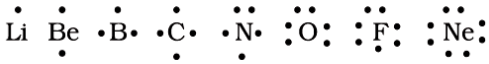
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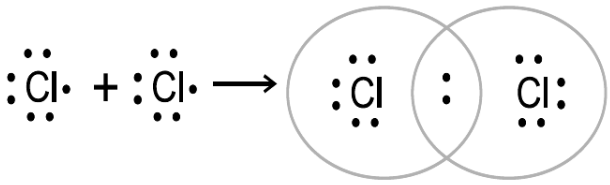
CHEMICAL BONDING AND MOLECULAR STRUCTURE

Kössel-Lewis Approach to Chemical Bonding:

- In 1916 Kössel and Lewis were the first to provide some logical explanation of **valence** which was based on the inertness of **noble gases** and gave electronic theory of chemical combination also called octet rule.
- Kössel's first insight into the mechanism of formation of electropositive and electronegative ions related the process to the attainment of noble gas configuration by respective ions.
- Electrostatic attraction between ions is the cause for their stability.
- Lewis postulated that atoms achieve the stable octet when they are linked by chemical bond.
- **Valence Electron:** In the formation of a molecule, only the outer shell electrons take part in chemical combination and they are known as valence electrons.
- **Lewis Symbol:** Notations to represent valence electrons in an atom are called Lewis symbols For example



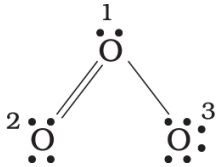
- The dots represent electrons. Such structures are referred to as Lewis dot structures.
- The Lewis dot structures can be written for other molecules also, in which the combining atoms may be identical or different. The important conditions being that:
- Each bond is formed as a result of sharing of an electron pair between the atoms.
- Each combining atom contributes at least one electron to the shared pair.
- The combining atoms attain the outer shell noble gas configuration as a result of the sharing of electrons.



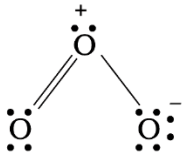
$8e^-$ $8e^-$

or Cl – Cl

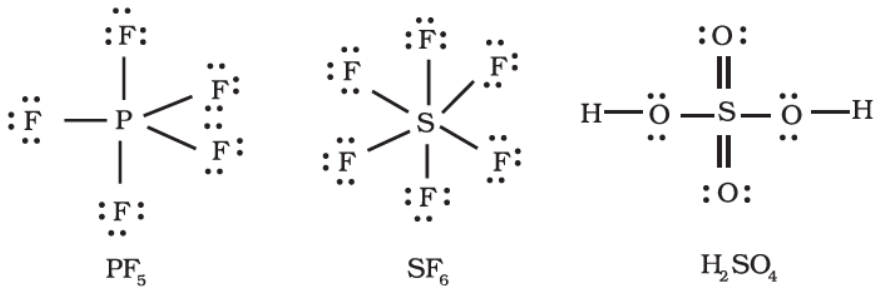
- Let us consider the ozone molecule (O_3). The Lewis structure of O_3 may be drawn as:



- The atoms have been numbered as 1,2 and 3. The formal charge on:
- The central O atom marked $1 = 6 - 2 - 12(6) = +1$
- The end O atom marked $2 = 6 - 4 - 12(4) = 0$
- The end O atom marked $3 = 6 - 6 - 12(2) = -1$
- Hence, we represent O_3 along with the formal charges as follows:



- **Limitation of the octet rule :**
  Incomplete octet of the central atom: Examples are $LiCl, BeH_2, BCl_3$ etc.
  Li: Cl H: Be: H Cl: B̈: Cl
  Odd-electron molecules: Examples are NO, NO_2 etc
- N̈ = Ö Ö = N̈ − Ö
- **Expanded octet:** Elements in and beyond the third period of the periodic table have, apart from 3s and 3p orbitals, 3d orbitals also available for bonding. Therefore they can expand their octets. Examples are PF_5, SF_6 , H_2SO_4 etc. The number of electrons around the central atom in PF_5, SF_6 and H_2SO_4 are 10,12, and 12 respectively.

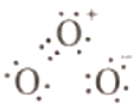
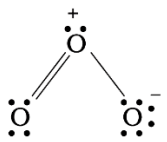


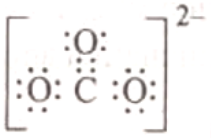
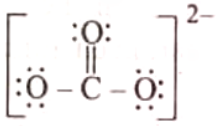
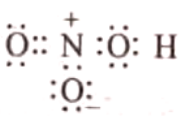
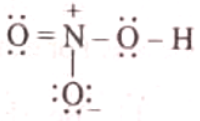
Drawbacks of octet theory:

- Some noble gases (for example xenon and krypton) also combine with oxygen and fluorine to form a number of compounds like $XeF_2, KrF_2, XeOF_2$ etc.,
- This theory does not account for the shape of molecules.
- This theory does not explain the relative stability of the molecules being totally silent about the energy of a molecule.

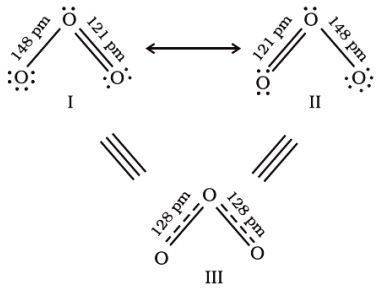
Ionic or Electrovalent Bond:

- Ionic bonds are formed between a positive ion (cation) and negative ion (anion). The positive and negative ions are formed either by removal of electron(s) or addition of electron(s) to the neutral atom, which is represented as **Covalent Bond:** These bonds are formed by the mutual sharing of electron between atoms. When two atoms share one electron pair they are said to be joined by a single **covalent bond**.
- When two atoms share two pairs of electrons, the covalent bond between them is called a double bond.
- When combining atoms share three electron pairs, they are said to be joined by a triple bond.
- The Lewis dot structures provide a picture of bonding in molecules and ions in terms of the shared pairs of electrons and the octet rule.
- Lewis representations of a few molecules/ions table:

Molecule/Ion		Lewis Representation
H_2	H: H*	H − H
O_2	:Ö::Ö:	Ö = Ö:
O_3		
NF_3		

CO_3^{2-}		
HNO_3		

- In case of polyatomic ions, the net charge is possessed by the ion as a whole and not by a particular atom. It is, however, feasible to assign a formal charge on each atom.
- **Formal charge:** It can be defined as the difference between the number of valence electrons of that atom in an isolated or free state and the number of electron assigned to that atom in the Lewis structure.
- Formal charge (F.C.) on an atom in a Lewis structure
- $\left[\begin{array}{l}\text{total number of}\\ \text{valence electrons}\\ \text{in the free atom}\end{array}\right] - \left[\begin{array}{l}\text{total number of non}\\ \text{bonding (lone pair)}\\ \text{electrons}\end{array}\right]$
  $-(1/2)\left[\begin{array}{l}\text{total number of}\\ \text{bonding (shared)}\\ \text{electrons}\end{array}\right]$
- **Bond angle:** It is defined as the angle between the orbitals containing bonding electron pairs around the central atom in a molecule/complex ion.
- **Bond enthalpy** is defined as the amount of energy required to break one mole of bonds of a particular type between two atoms in a gaseous state The unit of bond enthalpy is $kJmol^{-1}$.
- It is important that larger the bond dissociation enthalpy, stronger will be the bond in the molecule.
- In case of polyatomic molecules the term mean or average bond enthalpy is used.
- Bond order: It is given by number of bonds between the two atoms in a molecule.
- As the bond order increases, bond enthalpy also increases but bond length decreases.
- Isoelectronic molecules and ions have identical bond orders for example, F_2 and O_2^{2-} have bond order 1 .
- **Resonance:** A number of molecules and polyatomic ions can not be described accurately by a single lewis structure so, a number of representations based on the same skeletal structure are written, these together represent a molecule or ion, this phenomena is known as resonance.
- The contributing structure or cannonical forms constitute the resonance hybrid. For example ozone.



Polarity of Bonds: The covalent bonds may be either polar or non polar, depends upon the electro negativity difference of the atoms which forms the covalent bonds.

As a result of polarisation the molecule possesses dipole moment which can be defined as the product of the magnitude of the charge and the distance between the centres of positive and negative charge.

- Dipole moment is usually designated and by a Greek letter ' μ '. Mathematically, dipole momentis expressed as follows: $(\mu) =$ charge $(Q) \times$ distance of separation (r)

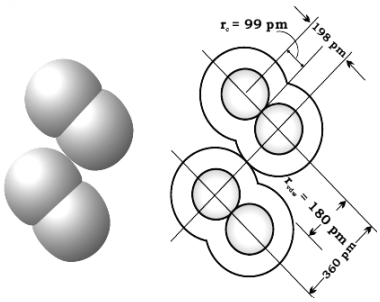
- Dipole moment is expressed in Debye units (D) I $D = 3.33564 \times 10^{-30} C$ m where C is coulomb and m is meter.
- Dipole moment is represented by the crossed arrow ($\leftrightarrow$). The cross is on positive endand arrow head is on negative end. e.g. $H \leftrightarrows F$
  In case of polyatomic molecule she dipole moment is the vector sum of the dipole moments of various bond.

- **Fajan's Rule:** The ionic bonds also have some partial covalent character which was discussed by Fajans in terms of the following rules:
  $M(g) \rightarrow M^+(g) + e^-$;
  Ionization enthalpy
- $X(g) + e^- \rightarrow X^-(g)$;
  Electron gain enthalpy
  $M^+(g) + X(g) \rightarrow MX(s)$

- The electron gain process may be exothermic or endothermic whereas ionization is always endothermic.
- The compound formed by ionic bonds are known as ionic compound.
- Most ionic compounds have cations derived from metallic elements and anions from non-metallic elements.
- Ionic compounds in the crystalline state consist of orderly three-dimensional arrangements of cations and anions held together by coulombic interaction energies.
- A qualitative measure of the stability of an ionic compound is provided by its enthalpy of lattice formation.
- **Lattice Enthalpy:** The Lattice Enthalpy of an ionic solid is defined as the energy required to completely separate one mole of a solid ionic compound into gaseous constituent ions.

Bond Parameters

- Bond length: It is defined as the equilibrium distance between the nuclei of two bonded atoms in a molecule. Each atom of the bonded pair contributes to bond length. It is measured in terms of covalent radius (for bonded atom) and van der Waal's radius (for non-bonded atoms).
- **Covalent radius (r_c):** It is half of the internuclear distance between two bonded atoms. **Van der Waal's radius (r_{vdv}):** It is half of the internuclear distance between two non-bonded atoms. Fig.: Covalent and Van der Waals radil in a chlorine molecule. The inner circles correspond to the size of the chlorine atom.



- The smaller the size of the cation and the larger the size of the anion, the greater the covalent character of an ionic bond.
- The greater the charge on the cation, the greater the covalent character of the ionic bond.
- For cations of the same size and charge, the one, with pseudo electronic configuration i.e., $(n-1)d^n ns^\circ$, typical of transition metals, are more polarising than the one with a noble gas configuration, ns^2np^6, typical of alkali andalkaline earth metal cations.
- The polarising power of the cation, the polarisability of the anion and the extent of distortion (polarisation) of anion are the factors, which determine the percent covalent character of the ionic bond.

The Valence Shell Electron Pair Repulsion (VSEPR) Theory

- Sidgwick and Powell in 1940, proposed a simple theory based on the repulsive interactions of the electron pairs in the valence shell of the atoms. It was further developed and redefined by Nyholm and Gillespie in (1957). The main postulates of VSEPR theory are as follows:
- The shape of a molecule depends upon the number of valence shell electron pairs (bonded or nonbonded) around the central atom.
- Pairs of electrons in the valence shell repel one another since their electron clouds are negatively charged.
- These pairs of electrons tend to occupy such positions in space that minimise repulsion and thus maximise distance between them.
- The VSEPR model is applicable to any such structure where two or more resonance structures can represent a molecule.
- The repulsive interaction of electron pairs decrease in the order: Lone pair (lp) Lonepair (lp) > Lonepair (lp)-Bond pair (bp) > Bond pair (bp)-Bond pair (bp).

Table: Shape of molecule containing bond pair and lone pair

Molecule	**No. of bonding**	**No. of Lone pair**	**Geometry**	**Shape**	
	2	0	Linear	Linear	$BeCl_2, H_8Cl_2$
AB_3	3	0	Trigonal planar	Trigonal planar	BF_3
	2	1	Trigonal planar	Bent	SO_2, O_3
AB_4	4	0	Tetrahedral	Tetrahedral	CH_4, NH_4
	3	1	Tetrahedral	Trigonal	

- The valence bond theory explains the slape, the formation and directions properties of bonds in polyatomic molecules in terms of overlap and hybridisation of atomic orbitals.

Types of overlapping and nature of covalent bond:

- **Sigma(σ) bond:** This type of covalent bond is formed by the end to end (head-on) overlap of bonding orbitals along the internuclear axis. This is called as head on overlap or axial overlap. This can be formed by s, $s-p$, and $p-p$ types of combinations of atomic orbitals.
- pi (π) bond : In the formation of π bond the atomic orbitals overlap in such a way that their axes remain parallel to each other and perpendicular to the internuclear axis.
- Basically the strength of a bond depends upon the extent of overlapping.
- In case of sigma bond, the overlapping of orbitals takes place to a larger extent. Hence, it is stronger as compared to the pi bond where the extent of overlapping occurs to a smaller extent.
- In the formation of multiple bonds between two atoms of a molecule, pi(π) bond (s) is formed in addition to a sigma bond.

Hybridisation

- In order to explain the characteristic geometrical shapes of polyatomic molecules like CH_4, NH_3 and H_2O etc., Pauling introduced the concept of hybridisation.
- Hybridisation can be defined as the process of intermixing of the orbitals of slightly different energies so as to redistribute their energies, resulting in the formation of new set of orbitals of equivalent energies and shape known as hybrid orbitals.

Main features of hybridisation: The number of hybrid orbitals is equal to the number of the atomic orbitals that get hybridised.

- The hybridised orbitals are always equivalent in energy and shape.
- The hybrid orbitals are more effective in forming stable bonds than the pure atomic orbitals. These hybrid orbitals are directed in space in some preferred direction to have minimum repulsion between electron pairs and thus a stable arrangement.

Important conditions for hybridisation:

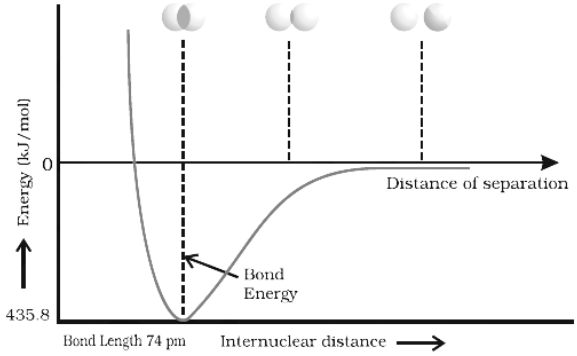
- The orbitals present in the valence shell of the atom are hybridised.
- The orbitals undergoing hybridisation should have almost equal energy Promotion of electron is not essential condition prior to hybridisation
- It is not necessary that only half filled orbitals participate in hybridisation. In some cases, even filled orbitals of valence shell take part in hybridisation.

$AB_2\ K_2$	2	2	Tetrahedral	Bent	H_2O
	5	0	Trigonal bi-	Trigonal bi-	PCl_5
$AB_2\ K$	4	1	Trigonal bi-	See-saw	SF_4
$AB_3\ K_2$	3	2	Trigonal bi-	T-shape	ClF_3
AB_6	6	0	Octahedral	Octahedral	SF_6
AB_2E	5	1	Octahedral	Square pyramidal	BrF_5
AB_4E_2	4	2	Octahedral	Square planar	XeF_4

- The VSEPR theory gives the geometry of simple molecules but theoretically, it does not explain them and also it has limited applications.

Valence Bond Theory

- Valence bond theory was introduced by Heitler and London (1927) and developed further by Pauling and others. When two atoms approach each other, attractive and repulsive forces begin to operate. If the magnitude of attractive forces is more than repulsive forces the atoms will come closer and a stage will be reached where net force of attraction. balances force of repulsion and system acquires minimum potential energy. Now atoms are said to be bonded and energy released during the course of time is called bond enthalpy. For example bond enthalphy for H_2.



- Fig: The potential energy curve for the formation of H_2 molecules as a function of internuclear distance of the H-atoms.

Orbital overlap concept:

- When two atoms are so close that their atomic orbitals undergo partial interpenetration, this is called over lapping of atomic. Orbital result in the pairing of valence electrons having opposite spin.
- Greater the overlap, the stronger is the bond formed between two atoms.

- Table: Important Hybridisation

Hybridisation	Atomic Orbital	Shape of Molecule	Bond Angle	Example
sp	$s(1)+p(1)$	Linear	180°	BeF_2, CO_2, C_2H_4
sp^2	$s(1)+p(2)$	Trigonal planar	120°	BCl_1, C_2H_4
s'	$s(1)+p(3)$	Tetrahedral	109.5	CH_4
sp^3d	$s(1)+p(3)+d(1)$	Trigonal Bipyramidal	90° (axial) 120° (eqatorial)	$PCl_{5.}PF_5$
sp^1d^2 d^2sp^2	$s+p(3)+d(2)$ $d(2)+s+p(3)$	Octahedral	90°	$SF_6[CrF_6]_{3.}$ $[Co(NH_3)_6]^-$

- **Molecular Orbital Theory:**
- Molecular orbital (MO) theory was developed by F. Hund and R.S. Mulliken in 1932. The salient features of this theory are:
- The electrons in a molecule are present in the various molecular orbitals as the electrons of atoms are present in the various atomic orbitals.
- The atomic orbitals of comparable energies and proper symmetry combine to form molecular orbitals.
- An atomic orbital is monocentric (only one nuclei) while a molecular orbital is polycentric (more than one nuclei). The number of molecular orbital formed is equal to the number of combining atomic orbitals. When two atomic orbitals combine, two molecular orbitals are formed. One is known as bonding molecular orbital while the other is called antibonding molecular orbital.
- The bonding molecular orbital has lower energy and hence greater stability than the corresponding antibonding molecular orbital.
- The electron probability distribution around a group of nuclei in a molecule is given by a molecular orbital. The molecular orbitals like atomic orbitals are filled in accordance with the aufbau principle, Pauli's exclusion principle and the Hund's rule.

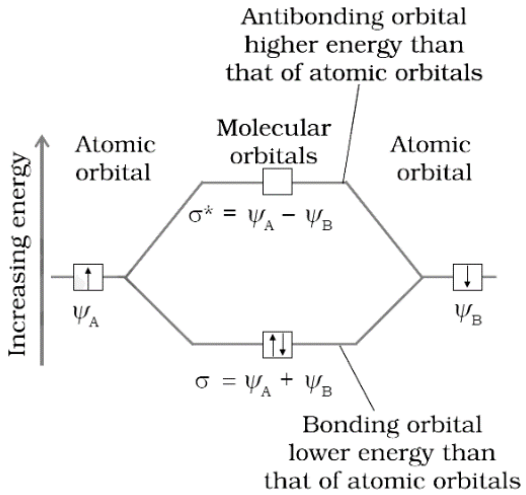
Formation of molecular orbitals Linear combination of atomic orbitals (L.CAO):

- According to wave mechanics, the atomic orbitals are expressed by wave functions (φ 's) which represent the amplitude of the electron waves. These are obtained from the solution of Schrödinger wave equation.
- Mathematically, the formation of molecular orbitals may be described by the linear combination of atomic orbitals that can take place by addition and by subtraction of wave functions of individual atomic orbitals as shown: $\psi_{MO} = \psi_A \pm \Psi_B$
- Hence, the two molecular orbitals σ and σ^* are formed as:
- $\sigma = \Psi_A + \psi_B$(Bonding M.O.)
- $\sigma^* = \Psi_A - \psi_B$ (Antibonding M.O.)
- For molecules such as B_2, C_2, N_2, etc. the increasing order of energies of various molecular orbitals is

- $\sigma 1s < \sigma^* 1s < \sigma 2s < \sigma^* 2s < (\pi 2p_x = \pi 2p_y) < \sigma 2p_z < (\pi^* 2p_x = \pi^* 2p_y) < \sigma^* 2p_z$
- The important characteristic feature of this order is that the energy of $\sigma 2p_z > \pi 2p_x$ and $\pi 2p_y$ The distribution of electrons among various molecular orbitals is called the electronic configuration of the molecule.
- **Stability of Molecules:** If N_b is the number of electrons occupying bonding. orbitals and N_a the number occupying the antibonding orbitals, then the molecule is stable if $N_b > N_a$ the molecule is unstable if $N_b < N_a$. **Bond order (B.O.) :** It is defined as one half the difference between the number of electrons present in the bonding and the antibonding orbitals.
- Bond order (B.O.) $= \frac{1}{2}(N_b - N_a)$
- A positive bond order means a stable molecule while a negative or zero bond order means an unstable molecule. Integral bond order values of 1,2 or 3 correspond to single, double or triple bonds respectively. **Bond order :** It can be taken as an approximate measure of the bond length between two atoms. The bond length decreases as bond order increases.
- **Magnetic nature :** If all the molecular orbitals in a molecule are doubly occupied, the substance is diamagnetic. However if one or more molecular orbitals are singly occupied it is paramagnetic e.g., O_2 molecule.
- Bonding in Some Homonuclear Diatomic Molecules. Bonding in some homonuclear diatomic molecules is exhibited as follows

Types of Molecules	Molecular orbital configuration	Bond order	Nature of Molecules
H_2	$(\sigma \mid s)^2$	1	Diamagnetic
He_2	$(\sigma \mid s)^2(\sigma^* \mid s)^2$	0	Does not exist
Li_2	$(\sigma 1s)^2(\sigma^* 1s)^2(\sigma 2s)^2$	1	Diamagnetic
C_2	$(\sigma \mid s)^2(\sigma^* 1s)^2(\sigma * 2s)^2(\pi 2p_x^2 = \pi 2p_y^2)$	2	Diamagnetic (both π bonds)
O_2	$(\sigma 1s)^2(\sigma^{*1} 1s)^2(\sigma 2s)^2(\sigma^* 2s)^2(\sigma 2p_x)^2$ $(\pi 2p_x^2\ {}^2\pi 2p_y^2)(\pi^* 2p_x^1 \pi^* 2p_y')$	2	Paramagnetic (1σ, 1π bond)

- Antibonding orbital higher energy than that of atomic orbitals Bonding orbital lower energy than that of atomic orbitals.



- The electron density in a bonding molecular orbital is located between the nuclei of the bonded atoms because of which the repulsion between the nuclei is very less while in case of an antibonding molecular orbital, most of the electron density is located away from the space between the nuclei.

Conditions for the Combination of Atomic Orbital:

- The linear combination of atomic orbitals to form molecular orbitals takes place only if the following conditions are satisfied:
  The combining atomic orbitals must have the same or nearly the same energy.
- The combining atomic orbitals must have the same symmetry about the molecular axis. By convention z-axis is taken as the molecular axis. The combining atomic orbitals must overlap to the maximum extent.

Types of Molecular Orbital :

- Molecular orbitals of diatomic molecules are designated as σ (sigma), π (pi), δ (delta), etc. The sigma (σ) molecular orbitals are symmetrical around the bondaxis while pi (π) molecular orbitals are not symmetrical.
  If internuclear axis is taken to be in the z-direction, it can be seen that a linear combination of $2p_z$ - orbitals of two atoms also produces two sigma molecular orbitals designated as $\sigma_2 p_z$ and $\sigma^{\star} 2p_z$.
- A π bonding MO has larger electron density above and below the inter -nuclear axis. The π^* antibonding MO has a node between the nuclei.
- The increasing order of energies of various molecular orbitals for O_2 and F_2 is given below: $\sigma]s < \sigma^*1s < \sigma 2s < \sigma^*2s < \sigma 2p_z < (\pi 2p_x = \pi 2p_y) < (\pi^*2p_x = \pi^*2p_x) < \sigma^*2p_2$
- Isoelectronic means equal number of total electrons or valence electron and that species which have isoelectron will also be isostructural. When the number of total electrons in species is same, then the species will exhibit same bond order.
- The repulsive interaction of electron pairs decrease in the order: Lone pair - Lone pair > Lone pair - Bond pair > Bond pair - Bond pair. Bond angle in any molecules depends on its geometry or shape. Like in CH_4 the both geometry and shape are equivalent and shows bond angle 109 degree as sp^3 hybridization shown by CH_4. But in case of NH_3 and H_2O, geometry and shape are different and only shape becomes the responsible factor for bond angle. Since shape depends upon the presence of lone pair, so bond angle decreases as the number of lone pair increases.
- Therefore, the decreasing order of bond angle is $CH_4 > NH_3 > H_2O$.
- Maximum bond angle is shown by that species which have sp hybridization. Shape of molecules depends upon the presence of lone pair. Any species can show pyramidal shape if it has one lone pair with sp^3 hybridization. Both $Li_2^+$and Li_2^-have positive bond order, i.e,. $+1/2$. So, both will be stable. Stability of any molecule increases only when bond order increases. So the species in which bond order increases during the formation of an ion would be more stable.
- In homo or hetro molecules, the number of electron in last orbital confirms that the species is paramagnetic or not.

Tips/Tricks/Techniques ONE-LINERS (Exam Special)

Hybridization $= \frac{1}{2}$[No. of valence electrons of Centre atom + number of monovalent atoms surrounding-charge on cation + charge on anion]

Applications of Intermolecular H-Bonding:

Water : Water has the lowest molecular weight among the hydrides of group 16 elements yet it has the highest melting and boiling points.

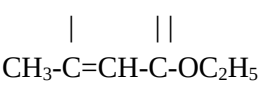
Ice has less density than water

- Amides associate and have higher melting and boiling points.
- Applications of Intramolecular H-Bonding:

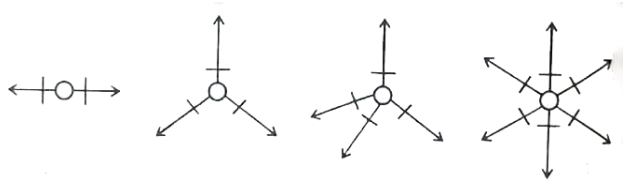
- Salicylic acid is stronger acid than 0 -methoxy benzoic acid.
- Ethyl aceto acetate - It exists in two forms-



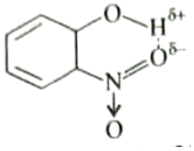
- Keto form Enol form
- Enolic form is more volatile due to chelation



- Maleic acid is stronger acid than fumaric acid: Maleate ion can be stabilised by Chelation because hydrogen and Hydrogen Bonding
- The hydrogen bond can be defined as the attractive force which bing hydrogen atom of one molecule with the electronegative atom (F, O or N) another molecule.
- Hydrogen bond acts as a bridge between two atoms which holds one atom by covalent bond and the other by hydrogen bond. Hydrogen bond is represented by a dotted line (....) while a solid line represent. the covalent bond.
- The magnitude of H-bonding depends on the physical state of the compound It is maximum in the solid state and minimum in the gaseous state.
- Intermolecular hydrogen bond is formed between two different molecule of the same or different compounds.
- Intramolecular hydrogen bond is formed when hydrogen atom is in between the two highly electronegative (F, O, N) atoms present within the same molecule,

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- **Molecular Dipole Moment:-**
- A molecule is polar if all the bond dipoles do not cancel out each other due to asymmetric structure. A molecule has zero net dipole moment in the following cases of symmetries:
- The simplest way to find the hybridization is to find steric number (σ bonds $+$lp of e^-). For BF_3, steric number is 3 . So the hybridization will be sp^2.
- 
- Stability of any molecules depend upon the bond order. If bond order, zero is exhibited by any molecule then that molecule will be least stable or do not exist.
- The molecules which have regular structure or symmetrical structure with the same bonded atoms, will always exhibit zero dipole moment.
- Among O_2, N_2, C_2 and Be_2 molecules, bónding in C_2 can be explained by MOT only whereas in other, bonding can be explained by VBT or MOT both.
- The π bonding involves overlap of p-orbitals of oxygen with d-orbitals of manganese.
- Bond order depends upon the number of electrons in bonding and antibonding orbitals. Higher the bonding electron, higher will be the bond order.
- Total valence electron in ClF_3 is 28. So sterie number will be 5 . By subtracting the number of fluorine atoms from steric number wo can get the number of lone pairs.
- oxygen responsible for forming hydrogen bond are very near to each other. On the other hand fumerate ion cannot stabilise by chelation because hydrogen and oxygen are on opposite sides to each other Hence formation of fumerate ion does not take place.

- Acid character of nitrophenols. It follows the following order: p-nitrophenol > o-nitrophenol > m-nitrophenol Acid character of 0 - nitrophenol is suppressed by chelation



- o-nitrophenol is more volatile (b.pt $214°C$) as compared to meta (b.pt $290°C$) and para (b.pt $279°C$). It is due to chelation
- $\Delta H_f = \Delta H_{sub} + \frac{1}{2}\Delta H_d + \text{I.E.} + \text{E.A.} + U$
- where,
- In order to determine the number of $p\pi - p\pi$ or $p\pi - d\pi$ bonds, the following formula can be used:
- Calculate the number of σ bonds = number of surrounding atoms
- Lone pair = $1/2$ (Valence electron - number of covalent bond)
- Hybridization = σ bonds + number of lone pairs
- $\therefore$ Now, number of $p\pi - p\pi$ bonds = number of unhybridized p-orbitals left
- $p\pi - d\pi$ bonds will be formed when π bonds are more than the number of unhybridized p-orbitals left.
- The more the electronegativity of atom involved in H-bonding, the more is the bond strength eg. $H\cdots F > H\cdots O > H\cdots N$
  $10kcal/mole > 7kcal/mole > 2.0kcal/mole$
- Formation of a chemical bond is always an exothermic process.
- $FeCl_3$ is more covalent than $FeCl_2$ becuase polarising power of Fe^{3+} is more than that of Fe^{2+}.
- Free rotation about a σ-bond is possible but it is not possible in π-bond.
- The hydrogen bond is very strong in HF and persist even in vapour state. Gaseous hydrogen fluoride is largely polymerised into the molecular species H_2F_2, H_3F_3, H_4F_4, ... etc.
- C_2 molecule is an exception which contains no σ-bond but two π-bonds.
- Ionic compounds consist of ions in solid state while covalent compound consist of molecules.
- The relative strength of various bonds is: ionic > coordinate > covalent > metallic > H-bond > van der Waal forces.

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PW Web/App - https://smart.link/7wwosivoicgd4

Library- https://smart.link/sdfez8ejd80if