UNIT-10 CHEMICAL BONDING

MY REVISION TIMELINE:-

SUMMARY:-

- Chemical bonding is the interatomic attractive forces which hold the constituent atoms/ions together in a molecule.
- The atoms share or transfer electrons so that all atoms involved in chemical bonding obtain 8 electrons in their outer shell (valence electrons)
- Types of covalent bonding:
    - Covalent
    - Ionic
- Covalent bond: Mutual sharing of one or more pairs of electrons between two combining atoms result in the formation of chemical bond called covalent bond.
- Ionic bond: Electrostatic attractive force.
- Bond length: The distance between the nuclei of the two covalently bonded atoms.
- Bond order: The number of bonds formed between the two bonded atoms in a molecule.
- Bond enthalpy: Minimum amount of energy required to break one mole of a particular bond in molecules in their gaseous state.
- Dipole moment: The polarity of a covalent bond can be measured in terms of dipole moment.

  $u = q \times 2d$

  u → dipole moment (vector quantity)
- Electronegativity difference ($X_A - X_B$)
    - Equal to 1.7, A-B bond has 50% ionic character.
    - Greater than 1.7, A-B bond has more than 50% ionic character.
    - Lesser than 1.7, A-B bond has less than 50% ionic charcter.
- $l_p\text{-}l_p > l_p\text{-}b_p > b_p\text{-}b_p$ where l_p is lone pair and b_p is bond pair.
- Sigma bond: When two atomic orbitals overlap linearly along the axis, the resultant bond is σ (sigma) bond.
- Pi bond: When two atomic orbitals overlap sideways, the resultant covalent bond is called π (pi) bond.
- Hybridisation: It is the process of mixing of atomic orbitals of the same atom with comparable energy to form equal number of new equivalent orbitals with same energy.
- Types of hybridisation:
    - sp
    - sp^2
    - sp^3
    - sp^3d

- sp^3d^2

➢ Bond order = $\frac{N_b - N_a}{2}$

N_b → Total number of electrons present in the bonding molecular orbitals.

N_a → Total number of electrons present in the antibonding molecular orbitals.

➢ Metallic bonding: The forces that keep the atoms of the metal so closely in a metallic crystal constitute metallic bond.

TEXTBOOK EVALUATION

Multiple choice questions:-

1. In which of the following compounds does the central atom obey the octet rule?

(a) XeF_4 (b) $AICI_3$

(c) SF_6 **(d) SCl_2**

Explanation:

Compound	No. of valance electron on the central atom
XeF_2	10
$AlCl_3$	6
SF_6	12
SCl_2	8

2. In the molecule $O_A = C = O_B$ the formal charge on O_A, C and O_B are respectively.

(a) – 1, 0, +1 (b) +1, 0, – 1

(c) – 2, 0, +2 **(d) 0, 0, 0**

Explanation:

Formal change on O_A or O_B $= N_v - (N_l + \frac{N_b}{2})$

$= 6 - (4 + \frac{4}{2})$

$= 0$

Formal change on carbon $= 4 - (0 + \frac{8}{2})$

$= 0$

3. Which of the following is electron deficient?

(a) PH_3 (b) $(CH_3)_2$

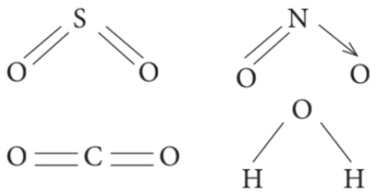
(c) BH_3 (d) NH_3

4. Which of the following molecule contain no π bond?

(a) SO_2 (b) NO_2

(c) CO_2 **(d) H_2O**

Explanation:



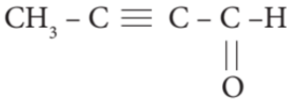
Water contains only sigma bonds and no pi bonds.

5. The ratio of number of sigma (σ) and pi (π) bonds in 2- butynal is

(a) 8/3 (b) 5/3

(c) 8/2 (d) 9/2

Explanation:



$$CH_3-C\equiv C-\underset{\underset{O}{||}}{C}-H$$

No. of sigma bonds = 4 C-H + 2 C-C + 1 C-O = 8

No. of pi bonds = 2 C-C + 1 C-O = 3

$\boxed{\textbf{\textit{Ratio = 8/3}}}$

6. Which one of the following is the likely bond angles of sulphur tetrafluoride molecule?

(a) 120°, 80° (b) 109°, 28

(c) 90° **(d) 89°, 117°**

Explanation:

Normal Trigonal bipyramidal angle is 90° and 120°, due to l.p-l.p repulsion, bond angle in SF_4 is 89° and 117°

7. Assertion: Oxygen molecule is paramagnetic.

Reason: It has two unpaired electron in its bonding molecular orbital.

(a) both assertion and reason are true and reason is the correct explanation of assertion

(b) both assertion and reason are true but reason is not the correct explanation of assertion

(c) assertion is true but reason is false

(d) Both assertion and reason are false

Explanation:

O_2 is **paramagnetic.** It has two unpaired electrons in its antibonding molecular orbital.

8. According to Valence bond theory, a bond between two atoms is formed when

(a) fully filled atomic orbitals overlap **(b) half filled atomic orbitals overlap**

(c) non-bonding atomic orbitals overlap (d) empty atomic orbitals overlap

9. In CIF_3, NF_3 and BF_3 molecules the chlorine, nitrogen and boron atoms are

(a) sp^3 hybridised (b) sp^3, sp^3 and sp^2 respectively

(c) sp^2 hybridised **(d) sp^3d, sp^3 and sp^2 hybridised respectively**

10. When one s and three p orbitals hybridise,

(a) four equivalent orbitals at 90^0 to each other will be formed

(b) four equivalent orbitals at 109^0 28' to each other will be formed.

(c) four equivalent orbitals, that are lying the same plane will be formed

(d) none of these

Explanation:

sp^3 hybridisation, geometry – Tetrahedral, bond angle 109° 28´

11. Which of these represents the correct order of their increasing bond order?

(a) $C_2 < C_2^{2-} < O_2$ (b) $C_2^{2-} < C_2^+ < O^2 < O_2^{2-}$

(c) $O_2^{2-} < O_2^+ < O_2 < C_2^{2-}$ **(d) $O_2^{2-} < C_2^+ < O_2 < C_2^{2-}$**

Explanation:

Bond order = (n_b – n_a)

Bond order of O_2^{2-} = ½ (8 – 6) = 1

Bond order of C_2^{+} = ½ (5 – 2) = 1.5

Bond order of O_2 = ½ (8 – 4) = 2

Bond order of C_2^{2-} = ½ (8 – 2) = 3

12. Hybridisation of central atom in PCl_5 involves the mixing of orbitais.

(a) s, p_x, p_y, d_x^2, d_x^2-y^2 (b) s, p_x, p_y, p_{xy}, d_x^2-y^2

(c) s, p_x, p_y, p_z, d_x^2-y^2 (d) p_x, p_y, p_{xy}, d_x^2-y^2

Explanation:

PCl_5 – sp^3d hybridisation

s, p_x, p_y, p_z and d_x^2-y^2

13. The correct order of O – O bond length in hydrogen peroxide, ozone and oxygen is

(a) $H_2O_2 > O_3 > O_2$ **(b) $O_2 > O_3 > H_2O$**

(c) $O_2 > H_2O_2 > O_3$ (d) $O_3 > O_2 > H_2O_2$

Explanation:

O_2 = 2

O_3 = 1.5

H_2O_2 = 1

14. Which one of the following is diamagnetic?

(a) O_2 **(b) O_2^{2-}**

(c) O_2^{2+} (d) None of these

15. Bond order of a species is 2.5 and the number of electrons are in its bonding molecular orbital is found to be 8. The no. of electrons in its anti-bonding molecular orbital is

(a) three

(b) four

(c) zero

(d) cannot be calculated form the given information.

Explanation:

Bond order = ½ (n_b – n_a)

2.5 = ½ (8 – n_a)

5 = 8 – n_a

$\boxed{n_a = 3}$

16. Shape and hybridisation of IF_5 are

(a) Trigonal bipyramidal, sp^3d^2 (b) Trigonal bipyramidal, sp^3d

(c) Square pyramidal, sp^3d^2 (d) Octahedral, sp^3d^2

17. Pick out the incorrect statement from the following.

(a) sp^3 hybrid orbitals are equivalent and are at an angle of 109°28’ with each other.

(b) dsp^2 hybrid orbitals are equivalent and bond angle between any two of them is 900.

(c) All five sp^3d hybrid orbitals arc not equivalent. Out of these five sp^3d hybrid orbitals, three are at an angle of 120°, remaining two are perpendicular to the plane containing the other three

(d) none of these

18. The molecules having same hybridisation, shape and number of lone pairs of electrons are

(a) SeF_4, XeO_2F_2 (b) SF_4, XeF_2

(c) $XeOF_4$, TeF_4 (d) $SeCI_4$, XeF_4

Explanation:

SeF_4, XeO_2F_2 – sp^3 hybridisation

19. In which of the following molecules / ions BF_3, NO_2^-, H_2O the central atom is sp^2 hybridised?

(a) NO_2^- and H_2O (b) NO_2^- and H_2O

(c) BF_3 and NO_2 (d) BF_3 and NH_2^-

Explanation:

H_2O – sp^3

NO_2^- - sp^2

BF_3 – sp^2

NH_2^- - sp^3

20. Some of the following properties of two species, NO_3^- and H_3O^+ are described below. Which one of them is correct?

(a) dissimilar in hybridisation for the central atom with different structure.

(b) isostructural with same hybridisation for the Central atom.

(c) different hybridisation for the central atom with same structure

(d) none of these

Explanation:

NO_3^- - sp^2 hybridisation, planar

H_3O^- - sp^3 hybridisation, pyramidal

21. The types of hybridisation on the five carbon atom from right to left in the, 2, 3 pentadiene.

(a) sp^3, sp^2, sp, sp^2, sp^3 (b) sp^3, sp, sp, sp, sp^3

(c) sp^2, sp, sp^2, sp , sp^3 (d) sp^3, sp^3, sp^2, sp^3, sp^3

Explanation:

$\overset{Sp^3}{CH_3} - \overset{Sp^2}{CH} = \overset{Sp}{C} = \overset{Sp^3}{CH} - \overset{Sp^3}{CH_3}$

22. XeF_2 is isostructural with

(a) $SbCl_2$ (b) $BaCl_2$

(c) TeF_2 **(d) ICl_2^-**

23. The percentage of s-character of the hybrid orbitals in methane, ethane, ethene and ethyne are respectively

(a) 25, 25, 33.3, 50 (b) 50, 50, 33.3, 25

(c) 50, 25, 33.3, 50 (d) 50, 25, 25, 50

Explanation:

CH_4 – sp^3	25%
CH_3-CH_3 – sp^3	25%
CH_2=CH_2 – sp^2	33.3%
$CH \equiv CH$ – sp	50%

24. Of the following molecules, which have shape similar to carbon dioxide?

(a) $SnCI_2$ (b) NO_2

(c) C_2H_2 (d) All of these

Explanation:

CO_2 is linear

C_2H_2 is also linear

25. According to VSEPR theory, the repulsion between different parts of electrons obey the order

(a) l.p – l.p > b.p – b.p > l.p – b.p (b) b.p – b.p > b.p – 1.p > l.p – b.p

(c) l.p – l.p > b.p – l.p > b.p – b.p (d) b.p – b.p > l.p – l.p > b.p – l.p

26. Shape of CIF_3 is

(a) Planar triangular (b) Pyramidal

(c) 'T' Shaped (d) none of these

Explanation:

ClF_3 is sp^3 hybridisation.

27. Non- Zero dipole moment is shown by

(a) CO_2 (b) p – dichlorobenzene

(c) carbon tetrachloride **(d) water**

28. Which of the following conditions is not correct for resonating structures?

(a) the contributing structure must have the same number of unpaired electrons.

(b) the contributing structures should have similar energies.

(c) the resonance hybrid should have higher energy than any of the contributing structure.

(d) none of these

29. Among the following, the compound that contains, ionic, covalent and coordinate linkage is

(a) NH_4Cl (b) NH_3

(c) NaCl (d) none of these

30. CaO and NaCl have the same crystal structure and approximately the same radii. If U is the lattice energy of NaCl, the approximate lattice energy of CaO is

(a) U (b) 2U

(c) U/2 **(d) 4U**

Write brief answers to the following questions:-

31. Define the following i) Bond order ii) Hybridisation iii) σ- bond

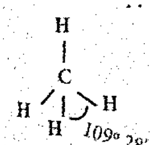
- **Bond order:** The number of bonds formed between the **two bonded atoms** in a molecule.
- **Hybridisation:** It is the process of **mixing of atomic orbitals** of same atom with comparable energy to form equal number of new equivalent orbitals with same energy.
- **σ- bond:** When two atomic orbitals **overlap linearly** along the axis, the resultant bond is called a σ-bond.

32. What is a pi bond?

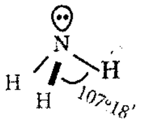
- When two atomic orbitals **overlaps sideways**, the resultant covalent bond is called a π-bond.
- When we consider x-axis as molecular axis, the $\mathbf{p_y}$**-**$\mathbf{p_y}$ and $\mathbf{p_z}$**-**$\mathbf{p_z}$ overlaps will result in the formation of a π-bond.

33. In CH_4, NH_3 and H_2O, the central atom undergoes sp^3 hybridisation - yet their bond angles are different. Why?

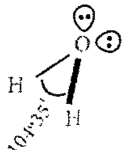
- The expected structure for all these compounds are **tetrahedron** and the bond angle is **109° 28′.**
- Due to the presence of **one lone pair of electron** on **nitrogen** atom on ammonia and the **two lone pairs of electrons** on **oxygen** atom of water the structure deviates and the bond angle varies.
- The lone pair of electrons occupy **more space** and they **repel the electrons** in the covalent bonds.



Tetrahedron
(no lone pair)



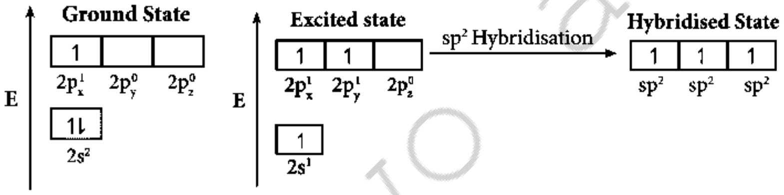
Trigonal pyramid
(one lone pair)



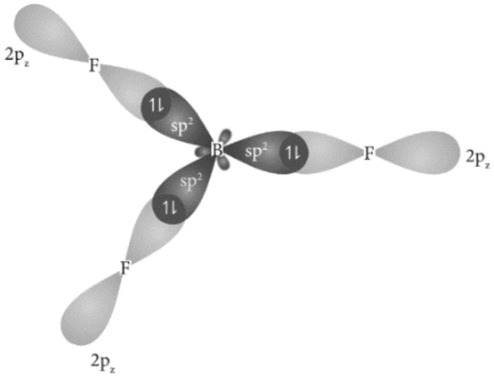
Bent 'V' shape
(two lone pair)

34. Explain sp^2 hybridisation in BF_3.

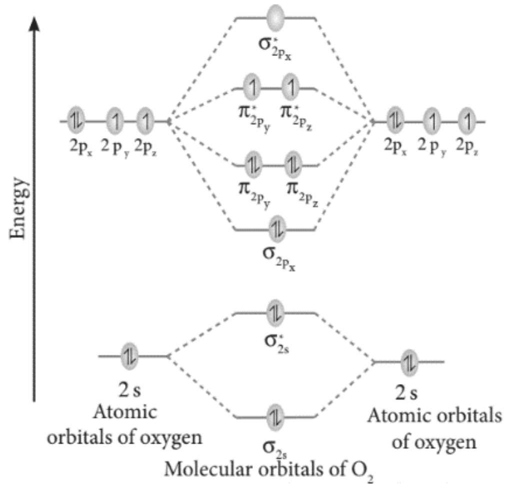
- The ground state valence shell electronic configuration of Boron atom is **[He]$2s^22p^1$**.



- In the ground state boron has only one unpaired electron in the valence shell.
- In order to form three covalent bonds with fluorine atoms, three unpaired electrons are required. To achieve this, one of the paired electrons in the 2s orbital is promoted to the $2p_x$ orbital in the excited state.
- In boron, the s orbital and two p orbitals (p_x and p_y) in the valence shell hybridises, to generate three equivalent sp^2 orbitals.
- These three orbitals lie in the same xy plane and the angle between any two orbitals is equal to 120°.
- The three sp^2 hybridised orbitals of boron now overlap with the $2p_x$ orbitals of fluorine (3 atoms). This overlap takes place along the axis:



35. Draw the M.O diagram for oxygen molecule calculate its bond order and show that O_2 is paramagnetic.



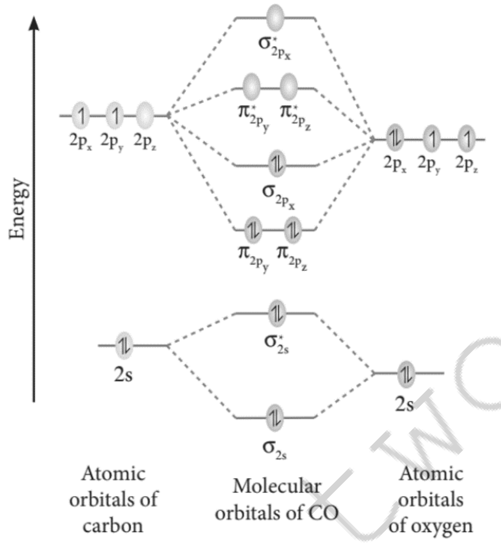
Electronic configuration of O atom is **$1s^22s^22p^4$**.

Electronic configuration of O_2 molecule is $\boldsymbol{\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \sigma_{2p_x}^2, \pi_{2p_y}^2, \pi_{2p_z}^2, \pi_{2p_y}^{*1}, \pi_{2p_z}^{*1}}$

Bond order = $\frac{N_b - N_a}{2} = \frac{10-6}{2} = 2$

Molecule has **two unpaired electrons**. Hence, it is **paramagnetic**.

36. Draw MO diagram of CO and calculate its bond order.



Electronic configuration of C atom is **$1s^22s^22p^2$**.

Electronic configuration of O atom is **$1s^22s^22p^4$**.

Electronic configuration of O_2 molecule is $\boldsymbol{\sigma_{1s}^2, \sigma_{1s}^{*2}, \sigma_{2s}^2, \sigma_{2s}^{*2}, \pi_{2p_y}^2, \pi_{2p_z}^2, \sigma_{2p_x}^2}$

Bond order = $\frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

Molecule has **no unpaired electrons**. Hence, it is **diamagnetic**.

37. What do you understand by Linear combination of atomic orbitals in MO theory?

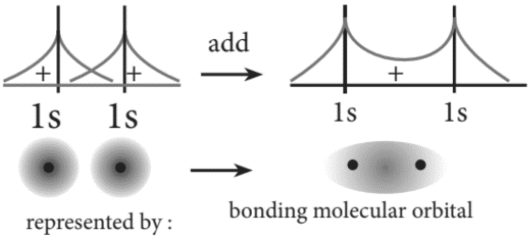
- The wave functions for the molecular orbitals can be obtained by solving **Schrodinger wave equation** for the molecule. Since solving **Schrodinger** wave equation is **too complex**, a most common method is the **linear combination of atomic orbitals** (LCAO) is used to obtain wave function for molecular orbitals.
- Atomic orbitals are represented by **wave functions ψ**. Consider two atomic orbitals represented by the wave functions ψ_A and ψ_B with comparable energy that combines to form two molecular orbitals.
- One is **bonding molecular orbital** (ψ bonding) and the other is **anti-bonding molecular orbital** (ψ anti-bonding).
- The wave function for molecular orbitals, ψ_A and ψ_B can be obtained by the LCAO as shown below:

  $\boldsymbol{\psi_{\text{bonding}} = \psi_A + \psi_B}$

  $\boldsymbol{\psi_{\text{anti-bonding}} = \psi_A - \psi_B}$
- The formation of **bonding molecular orbital** can be considered as the result of **constructive interference** of the atomic orbitals and the formation of **anti-bonding**

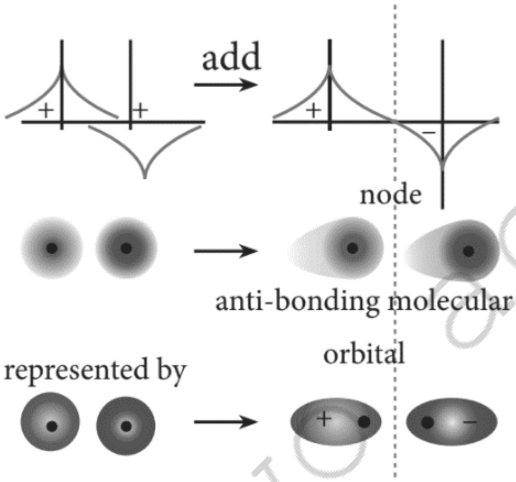
molecular orbital can be the result of the **destructive interference** of the atomic orbitals. The formation of two molecular orbitals from two 1s orbitals is show below.

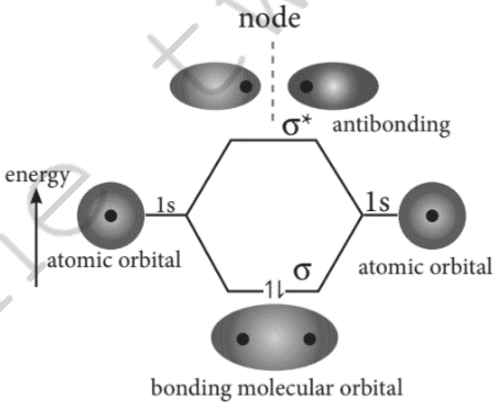
- **Constructive interaction:** The two 1s orbitals are in phase and have the same sign.



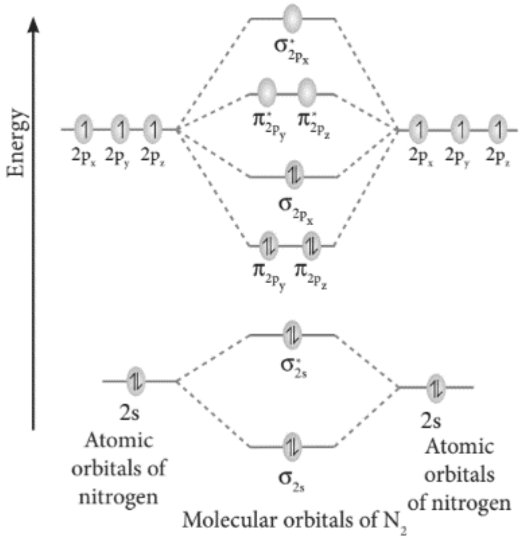
σ bonding MO

- **Destructive interaction:** The two 1s orbitals are out of phase and have opposite sign.





38. Discuss the formation of N_2 molecule using MO Theory.



Electronic configuration of N atom is **$1s^2 2s^2 2p^3$**.

Electronic configuration of N_2 molecule is $\boldsymbol{\sigma^2_{1s}, \sigma^{*2}_{1s}, \sigma^2_{2s}, \sigma^{*2}_{2s}, \pi^2_{2p_y}, \pi^2_{2p_z}, \sigma^2_{2p_x}}$

Bond order = $\frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

Molecule has **no unpaired electrons**. Hence, it is **diamagnetic**.

39. What is dipole moment?

The polarity of a covalent bond can be measured in terms of dipole moment which is defined as **u = q × 2d.**

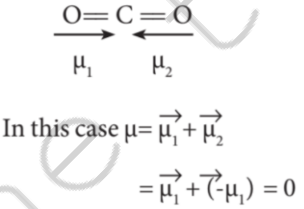
u → dipole moment

q → charge

2d → distance between two charges

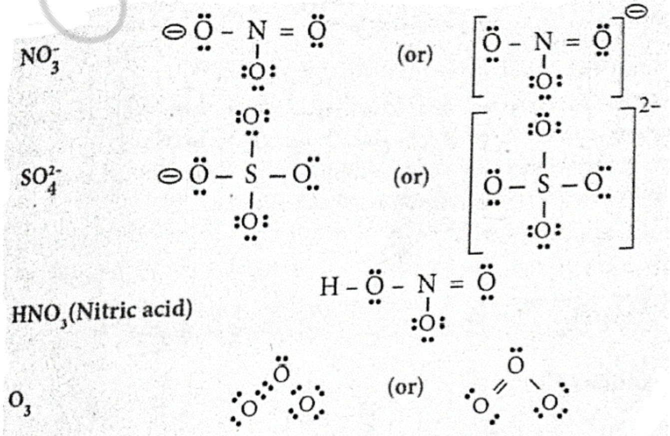
40. Linear form of carbon dioxide molecule has two polar bonds, yet the molecule has Zero dipole moment why?

- The linear form of carbon dioxide has **zero dipole moment**, even though it has two polar bonds.
- In CO_2, there are two polar bonds [C = O], which have dipole moments that are equal in magnitude but have opposite direction.
- Hence the net dipole moment of the CO_2 is **$\mu = \mu_1 + \mu_2 = \mu_1 + (-\mu_1) = 0$**



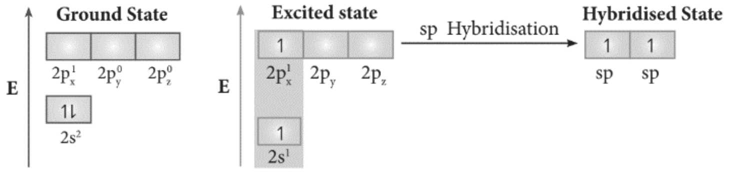
$$\text{In this case } \mu = \vec{\mu_1} + \vec{\mu_2}$$
$$= \vec{\mu_1} + \overrightarrow{(-\mu_1)} = 0$$

41. Draw the Lewis structures for the following species. i) NO_3^- ii) SO_4^{2-} iii) HNO_3 iv) O_3

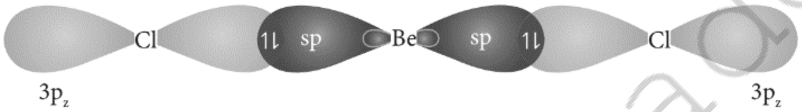


42. Explain the bond formation in $BeCl_2$ and $MgCl_2$.

- **$BeCl_2$:**
  - In $BeCl_2$ both the Be-Cl bonds are equivalent and it was observed that the molecule is **linear**.
  - It is **sp** hybridised.
  - One of the paired electron in the **2s** orbital gets excited to **2p** orbital and the electronic configuration at the **excited** state.
  - The 2s and 2p orbitals hybridise and produce two equivalent sp hybridised orbitals which have **50% s-character** and **50% p-character**.
  - These sp hybridised orbitals are oriented in **opposite** direction.



  - Each of the sp hybridised orbitals linearly overlap with **$3p_z$** orbital of chlorine to form covalent bond between Be and Cl.



- **$MgCl_2$:**
  - $$\underset{(2,8,2)}{Mg} + \underset{(2,8,7)}{Cl} + \underset{(2,8,7)}{Cl} \longrightarrow \underset{(2,8)\,(2,8,8)\,(2,8,8)}{MgCl_2}$$
  - Magnesium atom donates **one electron each to each of the chlorine atoms**. Mg atom donates 2 electrons in total.
  - Mg becomes **Mg^{2+} cation** attaining stable **Neon** configuration.
  - Two Cl atoms gain one electron each and forms Cl^- anion with stable configuration of **Argon**.
  - Mg^{2+} and $2Cl^-$ ions combine to form ionic crystal in which they are held together by **electrostatic attractive force**.

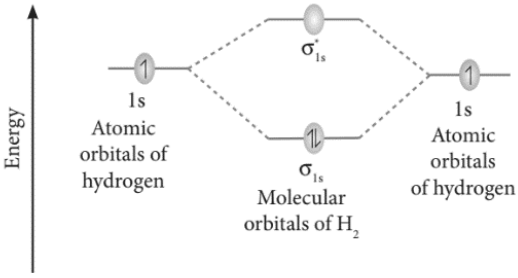
43. Which bond is stronger σ or π? Why?

σ bond is the **stronger** bond. The area of contact (overlap) between two orbitals in **π** bond is **less** whereas in **σ** bond the **overlapping is more.**

44. Define bond energy.

Minimum amount of energy required to **break one mole of a particular bond** in molecules in their **gaseous state** is called bond energy.

45. Hydrogen gas is diatomic whereas inert gases are monoatomic – explain on the basis of MO theory.



Electronic configuration of H atom is **$1s^1$.**
Electronic configuration of H_2 molecule is $\boldsymbol{\sigma_{1s}^2}$
Bond order = $\frac{N_b - N_a}{2} = \frac{2-0}{2} = 1$
Molecule has **no unpaired electrons**.
Hence, it is **diamagnetic**.

Helium molecule (i.e.) H_2:

Electronic configuration of H atom is **$1s^2$.**

Electronic configuration of H_2 molecule is $\boldsymbol{\sigma_{1s}^2 \sigma_{1s}^{*2}}$

Bond order = $\frac{N_b - N_a}{2} = \frac{2-2}{2} = 0$

46. What is Polar Covalent bond? Explain with example.

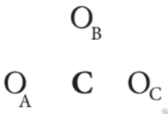
- If a covalent bound is formed between atoms having **different electro negativities**. The atom with **higher electronegativity will have greater tendency** to **attract** the shared pair of electrons towards itself than the other atom. As a result, the cloud of shared electron pair gets distorted and **polar covalent bond** is formed.
- Example – **HF** – Hydrogen fluoride:
  The electro negativities of hydrogen and fluorine on **Pauling's scale are 2.1 and 4** respectively. It means that **fluorine** attract the **shared pair of electrons** approximately twice as much as hydrogen which leads to **partial negative charge** on **fluorine** atom and **partial positive charge** on **hydrogen** atom. Hence, the **H – F bond** is said to be a **polar covalent bond**.

47. Considering x- axis as molecular axis, which out of the following will form a sigma bond.

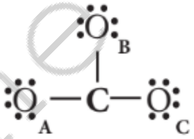
i) 1s and $2p_y$ - Sigma bond
ii) $2p_x$ and $2p_x$ - Sigma bond
iii) $2p_x$ and $2p_z$ - No sigma bond
iv) 1s and $2p_z$ - Sigma bond

48. Explain resonance with reference to carbonate ion?

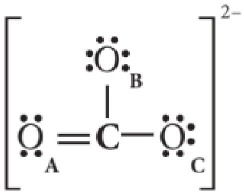
- The **skeletal structure** of carbonate ion (The oxygen atoms are denoted as O_A, O_B, O_C)



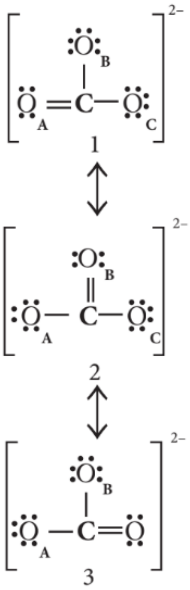
  Total number of valence electrons = [1×4(carbon)] + [3×6(oxygen)] + [2 (charge)] = **24 electrons**.
- Distribution of these valence electrons gives us the following structure:



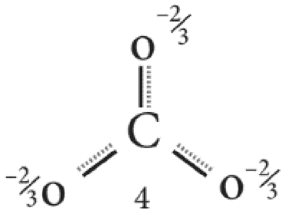
- Complete the octet for carbon by **moving a lone pair from one of the oxygens** (O_A) and write the charge of the ion (2-) on the upper right side.



- For the above structure, we can draw **two additional lewis structures** by moving the lone pairs from the other two oxygen atoms O_B and O_C. and thus creating **three similar structures** in which the relative positive of the atoms are same.



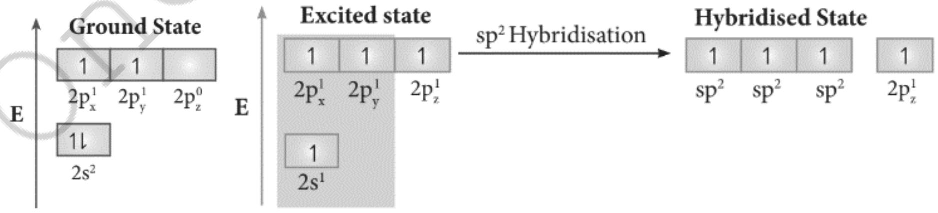
- They only **differ** in the **position of bonding and lone pair of electrons**. Such structures are called **resonance structures** and this phenomenon is called **resonance**.
- It is evident from the experimental results that all **carbon-oxygen bonds** in carbonate ion are equivalent. The actual structure of the molecule is said to be a resonance hybrid, an average of these **3 resonance forms**. The following structure gives a qualitative idea about the correct structure of CO_3^{2-} (carbonate) ion.



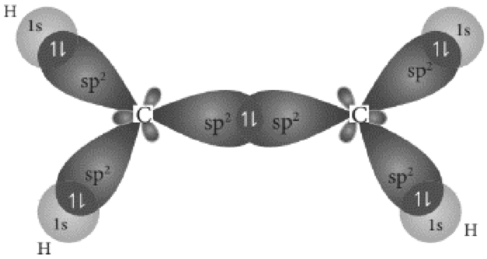
49. Explain the bond formation in ethylene and acetylene.

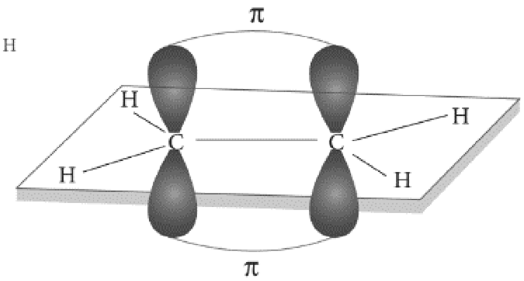
Bonding in Ethylene (C_2H_4):

- Bonding in ethylene can he explained by **hybridisation concept**.
- The **valency** of carbon is **4**. The electronic configuration of carbon is **$1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^0$**. One electron from 2s orbital is promoted to $2p_z$ orbital in the excited state to satisfy the valency of carbon.



- In ethylene both the carbon atoms undergo **sp^2 hybridisation** involving 2s, $2p_x$ and sp_y orbitals resulting in 3 equivalent sp^2 hybridised orbitals lying in the XY plane at an angle of **120^0** to each other. The unhybridised $2p_z$ orbital lies perpendicular to the xy plane.

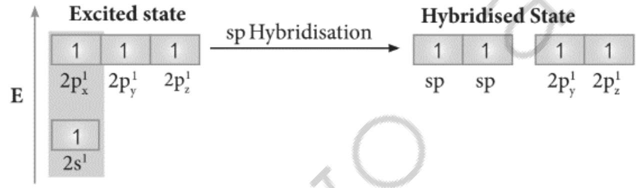


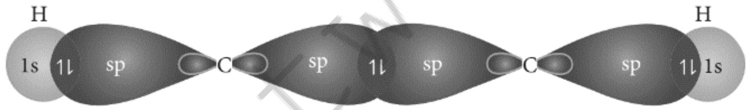


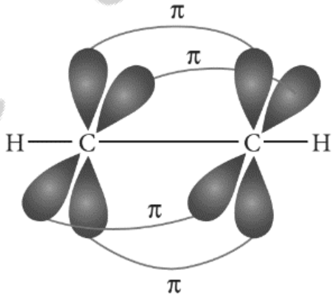
- One of the sp^2 hybridised orbitals of each carbon atoms lying along the X – axis **linearly overlaps** with each other resulting in the formation of **C – C sigma bond**. The other two sp^2 hybridised orbitals of both carbon atom **linearly overlap** with the four is orbitals of four hydrogen atoms leading to the formation of two **C – H sigma bonds** on each carbon atom.
- The unhybridised $2p_z$ orbital of both carbon atoms can **overlap only sideways** as they are not in the molecular axis. This **lateral overlap** results in the formation of a **pi bond** between the two carbon atoms.

Bonding in acetylene (C_2H_2):

- Bonding in acetylene can be explained by **hybridisation concept**.
- The electronic configuration of valence shell of carbon atom in the ground state is [He] **$2s^2$ $2p_x^1$ $2p_z^0$**. One electron from 2s orbital is promoted to $2p_z$ orbital in the excited state to satisfy the valency of carbon.







- In acetylene molecule, both the carbon atoms are in sp hybridised state. The 2s and $2p_x$ orbitals resulting in two equivalent sp hybridised orbitals are formed lying in a **straight line** along the X – axis. The unhybridised $2p_y$, and $2p_z$ orbitals lie **perpendicular** to the X-axis.
- One of the two sp hybridised orbitals of each carbon atom **linearly overlaps** with each other resulting in the formation of a **C – C sigma bond**. The other sp hybridised orbital of both carbon atoms **linearly overlap** with the two is orbitals of two hydrogen atoms leading to the formation of one **C – H sigma bond** on each carbon atom.

- The unhybridised $2p_y$ and $2p_z$ orbitals of each carbon atom **overlap sideways**. This lateral overlap results in the formation of **two pi bonds**. ($p_y – p_y$) and ($p_z – p_z$) between the two carbon atoms.

50. What type of hybridisations are possible in the following geometrics?

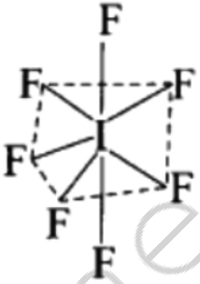
a) octahedral - sp^3d^2 or d^2sp^3
b) tetrahedral - sp^3
c) square planer - dsp^2

51. Explain VSEPR theory. Applying this theory to predict the shapes of IF_7, and SF_6.

- The **shape** of the molecules depend on the **number of valence shell electron pair around the central atom.**
- There are **two types of electron pairs** namely, **bond pairs** and **lone pairs**. The bond pair of electrons are those shared between two atoms while the lone pairs are the valence electron pairs that are not involved in bonding.
- Each pair of valence electrons around the central atom **repel each other** and hence they are located as far away as possible in **three dimensional space** to minimise the repulsion between them.
- The repulsive interaction between the different types of electron pairs is in the following order: $\mathbf{l_p\text{-}l_p > l_p\text{-}b_p > b_p\text{-}b_p}$ where l_p is lone pair and b_p is bond pair
- The **lone pair** of electrons are localised only on the **central atom and interact with only one nucleus** whereas the **bond pairs** are **shared between two atoms and they interact with two nuclei**. Because of this, the **lone pairs occupy more space** and have **greater repulsive power** than the **bond pairs** in a molecule.

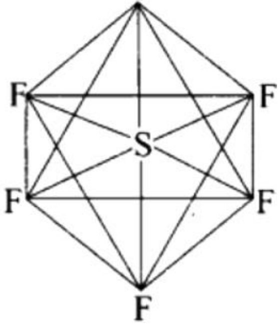
IF_7:

It is an **AB_7** type molecule. This molecule has **7 bond pairs** of electrons and **no lone pair** of electrons. Due to bond pair-bond pair interaction of electrons, IF_7 has **pentagonal bipyramidal shape.**



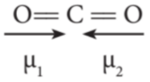
SF_6:

It is an **AB_6** type molecule. This molecule has **6 bond pairs** of electrons and **no lone pair** of electrons. Due to bond pair-bond pair interaction of electrons, SF_6 has **octahedral shape.**



52. CO_2 and H_2O both are triatomic molecule but their dipole moment values are different. Why?

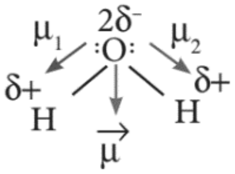
- Molecules having **polar bonds** will not necessarily have a **dipole moment.**
- For example: The linear form of carbon dioxide has **zero dipole moment**, even though it has two polar bonds.
- In CO_2, there are two polar bonds [C = O], which have dipole moments that are equal in magnitude but have opposite direction.
- Hence the net dipole moment of the CO_2 is $\mu = \mu_1 + \mu_2 = \mu_1 + (-\mu_1) = 0$

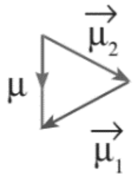


In this case $\mu = \vec{\mu_1} + \vec{\mu_2}$

$= \vec{\mu_1} + \vec{(-\mu_1)} = 0$

- In case of water, net dipole moment is the vector sum of $\mu_1 + \mu_2$





- Dipole moment in water is found to be **1.85D.**

53. Which one of the following has highest bond order?

N_2, N_2^+ or N_2^-

N_2 (14 electrons)

Bond order = $\frac{N_b - N_a}{2} = \frac{10-4}{2} = 3$

N_2^+ (13 electrons)

Bond order = $\frac{N_b - N_a}{2} = \frac{9-4}{2} = 2.5$

N_2^- (15 electrons)

Bond order = $\frac{N_b - N_a}{2} = \frac{10-5}{2} = 2.5$

So **N_2** has **highest bond order**.

54. Explain the covalent character in ionic bond.

- Ionic compounds like **lithium chloride** shows **covalent character** and it is soluble in organic solvents such as ethanol.
- The partial covalent character in ionic compounds can be explained on the basis of a phenomenon called **polarisation**.
- In an ionic compound, there is an **electrostatic attractive force** between the cation and anion. The positively charged cation attract the valence electrons of anion while **repelling the nucleus**.
- This cause a **distortion** in the electron cloud of the anion and its ele**ctron density drills** towards the cation, which results in some sharing of valence electrons between these ions. Thus, a **partial covalent character** is developed between them. This phenomenon is called **polarisation**.
- Thus due to **polarisation**, **ionic compounds shows covalent character**.

55. Describe Fajan's rule.

- The **ability of a cation to polarise** an **anion** is called its **polarising ability** and the tendency of an anion to get polarised is called its polarisibility. The extent of polarisation in an ionic compound is given by the Fajan’s rule.
- To show **greater covalent character**, both the **cation** and **anion** should have **high charge** on them. Higher the positive charge on the Cation greater will be the **attraction** on the electron cloud of the anion.
- Similarly higher the magnitude of negative charge on anion, greater is its polansability. For example, $Na^+ < Mg^{2+} < Al^{3+}$, the covalent character also follows the order – **$NaCI < MgCl_2 < AICI_3$**
- The **smaller cation** and **larger anion** show greater covalent character due to the **greater extent of polarisation**. e.g., LiCl is more covalent than NaCI.
- Cation having **$ns^2np^6nd^{10}$** configuration shows **greater polarising power than the cations with ns^2np^6** configuration. e.g., CuCI is more covalent than NaCl.