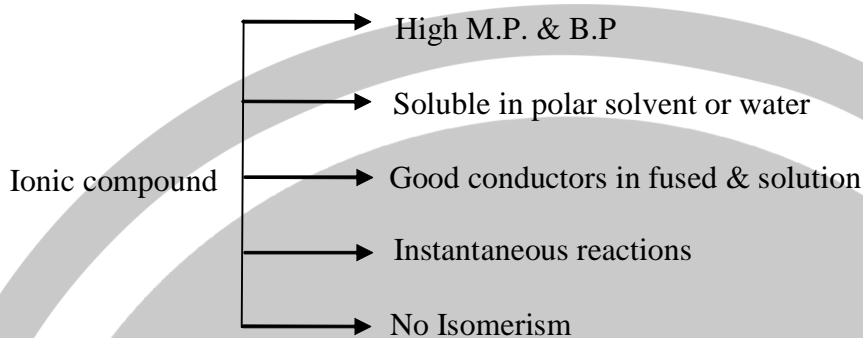- $$\cup = \frac{AZ^+ Z^- e^2 No}{r_0} \left[1 - \frac{1}{n} \right]$$

$$\left(Born - Lande\ equation \right)$$
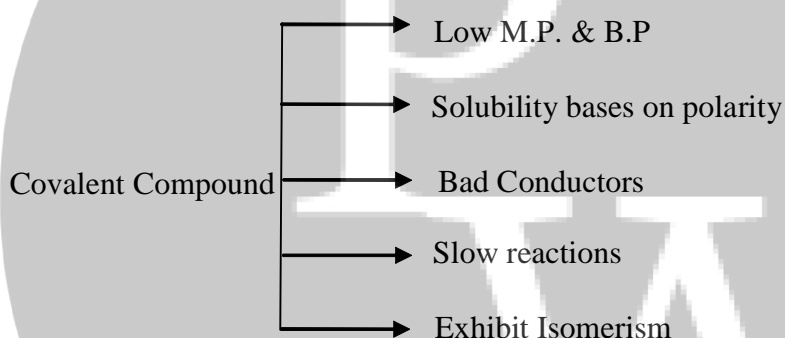
Coordination number of Na^+ and Cl^- in NaCl crystal is (6:6)

Coordination number of Cs^+ and Cl^- in NaCl crystal is (8:8)

- The total amount of energy change in a reaction remains the same whether the reaction take place in one (or) several stages is known as Hess law of constant heat summation.

- The maximum number of nearest oppositely charged ions surrounding any particular ion in an ionic
- Crystal is called the co-ordination number of that ion.

- Crystal structure of NaCl is face centered cubic lattice structre.
- Crystal structure of CsCl is body centered cubic lattice structure.

- **PROPERTIES OF IONIC COMPOUNDS**

Ionic compound
→ High M.P. & B.P
→ Soluble in polar solvent or water
→ Good conductors in fused & solution
→ Instantaneous reactions
→ No Isomerism

- Properties of covalent compounds

Covalent Compound
→ Low M.P. & B.P
→ Solubility bases on polarity
→ Bad Conductors
→ Slow reactions
→ Exhibit Isomerism

- The number of electrons an atom of an element contributes in the formation of covalent compound is called covalency of the element.
- The charge possible that an atom in a molecule would have if all the atoms had the same electronegativity is called formal charge.

Formal charge

$$\left(Q_f\right) = \left[N_A - N_M\right]$$

$$= \left[N_A - N_{L.P.} - \frac{1}{2}N_{B.P.}\right]$$

Where
N_A = No. of electrons in the valence shell in free atom

N_M = No. of electrons belonging to the atom in molecule
$N_{L.P.}$ = No. of electrons in unshared pairs is lone pairs.
$N_{B.P.}$ = No. of electrons in bond pairs.

- The molecule with a smallest formal charges on the atoms, is considered as lower energy structure of a molecule.

- V.S.E.P.R. theory

Number of bond pairs	Number of lone pairs	Formula	Molecular Shape	Hibridi zation	Bond Angle	Example
2	0	AB_2	Linear	Sp	180^0	$BeCl_2, BeF_2, C_2H_2$
3	0	AB_3	Plane Triangle	sp^2	120^0	BF_3, BCl_3
4	0	AB_4	Tetrahedron	sp^3	$109^0, 28^1$	CH_4, NH_4^+, CCl_4
3	1	AB_3E	Trigonal pyramid	sp^3	$107^0, 102^0, 30^1$	NH_3, H_3O^+, NF_3
2	2	AB_3E_2	Angular V-Shaped	sp^3	$104^0 40^1, 102^0$	H_2O, F_2O
5	0	AB_5	Trigonal Bipyramid	sp^3d	$120^0, 90^0$	PCl_5, PF_5
2	3	AB_2E_3 $(or) A_3E_3$	Linear	sp^3d	180^0	XeF_2, I_3^-
6	0	AB_6	Octahedron	sp^3d^2	90^0	SF_6

A = Central atom in the compound
B = atom linked to the central atom
E = lone pairs of electron.

- A covalent bond formed between two atoms the electron clouds concentrated along the internuclear axis having a cylindrical symmetry is known as a sigma bond.
- A covalent bond formed by a side wise overlap of p orbitals (or) p and d orbitals of atoms, in which the electron clouds are present above and below the internuclear axis is known as pi-bond.
- Strength of bonds $\sigma_{p-p} > \sigma_{s-p} > \sigma_{s-s} > \pi_{p-p}$
- p_x and p_y - orbitals will not overlap because their orientations and symmetries are not same.
- The intermixing of atomic orbitals of almost equal energies of an atom and their redistribution into equal number of identical orbitals.
- In PCl_5 the axal bonds of P-Cl are longer than equatorial bonds because the repulsions of the axial bond pairs stronger than equatorial bond pairs.

- **<u>Molecular orbital theory</u> :-**
  Region around the Nuclei of the bonded atoms in a molecule where the probability of finding electrons is maximum

Anti bonding orbitals

Non – bonding orbitals

AO's   AO's
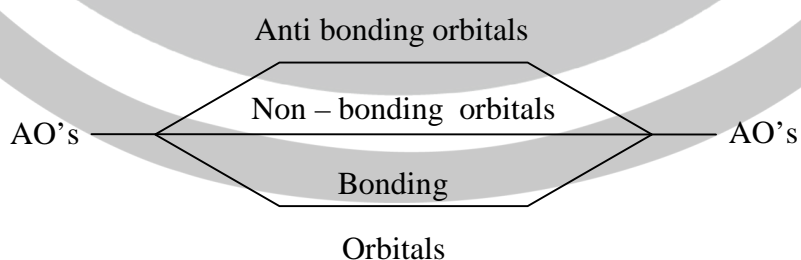
Bonding

Orbitals

Order of energies :-

- **Bonding orbitals < Non - bonding orbitals < anti - bonding orbitals**

Sl. No.	AO of 1st atom	AO of 2nd atom	Designation of bonding Mo	Designation of Antibonding Mo
1	S	S or P_x	σ	σ^*
2	P_x	S or P_x having end on over lap	σ	σ^*
3	P_y	P_y side wise over lap	π	πP_y^*
4	P_z	P_z side wise over lap	π	πP_z^*

- Sequence of energy levels of molecular orbitals
  a) For lighter elements (B,C or N)

$$\sigma 1S < \sigma 1S^* < \sigma 2S < \sigma^* 2S < \begin{Bmatrix} \pi 2P_y \\ \pi 2P_z \end{Bmatrix} < \sigma 2P_x < \begin{Bmatrix} \pi^* 2P_y \\ \pi^* 2P_z \end{Bmatrix} < \sigma^* 2P_x$$

  b) For oxygen and heavier elements :

$$\sigma 1S < \sigma 1S^* < \sigma 2S < \sigma^* 2S < \sigma 2P_x < \begin{Bmatrix} \pi 2P_y \\ \pi 2P_z \end{Bmatrix} < \begin{Bmatrix} \pi^* 2P_y \\ \pi^* 2P_z \end{Bmatrix} < \sigma^* 2P_x$$

- $$\text{Bond order} = \frac{\text{Bonding electron} - \text{Antibonding electrons}}{2}$$

- Hydrogen bond formed between the atoms of the same molecule is known as intramolecular hydrogen bond.
- Hydrogen bond formed betwen two different polar molecules (the same (or) different substance) is known as inter molecular hydrogen bond.
- As the temperature decreases thermal energy of gases decreases. The inter molecular repulsions decreases while attractive forces increases and therefore gases changes to liquid.
- **Chemical Bond.** It is the force of attraction which holds or binds the constituent atoms together in a molecule.
- **Octet Rule.** An atom tends to gain, lose or share electrons during molecule formation such that there are eight electrons surrounding it.
- **Valence Electrons.** The electrons present in outermost shell are called valence electrons.
- **Valence Bond Theory.**

(i) Covalent bond is formed by overlapping of half filled atomic orbitals.
(ii) Lowering of energy must take place during covalent bond formation. Attractive forces should be more than repulsive forces.
(iii) Strength of bond depends upon extent of overlapping. Greater the extend of overlapping stronger will be the bond.

- **Ionic** Bond. It is the electrostatic force of attraction which holds the array of oppositely charged ions together. It is formed by complete transfer of electrons from one atom to another.
- **Covalent Bond.** It is the force of attraction which arises due to the mutual sharing of electrons between two atoms. This definition is according to electronic theory proposed by G. N. Lewis in 1916.
- **Lone Pair of Electrons.** The pair of electrons which do not take part in bond formation are called !one pair of electrons or non-bonding pair or unshared pairs.

+ **Double Bond.** When a bond is formed by sharing two electrons of each of the atom, it is called i double bond, *e.g.,*: $O :: O :$ or $: O = O:$

+ **Triple Bond.** It is formed by sharing of three electrons of each atom, *e.g.,* $:N \vdots N:$ or $: N \equiv N:$

+ **VSEPR Theory (Valence-Shell Electron-Pair RepulsionTheory).** According to this theory, the direction of the bonds around an atom in a molecule depends upon the total number of electron pairs (bonding as well as non-bonding) in the valence shell of an atom. Since the electron pairs repel each other, therefore that geometrical arrangement will be favoured in which repulsion is minimum.

+ Covalent Bond. It is formed by overlapping of atomic orbitals having unpaired electrons according to theory based on modern wave mechanical model.

+ **Bond Length.** It is defined as the average distance between the centres of the nuclei of the two bonded atoms in a molecule corresponding to minimum energy and maximum stability. It is measured in pm (picomctre).

+ **Bond Energy.** The bond energy of a particular bond in a compound is defined as the average amount of energy produced when one mole of bonds is formed from isolated gaseous atoms or the amount of energy required when 1 mole of bonds is broken to get isolated gaseous atoms. It is measured in kilo Joules.

+ **Bond Angle.** It is the average angle between two (molecular orbitals) bonds meeting at the same nucleus in a poly-atomic molecule. It is measured in degrees and minutes.

+ **Hybridization.** It is the phenomenon of intermixing of atomic orbitals of slightly different energies to give rise to new hybridized orbitals having equivalent energy and identical shapes. Hybridized orbitals are more stable than atomic orbitals.

+ **sp^3 Hybridization. If** one s and three p-orbitals are intermixed or hybridized, this is known as sp^3-hybridization. Four equivalent orbitals will he formed which are called sp^3 hybridized orbitals which are directed along the four corners of a tetrahedron with bond angle 109° 28', *e.g.,* in CH_4, carbon undergoes sp^3 hybridization and it has tetrahedral shape.

+ **sp^3d Hybridisation.** When one *s,* three *p* and one d-orbitals intermix together and give rise to five sp^3d hybridised orbitals. The shape of molecule is *trigonal bipyramidal.* Bond angle is 90°, 120°, e.g., PF_3, PCl_3 involves . sp^3dhybridisation.

+ **$sp^3 d^2$ Hybridisation.** When one *s,* three *p* and two d-orbitals intermix to give rise to six sp^3d^2-hybridised orbitals. The shape of molecule having.$sp^3 d^2$ hybridisation will be octahedral with bond angle 90°, *e.g.,*$SF_6[CrF_6]^{2-}$ involve sp^3d^2 hybridisation and possess octahedral shape.

+ **sp^2d or dsp^2 Hybridisation.** When one *s,* two p and one d-orbitals intermix to give rise to four sp^2d or *dsp^2* hybridised orbitals having square planar shape with bond angle 90°, *e.g.,* [Ni $(CN)4]^{2-}$ undergo dsp^2 hybridisation and are square planar in shape.

+ **sp^2 Hybridization.** If one *s* and two p-orbitals are intermixed, it is called sp^2-hybridization. Three equivalent orbitals are called sp^2 hybridized orbitals which will be directly along the corners of equilateral triangle having bond angle 120°, *e.g.,* in BF_3, 'B' atom undergoes sp^2 hybridization and its shape is a planar.

+ **sp Hybridization.** If one s and one *p* orbital intermix, it is called sp-hybridization. The two *sp* hybridized orbitals are tar-apart. In BeF_2. Be atom undergoes sp-hybridization and it has linear shape with bond angle 180°.

+ **Sigma Bond.** The bond which is formed by the head on overlapping of half-filled atomic orbitals along the inter-nuclear axis, is called sigma bond.

+ **Pi Bond.** The bond which is formed by the sideways or lateral overlapping of half-filled atomic orbitals in a direction perpendicular to the inter-nuclear axis, is called pi bond.

+ **The Coordinate Covalent Bond.** It is a special type of covalcnt bond in which shared pair of electrons is supplied only by one of the atoms forming the bond. The atom which supplies the electrons is called donor while the other atom which only uses the shared pair of electrons is known as acceptor. It is shown by arrow-head (®)

+ **Polar Covalent Bond,** if covalent bond is formed between two dissimilar atoms having difference in electronegativity. then the shared pair of electrons is more towards more electronegative atom which acquired d^- charge and other atom acquires $d^{+|}$ charge. This type of bond is called polar covalent bond. The compounds having polar covalent bond form ions in aqueous solution. It is also called ionic character of the covalent bond.

+ **Dipole Moment.** Dipole moment of a molecule is the product of its net positive or negative charge and distance between the atoms.

  Mathematically, $m = e \times d$, where 2 = net + ve or -ve charge, d is distance between the atoms. Its unit is Debye, where 1 Debye = 10^{18} e. u. cm. Its S.I. unit is Cm (Coulomb meter).

+ **Ionic Solids.** The solids in which there is three dimensional array of cations and anions are called ionic solids, *e.g.,* LiCl, LiF, NaCl, NaF, etc.

+ **Covalent Solids.** In covalent solids, atoms are linked to one another by covalent bonds forming a giant network, e.g., SiC (silicon carbide), diamond are examples of covalent solids.

+ **Metallic Bond.** It is formed by the simultaneous force of attraction between the Kernels (+vely charged atoms) and mobile electrons which hold the metal atoms together. All metals and alloys have metallic bond.

+ **Resonance.** In case of some of the molecules, all of their properties can't be explained by a single structure. The molecule can be represented by two or more structures. These structures are called canonical forms. The molecule is said to be resonance hybrid of all these structures. These structures have no physical reality in the sense that they cannot be prepared in the laboratory. They exist in imagination. This phenomenon of existing in more than one structure is called resonance, e.g.,.

Resonating Structures of Benzene
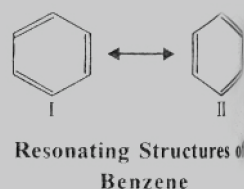
+ **Resonance Energy.** It is defined as difference in energy of actual structure of compound and most stable resonating structure, *e.g.,* benzene has resonance energy 150,3 kJ mol^{-1} or 36 kcal rnol^{-1}.

+ **Predicting the Hybrid State of Atom in Different Species:**

  (i) Write the valence electrons of central atom (V).

  (ii) Add number of surrounding atoms except oxygen (S A).

  (iii) If there is positive charge on species, subtract the charge, if there is -ve charge, add the charge.

  (iv) Divide the sum by 2 to get value of X.

  $$X = \frac{1}{2}(V + SA \pm E)$$

Value of X	2	3		4	5	6	7
HybridSiate of Central Atom	sp			sp^2	sp^3	sp^3d	sp^3d^2	sp^3d^2

Note. The above formula is rot applicable to predict hybrid state of metal in complexes and species

+ **Molecular Orbital Theory.** Molecular orbital gives electron probability distribution around a group of nuclei in a molecule. They are filled in the same way as atomic orbitals. Molecular orbitals are formed by linear combination of atomic orbitals.

+ **Bonding molecular orbital.** A molecular orbital that *is* formed by addition overlap (*i.e.,* when the lobes of atomic orbitals overlap with the same sign) of two atomic orbitals is known as *bonding molecular orbital.* It is represented as

$$y_{AB} = y_A + y_B$$

Its energy is lower than the atomic orbitals from which it is formed. It favours bonding.

+ **Anti-bonding molecular orbital.** A molecular orbital that is obtained by the subtraction overlap *(i.e.,* when the lobes of atomic orbitals overlap with the opposite sign) of two atomic orbitals is known as *anti-bonding molecular orbital.* It is represented as

$$y^*_{AB} = y_A - y_B$$

It is energy is higher than the atomic orbitals from which it is formed. It does not favour bonding.

+ **Bond order. It** is defined as half of the difference between number of electrons in bonding and anti-bonding orbitals, *i.e.,*

Bond order = 1/2 (N_b - N_a) where N_b are number of electrons in bonding orbitals and N_a are number of electrons in anti-bonding orbitals.

Bond order helps in estimating stability of atom.

+ **Relationship between electronic configuration and molecular behaviour:**

    (a) If Nb is greater than Na, the molecule is stable.

    (b) The molecule is unstable if Na, is greater than Nb.

    (c) The molecule is also unstable if Na is equal to Nb because anti-bonding effect is stronger than bonding effect.

+ **Sigma** (s) **molecular orbitals.** A molecular orbital which is formed from the overlap of two.s atomic orbitals or head to head overlap of one *s* and p-atomic orbitals or head to head overlap of two p-atomic orbitals, is known as *sigma molecular orbital.*

+ **Pi** (p) **molecular orbitals.** A molecular orbital which is formed by lateral overlap of two parallel *p*-orbitals is known as *pi* (p) *molecular orbital.*

+ **Paramagnetism.** The presence of one or more unpaired electrons is the cause of paramagnetism. A paramagnetic substance is attracted by an applied magnetic field.

+ **Diamagnetism.** The absence of unpaired electrons is the cause of diamagnetism. A diamagnetic substance will be repelled by an applied magnetic field.

+ **Ferromagnetism.** It is due to large number of unpaired electrons. The substance remains magnetised even after applied magnetic field is removed.

+ **Conditions for the Combination of Atomic Orbitals.** The linear combination of atomic orbitals takes place only if the following conditions are satisfied:

(/) The combining atomic orbitals must have same or nearly same energy. It means 1s orbiial combines with 1s orbital but not with 2s orbital because energy of 2s orbital is appreciably higher than 1s orbital. 2s orbital can overlap with 2s orbital to form s 2s and s 2s as shown in . Figure.

+ **Energy level Diagrams for Molecular Orbitals.** We have observed that 1s atomic orbitals on two atoms form two molecular orbitals designated as σ1s and a"s *1s.

In the same manner, the *2s* and *2p* atomic orbitals (eight atomic orbitals) give rise to the following eights molecular orbitals.

Antibonding	MO's	$s*2s$	$s*2p_z$	$p*2p_x$	$p*2p_y$
Bonding	MO's	$s\,2s$	$s\,2p_z$	$p2p_x$	$p2p_y$

+ The energy levels of these molecular orbitals have been determined experimentally from the spectroscopic data for homonuclear diatomic molecules of second row (period) elements of periodic table.

The increasing order of energies of various molecular orbitals for O_2 and F_2 is given below.

$$s\,1s < s*1s < s\,2s < s*2s < s\,2p_z < p2p_x = p2p_y < p*2p_x = p*2p_y < p*2p_z$$

+ However, this sequence of energy levels of molecular orbitals is not correct for remaining mofeculi Li_2, Be_2, B_2, C_2, N_2. For instance, it has been observed experimentally that for molecules such as B_2, C_2, N_2 etc., the increasing order of energies of various molecular orbitals is

$$s\,1s < s*1s < s\,2s < s*2s < p2p_x = p2p_y < s\,2p_z < p*2p_x = p*2p_y < s*2p_z$$

The important characteristic feature of this order is that the energy of $s\,2p_z$ **molecular orbital is night than that of** $p2p_x$ **and** $p2p_y$ **molecular orbitals** in these molecules.

Table Some Homonuclear molecules and their ions, Heteronuclear molecules and ions

S. No.	Molecule/ ion	No. of electrons	Electronic Configuration	Bond Order $\frac{1}{2}(N_b - N_a)$	Magnetic Property
1.	H_2	2	$(\sigma 1s)^2$	$\frac{1}{2}(2-0)=1$	Diamagnetic
2.	H_2^+	1	$(\sigma 1s)^1$	$\frac{1}{2}(1-0)=1/2$	Paramagnetic
3.	H_2^-	3	$(\sigma 1s)^2(\sigma^*1s)^1$	$\frac{1}{2}(2-1)=1/2$	Paramagnetic
4.	He_2	4	$(\sigma 1s)^2(\sigma^*1s)^2$	$\frac{1}{2}(2-2)=0$	Does not exist
5.	He_2^+	3	$(\sigma 1s)^2(\sigma^*1s)^1$	$\frac{1}{2}(2-1)=1/2$	Paramagnetic
6.	Li_2	6	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2$	$\frac{1}{2}(4-2)=1$	Diamagnetic
7.	Be_2	8	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2$	$\frac{1}{2}(4-4)=0$	Does not exist
8.	B_2	10	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\pi 2p_x^1 \pi 2p_y^1)$	$\frac{1}{2}(6-4)=1$	Paramagnetic
9.	C_2	12	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\pi 2p_x^2 \pi 2p_y^2)$	$\frac{1}{2}(8-4)=2$	Diamagnetic
10.	C_2^+	11	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\pi 2p_x^2 \pi 2p_y^1)$	$\frac{1}{2}(7-4)=3/2$	Paramagnetic
11.	C_2^-	13	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\pi 2p_x^2 \pi 2p_y^2)(\sigma 2p_z^1)$	$\frac{1}{2}(9-4)=5/2$	Paramagnetic
12.	N_2	14	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\pi 2p_x^2 \pi 2p_y^2)(\sigma 2p_z^2)$	$\frac{1}{2}(10-4)=3$	Diamagnetic
13.	N_2^+	13	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\pi 2p_x^2 \pi 2p_y^2)(\sigma 2p_z^1)$	$\frac{1}{2}(9-4)=5/2$	Paramagnetic
14.	O_2	16	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\sigma 2p_z^2)(\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^*2p_x^1 \pi^*2p_y^1)$	$\frac{1}{2}(10-6)=2$	Paramagnetic
15.	O_2^+	15	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\sigma 2p_z^2)(\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^*2p_x^1)$	$\frac{1}{2}(10-5)=5/2$	Paramagnetic
16.	O_2^-	17	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\sigma 2p_z^2)(\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^*2p_x^2 \pi^*2p_y^1)$	$\frac{1}{2}(10-7)=3/2$	Paramagnetic
17.	O_2^{2-}	18	$(\sigma 1s)^2(\sigma^*1s)^2(\sigma 2s)^2(\sigma^*2s)^2(\sigma 2p_z^2)(\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^*2p_x^2 \pi^*2p_y^2)$	$\frac{1}{2}(10-8)=1$	Diamagnetic

No.	Molecule	e⁻	Electronic configuration	Bond order	Magnetic nature
18.	F_2	18	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^* 2p_x^2 \pi^* 2p_y^2)$	½ (10 – 8) = 1	Diamagnetic
19.	F_2^+	17	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^* 2p_x^2 \pi^* 2p_y^1)$	½ (10 – 7) = 3/2	Paramagnetic
20.	F_2^-	19	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^* 2p_x^2 \pi^* 2p_y^2) (\sigma^* 2p_z^1)$	½ (10 – 9) = 1/2	Paramagnetic
21.	Ne_2	20	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^* 2p_x^2 \pi^* 2p_y^2) (\sigma^* 2p_z^2)$	½ (10 – 10) = 0	Does not exist
22.	Ne_2^+	19	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^* 2p_x^2 \pi^* 2p_y^2) (\sigma^* 2p_z^1)$	½ (10 – 9) = 1/2	Paramagnetic
23.	NO	15	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z^2) (\pi 2p_x^2 \pi 2p_y^2)$ $(\pi^* 2p_x^1)$	½(10 – 5)=5/2	Paramagnetic
24.	(NO)⁺	14	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z^2) (\pi 2p_x^2 \pi 2p_y^2)$	½ (10 – 4) = 3	Diamagnetic
25.	CO	14	$(\sigma 1s)^2 (\sigma^* 1s)^2 (\sigma 2s)^2 (\sigma^* 2s)^2 (\sigma 2p_z)^2 (\pi 2p_x^2 \pi 2p_y^2)$	½ (10 – 4) = 3	Diamagnetic

+ **The Hydrogen Bond.** It is intermolecular force of attraction between hydrogen atom of one molecule and a highly electronegative element (such as N, O or F) within the same molecule or another molecule of same or different compounds.

The H atom attached to nitrogen, oxygen or fluorine is able to interpose itself between two such electronegative atoms bonding them together. This is called hydrogen bond.

The essential requirements for an H-bond are :

(i) a hydrogen atom attached to highly electronegative atom

(ii) a lone pair of electrons on electronegative atom

The hydrogen bond is written as

A-H............B or A-H.............A.